



CHEMISTRY

For second year secondary

Second term

2025 - 2026



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Introduction

In the context of educational development to keep pace with global and local changes, and as a continuation of the diligent efforts made by the Ministry of Education to raise the level of curriculum content and link it to society and the environment, His Excellency the Minister of Education has assigned the re-evaluation and review of the scientific content of chemistry for the second year of secondary school.

The assigned committee has made the necessary amendments and additions, which led to:

- (1) Eliminating unnecessary repetition and redundancy, excluding parts previously studied by the student, and rephrasing some sections of the book in a logical, sequential, and organized manner.
- (2) Adding some concepts and applications to keep up with modern scientific trends.
- (3) Linking study topics to daily life, their environmental impacts, and industrial applications.
- (4) Focusing on mathematical treatments in understanding some topics to strengthen intellectual aspects.
- (5) Preparing some illustrative figures and employing them to serve scientific concepts.
- (6) The desired objectives for studying each chapter of the book have been set out in its introduction to provide an indicator for the student and teacher of what has been achieved.
- (7) Diversifying assessment to include measuring different levels of education.

In its current form, the book contains three chapters that are integrated and interconnected, reflecting harmony with the contents of chemistry books in international curricula and including useful industrial and environmental applications.

It clearly focuses on developing understanding, analytical, and innovative abilities, and aligns with the national standards set by the Ministry for the development of the chemistry curriculum.

We hope that this book in its new form will be a useful source of science and knowledge in the field of chemistry, achieve the desired goal, and be a great aid to our students, to whom we wish success and prosperity.

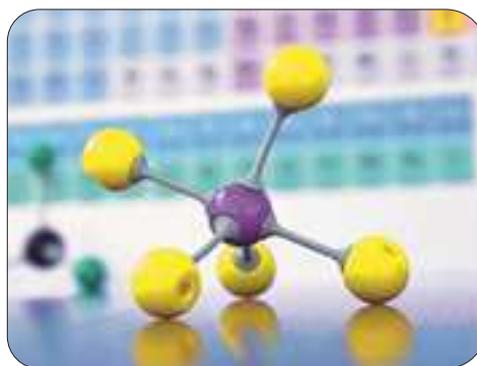
Development Committee

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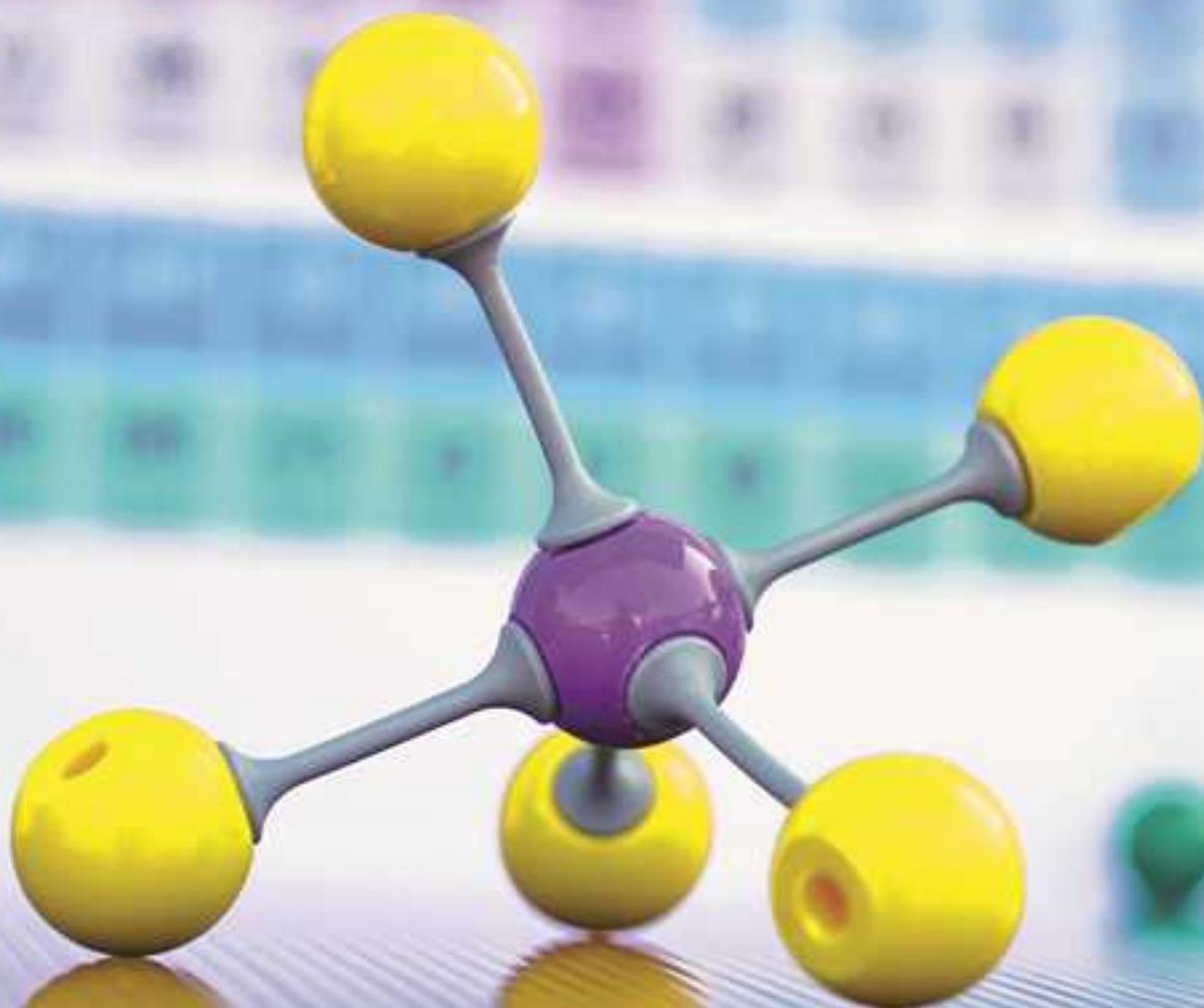
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Chapter One

Bonds and spatial shapes of molecules



Objectives

At the end of the student's study of the chapter on bonds and spatial shapes of molecules, he should be able to:

- 1- Explains why most atoms form chemical bonds.
- 2- Describes both ionic bonds and covalent bonds.
- 3- Determines the type of bond based on electronegativity differences.
- 4- Identifies applications in science, technology, and society.
- 5- Explains the electronic theory of valence (the Octet Rule).
- 6- Identifies the drawbacks of the Octet Rule.
- 7- Explains the reason for covalent bond formation in hydrogen and hydrogen fluoride molecules based on Valence Bond Theory.
- 8- Identifies the concept of hybridization and how it occurs.
- 9- Explains how hydrogen atoms bond with a carbon atom to form a methane molecule
- 10- Compare between sigma and pi bonds
- 11- Determines hybridization type in each of methane, ethylene, and acetylene.
- 12- Determines molecular shapes of molecules in light of the valence shell electron pair repulsion theory.
- 13- Determines the donor and acceptor atoms when forming a coordinate bond.
- 14- Defines the hydrogen bond.
- 15- Explains the reason for the high boiling point of water.
- 16- Infers metal properties, such as hardness and high melting points, from the number of free valence electron.

Chemical combination

- You have previously studied the detailed structure of the atom, and you have learned that the most stable atoms of the elements are the atoms of noble gases, such as: helium, neon, and argon ... etc., due to the stability of their electronic configuration, **as illustrated in Table (1)**:

Table (1)

Noble gas	Electronic configuration
${}^2\text{He}$	$1s^2$
${}^{10}\text{Ne}$	$[\text{He}], 2s^2, 2p^6$
${}^{18}\text{Ar}$	$[\text{Ne}], 3s^2, 3p^6$
${}^{36}\text{Kr}$	$[\text{Ar}], 4s^2, 3d^{10}, 4p^6$
${}^{54}\text{Xe}$	$[\text{Kr}], 5s^2, 4d^{10}, 5p^6$
${}^{86}\text{Rn}$	$[\text{Xe}], 6s^2, 4f^{14}, 5d^{10}, 6p^6$

In light of the above, it is clear that for the stability of the remaining elements, it is necessary to complete their outer energy level with electrons, through losing, gaining, or sharing electrons in chemical reactions so that their electronic configuration becomes similar to the electronic configuration of the nearest noble gas to them in atomic number.

