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Course Handout (Tutorials)

Intended for students of: LMD ST. Level: 1st year

Problems and Solutions in the Structure of Matter

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Summary

General Introduction

Researchers, engineers, and technicians must solve important problems, and purely theoretical and book-based knowledge does not meet the requirements of their professional practice. But not all students who are required to undergo this learning, for example, in undergraduate science programs, will become chemistry professionals.

Moreover, the problems they may be asked to solve are not of the same nature as those that arise in real life. They generally contain all the information necessary for solving them, and only that information, excluding any extraneous information or data. Therefore, the goal is not to acquire know-how that can be used in real-life situations.

Additionally, we might think of problem-solving as a means of learning the discipline in question (and therefore, for us, chemistry).

This manuscript is a Problems and Solutions in the Structure of Matter, module “Chemistry 1”. It is intended for undergraduate students in the first year of their undergraduate degree program (L1) in Science and Technology (ST). It is presented according to the common core curriculum of the Science and Technology field. It covers the structure of matter (description of the atom and chemical bonding), chemical elements, and the periodic table with energy quantification. It is reinforced with practical exercises to help students better assimilate new concepts for successful exam completion and improve their ability to solve chemistry problems.

The exercises are judiciously accompanied by detailed answer keys, allowing students to practice for the competitive exam and to self-assess.

This handout, "Problems and Solutions in the Structure of Matter ", module "Chemistry 1", includes six chapters presented according to the common core curriculum for the Science and Technology field. Each chapter includes:

• Knowledge and Skills (learning objectives), reproducing line by line the current framework. They will help you understand what the acquired knowledge can be used for and in what form.

• Exercises with detailed answers for each exercise. They will give you the opportunity to verify, apply, or deepen your knowledge of the chapter's main concepts.

I hope this manual will be of great use to all students, regardless of their study goals.

Chapter 1 :

Basic Concepts

Exercises and correction

Introduction:

Chemistry, the study of the transformation of matter, is based on a deep understanding of the internal structure of the substances around us.

This excursion into the infinite world leads us to examine atoms, these fundamental components of matter, changes in the states of matter, the concepts of atom, molecule, mole, and Avogadro's number, the atomic mass unit, atomic and molecular molar mass, molar volume, as well as the forces that bind them together to form more complex molecules and structures.

Chemistry is therefore the science that studies the nature, composition, and properties of the various substances that make up the physical universe. Each substance has a name that summarizes a set of properties of that substance that our senses reveal to us: appearance, color, hardness, specific gravity, etc. These are the properties of that substance.

Knowledge and Skills:

- ✓ Use scientific notations and determine significant figures;
- ✓ Define SI base units and convert physical quantities from one system of units to another;
- ✓ Knowing how to calculate a quantity of matter from its mass or volume;
- ✓ Appreciate significance of atomic mass, average atomic mass, molecular mass and formula mass;
- ✓ Describe the terms – mole and molar mass.

Exercise 1:

In the MKSA (IS) system, the fundamental quantities are the meter, the kilogram, the second, and the ampere, corresponding respectively to length (L), mass (M), time (T), and electrical current (I).

a) Give the dimensional equations and units in the IS (MKSA) system for:
Speed, acceleration, force, energy, pressure, electrical current, and potential difference.
The CGS system (centimeter, gram, second) is still frequently used, as its units are more suited to measuring small quantities.

b) Define the conversion factors for the IS -CGS conversion for all the above quantities. Present the results in table form.

Exercise 2:

What is the quantity of matter corresponding to 1.5 kg of sodium chloride (NaCl)?

Exercise 3:

How much matter do the following samples represent?

a) 11.2 g of iron (Fe)

b) 1.6 x 10^{-3} g of sulfur (S_8)

c) 10 g of sugar ($C_{12}H_{22}O_{11}$)

d) 1.5 x 10^{2} kg of (CaO)

e) 0.8 L of chlorine gas (Cl2) (at 0°C and 1 atm)

f) 0.02 L of liquid tetrachloromethane (CCl4) (ρ = 1.595 g.mL^{-1})

Exercise 4 :

What is the mass, in grams, of each of the following samples?

a) 0.2 mol of copper (Cu)

b) 11.2 L (0 C°, 1 atm) of methane (CH_4) gas

c) 1.4.10^{-2} mol of silica (SiO_2)

d) 4.5 mol of potassium bromide (KBr)

e) 2.103 mol of calcium sulfate ($CaSO_4$)

f) 1 m^3 (0 C°, 1 atm) of oxygen (O_2)

Exercise 5:

Calculate the number of moles and the number of atoms in the following cases:

1) An iron nail of mass 6,3 g (M_{Fe} = 56 g/mol).

2) 0,5 kg of silicon (M_{Si} = 28 g/mol).

3) 4,48 Liters of dinitrogen (M_N = 14 g/mol)

Exercise 6:

12 g of potassium hydroxide (KOH) are dissolved in 250 mL of water (ρ_{water} = 1 kg/L).

1) Calculate the moles number of dissolved KOH (M_K = 39 g/mol).

2) Calculate the molality, molarity and normality of KOH

Exercise 7:

The density of 14.5 mol.dm^{-3} nitric acid, HNO_3 is 1.42 g cm^{-3}.

- Calculate the molality of the nitric acid (HNO_3) solution.

Exercise 8:

A solution is prepared by dissolving 30 g of methanol (CH_3OH) in 70 g of water (H_2O).

- Calculate the mole fraction of methanol in the solution.

Exercise 9:

A solution is prepared by adding 2 g of a substance A to 18 g of water.

- Calculate the mass per cent of the solute.

Exercise 10:

The mass of the proton, neutron and electron are 1.6726485.10^{-24} g, 1.6749543.10^{-24} g and 9.109534×10^{-28} g, respectively.

1) Define the atomic mass unit (a.m.u) and give its value in grams (g)

2) Calculate in a.m.u the mass of the proton, neutron and electron.

3) Calculate according to Einstein's relationship (mass-energy equivalence) the energy content of 1 a.m.u expressed in MeV. (Avogadro number: 6.022045×10^{23} mol^{-1}).

Correction of Exercises (Chapter 1)

Exercise 1:

a) Units in the international system (IS) of fundamental quantities:

Base Physical Quantity	Units in the IS (MKSA)	Units in the IS (CGS)
Length (L)	metre m	cm
Mass (M)	kilogram kg	g
Time (T)	second s	s
Temperature (θ)	kelvin K	K
Amount of matter (n)	mole mol	mol
Electrical intensity (I)	ampere A	A

Where: MKSA: (meter, kilogram, second, and ampere); CGS: (centimeter, gram, second)

Dimensional equations and units in the IS-MKSA system:

Base Physical Quantity	Equation	Dimension	Units in the IS MKSA
Speed (v)	v = x/t	$L.T^{-1}$	$m\ s^{-1}$
Acceleration (a)	a= v/t	$L.T^{-2}$	$m\ s^{-2}$
Force (F)	F= a .m	$M.\ L.T^{-2}.$	$kg\ m\ s^{-2}$
Energy (E)	E= F. x	$M\ L^2.T^{-2}$	$kg\ m^2\ s^{-2}$
Pressure (P)	P= F /S	$M.L^{-1}.T^{-2}$	$kg\ m^{-1}\ s^{-2}$
Quantity of electricity (Q)	Q= I. t	I. T	A s

Where: x: distance; t: time; S: surface; h: height;

b) Units of quantities in the IS-CGS:

Grandeur	Unités dans le SI MKSA	Unités dans le SI-CGS
Vitesse (v)	$m\ s^{-1}$	$10^2\ cm\ s^{-1}$
Accélération (a)	$m\ s^{-2}$	$10^2\ cm\ s^{-2}$
Force (F)	$kg\ m\ s^{-2}$	$10^3\ g\ 10^2\ cm\ s^{-2} = 10^5\ g\ cm\ s^{-2}$
Energie (E)	$kg\ m^2\ s^{-2}$	$10^3\ g\ (10^2)^2\ cm^2\ s^{-2} = 10^7\ g\ cm^2\ s^{-2}$

Pression (P)	kg m^{-1} s^{-2}	10^3 g 10^{-2} cm s^{-2} = 10 g cm^{-1} s^{-2}
Quantité d'électricité (Q)	A s	A s

Exercise 2:

The molar mass of NaCl is 58.5 g/mol, and 1.5 kg (1500 g)

therefore n = m/M

n = 1500 g / 58.5 g/mol = 25.64 mol of this compound.

Exercise 3:

n = m/M

a) n = m/M = 0,20 mol
b) n = m/M = 1,6.10^{-3}/ (32,07 x8) = 6,24. 10^{-6} mol
c) n = m/M = 2,92. 10^{-2} mol
d) n = m/M = 2,67. 10^{3} mol
e) n = m/M = 3,57. 10^{-2} mol
f) n = m/M = 2,07. 10^{-1} mol

Exercise 4:

a) m = n x M = 12,71 g
b) m = M.P.V/R.T = 8,015 g
c) m = n x M = 0,84 g
d) m = n x M = 535,55 g
e) m = n x M = 286,32 g
f) m = M.P.V/R.T = 1427,68 g

Exercise 5:

The number of moles (n) and the number of atoms (N):

1) An iron nail of mass 6.3 g (M_{Fe} = 56 g/mol).

n = m / M_{fe} = 6,3 / 56= 0,1125 moles.

N = n × N_A = 0,1125 × 6,023 ×10^{23} = 0,677875 ×10^{23} atoms

N_A : Avogadro Number: 6.023×10^{23}.

2) 0,5 kg of silicon (M_{Si} = 28 g/mol):

n = m / M_{Si} = 0,5 / 28 = 0,0178 moles.

N = n × N_A = 0, 0178 × 6,023 ×10^{23} = 0, 1072094 ×10^{23} atoms

3) 4,48 Liters of dinitrogen (M_N = 14 g/mol):

n = V/ V_m = 4, 48 / 22, 4 = 0, 2 mols.

N = n × N_A = 0, 2 × 6,023 ×10^{23} = 1, 2046 ×10^{23} atoms

V_m: molar volume.

Exercise 6:

The moles number of dissolved KOH (M_K = 39 g/mol).

n= m / M_{KOH} = 12 / (39 + 16 + 1) = 0, 21 mols.

Molality (C_m) :

$$\boldsymbol{C_m} = \frac{\boldsymbol{n}(\boldsymbol{solute})}{\boldsymbol{m}(\boldsymbol{solvent})} = \frac{0{,}21}{250 \times 10^{-3}} = 0{,}85\,\frac{mol}{Kg}\ , where: \boldsymbol{\rho} = \frac{\boldsymbol{m}}{\boldsymbol{V}}$$

Molarity (C_M) :

$$\boldsymbol{C_M} = \frac{\boldsymbol{n}(\boldsymbol{solute})}{\boldsymbol{V}(\boldsymbol{solution})} = \frac{0{,}21}{250 \times 10^{-3}} = 0{,}85\ mol/L$$

Normality (C_N) :

$$\boldsymbol{C_N} = \boldsymbol{C_M} \times \boldsymbol{n_{eq}}$$

The number n_{eq} of KOH:

KOH → K^+ + OH^-

n_{eq} = 1

So: C_N = C_M = 0,85 N

Exercise 7:

$$\text{Molality } = \frac{n\ of\ moles\ of\ solute}{mass\ of\ solvent\ (Kg)}$$

$$C_m = C_M \times M$$

$$C_m = 14{,}5\ \frac{mol}{l} \times 63\ \frac{g}{mol} = 913{,}5\ g/mol$$

913,5 g of solute → in 1L (1000mL)

Masse of solution 1,42 g → in 1 mL

1420 g → in 1000 mL

$m_{solution} = m_{solute} + m_{solvent}$

$m_{solvent} = m_{solution} - m_{solute}$ = 1420 - 913,5 = 506, 5g = 0;5065 Kg

So: Molality of HNO_3 = n/ $m_{of\ solvent}$

Exercise 8:

Calculation of the mole fraction of methanol:

$X_i = n_i / n_1 + n_2$

$x_{CH3OH} = n_{CH3OH} / n_{CH3OH} + n_{H2O}$

n = m/ M

x_{CH3OH} = 0,12

Exercise 9:

Mass per cent of $\boldsymbol{A = \frac{Mass\ of\ A}{Mass\ of\ solution} \times 100}$

$$= \frac{2g}{2g\ of\ A + 18g\ of\ water} \times 100$$

$$= \frac{2g}{20g} \times 100$$

$$= 10\%$$

Exercise 10:

1) Definition of the atomic mass unit (a.m.u):

1 a.m.u = 1/12 x (12 / NA) = 1/ N = $1,66030217.10^{-24}$g = $1,66030217.10^{-27}$ Kg

2) Calculation in a.m.u the mass of the proton, neutron and electron

1 a.m.u → $1,66030217.10^{-24}$g

mp →$1,67.10^{-24}$g

mp = $1,67.10^{-24}/1,66030217.10^{-24}$ = 1,007277 u.m.a.

mn = 1,008665 u.m.a.

me = 0,000549 u.m.a.

3) The energy content of 1 a.m.u expressed in MeV:

Einstein's relationship: E (1 u.m.a) = Δmc^2

E = $1, 66030217.10^{-24}.10^{-3}$ x $(3.108)^2$

Conclusion:

Since the study of chemistry touches on all aspects of life, it is crucial. Chemists investigate the characteristics, composition, and transformations of substances. Matter, which can be solid, liquid, or gas, is a component of all substances. In states of matter, the constituent particles are retained in various ways and display their distinctive characteristics. Elements, compounds, and mixes are more classifications for matter. Mixtures are common, and a lot of the materials we encounter are mixtures. An element is made up of only one kind of particle, which could be an atom or a molecule.

Chaptre 2 :

Main constituents of matter

Exercises and correction

Introduction:

The atom is the smallest entity of matter that is preserved during a chemical phenomenon. The word “atom” was already used in ancient Greece.

Atoms are the first differentiated corpuscles of matter. They consist of a nucleus made up of nucleons (protons and neutrons) and an electronic procession made up of electrons. It is designated by its name and symbol. An atom being electrically neutral, the number of electrons is equal to the number of protons.

Knowledge and Skills:

- ✓ Determine the number of protons, neutrons, and electrons in an atom or ion, given the numbers A and Z;
- ✓ Calculate the apparent atomic mass of a natural element based on its isotopic composition and vice versa;
- ✓ Deduce and calculate the isotopic composition of an element, given its apparent atomic mass and the atomic masses of its isotopes;
- ✓ Calculate a molecular mass;
- ✓ Calculate a binding energy and cohesion energy (in J or MeV) from a mass defect.

Exercise 1:

What words are replaced by the letters **(a), (b), … (m)** in the following text?

- The mass number A is the number of ……… (a), that is, the sum of the numbers of ……… (b) and ……… (c).
- The atomic number Z is the number of …….. (d), equal to the number of ……… (e) in an atom, but different from the latter in a …… (f).
- N (equal to A-Z) is the number of …….. (g).
- A ……… (h) is defined by the value of A and Z.
- A ……… (i) is defined by the value of Z.
- Two ……. (j) are two ……… (k) that have the same value of Z, and therefore belong to the same …… (l), but have different values of N and …….. (m).