Lewis dot representation for valence electrons

Valence electrons play an important role in the nature of the bonds formed by atoms, and the scientist Lewis devised a simplified method in which he used dots to represent (valence electrons).

As shown in Table (2):

Table (2)

Group	1A	2A	3A	4A	5A	6A	7A	0
The third period	${}^{11}\text{Na}$	${}^{12}\text{Mg}$	${}^{13}\text{Al}$	${}^{14}\text{Si}$	${}^{15}\text{P}$	${}^{16}\text{S}$	${}^{17}\text{Cl}$	${}^{18}\text{Ar}$
Electronic configuration	$[\text{Ne}], 3s^1$	$[\text{Ne}], 3s^2$	$[\text{Ne}], 3s^2, 3p^1$	$[\text{Ne}], 3s^2, 3p^2$	$[\text{Ne}], 3s^2, 3p^3$	$[\text{Ne}], 3s^2, 3p^4$	$[\text{Ne}], 3s^2, 3p^5$	$[\text{Ne}], 3s^2, 3p^6$
Lewis dot model	$\cdot\text{Na}$	$\cdot\text{Mg}\cdot$	$\cdot\text{Al}\cdot$	$\cdot\text{Si}\cdot$	$\cdot\ddot{\text{P}}\cdot$	$\cdot\ddot{\text{S}}\cdot$	$:\ddot{\text{Cl}}\cdot$	$:\ddot{\text{Ar}}:$

Lewis called the pair of electrons present in one of the orbitals of the outer level, which did not participate in forming bonds, the term **lone pair** ; and the pair of electrons responsible for forming the bond, the term **bond pair** .

And in light of our knowledge about the structure of the atom, we will study two types of bonds, **which are:**

First: chemical bonds.

Second: physical bonds.

Firstly Chemical bonds

The type and strength of the chemical bond depend on the electronic structure of the atoms forming it.

We will study three types of chemical bonds, **which are:**

- ① The ionic bond.
- ② The covalent bond.
- ③ The coordinate bond.

① The ionic bond

- **The ionic bond** is an electrostatic attraction between a positively charged ion (cation) and a negatively charged ion (anion), as illustrated by the ionic bonding in the molecule of **sodium chloride** (Figures 2,1).

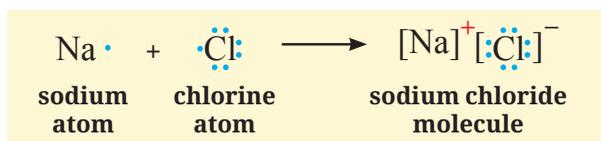


Figure (2) Ionic bonding by the Lewis dot method

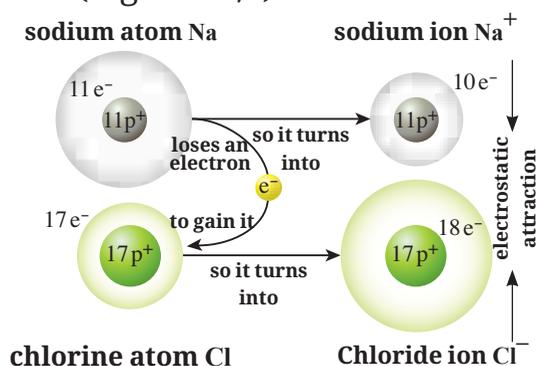


Figure (1) Ionic bonding in the sodium chloride molecule

- The difference in electronegativity between the bonded elements plays a fundamental role in determining the percentage of ionic character in their compounds, and when the percentage of ionic character in the compound **is greater than 50%**, it is described as an **ionic compound**. Experimental evidence has shown that the greater the difference in electronegativity between the two bonded elements exceeds (1.7), the ionic character increases in the resulting compound (Figure 3).

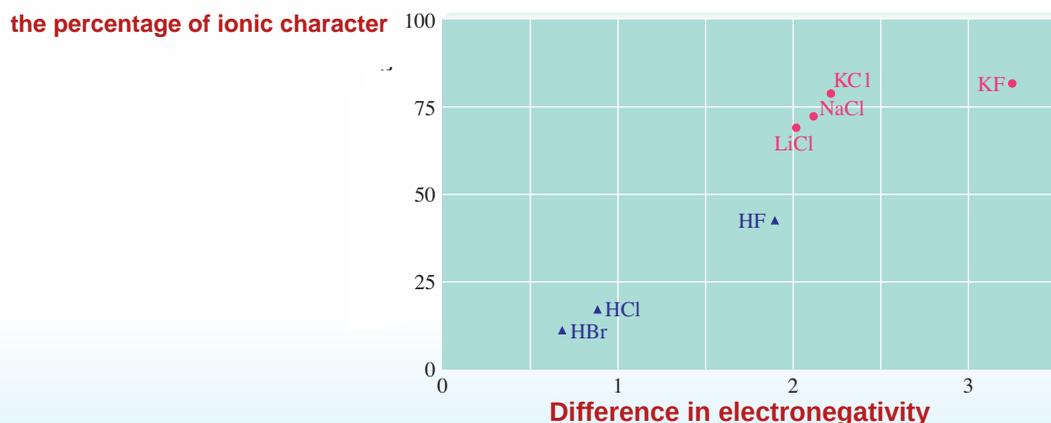


Figure (3)

- The ionic character in the compound KCl is stronger than in the compound LiCl , because the difference in electronegativity between potassium and chlorine (2.2) is greater than that between lithium and chlorine (2). Ionic bonds are strong when any of the metals of groups (1A) and (2A) bond with any of the nonmetals of groups (6A) and (7A).

Properties of ionic compounds

The difference in electronegativity of the elements forming the ionic bond affects the ionic properties of their compounds, and the arrangement of ions in the crystal lattice of ionic compounds (Figure 4) also affects their following properties:

- ① Solids, most of which are soluble in polar solvents such as water, forming aqueous solutions, and do not dissolve in organic (nonpolar) solvents such as benzene.
- ② They conduct electric current when in the form of a melt or aqueous solution due to the ease of movement of free or hydrated ions, and do not conduct electric current when in the solid state, due to the difficulty of ion movement inside the crystal lattice, because of the strong attractive forces between ions of opposite charge.
- ③ High melting and boiling points, to overcome the attractive forces between cations and anions in the crystal lattice.
- ④ Most of their reactions occur instantaneously.



Figure (4) the crystal lattice in sodium chloride

Science, Technology, and Society (STS)

the relation between ions and heart health:

It arises from the movement of sodium ions Na^+ , potassium (K^+), and calcium (Ca^{2+}) through the heart membranes, **ion currents are converted into measurable electronic current in the electrocardiogram (ECG)** (Figure 5). An imbalance in the concentration of these ions in the body's cells leads to a change in the electrical signals produced by the heart muscle.

The results of this test help diagnose heart attacks and arrhythmias.

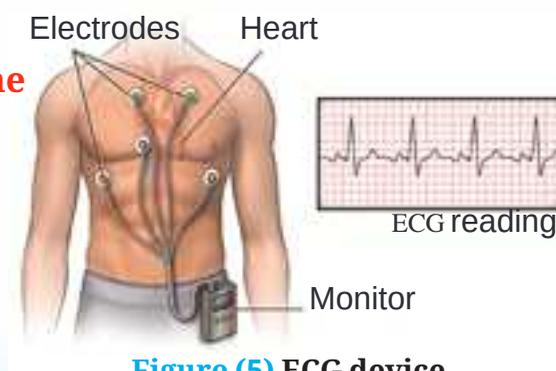


Figure (5) ECG device "electrocardiograph"

2 Covalent bond

The covalent bond is formed between two identical or different atoms where the difference in electronegativity between them is **less than 1.7** usually, and the covalent bond arises through **sharing of electrons** between the bonded atoms. The difference in electronegativity of the two atoms bonded covalently leads to a certain percentage of **ionic character** in the covalent compound and the **covalent bond is classified into:**

- (A) **Pure covalent bond.**
- (B) **Nonpolar covalent bond.**
- (C) **Polar covalent bond.**

(A) Pure covalent bond:

The **pure** covalent bond is formed between two atoms with the same electronegativity, so it is often formed between two identical atoms of a single nonmetal element, **as illustrated by the examples shown in Figure (6)**

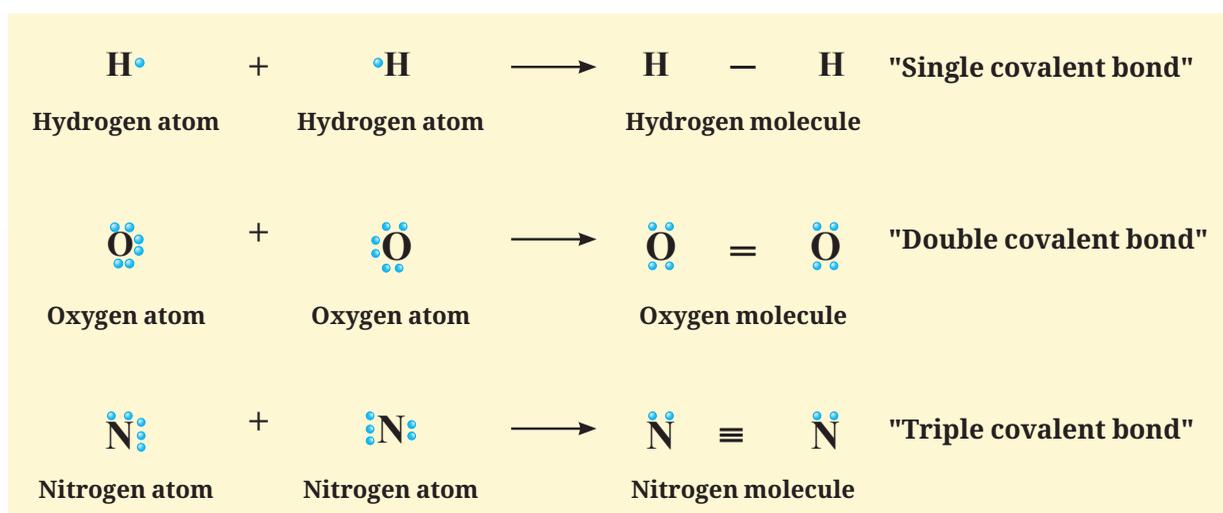
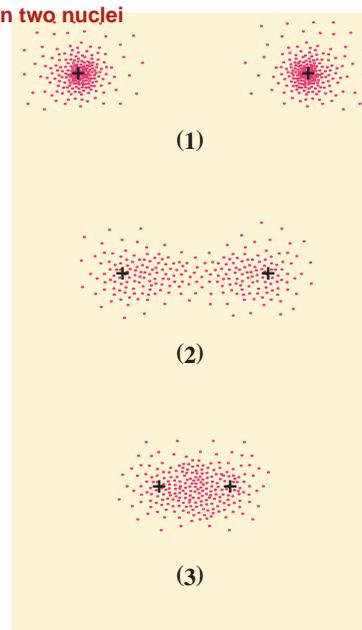


Figure (6) Formation of the pure covalent bond

Explanation of bond formation in the hydrogen molecule H_2

- ① Attraction occurs between the electrons and protons of the two atoms, leading to their approach.
- ② The change in distance between the two atoms as a result of their attraction leads to a change in the shape of the electron density between them, and the electron density increases as the distance between them decreases.
- ③ The electron density becomes maximum when the two atoms are at a certain distance where the potential energy of the molecule is **at its minimum**, and the energy of the H_2 molecule is **less than** the sum of the energies of the two unbonded H atoms (**Figure 7**).

Decrease the distance between two nuclei



(Figure 7)
Formation of the bond
in the molecule H_2

B Nonpolar covalent bond:

The **nonpolar** covalent bond is formed between two different atoms of two nonmetal elements where the difference in electronegativity between them **does not exceed 0.4** as illustrated by the two examples shown in **Figure (8)**.

Si	—	P		C	—	H
1.8		2.1	electronegativity	2.5		2.1
	0.3		electronegativity difference		0.4	

Figure (8)

Hydrocarbons (compounds of carbon and hydrogen only) such as methane CH_4 (**Figure 9**) are described as **nonpolar compounds**.

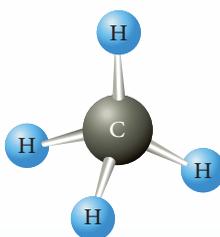


Figure (9) Nonpolar
methane molecule

C Polar covalent bond:

The **polar** covalent bond is formed between two different atoms of two elements, which may both be nonmetals or one may be a metal and the other a nonmetal, and the difference in electronegativity between them is **greater than 0.4 and less than 1.7** such as the bond formed in the hydrogen chloride molecule between a chlorine atom and a hydrogen atom (Figure 10).

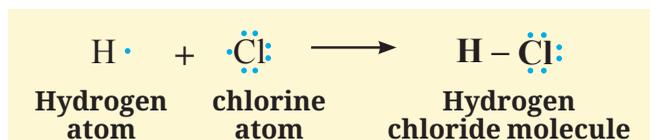


Figure (10) Formation of the polar covalent bond in the hydrogen chloride molecule

And due to the high electronegativity of chlorine (3) compared to the electronegativity of hydrogen (2.1), the electrons of the bond formed between them are displaced toward chlorine (i.e., they spend more time around the chlorine atom), resulting in a partial negative charge ($-\delta$) on the chlorine atom and a partial positive charge ($+\delta$) on the hydrogen atom (Figure 11). Therefore, polar molecules are attracted to electric fields, and the polarity of the molecules increases with the increase in the difference in electronegativity between their elements.

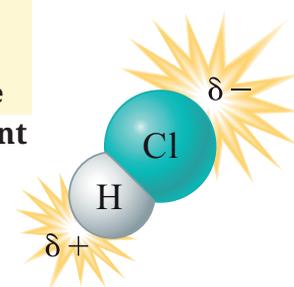


Figure (11) molecule HCl polar

The polar covalent bond is evident in many molecules, such as:

Hydrogen halides HX, water H_2O , ammonia NH_3 , aluminium chloride AlCl_3 , and titanium tetrachloride TiCl_4

And some molecules are **nonpolar** despite containing polar bonds, because their spatial shape leads to the cancellation of the polar effect of these bonds (i.e., the resultant dipole moment equals zero).

This is evident in the linear shape of the carbon dioxide molecule CO_2 and the trigonal planar shape of the BF_3 molecule

Properties of covalent compounds

- ① Most of them dissolve in organic solvents, such as benzene and carbon tetrachloride, and most do not dissolve in polar solvents.
- ② Solid, liquid, or gas substances do not conduct electricity because they are in a non-ionized form, except that some of their aqueous solutions conduct electricity, such as: $\text{HCl}_{(\text{aq})}$
- ③ Their melting and boiling points are often low due to the weak attractive forces between their molecules.
- ④ Most of their chemical reactions are relatively slow.

Science, Technology, and Society (STS)

Covalent drugs:

- They are a type of drug designed to form permanent **covalent** bonds with the DNA or the protein of the disease-causing virus, thereby inhibiting its function (Figure 12).



Figure (12)

The covalent bonding between the viral DNA and the covalent drug

- The **mass spectrometer** is an important tool used in the development of covalent drugs used to treat viral diseases such as COVID-19, as it determines the molecular mass of the covalent drug.

Before and after its binding to the viral protein to determine the active part of the drug and the atom to which it binds in the viral protein.

- It is also used to verify the possibility of the drug binding covalently with other non-target proteins in the body, which reduces potential risks when using the drug and helps in designing drugs that are safe for health.

Theories explaining the covalent bond

- There is more than one theory developed to explain the covalent bond according to our changing understanding of the properties of the electron, and we will briefly present some of the theories that have been proposed to explain the covalent bond.

1 The electronic theory of valence (octet theory)

- The **octet theory** is also known as the electronic theory of valence, or the **Kossel–Lewis theory**, and states that all atoms of elements—except hydrogen, lithium, and beryllium—tend to achieve the stable octet configuration of the outer energy level (Figures 13 - 15).