Exercise 2:

Which of the following ions have the same number of electrons as the sulfur atom $^{32}_{16}S$ **?**

a) $^{35}_{17}Cl^{+}$

b) $^{34}_{16}S^{+}$

c) $^{40}_{18}Ar^{2+}$

d) $^{35}_{16}S^{2-}$

Exercise 3:

Determine the number of neutrons, protons and electrons present in each of the following atoms or ions:

$^{19}_{9}F$, $^{40}_{20}Ca$, $^{24}_{12}Mg^{2+}$, $^{52}_{24}Cr^{3+}$, $^{127}_{53}I^{-}$, $^{108}_{47}Ag^{+}$, $^{79}_{34}Se^{2-}$, $^{222}_{86}Rn$, $^{238}_{92}U$

Exercise 4:

Find the missing data in the following table.

Élément	Symbole	Protons	Neutrons	Electrons	Z	A
a) Sodium	-	-	12	-	-	-
b) -	^{40}K	-	-	-	-	-
c) Silicium	-	-	14	-	-	28
d) Rubidium	-	37	-	-	-	85
e) Arsenic	-	-	42	33	-	-
f) -	-	-	-	-	47	108
g) -	-	53	74	-	-	-
h) -	Au	79	-	-	-	197
i) -	-	-	138	-	88	-
j) Cadmium	-	-	64	46	-	-
k) -	-	26	-	23	-	-
l) -	Se^{2-}	-	-	-	34	79
m) Chlore	-	-	-	18	-	35

Exercise 5:

I) determine the number of protons, neutrons and electrons of the following atoms or ions:

$$^{10}_{5}B\ ;\ ^{24}_{12}Mg\ ;\ ^{35}_{17}Cl^{-}\ ;\ ^{56}_{26}Fe\ ;\ ^{63}_{29}Cu\ ;\ ^{63}_{29}Cu^{2+}$$

II) Natural boron (B) has an atomic molar mass of 10.811 g/mol. Knowing that it is composed of ^{10}B and ^{11}B with relative atomic masses equal to 10.0124 g/mol and 11.00931 g/mol, respectively. Determine its isotopic composition (percentage abundance of each isotope).

III) The elements below are found in the periodic table.

$$^{20}_{9}F\ ;\ ^{3}_{1}H\ ;\ ^{235}_{92}U\ .$$

1- What is the constitution of each nucleus?
2- What is the corresponding mass fault? Deduce the binding energy per nucleon.
3- Compare the nucleus stabilities of these elements.

Data:

m_p = 1.00758 u , m_n = 1.00866 u

^{20}F = 20.00165 u ; ^{3}H = 3.0165 u ; ^{235}U = 235.12 u

Exercise 6:

A sample of magnesium (Mg) is analyzed using a Bainbridge spectrograph. The monatomic ions carrying two elementary charges enter the analyzer through a slot F at a velocity of 10^{5} m/s and under the influence of the magnetic field of 1 Tesla. Three spots S_1, S_2 and S_3 are observed on the ion detector, the characteristics of which are summarized in the following table:

Spot number	S_1	S_2	S_3
Number of ions detected per second	1572	202	226
Distance between slot and spot: d (cm)	2.5	2.6	2.7

Knowing that the intensities of the spots are proportional to the number of ions detected per second, determine:

1. The number of isotopes of natural magnesium.

2. The mass in m.a.u. of the lightest isotope.

3. The percentage abundance of each isotope and the average atomic mass of Mg in m.a.u.

Exercise 7:

The wavelength of sodium vapor is equal to 5900A°, the speed of light c: 3.10^{8} m/s; Planck's constant h = 6.6210^{-34} J.s. Calculate:

1) The associated wavenumber in cm^{-1}.

2) The frequency as well as the period of the wave.

3) The energy of the photons emitted.

Exercise 8:

When electromagnetic radiation of wavelength 300 nm falls on the surface of sodium, electrons are emitted with a kinetic energy of 1.68 ×105 J mol^{-1}.

1) What is the minimum energy needed to remove an electron from sodium?

2) What is the maximum wavelength that will cause a photoelectron to be emitted?

Exercise 9:

For a second line in the Balmer series of the hydrogen spectrum, calculate:

1) The wavelength (in nm).

2) The frequency.

3) The energy.

Exercise 10:

A photon of light with a wavelength of 2166 nm is produced in the Brackett series when an electron makes a transition from higher energy level to lower energy level. Calculate:

1) Calculate the energy released (in J).

2) Indicate the transition of electron.

Exercise corrections (Chapter 2)

Exercise 1:

(a): nucleons

(b): protons

(c): neutrons

(d): protons

(e): electrons

(f): ion

(g): neutrons

(h): nuclide

(i): element

(j): isotopes

(k): nuclides

(l) element

(m): A.

Exercise 2:

a) and **c)**, They have the same number of electrons as the sulfur atom $^{32}_{16}S$

Exercise 3:

Atoms/ ions	neutrons	protons	electrons
$^{19}_{9}F$	**10**	**9**	**9**
$^{40}_{20}Ca$	**20**	**20**	**20**

$^{24}_{12}Mg^{2+}$	**12**	**12**	**10**
$^{52}_{24}Cr^{3+}$	**28**	**24**	**21**
$^{127}_{53}I^{-}$	**74**	**53**	**54**
$^{108}_{47}Ag^{+}$	**61**	**47**	**46**
$^{79}_{34}Se^{2-}$	**45**	**34**	**36**
$^{222}_{86}Rn$	**136**	**86**	**86**
$^{238}_{92}U$	**146**	**92**	**92**

Exercise 4:

Missing data are given in the order they appear on each line:

a) Na - 11- 11 - 11 - 23

b) Potassium - 19 - 21 - 19 - 19 - 40

c) Si - 14 - 14 – 14

d) Rb - 48 - 37 - 37

e) As - 33 - 33 – 75

f) Argent - Ag - 47 - 61 – 47

g) Iode - I - 53 - 53 - 127.

h) Or - 118 - 79 – 79

i) Radium - Ra - 88 - 88 – 226

j) Cd^{2+} - 48 - 48 – 112

k) Fer – Fe^{3+} - 30 - 26 – 56

l) Selenium - 34 - 45 – 36

m) Cl^{-} - 17 - 18 - 17

Exercise 5:

I) determine the number of protons, neutrons and electrons of the following atoms or ions:

Atoms or ions	protons	neutrons	electrons
$^{10}_{5}B$	**5**	**5**	**5**
$^{24}_{12}Mg$	**12**	**12**	**12**
$^{35}_{17}Cl^{-}$;	**17**	**18**	**18**
$^{56}_{26}Fe$	**26**	**30**	**26**

$^{63}_{29}Cu$	**29**	**34**	**29**
$^{63}_{29}Cu^{2+}$	**29**	**34**	**27**

II) The isotopic composition (percentage abundance of each isotope):

$$M_{moy} = \frac{\sum Mi\ Xi}{100} = \frac{M_1 M_1 + M_2 M_2}{100}$$

Withe : $x_1 + x_2 = 100$ or

$M = M_1\ x_1 + M_2\ x_2$

Withe : $x_1 + x_2 = 1$

If: $M = M_1\ x_1 + M_2\ x_2$

$10,811 = 10,0124\ x_1 + 11,00931\ x_2$

$x_1 + x_2 = 1 \Rightarrow x_1 = 1 - x_2$

x_2 = 0,80% and x_1 = 0,20%

The natural boron is composed of 80% of the ^{11}B isotope and 20% of the ^{10}B isotope.

III)

1) The constitution of each nucleus

nucleus	protons	neutrons
$^{3}_{1}H$	**1**	**2**
$^{20}_{9}F$	**9**	**11**
$^{235}_{92}U$	**92**	**143**

2) the mass fault and the binding energy per nucleon:
   - for $^{235}_{92}U$

$\Delta m = m_{the} - m_{exp}$

m_{the} = P.m_P = N. m_N = 236,93 a.m.u

Δm = 236,93 – 235, 12 = **1,81574 a.m.u**

energy per nucleon:

$E_l = E/A$

1 uma → 933,75 MeV

1,81574 a.m.u → E

E = 1695,90116 MeV

E_l = E/A= 7,2166 MeV/ nucleon

➢ **For** $^{20}_{9}F$

$\mathbf{m_{the} = P.m_P = N.\ m_N}$ = 20,16345 a.m.u

$\mathbf{\Delta m = m_{the} - m_{exp} = 0,16183}$ a.m.u

1 uma → 933,75 MeV

0,16183 a.m.u → E

E = 151,14922 MeV

$\mathbf{E_l = E/A = 7,5574}$ **MeV/ nucleon**

➢ **For** $^{3}_{1}H$

$\mathbf{m_{the} = P.m_P = N.\ m_N}$ = 3,0249 a.m.u

$\mathbf{\Delta m = m_{the} - m_{exp} = 0,0084}$ a.m.u

1 uma → 933,75 MeV

0,0084 a.m.u → E

E = 7,8456 MeV

$\mathbf{E_l = E/A = 2,6152}$ **MeV/ nucleon**

3) The stabilities of these elements.

ST (F) > ST (U) > ST (H)

Exercise 6:

1. The number of isotopes of natural magnesium.

Three spots: T1, T2, and T3, are therefore 3 isotopes of Mg.

2. The mass in a.u.m. of the lightest isotope.

analyseur:

$$\sum \vec{F} = m\vec{a} \begin{cases} a_t = \frac{dV}{dt} = 0 (V = cte) \\ a_N = \frac{\vartheta^2}{r} \end{cases}$$

Magnetic force: $\mathbf{F_m}$

$F_m = m \cdot a_N$

$q \cdot V \cdot B = m V^2/r$

$m = (q \cdot B \cdot r)/V = (q \cdot B \cdot d)/2V$

The lightest isotope corresponds to the smallest radius

SO: $m_1 = \frac{q \times B \times d_1}{2 \times V}$

For Mg^{2+} therefore q= 2 e

$m_1 = \frac{2\times 1{,}6\ 10^{-19}\times 1\times 2{,}5\ 10^{-2}}{2\times\ 10^{5}}$= 4 10^{-26} g/ (1,66 10^{-24})= **24,0964 a.m.u**

q , B et v = cte so:

$$\frac{m_1}{r1} = \frac{m_2}{r2} = \frac{m_3}{r3}$$

$$m_2 = m_1 \frac{r2}{r1} = m_1 \frac{d2}{d1}$$

m_2= 24.0964x 2,6/2,5= 25,0602 a.m.u

$$m_3 = m_1 \frac{r3}{r1} = m_1 \frac{d3}{d1}$$

M_3= 24.0964 x 2,7 / 2,5= 26,0421 a.m.u

3. The percentage abundance of each isotope and the average atomic mass of Mg in m.a.u.

$$\boldsymbol{M = \frac{\sum M_i x_i}{100}}$$

The total number:

n_t = 1572+202+226 = 2000

X_1=1572 x 100/2000 = 78,6 %

X_2= 10,1 %

X_3= 11,3 %

M= [(78,6x24.0964) + (10,1 x25,0602)+ (11,3 x26,0421)]/100 = 24,414 u.m.a.

Exercise 7:

1) The wavenumber in cm^{-1}

$\sigma = 1/\lambda$ (cm^{-1}) = $1/\ 5900.10^{-8}$ $= 1{,}69.10^{4}\ J.s$

2) The frequency

$\nu = c / \lambda = 3.10^{8} / 5900.10^{-10} = 5{,}08.\ 10^{14}\ S^{-1}$ or Hz

The period:

$T = 1/\nu = 1/\ 5{,}08\ .10^{14} = 1{,}97\ .\ 10^{-15}$ S

3) The energy of the photons emitted

$E = h\,\nu = 6.626 \times 10^{-34} \times 5{,}08.\ 10^{14} = 3.36 \times 10^{-19}$ J

Exercise 8:

The energy (E) of a 300 nm photon is given by

$$\boldsymbol{E_n = \frac{hc}{\lambda}}$$

$$= \frac{6.626 \times 10^{-34} Js \times 3 \times 10^{8} ms^{-1}}{300 \times 10^{-9} m}$$

$$= 6.626 \times 10^{-19} J$$

The energy of one mole of photons = $6.626 \cdot 10^{-19}$ J · $6.022 \cdot 10^{23}$ mol^{-1} = $3.99 \cdot 10^{5}$ J mol^{-1}
The minimum energy needed to remove one mole of electrons from sodium

= (3.99 –1.68) 10^{5} J mol^{-1} = $2.31 \cdot 10^{5}$ J mol^{-1}
The minimum energy for one electron:

$$= \frac{2.31 \times 10^{5} Jmol^{-1}}{6.022 \times 10^{23} electrons\ mol^{-1}}$$

$$= 3.84 \times 10^{-19} J$$

This corresponds to the wavelength:

$$\boldsymbol{\lambda = \frac{hc}{E}}$$

$$= \frac{6.626 \times 10^{-34} Js \times 3 \times 10^{8} ms^{-1}}{3.84 \times 10^{-19} J}$$

$$\boldsymbol{= 517\ nm}$$

(This corresponds to green light)

Exercise 9:

For a second line of Balmer series is produced when electron falls from n= 4 to n= 2

1) The wavelength (in nm)

$$\frac{1}{\lambda} = R_H \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$$

$1/\lambda = R_H (1/2^2 - 1/4^2) = 1,097 . 10^7 . (1/2^2 - 1/4^2)$

$\lambda = 4,8617 \cdot 10^{-7}$ m

2) The frequency.

$\nu = c / \lambda = 3.10^{8} / \ 4,8617 \cdot 10^{-7}$

$\nu = 6,17 \cdot 10^{14}\ S^{-1}$

3) The energy.

$E = h\nu = 6.626 \times 10^{-34} \times 6,17 \cdot 10^{14} = 4,084 \times 10^{-19}$ J

Emission $\Rightarrow E<0 \Rightarrow E = -4,084 \times 10^{-19}$ J

Exercise 10:

1) Calculation of the energy released (in J)

$$\Delta \boldsymbol{E} = \frac{\boldsymbol{hc}}{\boldsymbol{\lambda}}$$

$= 6.626 \times 10^{-34} \times 3.10^{8} / 2166 \times 10^{-9} = 9,1689 \times 10^{-20}$ J

Emission $\Rightarrow E<0 \Rightarrow$ **$E = -9,1689 \times 10^{-20}$ J**

2) Indicate the transition of electron.

$$\frac{1}{\lambda} = R_H \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$1/\lambda = R_H\ (1/4^2 - 1/x^2) \Rightarrow$

X= 7 $\Rightarrow$ electron falls from: **n= 7 to n=4**

Conclusion:

The fundamental units of all elements are atoms. They are an element's smallest chemically reactive components. According to John Dalton's 1808 initial atomic hypothesis, the atom is the ultimate, indivisible particle of matter. Atoms are made up of three basic particles: protons, neutrons, and electrons. This fact was empirically demonstrated by the end of the nineteenth century. Several atomic models were proposed to describe the structure of the atom after the discovery of subatomic particles. Rutherford came to the conclusion that an atom is made up of a small, positively charged nucleus at its core, encircled by electrons that orbit in circles.