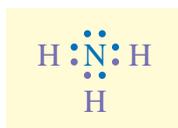


Figure (15)
Ammonia molecule NH_3



Figure (14)
Water molecule H_2O



Figure (13)
Chlorine molecule Cl_2

Drawbacks of the electronic theory of valence

- It does not apply to many molecules when explained based on the octet rule, **such as**:
 - Boron trifluoride molecule BF_3 , in which the boron atom is surrounded by only six electrons (Figure 16).
 - Phosphorus pentachloride molecule PCl_5 , in which the phosphorus atom is surrounded by ten electrons (Figure 17).
 - Sulfur hexafluoride molecule SF_6 , in which the sulfur atom is surrounded by twelve electrons (Figure 18).

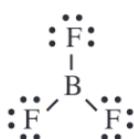


Figure (16) Boron trifluoride molecule

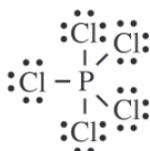


Figure (17) Phosphorus pentachloride molecule

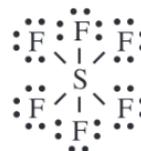
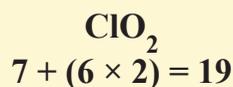


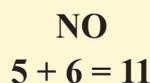
Figure (18) Sulfur hexafluoride molecule

- It does not apply to molecules whose total number of valence electrons of their atoms is **odd**, such as:

Chlorine dioxide



Nitric oxide



- Considering the covalent bond as merely a pair of shared electrons is no longer sufficient to explain many properties of molecules, such as the spatial shape of the molecule and the angles between the bonds in it.

Rules for drawing Lewis structures for molecules and ions with a single central atom:

The following steps should be observed:

- 1- Determine the valence electrons of each element in the molecule or atomic group, then calculate their total, let it be (A), noting that in the case of **negative** atomic groups, **add** to the total a number representing the magnitude of their charge, and in the case of **positive** atomic groups, **subtract** from the total a number representing the magnitude of their charge.
- 2- Calculate the total number of electrons needed to fill the valence shell of each atom (2 for hydrogen, 8 for the other elements), let it be (B).
- 3- Calculate the number of electrons shared in forming bonds, let it be (C), from the relation: $C = B - A$
- 4- Calculate the number of bonds, let it be (D), from the relation: $D = \frac{C}{2}$
- 5- Calculate the number of unshared (lone) pairs of electrons, let it be (E), From the relation: $E = \frac{A - C}{2}$
- 6- Draw the central atom that forms the largest number of covalent bonds with the terminal atoms, then connect them with the number (D) of covalent bonds.
- 7- Distribute the lone pairs of electrons (E) on the different atoms, so that each is surrounded by eight electrons (and hydrogen by 2 electrons), taking into account adding the amount and type of charge on the **atomic groups**.

Example (1)

$$N = 7$$

$$F = 9$$

Draw the Lewis structure for the nitrogen trifluoride molecule NF_3

Solution

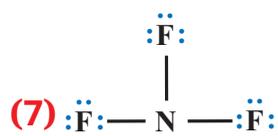
$$(1) A = 5 + (7 \times 3) = 26 e^-$$

$$(2) B = (1 \times 8) + (3 \times 8) = 32 e^-$$

$$(3) C = 32 - 26 = 6 e^-$$

$$(4) D = \frac{6}{2} = 3$$

$$(5) E = \frac{26 - 6}{2} = 10 e^-$$



Example (2) Draw the Lewis structure for the nitrite ion NO_2^-

$$\text{N} = 7$$

$$\text{O} = 8$$

$$(1) A = 5 + (6 \times 2) + 1 = 18 e^-$$

$$(2) B = 8 + (8 \times 2) = 24 e^-$$

$$(3) C = 24 - 18 = 6 e^-$$

$$(4) D = \frac{6}{2} = 3$$

$$(5) E = \frac{18 - 6}{2} = 6 e^-$$



2 Valence bond theory

- The valence bond theory states that the covalent bond is formed when two atomic orbitals, each containing a single electron, overlap, whether these atomic orbitals are pure, hybridized, or both.



Application (1): Bond formation in the chlorine molecule Cl₂

The chlorine molecule is formed as a result of the overlap of the two chlorine atoms by the single electron in the 2p orbital of each (Figure 19).

An overlap region is formed between the nuclei of the two chlorine atoms, where the electron density increases and the electron pair is attracted by the nuclei of both atoms simultaneously, and the electron density is distributed symmetrically along the axis connecting the two nuclei.

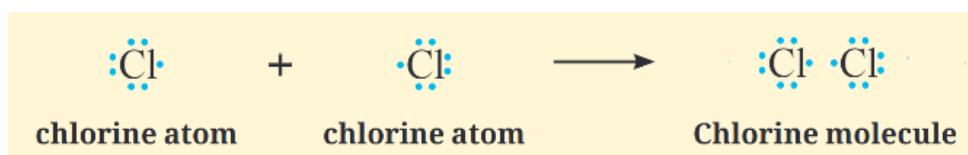


Figure (19) Formation of the chlorine molecule according to the valence bond theory



Application (2): Bond formation in the hydrogen fluoride molecule HF

The hydrogen fluoride molecule is formed as a result of the overlap of one of the 2p sublevel orbitals containing a single electron from the fluorine atom with the 1s orbital containing a single electron from the hydrogen atom (Figure 20).

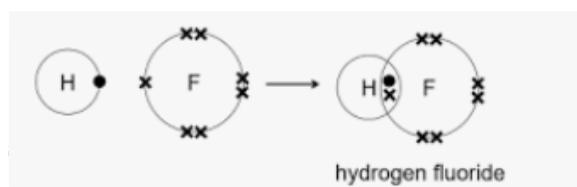


Figure (20) Formation of the hydrogen fluoride molecule according to the valence bond theory

Hybrid orbitals

The valence bond theory in its simple form (just orbital overlap) failed to explain covalent bonding in many molecules, **such as:** the methane molecule CH₄, the ethene molecule C₂H₄, the ethyne molecule C₂H₂, and the water molecule H₂O. However, it explained the forming bonds in these molecules and others by the occurrence of hybridization (mixing) of atomic orbitals before their union (the concept of hybrid orbitals).

Hybridization

It is the overlap between two or more different orbitals in the same atom, resulting in new orbitals known as **hybrid orbitals**.

It is noted that: hybridization occurs in the central atom of the molecule between the orbitals of the same atom that are close in energy, such as: (2s) with (2p) or (3s) with (3p) to produce a number of hybrid orbitals equal to the number of pure orbitals that entered the hybridization, and the energy of the hybrid orbitals **is equal**.

The shape of the hybrid orbitals differs from the pure orbitals that formed them, as they are more protruding outward, making their ability to overlap stronger than that of pure orbitals.

Types of hybridization

There are various types of atomic hybridization, and we will suffice with studying its types that occur between an orbital from the sublevel (ns) and one or more orbitals from the sublevel (np),

Hybridization may occur between orbitals with single electrons as in the carbon atom of methane, or between orbitals with a pair of electrons and orbitals with single electrons as in the nitrogen atom of ammonia.

This will become clear in the applications of hybridization of the sp^3 type.

The type of hybridization occurring in the molecule is determined from its Lewis structure and from the number of regions of electron density around its central atom, which represent the total number of bonding electron pairs and lone electrons.

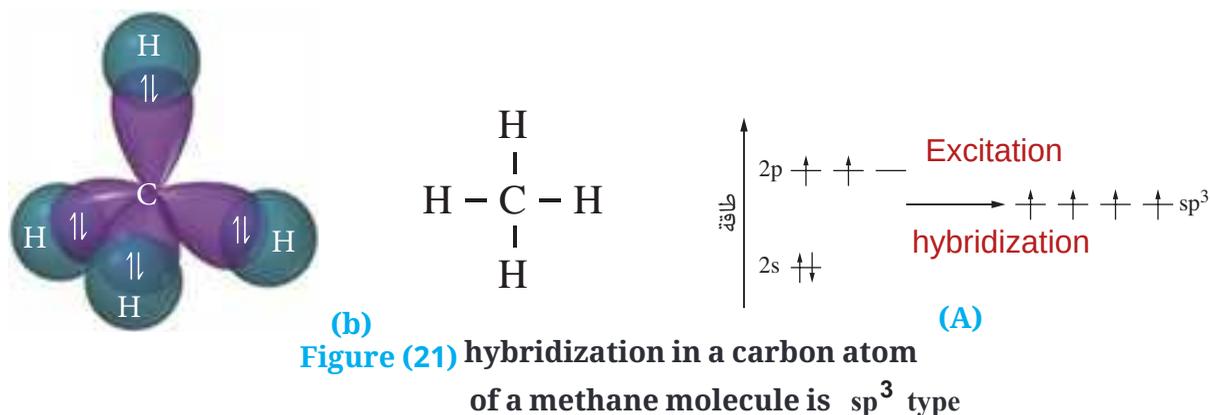
1 Hybridization sp^3

Hybridization occurs between an orbital (ns) and three orbitals (np) in the same atom to form **four** hybrid orbitals of the type (sp^3).