Chapter 3 :

Radioactivity – Nuclear reactions

Exercises and correction

Introduction:

Radioactivity, property exhibited by certain types of matter of emitting energy and subatomic particles spontaneously. It is, in essence, an attribute of individual atomic nuclei.

If the atom is unstable, the nucleus will transform over time to acquire a more stable state. This transformation is called disintegration. This disintegration is generally accompanied by radiation emissions, a process called radioactivity.

The transformation of one nucleus into another nucleus is called radioactive transmutation. There are two types of transmutation:

- Spontaneous: natural radioactive disintegration (radioactivity manifested by natural nuclides).
- Induced: artificial radioactivity.

- Natural radioactivity is that which exists naturally in nature.

- Artificial radioactivity is that obtained by bombarding atomic nuclei with particles (neutrons, protons, α-particles, electrons, positrons, etc.).

The nuclear fission: under the effect of neutron bombardment, a heavy nucleus can break into two lighter nuclei of comparable masses, with the emission of neutrons. The nuclear fusion: this type of reaction is based on the fusion of light nuclides to form heavier nuclides (target nucleus, $Z = 4$), with the release of a considerable amount of energy.

Knowledge and Skills:

- ✓ Describe the phenomenon of radioactivity;
- ✓ Write the equation for a nuclear reaction;
- ✓ Apply the principles of nuclear fission and fusion;
- ✓ Write a stoichiometric equation for a nuclear reaction;
- ✓ Calculate the energy exchanged during a nuclear reaction;

Exercise 1:

State whether the following statements are correct:

a) Natural radioactivity produces protons and neutrons.

b) Natural radioactivity is always accompanied by the emission of gamma radiation.

c) Natural radioactivity only produces fission products.

d) A nuclear reaction can be triggered by bombarding a target nucleus with neutrons.

Exercise 2:

Complete the following stoichiometric equations:

a) ${}^{58}_{26}\text{Fe} + 2{}^{1}_{0}n \rightarrow {}^{60}_{27}\text{Co} + \cdots$

b) ${}^{40}_{20}\text{Ca} + {}^{2}_{1}d \rightarrow \ldots + {}^{1}_{1}P$

c) ${}^{60}_{28}\text{Ni} + \cdots \rightarrow {}^{60}_{27}\text{Co} + {}^{1}_{1}P$

d) ${}^{242}_{96}\text{Cm} + {}^{4}_{2}\alpha \rightarrow \ldots + {}^{1}_{0}n$

e) ${}^{30}_{15}\text{P} \rightarrow \ldots + {}^{0}_{+1}e$

f) ${}^{235}_{92}\text{U} + {}^{1}_{0}n \rightarrow {}^{142}_{56}\text{Ba} + \cdots + 2{}^{1}_{0}n$

g) ${}^{35}_{92}\text{Cl} + n \rightarrow {}^{35}\text{S} + \ldots$

h) ${}^{209}\text{Bi} + \ldots \rightarrow {}^{210}\text{Bi} + \text{P}$

i) ${}^{58}\text{F} + \text{n} \rightarrow {}^{59}\text{Co} + \ldots$

j) ${}^{59}\text{Co} + \ldots \rightarrow {}^{56}\text{Mn} + \alpha$

k) ${}^{214}\text{Pb} \rightarrow {}^{214}\text{Bi} + \ldots.$

Exercise 3:

What nuclide is formed in the following reactions?

a) ${}^{9}\text{Be}$ (d,n)

b) ${}^{25}\text{Mg}$ (α,p)

c) ${}^{63}\text{Cu}$ (p,n)

d) ${}^{14}\text{N}$ (n,p)

e) ${}^{241}\text{Am}$ (α,2n)

f) ${}^{19}\text{F}$ (p, α)

Exercise 4:

Explain, using a complete stoichiometric equation, the meaning of the following expressions:

1) $^{27}Al(\alpha, n)^{30}P$

2) $^{6}Li(d, p)^{7}Li$

3) $^{23}Na(p, n)^{23}Mg$

4) $^{15}N(p, \alpha)^{12}C$

Exercise 5:

Complete the following nuclear reactions. For each equation, indicate the type of reaction:

1) $^{131}_{53}I \rightarrow \ ^{131}_{52}Te + \cdots$
2) $^{124}_{53}I \rightarrow \cdots + \ \beta^-$
3) $^{3}_{1}H \ + \ ^{2}_{1}H \ \rightarrow \ ^{1}_{0}n + \cdots$
4) $^{14}_{7}N \ + \ ^{4}_{2}He \ \rightarrow \ ^{16}_{8}O + \cdots$
5) $^{215}_{84}Po \rightarrow \ ^{211}_{82}Pb + \cdots$
6) $^{1}_{0}n + \ ^{235}_{92}U \ \rightarrow \cdots + \ ^{139}_{53}I + \ 3\,^{1}_{0}n$
7) $^{9}_{4}Be \ (\beta^+, \alpha)$ ...

Exercise 6:

The radioactive isotope of Cobalt $^{60}_{27}Co$ decays by emitting β^- radiation.

At the initial instant, its activity is 7,805 .10^{-6} Ci. After two years, it is 6. 10^{-6} Ci.

a) Write the nuclear decay reaction, specifying the composition of the nucleus formed.
b) Determine the radioactive constant λ of Cobalt.
c) Calculate the mass of Cobalt, undisintegrated, after two years of activity.

Data: 1 year = 365 days

Mass of Cobalt = 59.934

Avogadro's number = = 6,023. 10^{23}

1Ci (Curie) = 3,7. 10^{10} d.p.s

Exercise 7:

When the isotope $^{27}_{13}Al$ is bombarded by α particles, the isotope $^{30}_{15}P$ is produced.

a) Write the equation for the corresponding nuclear reaction.
b) The isotope $^{30}_{15}P$ is also unstable and emits β^+ positrons. How many new elements are formed?
c) The half-life of phosphorus $^{30}_{15}P$ is 3 minutes. How many β^+ particles will be emitted after 12 minutes per 16 grams of phosphorus?

Exercise 8:

What is the time required to decay:

1- 80% of the atoms of radioactive material if T = 35 years
2- 10% of a radioactive element if its half-life = 20 minutes

If the initial number of nuclei is 100%,

3- what is the number of atoms (X%) that decays in 3 revolutions?

Given that 90% of the isotope $^{9}_{19}F$ decays radioactively in 366 minutes,

4- what is its half-life?

Exercise corrections (Chapter 3)

Exercise 1:

The correct statements are: b, c, and d.

Natural radioactivity produces electrons, γ particles (helium nucleus), and energetic radiation.

Exercise 2:

Particle	symbols
Neutron	${}_{0}^{1}n$, or n
Proton	${}_{1}^{1}P$, or p, or ${}_{1}^{1}H$
Deuteron (deuterium nucleus)	${}_{1}^{2}d$, or d, or ${}_{1}^{2}H$
Particle α (helium nucleus)	${}_{1}^{1}\alpha$, or α, or ${}_{2}^{4}H$
Electron	${}_{-1}^{0}e$, or e^{-1} , or β^{-}
Positron (positive electron)	${}_{+1}^{0}e$, or e^{+}, or β^{+}

For mass number A, the sum on the first member is 58 + (2 X 1) = 60; the term to be identified therefore has zero mass (60 – 60 = 0).

For Z, the sum on the first member is 26 + (2 X 0) = 26; the particle to be identified must therefore have a charge of –1 (26 – 27 = –1). This is an electron.

a) e–

b) ^{41}Ca

c) n

d) ^{245}Cm

e) ^{30}Si

f) ^{92}Kr

g) p

h) d

i) e–

j) n

k) e–.

Exercise 3:

The symbolic representation used can be translated as:

${}^{9}_{4}\text{Be} + {}^{2}_{1}d \quad \rightarrow \; \dots . + {}^{1}_{0}n$

And we then easily find that this reaction gives:

a) ^{10}B

b) ^{28}Al

c) ^{63}Zn

d) ^{14}C

e) ^{243}Bk

f) ^{16}O

Exercise 4:

1) ${}^{27}_{13}Al + {}^{4}_{2}He \rightarrow {}^{30}_{15}P + {}^{1}_{0}n$
2) ${}^{6}_{3}Li + {}^{2}_{1}H \rightarrow {}^{7}_{3}Li + {}^{1}_{1}H$
3) ${}^{23}_{11}Na + {}^{1}_{1}H \rightarrow {}^{23}_{12}Mg + {}^{1}_{0}n$
4) ${}^{15}_{7}N + {}^{1}_{1}H \rightarrow {}^{12}_{6}C + {}^{4}_{2}He$

Exercise 5:

1) ${}^{131}_{53}I \rightarrow {}^{131}_{52}Te + {}^{0}_{1}e$ It is a disintegration β^{+}
2) ${}^{124}_{53}I \rightarrow {}^{124}_{54}I + {}^{0}_{-1}e$ It is a disintegration β^{-}
3) ${}^{3}_{1}H + {}^{2}_{1}H \rightarrow {}^{1}_{0}n + {}^{4}_{2}He$ This is a fusion reaction
4) ${}^{14}_{7}N + {}^{4}_{2}He \rightarrow {}^{16}_{8}O + {}^{1}_{1}H$ This is a transmutation reaction
5) ${}^{215}_{84}Po \rightarrow {}^{211}_{82}Pb + {}^{4}_{2}He$ It is a disintegration α
6) ${}^{1}_{0}n + {}^{235}_{92}U \rightarrow {}^{94}_{39}Y + {}^{139}_{53}I + 3{}^{1}_{0}n$ This is a fission reaction
7) ${}^{9}_{4}Be + {}^{0}_{1}e \rightarrow {}^{4}_{2}He + {}^{5}_{3}Li$ This is a transmutation reaction

Exercise 6:

a)

$^{60}_{27}Co \rightarrow \ ^{0}_{-1}e + ^{60}_{28}X$

The nucleus of the formed element (_28^60)X contains 28 protons and 32 neutrons.

b) Determination of the radioactive constant λ of Cobalt: $N_t = N_0 \,.\, e^{-\lambda t}$

$A_t = A_0 .\, e^{-\lambda t} \Rightarrow \lambda t = \ln A_0 / A_t \Rightarrow \lambda = 1/t \,.\, \ln A_0 / A_t = 4{,}17 \,.10^{-9}\ s^{-1}$

c) $A_t / \lambda = N_t = 5{,}3237 \,.10^{13}$ noyaux.

$m = M \,.\, N_t / N = 5{,}2975 \,.10^{-9}$ g

Exercise 7:

a) $^{27}_{13}Al + ^{4}_{2}\alpha \rightarrow \ ^{30}_{15}P + ^{1}_{0}X\ (^{1}_{0}n)$

b) $^{30}_{15}P \rightarrow \ ^{30}_{14}Y + ^{0}_{1}\beta +$

c) de particules β^+ émises :

$T\,^{30}_{15}P = 3$ min , $m_0 = 16g$ (at t = 0) ⟶ t = 12 min , $N\beta^+ = ?$

	$^{30}_{15}P$	→	$^{30}_{14}Y +$	$^{0}_{1}\beta +$
at t = 0	m_0			
	N_0			
at t = 12 min	m			
	N_t ($^{30}_{15}P$ no desint)		$N(^{30}_{15}P$ desint$)=N_{Xformed}$	$N(^{30}_{15}P$ desint$)=$

$N\beta^+$formed

$N\beta^+$formed = $N(^{30}_{15}P$ desint$) = N_0 - N_t$

$\ln N_0 / N_t = -\lambda t \Rightarrow N_t = N_0 \,.\, e^{-\lambda t}$

$N\beta^+$formeed = $N(^{30}_{15}P$ desint$) = N_0 - N_t = N_0 - N_0 \,.\, e^{-\lambda t} = N_0 \,.\, (1- e^{-\lambda t})$

$N\beta^+$formed = $N_0 \,.\, (1- e^{-\lambda t})$

$T = \ln 2/\lambda \Rightarrow \lambda = \ln 2/T = 0{,}69/ 3 \Rightarrow \lambda = 0{,}23\ min^{-1}$

$N_0 = m_0 \,.\, N/ M = 16.\ 6{,}023.\ 10^{23}/ 30 = 3{,}21 \,.10^{23}$

$N\beta^+$formed = $N_0 \,.\, (1- e^{-\lambda t}) = 3{,}21 \,.10^{23}.\ (1- e^{-0{,}23 \times 12}) = 3{,}21 \,.10^{23}.\ (1- e^{-2{,}76}) = 3{,}006.\ 10^{23}$

Exercise 8:

1) **$N_t = N_0 . e^{-\lambda t}$;** **T= ln 2/λ**

$$t = \frac{T}{0.69} ln \frac{N_0}{N_t}$$

N_0 = 100% ; $N_{disintegrated} = N_d$ = 80%

$N = N_0 - N_d$ = 100- 80 = 20%

$t = \frac{35}{0.69} ln \frac{100}{20} == 81.28\ years$

2) **$N_t = N_0 . e^{-\lambda t}$**

$N = N_0 - N_d$ = 100- 10 = 90%

$t = \frac{20}{0.69} ln \frac{100}{90} == 3.04\ min$

3) t = 3T

$N_t = N_0 . e^{-\lambda t}$; **λ = ln 2/T**

$$\boldsymbol{N_t = N_0 \times e^{-\frac{ln2}{T}3T}}$$

$$\boldsymbol{N_t = N_0 \times e^{ln(2).3}}$$

$$\boldsymbol{\frac{N_0}{N} = 2^3 \Rightarrow N = \frac{N_0}{8} = 12.5\%\ of N_0}$$

$$\boldsymbol{X = 100 - 12.5 = 87.5\%}$$

4) **$A_t = A_0 . e^{-\lambda t}$**

$$\boldsymbol{\frac{A}{A_0} = e^{-\lambda t} \Rightarrow \lambda t = ln \frac{A}{A_0}}$$

$$\boldsymbol{ln \frac{1}{t} = \frac{1}{\lambda} ln \frac{A}{A_0}; \qquad \lambda = ln \frac{2}{T}}$$

$$\boldsymbol{T = \frac{ln\ 2.t}{ln \frac{A}{A_0}}}$$

$$\boldsymbol{A_0 = 100\%; \quad A = 100 - 90 = 10\%}$$

T= 110,27min

Conclusion:

Unlike chemical reactions, which only alter the electrons without changing the atom's nucleus, radioactivity alters the atom's nucleus, hence the term "nuclear reaction."

It is crucial to remember that radioactivity has a wide range of useful applications, such as energy production, dating archaeological artifacts, and tracers in chemistry, biology, and medicine. If not handled appropriately, it could be harmful. Indeed, a nuclear particle collides with electrons and ionizes the atoms it comes into contact with when it travels through a material (flesh) via radiation. Tissues and other organic molecules are destroyed. Among these effects are:

- ✓ Pathological effects include leukemia, cancer development, and lesions.
- ✓ Genetic consequences include gene modification.