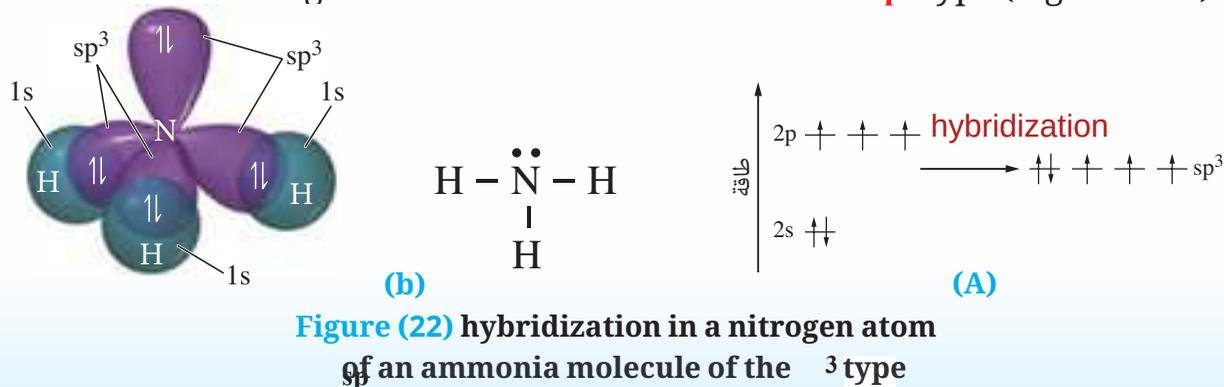
Application (1): Hybridization in the carbon atom of the methane molecule CH_4

The methane molecule contains CH_4 has four bonds that are identical in length and energy. This is achieved by exciting an electron from the 2s to the subshell 2p in a carbon atom, followed by hybridization between the orbitals 2s with a single electron and three orbitals 2p with single electrons to form four hybrid orbitals of the type sp^3 type (Figure 21), each of which is bonded to a hydrogen atom by a pair of bonding electrons, meaning that there are **four** areas of electron density around the carbon atom, so the hybridization that occurs in the carbon atom of the methane molecule is of the sp^3 type (Figure 21 b).



Application (2): hybridization in a nitrogen atom of an ammonia molecule NH_3

In the ammonia molecule NH_3 it has three identical bonds, which is achieved through a hybridization process between the orbitals 2s with a pair of electrons and three orbitals 2p, each with a single electron, to form four hybrid orbitals of the type sp^3 type (Figure 22 a, each of which is bonded to a hydrogen atom by a pair of bonding electrons, with a pair of free electrons remaining on the nitrogen atom, meaning that there are **four** areas of electron density around the nitrogen atom. Therefore, the hybridization that occurs in the nitrogen atom is considered to be of the sp^3 type (Figure 22 B).



Sigma bonds and pi bonds

• sigma bond (σ):

The bond formed by the **head-on** overlap between any two orbitals is described as a **sigma bond (σ)** (Figure 23).

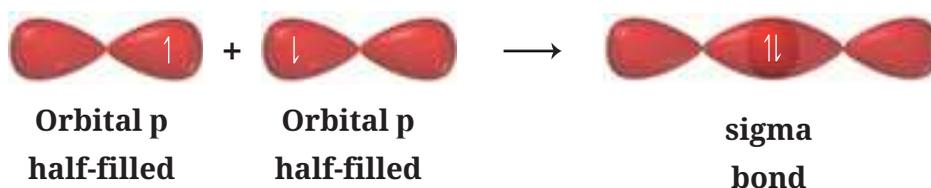


Figure (23) The sigma bond arises from the head-on overlap of two orbitals

The overlap may occur between two pure orbitals, or a pure orbital and a hybrid orbital (as in the CH_4 molecule and the NH_3 molecule), or between two hybrid orbitals (as will be discussed in hybridization of the sp^2 or sp type).

All **single** covalent bonds are of the **sigma** type, and they are strong bonds that are difficult to break.

• pi bond (π):

The bond formed by the **sideways** overlap between any two pure orbitals is called a **pi bond (π)** (Figure 24).

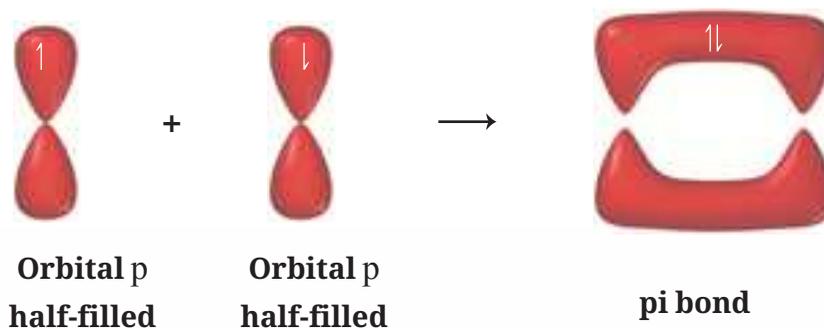


Figure (24) The pi bond arises from the sideways overlap of two orbitals, and the pi bond is weak and easy to break

2 hybridization sp^2

Hybridization occurs between the (ns) orbital with the (np_x) and the (np_y) orbitals to form **three** hybrid orbitals of the (sp^2) type.



Application: hybridization in a carbon atom of an ethene molecule C_2H_4

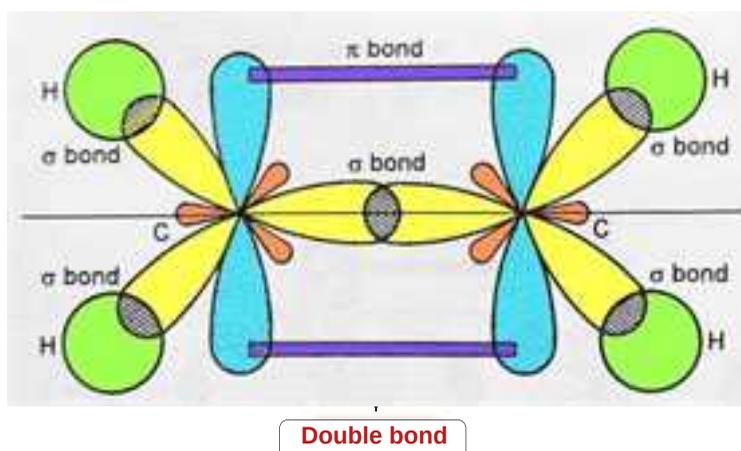
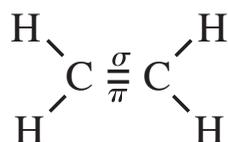


Figure (25) hybridization in a carbon atom of an ethene molecule of the sp^2 type

It is observed in Figure (25) that each carbon atom in the ethene molecule is surrounded **by three** regions of electron density, so the hybridization is of the sp^2 type, and there is a **double bond between the two carbon atoms of the ethene molecule**. All double bonds consist of a sigma bond that is difficult to break and a pi bond that is easy to break.

3 Hybridization sp

Hybridization occurs between the (ns) orbital and the (np_x) orbital to form **two** hybrid orbitals of the sp type



Application: hybridization in a carbon atom of an ethyne molecule C_2H_2

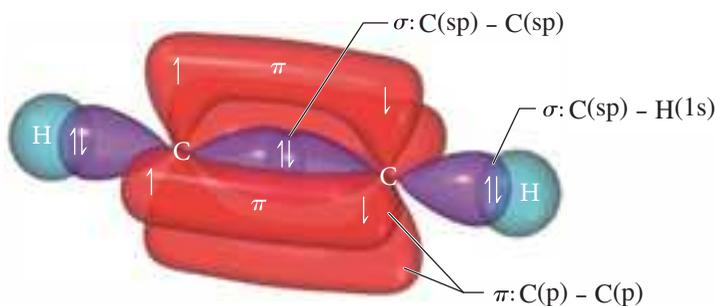


Figure (26) hybridization in a carbon atom of an ethyne molecule of the sp type

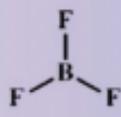
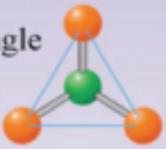
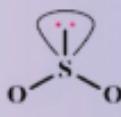
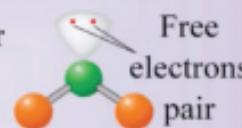
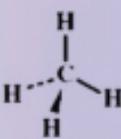
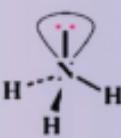
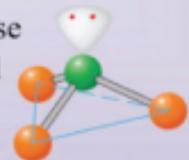
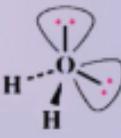
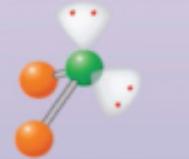
It is observed in Figure (26) that each carbon atom in the ethyne molecule is surrounded **by two** regions of electron density, so the hybridization is of the sp type, and there is a **triple bond** between the two carbon atoms of the ethyne molecule.

All triple bonds consist of a sigma bond and two pi bonds.

Valence Shell Electron Pair Repulsion Theory VSEPR

Table (3) illustrates the shapes of some molecules according to the Valence Shell Electron Pair Repulsion Theory.

Table (3)

Examples of particles	Shape of the solid part	Arrangement of the electron pairs	Electron pairs					
			Free	Bonded	Resultant			
$\text{F}-\text{Be}-\text{F}$ BeF_2	Linear AX_2 	Linear	0	2	2			
 BF_3	Plane triangle AX_3 	Planner Triangle	0	3	3			
 SO_2	Angular AX_2E  Free electrons pair					1	2	
 CH_4	Tetrahedron AX_4 	Tetrahedral	0	4				
 NH_3	three - base pyramid AX_3E 					1	3	4
 H_2O	angular AX_2E_2 					2	2	

A : represents the central atom ●

X : represents the atoms bonded to the central atom ●

E : represents the lone pairs of electrons.

Lone pairs of electrons control the values of the angles between the bonds in the molecule, because the lone pair is attached on one side to the nucleus of the central atom and is spatially spread out on the other side.

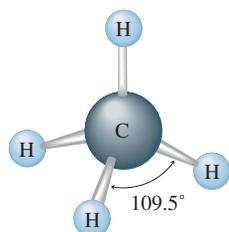
As for the bonding pair, it is attached on both sides to the nuclei of the two bonded atoms.

An increase in the number of lone pairs of electrons in the central atom of the molecule leads to an increase in the repulsive forces between them, which results in a decrease in the bond angles between the covalent bonds in the molecule (Figures 29 - 27).

In general, the repulsion is as follows:

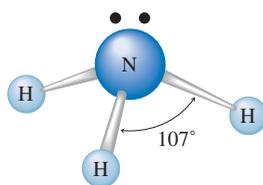
(lone pair, lone pair) > (lone pair, bonding pair) > (bonding pair, bonding pair)

How can you explain the trend in the values of the bond angles in water, ammonia, and methane in light of the above?



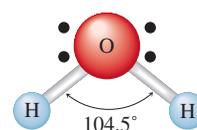
molecule CH_4

Figure (27)



molecule NH_3

Figure (28)



molecule H_2O

Figure (29)

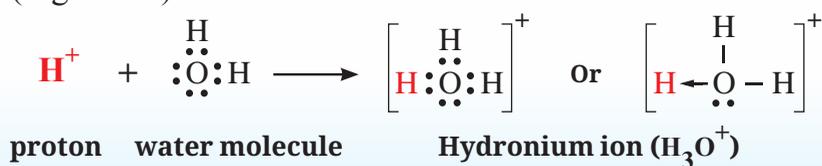
Lone pairs of electrons control the values of the angles between the bonds in covalent molecules

3 coordinate bond

The coordinate bond is considered a special type of covalent bond, and the electron pair forming either bond differs only in origin. While the electron pair forming the covalent bond comes from each atom contributing one electron, the electron pair forming the coordinate bond is a lone pair. Lone Pair, meaning it occupies a single orbital present in one of the atoms, called the **donor atom**. This free electron pair is donated to another atom that has an empty orbital and needs this electron pair to achieve a stable electronic configuration, and this atom is called the **acceptor atom**.

Application (1): Hydronium ion (H_3O^+)

The hydronium ion (H_3O^+) is formed when acids are dissolved in water, where the positive hydrogen ion (proton) bonds with the neutral water molecule through a coordinate bond, which is indicated by an arrow to distinguish it from the covalent bonds in the molecule (Figure 30).



proton water molecule

Hydronium ion (H_3O^+)

Figure (30) the coordinate bond in the hydronium ion



Application (2): ammonium ion (NH_4^+)

The ammonium ion (NH_4^+) is formed when ammonia gas is dissolved in water, where the water proton binds with the neutral ammonia molecule through a coordinate bond (Figure 31).

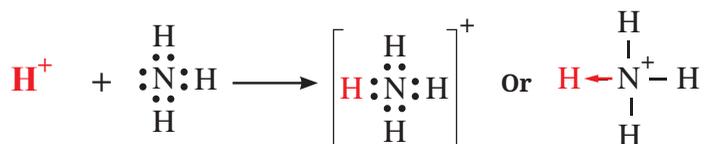


Figure (31) the coordinate bond in the ammonium ion



Application (3): liquid aluminium chloride $\text{Al}_2\text{Cl}_6(\text{l})$

The molecule of liquid aluminium chloride is formed from the connection of two molecules of aluminium chloride (AlCl_3) by two coordinate bonds (Figure 32).

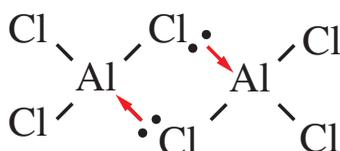


Figure (32) the coordinate bonds in liquid aluminium chloride

Secondly Physical bonds

The hydrogen bond affects the physical properties of some compounds, and the metallic bond affects the physical properties of metals, so both are considered physical bonds.

Hydrogen bond 1

Water boils at 100°C , and this temperature is considered very high for its molar mass (18 g/mol only). In comparison, the boiling point of a compound such as hydrogen sulfide gas, whose molar mass is (34 g/mol), boils at -61°C (Figure 33), and it was supposed that the boiling point of water would be lower than that.

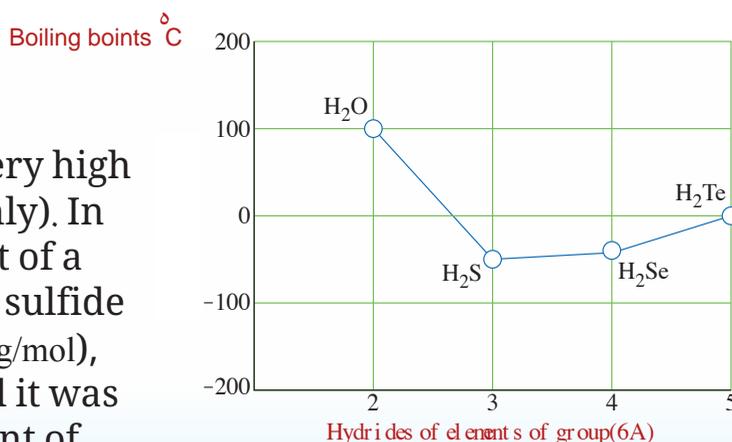


Figure (33)

The anomaly in the boiling point of water is due to the fact that the oxygen atom is characterized by a high electronegativity (3.5) compared to the electronegativity of the hydrogen atom (2.1).

As a result of this difference in electronegativity, the water molecule is polar; the oxygen atom carries a partial negative charge, while each of the two hydrogen atoms carries a partial positive charge. Due to the difference in charge on the oxygen atom and the two hydrogen atoms in the water molecule, the molecules attract each other through what is called the hydrogen bond or the hydrogen bridge according to modern terminology.

That is, the hydrogen atom becomes like a bridge or link between two oxygen atoms with high electronegativity, so the molecules come close enough to each other that the hydrogen atom can be considered a bond that connects the molecules together (Figure 34).