Chapter 4:

Electronic structure of the atom

Exercises and correction

Introduction:

Every atom is made up of a core, positively charged nucleus around by negatively charged electrons. The precise location of an electron as it travels around the nucleus within an atomic orbital—a mathematical function that denotes the physical region where the electron can be calculated to be—can never be known, according to quantum theory. There are s, p, d, and f orbitals based on their shape. These designations are used to characterize an atom's electron configurations and show the orbital shape. One s orbital, three p orbitals, five d orbitals, and seven f orbitals may occur. There can be no more than two electrons in each orbital.

The electron distribution of an atom with many electrons is separated into shells based on the quantum mechanical description of the atom. Subshells are expected to be made up of one or more orbitals that the electrons occupy, and the shells themselves are believed to be made up of one or more subshells.

According to Hund's rule of maximum multiplicity and the Pauli exclusion principle, an atom can have a large number of these orbitals, and electrons are arranged in those orbitals in ascending order of energy. This forms the basis of the electronic structure of atoms.

Knowledge and Skills:

- ✓ Calculate transition energies between levels;
- ✓ Determine the number or characteristics (frequency or wavelength) of spectral lines;
- ✓ List the four quantum numbers and indicate their possible values;
- ✓ Establish the electronic configuration of an atom in the ground state;
- ✓ Understand the Klechkowski, Hund, and Slater rule.

Exercise 1:

1. A hydrogen atom initially in the ground state absorbs 10.2 eV of energy. At what level is the electron located?

2. The electron of a hydrogen atom initially at the n=3 level emits radiation with a wavelength of λ = 1027 Å. At what level is the electron located?

Exercise 2:

Determine whether the following sets of quantum numbers are allowed in an atom. If not allowed, explain your answer.

(a) $n = 2$ $l = 2$ $m = +1$ $s = -½$

(b) $n = 3$ $l = 0$ $m = -1$ $s = +½$

(c) $n = 1$ $l = 0$ $m = 0$ $s = +½$

(d) $n = 2$ $l = 1$ $m = +1$ $s = 0$

Exercise 3:

Are the following statements true or false?

A) If l = 1, the electron is in a d orbital.

B) If n = 2, ml can be equal to –1.

C) For a d electron, ml can have the value 3.

D) If l = 2, the corresponding subshell can accommodate at most 10 electrons.

E) The number n of an electron in an f subshell can be equal to 3.

Exercise 4:

Can the quantum numbers n, l, and ml together have the following values?

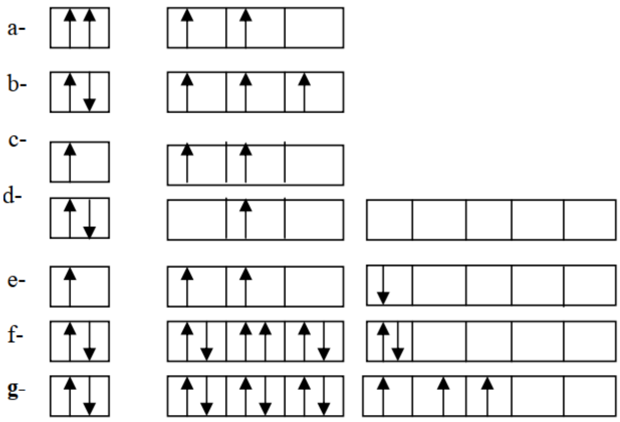
If yes, which subshell do they characterize?

n	l	*ml*	n	l	*ml*
a) 2	0	0	f) 5	3	-3
b) 4	1	-2	g) 4	2	2
c) 3	1	-1	h) 2	3	3
d) 4	-1	0	i) 3	0	0

e)	**2**	0	-1	j)	5	2	-2

Exercise 5:

Which of the following electronic structures do not respect the filling rules? Explain.



Exercise 6:

If the electronic configurations:

a) $1s^2, 2s^2, 2p^6, 3s^1$

b) $1s^2, 2s^2, 2p^6, 4s^1$

Are those of two neutral atoms, are these three statements correct?

If not, which one is false?

A) (a) and (b) represent two different elements.

B) (a) represents sodium.

C) Energy must be supplied to the atom to move from (a) to (b).

Exercise 7:

What is the electron configuration and orbital diagram for a phosphorus (P) atom? What are the four quantum numbers for the last electron added?

Exercise 8:

Determine the electronic configuration of the following species using Klechkowski's rule. Are any isoelectronic species present?

Atoms : B, Sr, Te, Os, Fr, Nb, Fe, Ge, Si, Ar.
Ions: Rb^{+}, Bi^{+3}, Ca^{+2}, S^{2-}, Br^{-}.

Exercise 9:

Establish the electronic configuration, in the ground state, of the following atoms or ions: Fluorine F (9), Cadmium Cd (48), Tin Sn (50), Strontium ion Sr^{2+} (38), Cobalt Co (27), Potassium K (19), Bromine ion Br- (35), Cesium ion Cs^{+} (55), Scandium Sc (21), Iodine I (53), Sulfur ion S-2 (16), Aluminum ion Al^{3+} (13).

Exercise 10:

Given the following elements:
Na (Z=11), Al^{3+}(Z=13), S^{2-}(Z=16),Cl (Z=17), Ar (Z=18), Ca^{2+}(Z=20), Cr (Z=24), Fe(Z=26), Co^{3+}(Z=27), As^{3-}(Z=33), Br (Z=35), I^{-} (Z=53), Cs(Z=55), Pb(Z=82), Bi^{3+} (Z=83).
Give their electronic structures as well as the filling of the quantum boxes of their outer shell.

Exercise 11:

Calculate the effective nuclear charge:

1- of an electron on the 4s orbital, then that of the electron on the 3d orbital of Cu (Z = 29)

2- an electron in the 4p orbital of Se (Z = 34).

Exercise corrections (Chapter 4)

Exercise 1:

1. Absorbed energy:

$$\Delta \boldsymbol{E}_{n_j \to n_i} = E_H \left(\frac{1}{{n_i}^2} - \frac{1}{{n_j}^2} \right)$$

Ground state: nj=1 $\Rightarrow \frac{\Delta E_{n_j \to n_i}}{E_H} = \frac{1}{{n_i}^2} - 1$

$$\frac{1}{{n_i}^2} = \frac{\Delta E_{n_j \to n_i}}{E_H} + 1 = 1 - \frac{10{,}2}{13{,}6} = 0{,}25$$

${n_i}^2 = 4$ ⇒ ni = 2 ⇒ The electron is at level 2

2. Wavelength of radiation emitted: λ = 1027 Å= 1027. 10^{-10} m

$$|\Delta E| = \frac{hc}{\lambda} = \frac{6{,}62.\,10^{-34} x\ 3.\,10^{8}}{1027.\,10^{-10}} = 1{,}934.\,10^{-18} = 12{,}086\ eV$$

$$\left|\Delta E_{n_i \to n_j}\right| = \frac{hc}{\lambda_2} = |E_H| \left(\frac{1}{{n_j}^2} - \frac{1}{{n_i}^2} \right)$$

$$\frac{1}{\lambda_2} = \frac{\left|\Delta E_{n_i \to n_j}\right|}{hc} = \frac{|E_H|}{hc} \left(\frac{1}{{n_j}^2} - \frac{1}{{n_i}^2} \right) = R_H \left(\frac{1}{{n_j}^2} - \frac{1}{{n_i}^2} \right)$$

$n_i = 3$ et $n_j = 1$ ⇒The electron falls back to the ground level.

Exercise 2:

(a) *n* = 2 *l* = 2 *m* = +1 *s* = -½ **F**;

l ≠ 2, 0 ≤ *l* ≤ n-1 ⇒ 0 ≤ *l* ≤ 1 ⇒ *l= 0, 1*

(b) *n* = 3 *l* = 0 *m* = -1 *s* = +½ **F**;

m ≠ -1, *l= 0* ⇒ *m = 0*

(c) *n* = 1 *l* = 0 *m* = 0 *s* = +½ **T ;**

0 ≤ *l* ≤ 0 ⇒ *l= 0*, *m* = 0 *s* = +½

(d) *n* = 2 *l* = 1 *m* = +1 *s* = 0 **F**;

s ≠ 0 ⇒ *s* = ±½

Exercise 3:

A) False: l = 1, p subshell.

B) True: n = 2 implies l = 0.1, hence ml can have the value –1.

C) False: d corresponds to l = 2, so it is impossible for ml to be equal to 3.

D) True: l = 2, (4l + 2) = 10.

E) False: f subshell, l = 3, n would have to be at least equal to 4.

Exercise 4:

The impossibilities are as follows:

b) ml cannot be less than –1 since l = 1.

d) l can never have the value –1 (minimum: 0).

e) ml can only be zero, since 1 = 0.

h) l cannot be greater than 1 (maximum: n – 1) and, consequently, ml cannot be equal to 3.

In conclusion: a) 2s - b) impossible - c) 3p - d) impossible - e) impossible - f) 5f - g) 4d - h) impossible - i) 3s - j) 5d.

Exercise 5:

a) not respect the filling rules (Inexact state): the two spins must be opposite (rule of Pauli).

b) ground state

c) not respect the filling rules (Excited state)

d) not respect the filling rules

e) not respect the filling rules (Excited state)

f) not respect the filling rules (Inaccurate condition). Hund's rule and Pauli's principle are not respected.

g) Fundamental state

Exercise 6:

A) Incorrect: since this is a neutral atom, the number of electrons is equal to the atomic number Z.

Z state is the same for both a and b, meaning they are the same element.

B) Correct: Z = 11, meaning it is sodium.

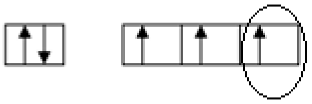
C) Correct: to go from the ground state (a) to the excited state (b).

Exercise 7:

$_{15}P$: $1s^2\ 2s^2\ 2p^6$ $\mathbf{\underline{3s^2\ 3p^3}}$

$n = 3$, $l = 1$, $m = -1, 0, +1$

p orbital



So: $n = 3$, $l = 1$, $m = -1, 0, +1$, $s = +½$

Exercise 8:

B (bore, Z = 5) : K, $2s^2$, $2p^1$

Sr (strontium, Z = 38) : K, L, M, $4s^2$, $4p^6$, $5s^2$

Te (tellure, Z = 52) : K, L, M, $4s^2$, $4p^6$, $4d^{10}$, $5s^2$, $5p^4$

Os (osmium, Z = 76) : K, L, M, N, $5s^2$, $5p^6$, $5d^6$, $6s^2$

Fr (francium, Z = 87) : K, L, M, N, $5s^2$, $5p^6$, $5d^{10}$, $6s^2$, $6p^6$, $7s^1$

Nb (niobium, Z = 41) : K, L, M, $4s^2$, $4p^6$, $4d^3$, $5s^2$

Fe (Iron, Z = 26) : K, L, $3s^2$, $3p^6$, $3d^6$, $4s^2$

Ge (germanium, Z = 32) : K, L, M, $4s^2$, $4p^2$

Si (silicon, Z = 14) : K, L, $3s^2$, $3p^2$

Ar (argon, Z = 18) : K, L, $3s^2$, $3p^6$

Rb (rubidium, Z = 37) : K, L, M, $4s^2$, $4p^6$

Bi^{+3} (bismuth, Z = 83) : K, L, M, N, $5s^2$, $5p^6$, $5d^{10}$, $6s^2$

Ca^{+2} (calcium, Z = 20) : K, L, $3s^2$, $3p^6$

S^{2-} (sulfur, Z = 16) : K, L, $3s^2$, $3p^6$

Br^- (bromine, Z = 35) : K, L, M, $4s^2$, $4p^6$

Isoelectronic (same configuration): Ar, Ca^{+2}, S^{2-}

Exercise 9:

F : [He] $2s^2\ 2p^5$

Cd : [Kr] $4d^{10}\ 5s^2$

Sn : [Kr] $4d^{10}\ 5s^2\ 5p^2$

Co : [Ar] $3d^7\ 4s^2$

K : [Ar] $4s^1$

Sc : [Ar] $3d^1\ 4s^2$

I : [Kr] $4d^{10}\ 5s^2\ 5p^5$

Sr^{2+} : [Kr]

Br^- : [Ar] $3d^{10}\ 4s^2\ 4p^6$

Cs^+ : [Xe]

S^{-2} : [Ne] $3s^2\ 3p^6$

Al^{3+} : [Ne]

Exercise 10:

Na (Z=11) : $1s^2\ 2s^2\ 2p^6\ \underline{3s^1}$ [↑]

Al^{3+}(Z=13) : $1s^2\ \underline{2s^2\ 2p^6}$ [↑↓] [↑↓][↑↓][↑↓]

S^{2-}(Z=16): $1s^2\ 2s^2\ 2p^6\ \underline{3s^2\ 3p^6}$ [↑↓] [↑↓][↑↓][↑↓]

Cl (Z=17) : $1s^2\ 2s^2\ 2p^6\ \underline{3s^2\ 3p^5}$ [↑↓] [↑↓][↑↓][↑]

Ar (Z=18) : $1s^2\ 2s^2\ 2p^6\ \underline{3s^2\ 3p^6}$ [↑↓] [↑↓][↑↓][↑↓]

Ca^{2+}(Z=20) : $1s^2\ 2s^2\ 2p^6\ \underline{3s^2\ 3p^6}$ [↑↓] [↑↓][↑↓][↑↓]

Cr (Z=24) : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\underline{4s^1 3d^5}$ [↑] [↑][↑][↑][↑][↑]

Fe(Z=26) : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\underline{4s^2 3d^6}$ [↑↓] [↑↓][↑][↑][↑][↑]

Co^{3+}(Z=27): $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\underline{3d^6}$ [↑↓][↑][↑][↑][↑]

As^{3-}(Z=33) : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\underline{4s^2}3d^{10}\ \underline{4p^6}$ [↑↓] [↑↓][↑↓][↑↓]

Br (Z=35): $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\underline{4s^2}3d^{10}\ \underline{4p^5}$ [↑↓] [↑↓][↑↓][↑]

I^- (Z=53): $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6 4s^2 3d^{10}\ 4p^6\underline{5s^2}\ 4d^{10}\underline{5p^6}$ [↑↓] [↑↓][↑↓][↑↓]

Cs(Z=55): $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6 4s^2 3d^{10}\ 4p^6 5s^2\ 4d^{10} 5p^6\underline{6s^1}$ [↑]

Pb(Z=82): $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6 4s^2 3d^{10}\ 4p^6 5s^2\ 4d^{10} 5p^6\underline{6s^2}4f^{14}\ 5d^{10}\underline{6p^2}$ [↑↓] [↑][↑][ ]

Bi^{3+} (Z=83) : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6 4s^2 3d^{10}\ 4p^6 5s^2\ 4d^{10} 5p^6\underline{6s^2}4f^{14}\ \underline{5d^{10}}$ [↑↓] [↑↓][↑↓][↑↓][↑↓][↑↓]

Exercise 11:

The expression for the effective charge of an electron is: $\mathbf{Z^* = Z - \sum \sigma_{i\rightarrow j}}$

1. Cu (Z=29) : **(**$1s^2$**)** **(**$2s^2\ 2p^6$**)** ($3s^2\ 3p^6$**)** **($3d^{10}$) ($4s^1$)**

Spatial aspect

(Parenthesis indicates different Slater groups)

$Z^*_{4s} = Z - (10\sigma_{3d\rightarrow 4s} + 8\ \sigma_{3s,3p\rightarrow 4s} + 8\sigma_{2s,2p\rightarrow 4s} + 2\sigma_{1s\rightarrow 4s})$

$Z^*_{4s} = 29 - [(18.\ 0.85) + 10] = 3,7$

$Z^*_{3d} = Z - (9\sigma_{3d\rightarrow 3d} + 8\sigma_{3s,3p\rightarrow 3d} + 8\sigma_{2s,2p\rightarrow 3d} + 2\sigma_{1s\rightarrow 3d})$

$Z^*_{3d} = 29 - (9.\ 0.35 + 8 + 10) = 7,85$

2. Se **(**Z = 34**):** **(**$1s^2$**)** **(**$2s^2\ 2p^6$**)** **(**$3s^2\ 3p^6$ **)** **(**$3d^{10}$**)** **($4s^2\ 4p^4$)**

Spatial aspect

(Parenthesis indicates different Slater groups)

$Z^*_{4p} = Z - (5\ \sigma_{4s,4p\rightarrow 4p} + 10\ \sigma_{3d\rightarrow 4p} + 8\ \sigma_{3s,3p\rightarrow 4p} + 8\ \sigma_{2s,2p\rightarrow 4p} + 2\ \sigma_{1s\rightarrow 4p})$

$Z^*_{4p} = 34 - [5.\ 0,35 + (18.\ 0,85) + 8 + 2] = 6,95$

Conclusion:

In classical mechanics, a particle is characterized by six parameters: three position parameters (x, y, z) and three velocity parameters (vx, vy, vz).
In quantum mechanics, a particle like the electron is characterized by four parameters; these are the four quantum numbers: n, l, m_l, and m_s.

Each electron in an atom can be characterized by four quantum numbers: n, l, ml, and ms. The combination of these four numbers is unique for a given electron. Therefore, two electrons in the same atom will always differ by at least one quantum number.

Electrons are arranged around the nucleus so that their energy level is the lowest. For the lowest levels, they are filled in the intuitive order (1s, 2s, 2p, etc.). But from the 3p level onward, the energy gap between the orbitals decreases and the filling order differs from the intuitive order. The filling of subshells and shells follows criteria based on:
Stability principle (Kleichkowski's law), Hund's rule and Pauli's exclusion principle.

Chapter 5:

Periodic table of elements

Exercises and correction

Introduction:

Mendeleev's table, or periodic table of elements, was classified in 1869 by D. Mendeleev. He realized that by placing the elements in ascending order of their atomic mass, a pattern emerged regarding their properties. Mendeleev arranged elements in horizontal rows and vertical columns of a table in order of their increasing atomic weights in such a way that the elements with similar properties occupied the same vertical column or group. He fully recognized the significance of periodicity and used broader range of physical and chemical properties to classify the elements.

The modern classification represents all chemical elements, classified according to their ascending atomic number Z and organized according to Klechkowski's rule. The periodic table contains 7 rows (7 periods) and 18 groups (columns).

The rows of the periodic table. They increase in atomic number from left-right, and each period corresponds to the number of electron shells of the elements in that period. The columns of the periodic table. Elements in the same group share similar chemical properties, as they have the same number of valence electrons.

The three main periodic properties are: Atomic Radius, Ionisation Energy and Electronegativity:

- Atomic Radius: Half the distance between the centers of two atoms of an element that are touching.
- Ionisation Energy: The energy required to remove one valence electron from a gaseous atom.
- Electronegativity: The measure of the ability of an atom to attract electrons for chemical bonding.

Knowledge and Skills:

- ✓ Recognize the importance of the electronic configuration and atomic number as the foundation for periodic classification;
- ✓ Classify elements into s, p, d, f blocks and learn their main characteristics;
- ✓ Recognise the periodic trends in physical and chemical properties of elements;
- ✓ Utilize scientific terminology appropriately, such as atomic/ionic radii, ionization enthalpy, electron gain enthalpy, electronegativity, and elemental valence;
- ✓ Establish the ionization energy and electronegativity, and describe their variations as a function of atomic number in the periodic table;
- ✓ Describe the evolution of the atomic radius in the columns and rows of the periodic table.

Exercise 1:

Consider the following atoms:

$_{7}N$, $_{21}Sc$, $_{24}Cr$,$_{25}Mn$, $_{26}Fe$, $_{29}Cu$, $_{30}Zn$, $_{47}Ag$, $_{79}Au$

1- Give the electronic configurations of the atoms.

2- Show the valence electrons for each atom. Deduce the number of valence electrons.

3- Classify these elements into its appropriate group, period and block.

Exercise 2:

We consider two elements of the fourth period whose external electronic structure includes three single electrons.

1. Write the complete electronic structures of each of these elements and determine their atomic number.

2. Determine the atomic number and give the electronic configuration of the element located in the same period as iron (Z = 26) and belonging to the same group as carbon (Z = 6).

Exercise 3:

Find the electronic configuration of the following elements and give the possible ions they can form;

$_{19}K$, $_{12}Mg$, $_{17}Cl$, $_{18}Ar$, $_{35}Br$, $_{22}Ti$, $_{37}Rb$.

Exercise 4:

Determine the electron configuration of the following species by reading it from the periodic table.

Atoms : Mn, Tc, P, Gd, Ac, Zr, Rn, Kr, Ne

Ions : Sn^{+2}, K^{+}, Se^{2-} , I^{-}, As^{3-}

Exercise 5:

Which of these statements is correct?

1- Element X belongs to the $_{14}Si$ group and has the same period as $_{26}Fe$. Its electronic configuration is:

a) [Ar] $4s^2\ 3d^{10}\ 4p^2$
b) [Ne] $3s^2\ 4s^2\ 3p^6\ 3d^{10}\ 4p^2$
c) [Ne] $3s^2\ 3d^{10}\ 3p^6\ 4s^2\ 4p^2$

2- The element whose Z= 50 belongs to:

a) The 5^{th} period and the 3^{rd} group

b) The 6^{th} period and the 4^{th} group

c) The 5^{th} period and the 5^{th} group

d) All answers are false.

3- Consider the following elements: F (Z=9), Na (Z=11), K (Z=19)

Classify these elements by increasing atomic radii,

a) rF< rNa < rk
b) rF > rNa > rK
c) rNa > rF > rK
d) All answers are wrong

4- As the ionization energy moves from left to right, the ionization energy increases.

a) In the same direction as the atomic radius

b) In the opposite direction to the atomic radius

c) In the direction of decreasing Z

5- The classification of the ionization energies of $_7N$, $_{11}Na$ and $_{15}P$ is given by $Ei_{(N)}$, $Ei_{(Na)}$ and $Ei_{(P)}$ respectively:

a) $Ei_{(N)} < Ei_{(Na)} < Ei_{(P)}$
b) $Ei_{(N)} < Ei_{(P)} < Ei_{(Na)}$

c) $Ei_{(P)} < Ei_{(Na)} < Ei_{(N)}$

d) All answers are wrong

Exercise 6:

In each of the following four sets, arrange the elements in order of increasing radius of their atoms:

a) Cs, F, K, N, Li

b) Ba, Cl, I, Sn, Sr

c) Al, In, F, O, Si, S

d) Al, Ca, Fr, Mg, Rb, S

Exercise 7:

1.Give the electronic configuration of the following elements:

$_{17}Cl^{-}$, $_{11}Na^{+}$, $_{37}Rb$, $_{29}Cu$, $_{42}Mo$, $_{15}P^{-}$

2. Locate these elements in the periodic table, indicating the period, group, and column.

3. Indicate the most stable ion for each of these elements, justifying your choice.

4. Rank these elements in order:

•Descending electronegativity

•Descending first ionization energy

•Increasing covalent radius

Exercise 8:

Consider the following atoms and ions: A, D, E, G, and M, with:

• Element A is the fourth alkali metal

• Atom D forms a more stable ion D^{+2}, which has the same structure as the second noble gas

• Atoms E and G belong to the same period as bromine ($_{35}Br$) and each have four valence electrons, and the mass of E is greater than that of G.

• Atom M is a halogen that belongs to the same period as copper ($_{29}Cu$).

Find the electronic configuration of these elements

Exercise 9:

1) Classify in each series the following elements according to their increasing radius:

   $_{11}Na$; $_{19}K$; $_{37}Rb$

   $_{6}C$; $_{7}N$; $_{8}O$

   $_{26}Fe$; $_{26}Fe^{2+}$; $_{26}Fe^{3+}$

   $_{17}Cl^{-}$; $_{18}Ar$; $_{20}Ca^{2+}$

2) Let be given $_{11}Na^{+}$, $_{12}Mg^{2+}$ et $_{13}Al^{3+}$. What is special about these ions? Which of these ions has the smallest ionic radius?

Exercise 10:

1- Arrange the elements below in order of increasing atomic radius

a) $_{5}B$, $_{6}C$, $_{7}N$, $_{8}O$, $_{9}F$

b) $_{9}F$, $_{17}Cl$, $_{35}Br$, $_{53}I$

2- Arrange the elements below in order of increasing ionization and affinity energies

a) $_{3}Li$, $_{1}H$, $_{11}Na$, $_{19}K$, $_{55}Cs$

b) $_{7}N$, $_{9}F$, $_{11}Na$, $_{13}Al$, $_{37}Rb$

Exercise corrections (Chapter 5)

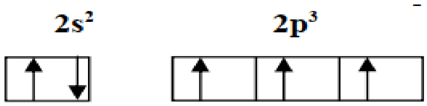
Exercise 1:

1 and 2)

Klechkowski's rule	spatial arrangement	Number of electrons
N (7) : $1s^2$ **$2s^2$ $2p^3$**	[He]**$2s^2$ $2p^3$**	**5**
K (19): $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$**$4s^1$**	[Ar]**$4s^1$**	**1**
Sc (21) $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ **$4s^2$ $3d^1$**	[Ar] **$3d^1$ $4s^2$**	**3**
Cr (24) $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ **$4s^1$ $3d^5$**	[Ar] **$3d^5$ $4s^1$**	**6**
Mn (25)$1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ **$4s^2$ $3d^5$**	[Ar]**$3d^5$ $4s^2$**	**7**
Fe (26) $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ **$4s^2$ $3d^6$**	[Ar]**$3d^6$ $4s^2$**	**8**
Cu (29) $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ **$4s^1$$3d^{10}$**	[Ar]**$3d^{10}$ $4s^1$**	**11**
Zn (30) $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ **$4s^2$$3d^{10}$**	[Ar]**$3d^{10}$ $4s^2$**	**2**
Ag (47) $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $3d^{10}$ $4s^2$ $4p^6$ **$5s^1$ $4d^{10}$**	[Kr]**$4d^{10}$ $5s^1$**	**11**
Au (79) $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $4s^2$ $3d^{10}$ $4p^6$ $5s^2$$4d^{10}$ $5p^6$ **$6s^1$** $4f^{14}$**$5d^{10}$**	[Xe]**$5d^{10}$ $6s^1$**	**11**

Example: N (7): $1s^2$ <u>**$2s^2$ $2p^3$**</u>

Representation of the valence layer using quantum boxes:



Note :

In the case of chromium Cr (Z=24), the structure of the valence layer must be according to Klechkowski's rule: **$4s^2$ $3d^4$**.

This structure is unstable. The most stable structure is therefore **$4s^1$ $3d^5$**. The electronic structure of the **3d** subshell is **half filled**.

In the case of copper Cu (Z=29), the structure of the valence layer must be according to Klechkowski's rule: **$4s^2$ $3d^9$**. This structure is unstable. The most stable structure is therefore **$4s^1$ $3d^{10}$**. The electronic structure of the 3d sublayer is **fully filled**.

“Half-filled or fully filled d orbitals are more stable”

3) A single element belongs to the period n = 2: N (Z=7) (group VA)

- The elements which belong to the period n=4 are: K (group IA), Sc (group IIIB), Cr (group VIB), Mn (group VIIB), Fe (group
VIIIB), Cu (group IB), Zn (group IIB)

- The elements which belong to the IB family are: Cu (4th period) Ag (5th period), Au (6th period)

- The elements which belong to the family of transition metals (their valence layer is of type **$(n-1)d^y\ ns^x$** where **$l \leq x \leq 2$** and **$1 \leq y \leq 10$**) are: Sc (group IIIB), Cr (group VIB), Mn (group VIIB), Fe (group VIIIB), Cu (group IB), Zn (group IIB)

Exercise 2:

1- The two elements are vanadium and arsenic.

Electronic structure of Vanadium **V**: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^3$ according to Klechkowski's rule.

The atomic number is: **Z = 23**

Electronic structure of arsenic:

Ace: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^3$ according to Klechkowski’s rule

The atomic number is **Z = 33**

2-

Electronic structure of iron Fe (Z=26):

[Ar] $3d^6\ 4s^2$; Iron belongs to the 4th period **n= 4**

Electronic structure of carbon C (Z=6) $1s^2\ 2s^2\ 2p^2$

Carbon belongs to the **$ns^2\ np^2$** valence shell electronic structure family.

So the electronic structure of germanium is: Ge [Ar] $3d^{10}\ 4s^2\ 4p^2$

Exercise 3:

1. K (19): [Ar] $4s^1$ only one possible ion K^+. K tends to have the stable structure of the inert gas argon.

2. Mg (12): [Ne] $3s^2$ two possible ions Mg^{2+} and Mg^+

3. Cl (17): [Ne] $3s^2\,3p^5$ only one possible ion Cl^- (structure of argon: inert gas

4. Ar (18): [Ne] $3s^2\,3p^6$ there is no ionization possible because its state is stable; it is an inert gas

5. Br (35): [Ar] $3d^{10}\,4s^2\,4p^5$ only one possible ion Br^- (structure of the inert gas krypton)

6. Ti (22): [Ar] $3d^2\,4s^2$ four possible ions Ti^{4+}, Ti^{3+}, Ti^{2+} and Ti^+. (Ti^{4+}, Ti^{3+} are the most stable)

7. Rb (37): [Kr] $5s^1$ only one possible ion Rb^+

Exercise 4:

Mn (Z = 25) : K, L, $3s^2$, $3p^6$, $3d^5$, $4s^2$

Tc (Z = 43) : K, L, M, $4s^2$, $4p^6$, $4d^5$, $5s^2$

P (Z = 15) : K, L, $3s^2$, $3p^3$

Gd (Z= 64) : K, L, M, $4s^2$, $4p^6$, $4d^{10}$, $4f^7$, $5s^2$, $5p^6$, $5d^1$, $6s^2$

Ac (Z = 89) : K, L, M, N, $5s^2$, $5p^6$, $5d^{10}$, $6s^2$, $6p^6$, $6d^1$, $7s^2$

Zr (Z= 40) K, L, M, $4s^2$, $4p^6$, $4d^2$, $5s^2$

Rn (Z = 86) : K, L, M, N, $5s^2$, $5p^6$, $5d^{10}$, $6s^2$, $6p^6$

Kr (Z= 36) K, L, M, $4s^2$, $4p^6$

Ne (Z = 10) : K,L

Sn^{+2} (Z = 50) : K, L, M, $4s^2$, $4p^6$, $4d^{10}$, $5s^2$

K (Z =19): K, L, $3s^2$, $3p^6$

Se^{2-} (Z= 34) : K, L, M, $4s^2$, $4p^6$

I^- (Z = 53) : K, L, M, $4s^2$, $4p^6$, $4d^{10}$, $5s^2$, $5p^6$

As^{3-} (Z = 33) : K, L, M, $4s^2$, $4p^6$

Exercise 5:

1) a
2) d
3) a
4) b
5) c

Exercise 6:

The atomic radius varies regularly within a period (except within a series d of transition elements), or within a column of the periodic table. However, it is not easy to compare the radii of atoms located on an oblique line. So try to connect the elements proposed in each series by a "path" formed only of verticals and horizontals. Example: for series a), Cs, K, Li (column 1), then Li, N, F (row 2).

a) F< N< Li <K< Cs

b) Cl <I< Sn <Sr <Ba

c) F< O <S< Si <Al<In

d) S< Al< Mg< Ca< Rb< Fr.