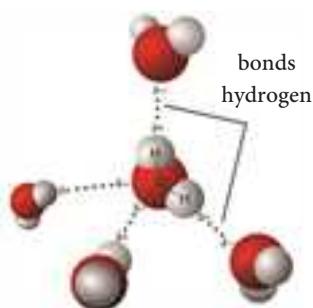


Figure (34)
Hydrogen bonds between water molecules

The hydrogen bond arises between a hydrogen atom bonded in a polar bond [such as (N-H), (O-H), (F-H)] and a pair of free electrons of another atom with high electronegativity, such as: (N, O, F).

The high boiling point of water is attributed to the fact that thermal energy is used to break the hydrogen bonds. Although the hydrogen bond has a clear effect on properties, its strength is much less than that of ordinary chemical bonds.

Table (4) illustrates the difference between the covalent bond and the hydrogen bond in liquid water.

Table (4)

	bond length	bond strength
Covalent bond	96 pm	460 kJ/mol
hydrogen bond	180 pm	20 kJ/mol

It is observed that the hydrogen bond is much weaker and longer than the covalent bond, and the strength of the hydrogen bond increases when the hydrogen bond is in a straight line with the polar covalent bond, as in the cases of water molecules H_2O and hydrogen fluoride HF , and when the difference in electronegativity between hydrogen and the other atom bonded by the polar bond increases.

2 Metallic bond

Each metal has a crystal lattice of a certain shape in which the valence electrons, free to move, gather as an electron cloud around the positive metal ions (cations) (Figure 35), reducing the repulsive forces between them in the crystal lattice.

The thermal and electrical conductivity in metals is attributed to the free valence electrons.

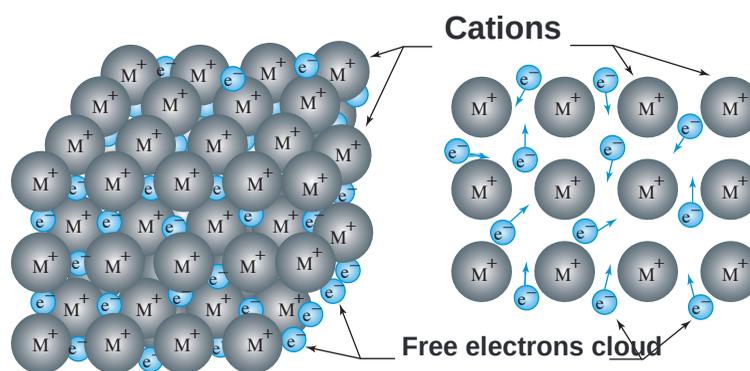


Figure (35) the metallic bond

The number of valence electrons in the metal atom plays an important role in the strength of the metallic bond. The greater the number of valence electrons in the metal atom, the more cohesive the atoms in the crystal become, and thus the metal becomes harder and its melting points higher.

Likewise, the strength of the metallic bond increases with the increase in charge density on the metal ions $\left(\frac{\text{ion charge}}{\text{ion size}}\right)$.

This is evident when comparing these properties for the metals sodium, magnesium, and aluminium from the third period elements as in Table (5).

Table (5)

Metal	Number of the outermost	Hardness mohw scale	Melting point °C
Na	1	(soft) 0.5	98 °C
Mg	2	(mild) 2.5	650 °C
Al	3	(hard) 2.75	660 °C

- (9) Which of the following compounds dissolves in carbon tetrachloride?
 (a) MgCl_2 (b) K_2CO_3 (c) C_2H_4 (d) NaCl
- (10) Which of the following compounds has the greatest ionic character?
 (a) NH_3 (b) H_2O (c) HF (d) H_2S
- (11) Which of the following conducts electric current?
 (a) $\text{KCl}_{(s)}$ (b) $\text{HCl}_{(g)}$ dissolved in benzene.
 (c) $\text{HCl}_{(aq)}$ (d) $\text{C}_2\text{H}_{4(g)}$ dissolved in carbon tetrachloride.
- (12) The charge of the ions that generates ionic currents is converted into electronic currents through the ECG heart recorder
 (a) 1+ only. (b) 2+ only. (c) 1+ and 2+ (d) 2+ and 3+
- (13) All of the following are properties of covalent compounds, **except**
 (a) Most of them dissolve in organic solvents.
 (b) The forces of attraction between its molecules are weak.
 (c) Most of their melts or solutions are good electrical conductors.
 (d) Most of their chemical reactions are slow.
- (14) Identifying the active molecule of the drug used in viral diseases requires
 (a) Estimating the molecular mass of the drug after it binds to the virus protein only.
 (b) Estimating the molecular mass of the drug before and after it binds to the virus protein.
 (c) Not estimating the molecular mass of the drug before it binds to the virus protein only.
 (d) Not estimating the molecular mass of the drug before and after it binds to the virus protein.

2 Give reasons for the following:

- (1) The boiling point of H_2O is higher than the boiling point of H_2S even though the molar mass of water is less than the molar mass of hydrogen sulfide.
- (2) The negative fluoride ion and the positive sodium ion have the same number of electrons despite the difference in their atomic numbers.
- (3) Formation of a coordinate bond in the ammonium ion.

- (4) The molecule CO_2 is nonpolar, although it contains two polar bonds.
- (5) The bond angle in the ammonia molecule is less than that in the methane molecule.
- (6) The nitrogen atom is considered the central atom in the molecule of nitric acid.
- (7) The number of non-bonding electron densities differs in methane compared to ammonia.
-

3 Answer the following:

(1) What is the type of chemical bond in the following compounds:



(2) Arrange the following bonds in order of increasing polarity:



4 Illustrate with a Lewis dot diagram the bonding method in each of:

(1) NaCl

(2) NH_3

(3) CN^-

5 Compare each of the following pairs in terms of the spatial shape of the molecule and the number of free and bonded electron pairs:

(1) BeF_2 - CH_4

(2) SO_2 - BF_3

6 Redraw the structure of the hydrazine molecule N_2H_4 opposite

Indicating on it the dot distribution of electron pairs (free and bonded).



7 Determine the spatial shape of the molecule that contains two atoms bonded to the central atom and one lone pair **with writing** its representative abbreviation.

8 Deduce the number of each of the atoms bonded to the central atom and the lone pairs **as well as the arrangement of** the electron pairs for the molecule with the abbreviation AX_2E

Chapter Two

The representative elements
of some regular groups
in the periodic table





Objectives

At the end of the student's study of the chapter on representative element in some regular groups in the periodic table, he should be able to:

- 1- Identify the elements of Group 1 (alkali metals) and their electronic configuration.
- 2- Identify the general properties of the elements of Group 1.
- 3- Identify the chemical activity of alkali metals.
- 4- Conclude methods of extraction alkali metals from their ores.
- 5- Identify the method of preparing sodium hydroxide in industry.
- 6- Identify the properties of sodium hydroxide.
- 7- Identify the method of preparing sodium carbonate in the laboratory and industry.
- 8- Identify applications in science, technology, and society for the elements of group (1A).
- 9- Identify the elements of Group 5 and their electronic configuration.
- 10- Determine the oxidation numbers of nitrogen in different compounds.
- 11- Identify the methods of preparing nitrogen.
- 12- Identify the physical and chemical properties of nitrogen gas.
- 13- Identify the method of preparing ammonia gas (ammonia) in industry and laboratory.
- 14- Conduct an experiment to detect ammonia gas (ammonia).
- 15- Compare different types of fertilizers.
- 16- Identify the method of preparing nitric acid in the laboratory.
- 17- Identify the chemical properties of nitric acid.
- 18- Identify applications in science, technology, and society for the elements of group (5A).
- 19- Observe safety and security rules in the laboratory.
- 20- Appreciate the efforts of scientists in serving and advancing humanity.

The representative elements of some regular groups

From our previous study of the periodic table, we learned that one of the most important objectives of its study is to classify the elements to facilitate their study in an organized manner. We will address the study of representative element in some regular groups, which means that their elements show a regular gradation not found in transition elements. In this chapter, we will present a more detailed study of the chemistry of these elements and the effect of the properties previously studied in the periodic table, such as atomic radius, ionization potential, and electronegativity, on the chemical and physical properties of these elements.