Exercise 7:

1) Electronic configuration

Elements	Structure of the atom	Structure of the ion
$_{17}Cl^-$	$[_{10}Ne]3s^2\ 3p^5$	$[_{10}Ne]3s^2\ 3p^6$
$_{11}Na^+$	$[_{10}Ne]3s^1$	$[_{10}Ne]3s^0$ or $[_2He]2s^2\ 2p^6$
$_{37}Rb$	$[_{36}Kr]5s^1$	
$_{29}Cu$	$[_{18}Ar]4s^1\ 3d^{10}$	
$_{42}Mo$	$[_{36}Kr]5s^1\ 3d^5$	
$_{15}P^-$	$[_{10}Ne]3s^2\ 3p^3$	$[_{10}Ne]3s^2\ 3p^4$

2) Place these elements in the periodic table

Elements	Period	Group	Column
$_{17}$Cl	3	VIIA	17
$_{11}$Na	3	IA	1
$_{37}$Rb	5	IA	1
$_{29}$Cu	4	IB	11
$_{42}$Mo	5	VIB	6
$_{15}$P	3	VA	15

Remember that to find the position or location of an element in the periodic table, you need to provide its horizontal and vertical coordinates, i.e., period and group (or column). In the case of an ion, you need to provide the position of its corresponding atom.

3) Most stable ion

Elements	More stable ion	Justification
$_{17}$Cl	Cl^-	$[_{10}Ne]\ 3s^2\ 3p^6$ ⇒rare gas therefore very stable structure
$_{11}$Na	Na^+	$[_2He]\ 2s^2\ 2p^6$ ⇒rare gas therefore very stable structure
$_{37}$Rb	Rb^+	$[_{18}Ar]\ 4s^2\ 3d^{10}\ 4p^6$ ⇒rare gas therefore very stable structure
$_{29}$Cu	$_{29}Cu^+$	$[_{18}Ar]\ 4s^0\ 3d^{10}$ ⇒is a very stable structure
$_{42}$Mo	$_{42}Mo^+$	$[_{36}Kr]\ 5s^0\ 4d^5$ ⇒is a very stable structure
$_{15}$P	$_{15}P^{3-}$	$[_{10}Ne]\ 3s^2\ 3p^6$ ⇒rare gas therefore very stable structure

4) Classification of elements according to their properties.

Period	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**	**11**	**12**	**13**	**14**	**15**	**16**	**17**	**18**
3	**Na**														**P**		**Cl**	
4											**Cu**							
5	**Rb**					**Mo**												

- For electronegativity:

Period 3 (Cl > P > Na) > Period 4 (Cu) > Period 5 (Mo > Rb)
Cl > P > Na > Cu > Mo > Rb

- For ionization energy:

Period 3 (Cl > P > Na) > Period 4 (Cu) > Period 5 (Mo > Rb)
Cl > P > Na > Cu > Mo > Rb

For the radius: it is the inverse of electronegativity

Period 3 (Cl < P < Na) < Period 4 (Cu) < Period 5 (Mo < Rb)

Cl < P < Na < Cu < Mo < Rb

Justification :

For electronegativity, within the same period, it increases with Z or from left to right, and within the same group, it increases as Z decreases or from bottom to top.

The first ionization energy evolves in the same way as electronegativity, while for the radius, it's practically the opposite.

Exercise 8:

•Element A is the fourth alkali metal, so it is: $\mathbf{_{37}A : [_{36}Kr]\ 5s^1}$

Attention !

Hydrogen is not an alkali metal, and alkalis begin in the second period, so the first alkali metal corresponds to n = 2 (Li :$[_2He]\ 2s^1$)

•The most stable ion, D^{2+}, = second noble gas = $[_{10}Ne]$ (10 electrons)

This means that the D atom has 12 electrons, so $_{12}D$: $[_{10}Ne]\ 3s^2$

•Atoms E and G belong to the same period as chlorine $_{35}Br$

$_{35}Br$: $[_{18}Ar]\ 4s^2\ 3d^{10}\ 4p^5$, so E and G belong to the fourth period. Since they each have 4 valence electrons, we have two possible structures: $[_{18}Ar]\ 4s^2\ 3d^2$ or $[_{18}Ar]\ 4s^2\ 3d^{10}\ 4p^2$ Since the mass of E is greater than that of G and we know that Z increases with mass, we assign the configurations as follows:

$_{22}G$: $[_{18}Ar]\ 4s^2\ 3d^2$ and $_{32}E$: $[_{18}Ar]\ 4s^2\ 3d^{10}\ 4p^2$

•The M atom is the halogen preceding copper $_{29}Cu$

$_{29}Cu$: $[_{18}Ar]\ 4s^1\ 3d^{10}$, so Cu belongs to the fourth period.

The halogen preceding copper must automatically belong to the period before the fourth, that is, the third. Halogen always ends in ….np^5

So, putting the two pieces of information together, we get:

$_{17}M$: [$_{10}Ne$] $3s^2$ $3p^5$

Exercise 9:

1) Within a given column of the periodic table, the valence shell number increases from top to bottom.
Consequently, the attraction between the outermost electron and the nucleus becomes increasingly weaker, and the atomic radius increases from sodium to rubidium.
rNa < rK < rRb
- C (Z=6), N (Z=7), O (Z=8) :
These atoms belong to the same period. The valence shell number remains the same, but the atomic number (Z) increases from carbon to oxygen.
The force of attraction becomes increasingly strong, and consequently, the radius decreases from carbon to oxygen.
rC > rN > rO
The electronic structures of Fe, Fe^{2+} et Fe^{3+} are:
Fe (Z=26) : [Ar] $3d^6$ $4s^2$
Fe^{2+}(Z=26) : [Ar] $3d^6$
Fe^{3+}(Z=26) : [Ar] $3d^5$
For these three elements, the number of protons remains constant, while the number of electrons decreases. This results in a decrease in the shielding effect exerted by the electrons on each other.
As a result, the attraction increases and the radius decreases:
rFe > rFe^{2+} > rFe^{3+}
- Cl^-(Z=17), Ar (Z=18), Ca^{2+}(Z=20)
These three elements have the same electronic structure: $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$.
They are isoelectronic (same number of electrons).
The attraction thus becomes stronger as the atomic number (Z) increases, and therefore the radius decreases:
rCl^- > rAr > rCa^{2+}

2) The Na^+, Mg^{2+} and Al^{3+} ions have the structure of Neon: $1s^2\ 2s^2\ 2p^6$.

They are isoelectronic.

Since the number of electrons is constant, the attraction becomes stronger as the atomic number (Z) increases, resulting in a decrease in radius.

Since ZAl > ZMg > ZNa, we have: $rNa^+ > rMg^{2+} > rAl^{3+}$

Exercise 10:

1- Arrange the elements below according to increasing atomic radius

a) **$_5B$, $_6C$, $_7N$, $_8O$, $_9F$**

Elements	Period	Group
$_5B$	2	IIIA or 13
$_6C$	2	IVA or 14
$_7N$	2	VA or 15
$_8O$	2	VIA or 16
$_9F$	2	VIIA or 17

$r_F < r_O < r_N < r_C < r_B$

b) **$_9F$, $_{17}Cl$, $_{35}Br$, $_{53}I$**

Elements	Period	Group
$_9F$	2	VIIA or 17
$_{17}Cl$	3	VIIA or 17
$_{35}Br$	4	VIIA or 17
$_{53}I$	5	VIIA or 17

$r_F < r_{Cl} < r_{Br} < r_I$

2- Arrange the elements below in order of increasing ionization and affinity energies

a) $_3Li$, $_1H$, $_{11}Na$, $_{19}K$, $_{55}Cs$

$E_I\ Cs < E_I\ K < E_I\ Na < E_I\ Li < E_I\ H$

$A_e\ Cs < A_e\ K < A_e\ Na < A_e\ Li < A_e\ H$

Conclusion:

Certainly the most significant idea in chemistry, both conceptually and practically, is the periodic table. It offers specialists new research directions, serves as a daily resource for students, and concisely organizes the entire field of chemistry.

He demonstrates in a remarkable way that chemical elements are not a random collection of entities, but that they exhibit tendencies and form families.

Anyone who wants to decipher the world and comprehend how it is made up of the basic components of chemistry must be familiar with the periodic table.

Chapter 6:

Chemical bonds

Introduction:

Chemical bond it is the attractive force which hold the constituent particles (atoms, ions etc) together in different chemical species is called a chemical bond.

Lewis postulated that atoms achieve the stable octet when they are linked by chemical bonds.

In the formation of a molecule, only valence shell electrons take part in chemical combination and they are known as valence electrons.

The bond formed between the two atoms by mutual sharing of electrons between them so as to complete their octets or duplets (in case of elements having only one shell) is called covalent bond or covalent linkage and the number of electrons contributed by each atom is known as covalency.

When a bond is formed by complete transfer of electrons from one atom to another so as to complete their outermost orbits by acquiring 8 electrons (i.e., octet) or 2 electrons (i.e., duplet) the bond formed is called ionic bond or electrovalent bond.

as well ascovalent bond is formed between two similar atoms, the shared pair electrons is equally attracted by the two atoms. As a result electron pair is situated exactly between the two identical nuclei. The bond so formed is called non polar covalent bond.

In the case of heterogeneous molecules, the shared pair of electrons between the two atoms is shifted more towards the more electronegative atom. The resultant covalent bond is a polar covalent bond. As result of polarisation the molecule possesses the dipole moment, dipole moment defined as the product of the magnitude of charge and the distance separating the centres of positive and nigative charges. Its direction is from positive end to nigative end.

Knowledge and Skills:

- ✓ understand KÖssel-Lewis approach to chemical bonding;
- ✓ Imagine and represent the electronic structure (Lewis structure);
- ✓ Recognize the octet rule and its limitations, draw Lewis structures of simple molecules;
- ✓ Recognize the formation of different types of bonds;
- ✓ Apply Gillespie's theory;
- ✓ Apply the VSEPR theory and predict the geometry of simple molecules;
- ✓ Determine the geometry of a simple molecule;
- ✓ Recognize the dipole moment

Exercise 1:

Establish the electronic structures (Lewis structures or formulas) of the following molecules or ions using the general construction rules.

For simplicity, the atom occupying a central position in the molecule has been identified by underlining it; all other atoms are directly bonded to it.

a) $\underline{Kr}F_6$

b) $\underline{C}OS$

c) $\underline{C}H_2Cl_2$

d) $Cl\underline{N}O$

e) $\underline{B}F_3$

f) $\underline{S}F_2$

g) $\underline{Se}F_4$

h) $H_3\underline{S}$

i) $H\underline{S}^-$

j) $\underline{Si}O_2$

k) $\underline{C}OCl_2$

l) $\underline{S}OCl_2$

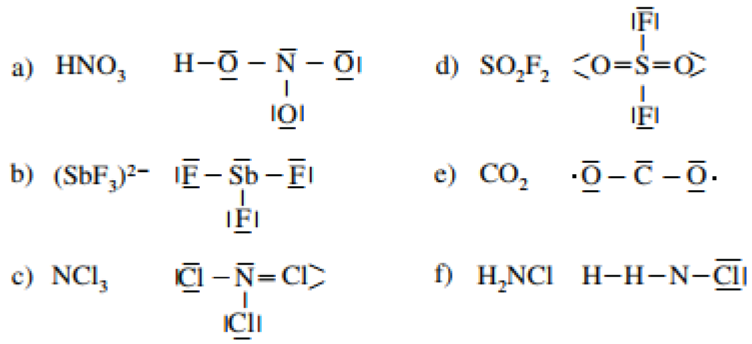
Exercise 2:

Write the Lewis structures of the following molecules or ions (the underlined symbol is that of the central atom).

a) $\underline{B}F^{4-}$	d) $\underline{Se}F_6$	g) $\underline{As}Cl_3$	j) $\underline{P}Cl_6^-$	m) $\underline{Xe}F_2$
b) XeF_4	e) $\underline{I}Cl_3$	h) $\underline{C}H3^-$	k) $\underline{S}O_2$	n) $\underline{C}H_3^+$
c) $H\underline{C}N$	f) $\underline{Sb}F_5$	i) $\underline{Te}Br_6$	l) $\underline{N}O_2F$	o) $\underline{I}F_5$

Exercise 3:

Do the following Lewis formulas actually correspond to the indicated molecular formulas (molecules or ions)? If they are incorrect, correct them.



Exercise 4:

Establish Lewis diagrams and predict the geometric shape of:

a) The anions: hypochlorite ClO^{-}, chlorite ClO^{2-}, chlorate ClO^{3-}, and perchlorate ClO^{4-};

b) The ammonium cation NH_4^{+}.

Exercise 5:

a) Knowing that both NF_3 and PF_3 exist, how can we explain why PF_5 exists and NF_5 does not?

b) Predict the basic geometry and shape of the following molecules using the VSEPR model and explain this by considering the hybridization of the central atom:

$SnCl_2$; H_2S ; PCl_3 ; GeF_4 ; $AsCl_5$; SF_4 ; SF_6 ; IF_7.

c) Which of the molecules studied in question b) are polar?

Exercise 6:

1. Give the Lewis notation for the following molecules and ions:

H_2 ; Cl_2 ; H_2O ; H_3O^{+} ; NH_3 ; NH_4^{+} ; CH_4 ; C_2H_6 ; SF_4 ; SF_6 ; PCl_3 ; PCl_5 ; NCl_3

2. Which of these compounds do not obey the Octet Rule?

3. Based on the electronic structures of sulfur and phosphorus atoms, explain the formation of the molecules SF_6 and PCl_5.

Exercise 7:

On the basis of individual bond polarity and orientation, determine whether each of the following molecules would be polar or nonpolar:

H_2O, I_2, CF_4, NH_3, CO_2

Exercise 8:

1- Draw the energy level diagram of the molecular orbitals (MO) of the N_2 and CN molecule.

2- Calculate the bond order of these molecules.

3- Deduce the magnetic properties of N_2 and CN.

Exercise 9:

Illustrate the hybridisation of the carbon and boron atoms in the following molecules.

BF_3 ; CO_2 ; CH_4 ; C_2H_4 et C_2H_2

Exercise 10:

In the water molecule, the angle HÔH has an experimental value of 105°.

1. Calculate the dipole moment of this molecule, assuming it is equal to the vector sum of the dipole moments of the two O-H bonds.

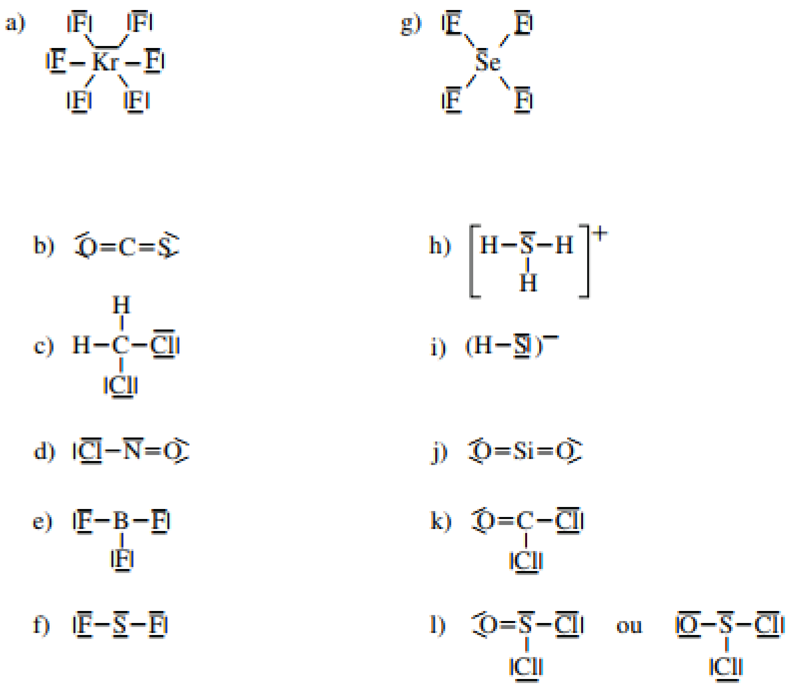
2. Calculate the ionic percentage of the O-H bond in H_2O.

Given: $\mu_{O\text{-}H}$ = 1,51D and $l_{O\text{-}H}$ = 0,96 Å.

Exercise corrections (Chapter 6)

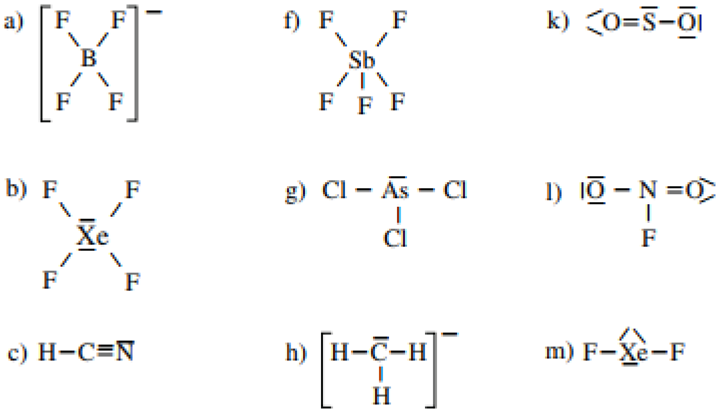
Exercise 1:

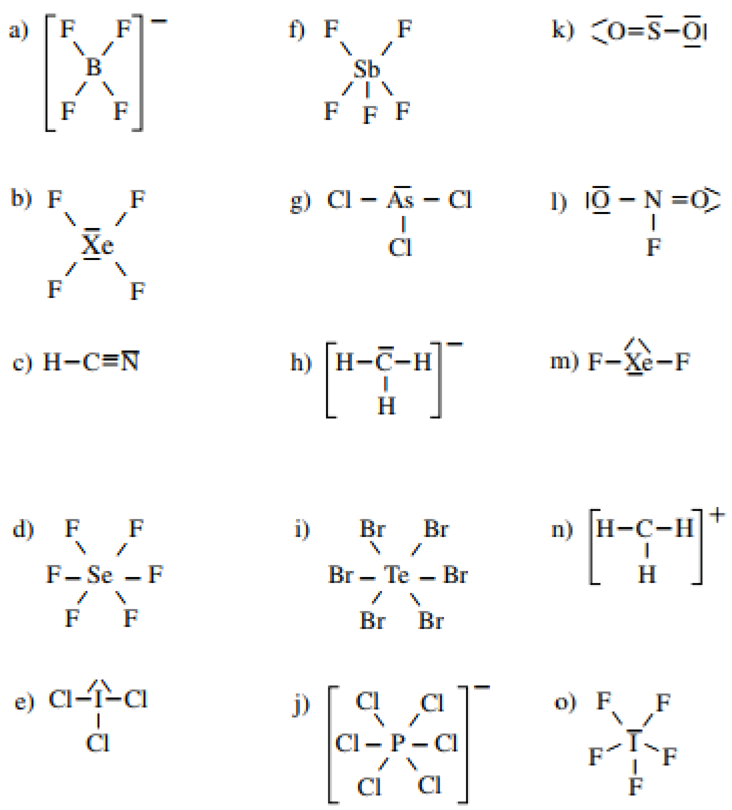
The following Lewis formulas represent all the doublets. You will sometimes find structures in which the 3 free doublets of terminal halogens are not given, or in which the two free doublets of doubly bonded oxygen are not represented: be careful, they must be included in the doublet count!



Exercise 2:

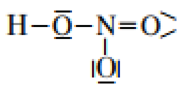
To simplify the formulas below, the three free pairs of halogens bound to the central atom are not shown.



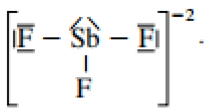


Exercise 3:

a) Incorrect. There are two extra electrons (this would be the formula for an HNO_3^{2-} ion). The correct formula is :



b) Two electrons are missing (this would be the formula for SbF_3). The correct formula is:

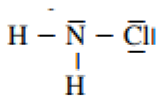


c) Incorrect. Nitrogen cannot have ten electrons in its outer shell; Chlorine does not have valence 2. The correct formula is :

d) Correct. No anomalies; this is one of two possible Lewis formulas.

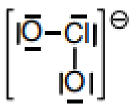
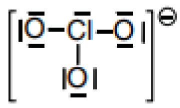
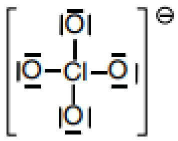
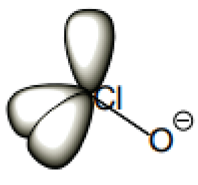
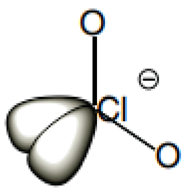
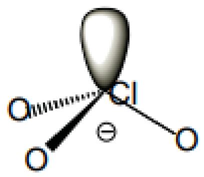
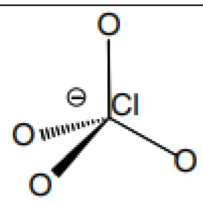
e) Incorrect. There should not be two unpaired electrons on the oxygen atoms. The correct formula is:

f) Incorrect. H cannot have four outer electrons (and therefore cannot form two bonds). The correct formula is:



Exercise 4:

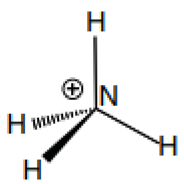
a) For all anions, the central atom is Cl, which has 7 valence electrons. The bonded atoms are always O atoms, which have 6 valence electrons.

Anion	ClO^-	ClO_2^-	ClO_3^-	ClO_4^-
Valence electrons	7 + 6 + 1 = 14	7 + 2 · 6 + 1 = 20	7 + 3 · 6 + 1 = 26	7 + 4 · 6 + 1 = 32
Pairs to distribute	7	10	13	16
Lewis				
VSEPR	AXE3	AX2E2	AX3E	AX
Anion	ClO^-	$ClO2^-$	$ClO3^-$	$ClO4^-$
Basic geometry	tetrahedron			
Shape	linear	triangular	pyramidal	tetrahedral
				

b) For the ammonium cation $NH4^+$, we proceed as before. There are:

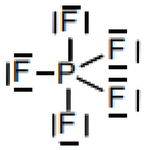
5 + 4 x 1 – 1 = 8 valence electrons, or 4 pairs to be distributed.

The VSEPR method yields AX4, and the ion's shape is that of a tetrahedron:



Exercise 5:

a) The Lewis representation of PF5 shows the hypervalence of phosphorus, that is, a derogation from the octet rule.



What is the physical meaning of this?

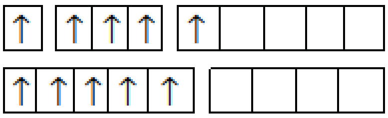
The outer electronic configuration of P (Z = 15, group 15) is:

$3s^2\ 3p^3$ 

Such a configuration does not allow the establishment of five single covalent bonds with respectively 5 fluorine atoms, each possessing a single electron . It must therefore be considered that, during the establishment of the bonds, the energy released allows the promotion of a 3s electron of P into a vacant 3d orbital, and the formation of 5 sp3d hybrid O.A., each describing an electron.

$3s^1\ 3p^3\ 3d^1$

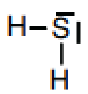
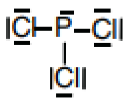
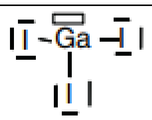
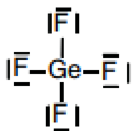
sp^3d

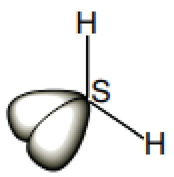
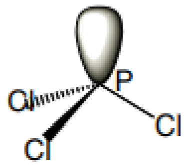
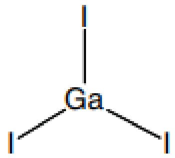
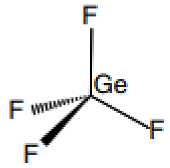


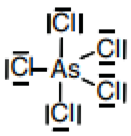
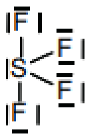
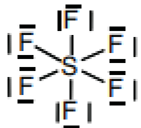
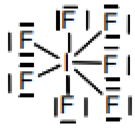
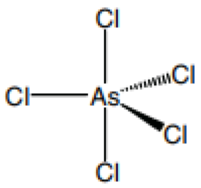
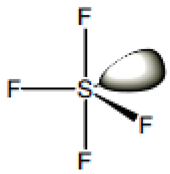
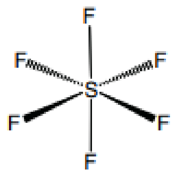
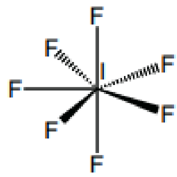
Nitrogen (Z = 7, group 15 like P), with an external configuration of $2s^2\ 2p^3$, belongs to the second period. It does not have a d subshell, and therefore cannot be in a hypervalent state like P. Consequently, NF_5 cannot exist.

Conclusion: Hypervalence can only exist for elements with a d subshell, i.e., elements from the third period or higher.

b) The results are summarized in the tables below.

Compound	H_2S	PCl_3	GaI_3	GeF_4
Atome central (Z, group)	S (Z = 16, Gr. 16)	P (Z = 15, Gr. 15)	Ga (Z = 31, Gr. 13)	Ge (Z = 32, Gr. 14)
Valence electrons	6 + 2 x 1 = 8	5 + 3 x 7 = 26	3 + 3 x 7 = 24	4 + 4 x 7 = 32
Pairs to distribute	4	13	12	16
Lewis			 Electron gap on Ga	
VSEPR	AX_2E_2	AX_3E	AX_3	AX_4
Central atom hybridization	sp^3	sp^3	sp^2	sp^3
Basic geometry	tetrahedron	tetrahedron	equilateral triangle	tetrahedron

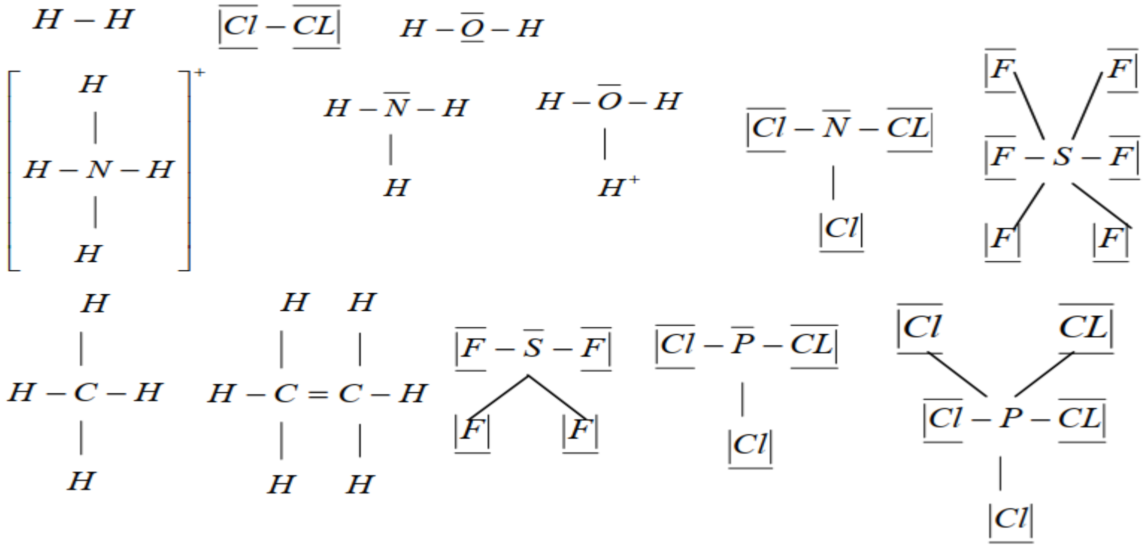
Shape	(isosceles) triangle	pyramid	equilateral triangle	tetrahedron
				

Compound	$AsCl_5$	SF_4	SF_6	IF_7
Atome central (Z, group)	As (Z = 33, Gr. 15)	S (Z = 16, Gr. 16)	S (Z = 16, Gr. 16)	I (Z = 53, Gr. 17)
Valence electrons	5 + 5 x 7 = 40	6 + 4 x 7 = 34	6 + 6 x 7 = 48	7 + 7 x 7 = 56
Pairs to distribute	20	17	24	28
Lewis	 Hypervalence for As	 Hypervalence for S	 Hypervalence for S	 Hypervalence for I
VSEPR	AX_5	AX_4E	AX_6	AX_7
Central atom hybridization	sp^3d	sp^3d	sp^3d^2	sp^3d^3
Basic geometry	triangular bipyramid	triangular bipyramid	square-based bipyramid (octahedron)	pentagonal bipyramid
Shape	triangular bipyramid	tetrahedron called "seesaw"	octahedron	pentagonal bipyramid
		 The bulky, non-bonding doublets are always in the equatorial plane.		

c) Molecules whose symmetry is not such that the vector sum of the bond moments is not zero are polar. These are H_2S; PCl_3; SF_4.

Exercise 6:

1. Lewis notation of the following molecules and ions:

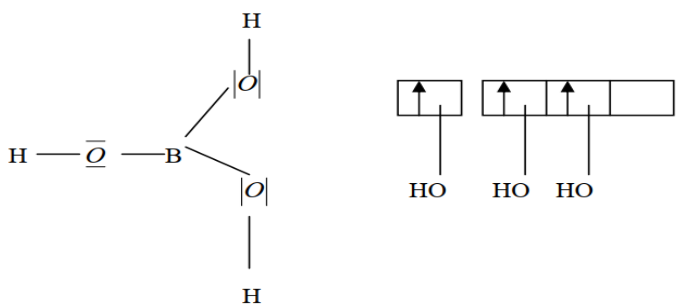


2. Octet Rule: Atoms characterized by Z>4 tend to have 8 electrons in their outer shell in order to acquire a stable state.Limite du modèle de l'octet

- The octet rule strictly applies only to C, N, O, and F atoms in the second period of the periodic table.
- These atoms can only have a maximum of eight electrons in their outermost shell (valence shell).
- However, an atom can have:
  - ➢ either fewer than eight electrons around it: this is the case for boron

    Example: H_3BO_3
  - ➢ or more than eight electrons in its M shell (third period): this is the case for phosphorus. *Exemple :* PCl_5

    B : $1s^2\ 2s^2\ 2p^1$ B*(état excité) : $1s^2\ 2s^1\ 2p^2$



Compounds that do not follow the Octet rule are: $\mathbf{SF_4}$ **;** $\mathbf{SF_6}$ **;** $\mathbf{PCl_5}$

3. Formation of molecules $\mathbf{SF_6}$ **et** $\mathbf{PCl_5}$**.**

S (Z = 16) **:** $1s^2\ 2s^2\ 2p^6\ \mathbf{3s^2\ 3p^4\ 3d^0}$

S*(Z = 16) **:** $1s^2\ 2s^2\ 2p^6\ \mathbf{3s^1\ 3p^3\ 3d^2}$

6 single electrons

P (Z = 15) **:** $1s^2\ 2s^2\ 2p^6\ \mathbf{3s^2\ 3p^3\ 3d^0}$

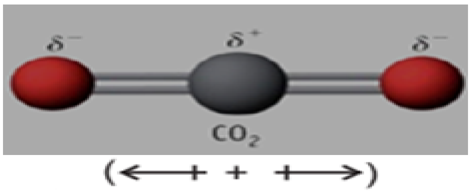
P*(Z = 15) **:** $1s^2\ 2s^2\ 2p^6\ \mathbf{3s^1\ 3p^3\ 3d^1}$

5 single electrons

Sulfur and phosphorus are 3rd period atoms, so they can accommodate more than 8 electrons using 3d atomic orbitals.

Exercise 7:

- **$\mathbf{CO_2}$:**
  nonpolar molecule: either has nonpolar bonds or bonds wose dipoles cancel to zero, for example: the dipole moment in case of CO_2 is zero.



The molecule is nonpolar.

- **$\mathbf{I_2}$:**

When covalent bond is formed between two similar atoms, for example in H2, O2, Cl2, N2, F2 or I_2, the shared pair of electrons is equally attracted by the two atoms. As a result electron pair is situated exactly between the two identical nuclei. The bond so formed is called nonpolar covalent bond.

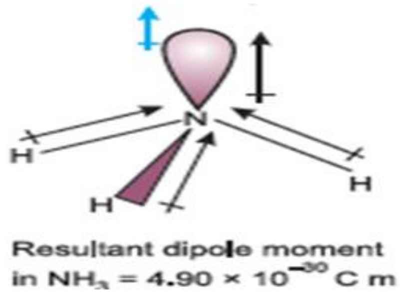
The molecule is nonpolar.

- **$\mathbf{NH_3}$:**

The molecule has pyramidal shape with a lone pair of electron on nitrogen atom.

The dipole moment of NH_3 molecule is 1.49 D.

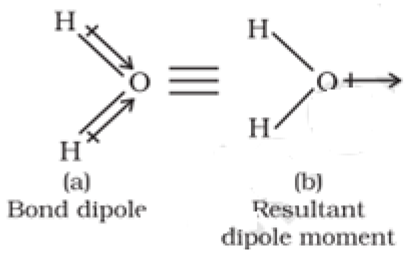
The molecule is polar.



- **H_2O :**

Water molecule has a bent structure with the two OH bonds oriented at an angle of 104.5°. The dipole moment of water is μ = 1.84D, which is the resultant of the dipole moments of two O-H bonds.

The molecule is polar.

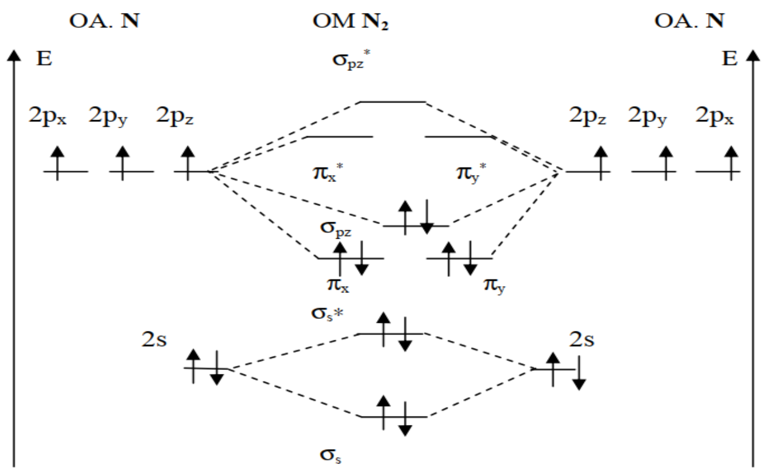


- **CF_4:** The molecule is nonpolar.

Exercise 8:

Energy diagram of the molecular orbitals (OM) of the homonuclear molecule N_2: (N: Z=7)

There are interactions between the s and p atomic orbitals which involve an inversion of the π_x and π_y molecular orbitals with the σ_{pz} molecular orbital.



The electronic structure of the N2 molecule is: $\sigma_s^2\ \sigma_s{*}^2\ (\pi_x^2 = \pi_y^2)\ \sigma_{pz}^{\ 2}$

Energy diagram of the molecular orbitals (MO) of the heteronuclear CN molecule:

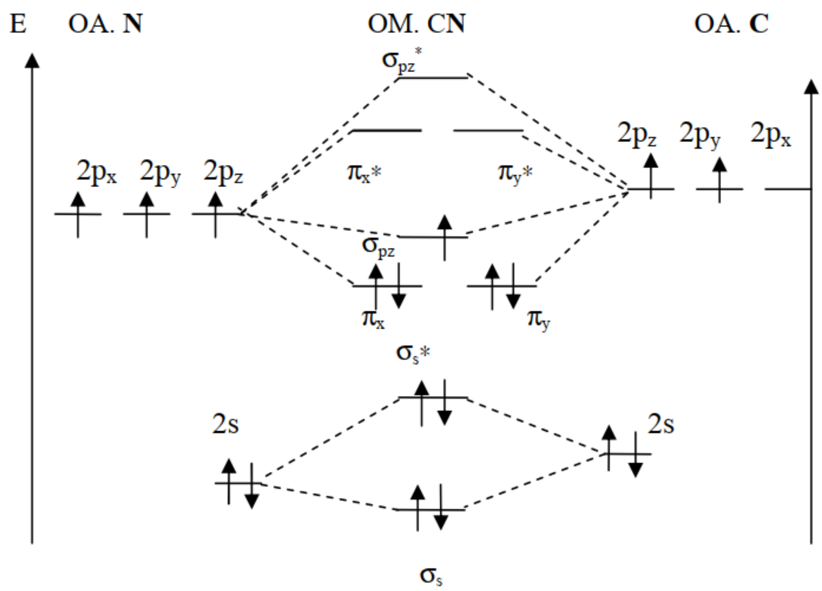
For heteronuclear molecules formed from atoms of the second period of the periodic table, there are interactions between the s and p atomic orbitals which involve an inversion of the πx and πy molecular orbitals with the σpz molecular orbital.

N(Z = 7) : 1s² $2s^2\ 2p^3$ **C(Z = 6) : 1s²** $2s^2\ 2p^2$

The nitrogen atom N is more electronegative than the carbon atom.

The values of the energies of the atomic orbitals of nitrogen are therefore lower than those of carbon.

The energy diagram, in this case, is asymmetrical.



The electronic structure of the CN molecule is: $\sigma_s^2\ \sigma_s{*}^2\ (\pi_x^2 = \pi_y^2)\ \sigma_{pz}^1$

2. Bond order:

Bond order (B.O.) = 1/2 [Nb-Na]

i(N_2) = ½(8-2) = 3

i(CN) = ½(7-2) = 2,5

3. The magnetic properties of N_2 and CN.

The presence of single electrons in the molecular orbitals of CN makes this molecule **paramagnetic**.

In the N_2 molecule, we find that the electrons are paired; which gives N_2 a **diamagnetic** character.

Exercise 9:

1. State of hybridization of carbon and boron atoms.

The overlap of hybridized atomic orbitals (sp^3, sp^2, sp) forms sigma σ bonds (axial overlap).

The overlap of p atomic orbitals forms π bonds (lateral overlap).

B(Z = 5) : $1s^2$ **$\underline{2s^2\ 2p^1}$**

B* (excité): $1s^2$ **$\underline{2s^1\ 2p^2}$**

Three valence electrons ($2s^1$ $2p^2$) provide three simple bonds but of different natures.

To ensure three identical bonds, the atomic orbitals of B (s and p) combine with each other to form three (identical) sp^2 hybridized orbitals.

If the bonds are formed from overlapping atomic orbitals pure (s and p), the angle between the bonds would then be 90°:

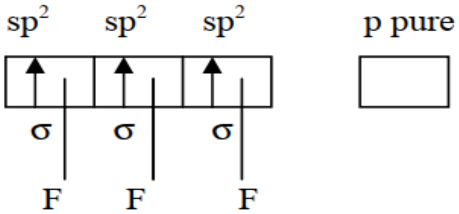
OA s (B) + OA p (F)

OA p (B) + OA p (F)

To ensure three identical bonds, the atomic orbitals of B (s and p) combine with each other to form three (identical) sp^2 hybridized orbitals.

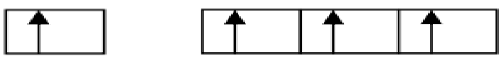
The angles between the connections become equal (120°). The molecule thus formed is planar.

BF_3: B sp^2 hybridized

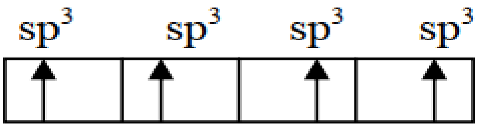


C (Z = 6): $1s^2$ **$2s^2$ $2p^2$**

C^*(excited): $1s^2$ **$2s^1$ $2p^3$**

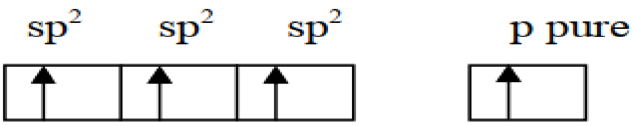


C sp^3 hybridized



The s orbital forms with the three p orbitals, four sp^3 hybridized orbitals which will ensure four σ bonds.

C sp^2 hybridized



The s orbital forms with the two p orbitals, three sp^2 hybridized orbitals which will provide three σ bonds, the "pure" p orbital forms a π bond.

C sp hybridized

The s orbital forms with the p orbital, two sp hybridized orbitals which will provide two σ bonds. The two "pure" p orbitals form two π bonds.

In the CO_2 molecule, the carbon is sp hybridized;

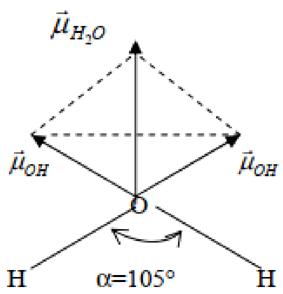
In the molecule; CH_4:, carbon is sp^3 hybridized.

In the C_2H_4 molecule:, the carbon is sp^2 hybridized.

In the C_2H_2: molecule, the carbon is sp hybridized.

Exercise 10:

1. Dipole moment of the H_2O molecule:



Since oxygen is more electronegative than hydrogen, the O-H bond is polarized.

Therefore, there is a dipole moment $\overrightarrow{\mu_{OH}}$ with the direction of each O-H bond, the direction being conventionally directed from positive charges to negative charges.

By adding the two $\overrightarrow{\mu_{OH}}$ vectors, we obtain the dipole moment $\overrightarrow{\mu_{H2O}}$of the molecule, which is directed along the bisector of the angle HÔH.

$$\vec{\mu}_{H_2O} = 2\vec{\mu}_{OH}$$

$$\mu_{H2O} = 2\mu_{OH} \, . \, \cos(\alpha / 2) = 2 \, . \, 1{,}51 \cos(105 / 2) = 1{,}84 \, Debye$$

2. Ionic percentage of the O-H bond in H_2O.

Ionic percentage = (experimental μ/theoretical μ).100%

The theoretical dipole moment: $\mu = \delta.e.d.\ 4.8$ (Debye) with $\delta = 1$

$\mu_{O-H} = 1.51D$ and $d_{O-H} = 0.96$ Å.

Ionic % = [1.51/(1.0.96 × 4.8)] × 100 = 32.8

The bond is 33% ionic.

Conclusion:

Lewis gave the first explanation of covalent bonding in terms of atoms exchanging electron pairs, and he connected the process to the creation of noble gas configurations by atoms reacting as a result of electron sharing.The Lewis dot symbols show the number of valence electrons of the atoms of a given element and Lewis dot structures show pictorial representations of bonding in molecules.

Maximum stability is achieved when the outer shell is filled, as in rare or noble gases (8 e- in the outer shell). With both ionic and covalent bonds, atoms tend to acquire a stable electronic configuration.

The VSEPR model used for predicting the geometrical shapes of molecules is based on the assumption that electron pairs repel each other and, therefore, tend to remain as far apart as possible. According to this model, molecular geometry is determined by repulsions between lone pairs and lone pairs; lone pairs and bonding pairs and bonding pairs and bonding pairs.

General Conclusion

Scientists are constantly discovering new compounds, orderly arranging the facts about them, trying to explain with the existing knowledge, organising to modify the earlier views or evolve theories for explaining the newly observed facts.

Understanding the structure of matter is crucial in many fields:

• Chemistry: It allows us to predict chemical reactions, design new molecules, and develop innovative materials.

• Biology: It sheds light on the fundamental mechanisms of life, from the structure of DNA to the interactions between proteins.

• Materials science: It guides the design of materials with specific properties, from semiconductors to polymers.

• Medicine: It is the basis for the creation of new drugs.

In summary, the structure of matter is a fundamental field of study that allows us to understand the properties and behaviors of the matter around us.

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