Firstly Elements of the (s) block

Elements of Group 1 (Alkali metals)

The (s) block of the periodic table includes groups (1A) and (2A). We will limit the study to the properties of the elements of group (1A), known as **metals alkali metals** (except hydrogen H), as shown in Table (1):

Table (1) «For illustration only»

Element	Level energy outer	First ionization potential (kJ/mol)	Ionic radius (pm)	Atomic radius (pm)	melting point (°C)	Density (g/cm ³)
Lithium ₃ Li	2s ¹	520	60	152	180	0.53
Sodium ₁₁ Na	3s ¹	495	95	186	98	0.97
Potassium ₁₉ K	4s ¹	419	133	227	64	0.86
Rubidium ₃₇ Rb	5s ¹	409	148	247	39	1.53
Cesium ₅₅ Cs	6s ¹	376	169	265	29	1.87

The presence of alkali elements in nature:

Sodium and potassium are among the elements available in the Earth's crust, where they occupy the sixth and seventh positions, in terms of the abundance of elements in the Earth's crust.

The most important sodium ore is halite (rock salt) (Figure 1), which is sodium chloride (NaCl), and the most important potassium ore is sylvite (Figure 2), which is potassium chloride (KCl), and carnallite (Figure 3), which is potassium chloride and magnesium chloride ($\text{KCl} \cdot \text{MgCl}_2 \cdot 6\text{H}_2\text{O}$).



Figure (1)
Carnallite



Figure (2)
Sylvite



Figure (3)
Halite

- As for the rest of the metals of the group, it is difficult to obtain them due to their low concentration in their ores and the high cost of extracting them from their ores. Francium is a radioactive element that was not discovered until 1946 as a result of the decay of the nucleus of actinium, as will be explained in Chapter Three.

General properties of group (1A) elements (alkali metals)

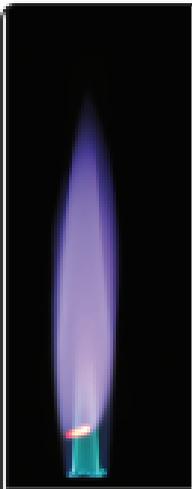
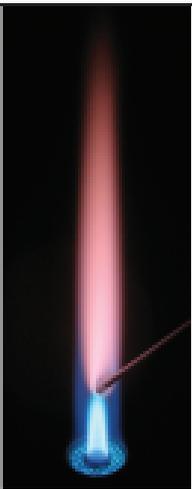
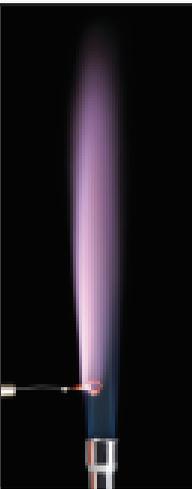
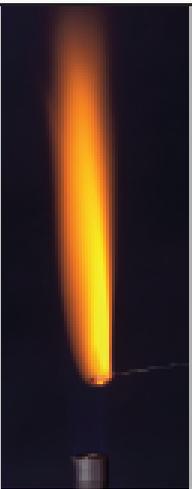
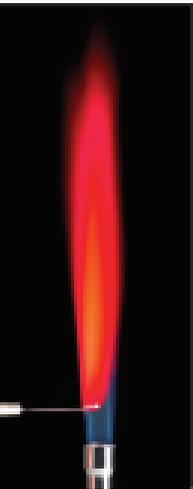
- (1) All the elements of the group are characterized by having a single electron in the outermost energy level (ns^1), and as a result, the following occurs:**
 - Each element of this group is located at the beginning of a new period in the periodic table.
 - All have an oxidation number of (1+) and the ion of each element has the same electronic configuration as the noble gas preceding it.
 - Their metallic bonds are weak, and therefore they are the softest metals and have the lowest melting and boiling points.
- (2) The large atomic volume of the elements of this group, and its increase with increasing atomic number, which results in:**
 - Low ionization potential (ease of losing the valence electron).
 - High chemical activity.
 - Strong reducing agents.
 - Their electronegativity is low compared to the electronegativity of other elements, so most of their compounds are characterized by a high degree of ionic character.
 - Low density.

- f) Their use in photoelectric cells, as in the case of potassium and cesium, where exposure of these materials to light leads to the release of electrons from the surface of the metal, which is known as the **photoelectric effect**.

Photoelectric cells are used to generate electricity and operate some types of calculators and doors that open automatically when approached.

- (3) **Exciting the electrons of the atoms of these elements to higher energy levels by heating, then their return to their ground energy levels, causes them to emit absorbed energy in the form of colored light with characteristic frequency and wavelength, as shown in Table (2):**

Table (2)

Flame test					
	Figure (4)	Figure (5)	Figure (6)	Figure (7)	Figure (8)
Element	Cesium	Rubidium	Potassium	Sodium	Lithium
color formed	Blue violet	Red violet	Light violet	Golden yellow	Crimson

This is used in the dry test (flame test) for these elements in their compounds as follows:

The platinum wire is dipped in concentrated hydrochloric acid to clean it, then dipped in the unknown salt and exposed to the non-luminous part of the Bunsen flame, so the flame acquires the characteristic color of the element cation (Figures 4–8).

Chemical activity of alkali metals

The metals of these elements rust easily in air due to their chemical activity, losing their metallic luster as a result of the formation of oxides of these metals. Therefore, they are kept immersed in liquid hydrocarbons such as mineral oils, away from air and moisture (Figure 9).



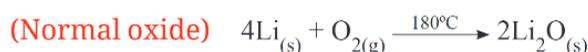
(Figure 9)
Sodium is stored under mineral oils

The following are some of the chemical reactions of these elements:

① Reaction with oxygen:

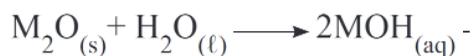
The product of the combustion reaction of alkali metals with oxygen varies according to the ionic radii of their ions, **as follows**:

- Only **lithium** metal reacts, because of its small ionic radius (it has the smallest ionic radius of any metal) with oxygen, forming **normal oxide** Li_2O in which the oxidation number of oxygen is (-2).

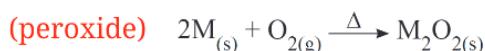


It is difficult to prepare Na_2O and K_2O by this method because they are quickly oxidized to Na_2O_2 and K_2O_2

The normal oxides of alkali metals M_2O are basic oxides that react with water to form the strongest strong alkalis, and the alkalinity of the group (1A) hydroxides increases with increasing atomic number of the element.



- Sodium, potassium, rubidium, and cesium metals react with pure oxygen by heating, forming **peroxide** M_2O_2 in which the oxidation number of oxygen is (-1), and this is due to the large ionic radii of these metals.



Sodium requires a relatively high temperature when reacting with oxygen.

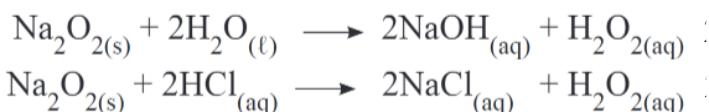


- Potassium, rubidium, and cesium metals react by heating with pure oxygen under a pressure greater than 1 mta forming **super oxide** MO_2 compounds in which the oxidation number of oxygen is $(-\frac{1}{2})$.



Peroxide and super oxide compounds are considered strong oxidizing agents, **where:**

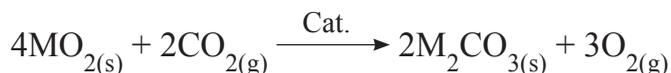
- Peroxide compounds react with water and acids to give hydrogen peroxide.



- Super oxide compounds react with water and acids to give both hydrogen peroxide and oxygen.

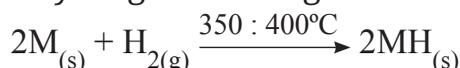


- The compounds of alkali super oxide react with carbon dioxide forming metal carbonate and oxygen.



② Reaction with hydrogen:

Alkali metals react with hydrogen forming ionic compounds called **hydrides**.



Where the metal loses an electron, turning into a positive ion M^+ (its oxidation number is +1), and the hydrogen atom gains this electron, turning into a negative ion H^- known as the hydride ion (its oxidation number is -1).

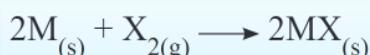
Hydrides resemble salts in terms of their electrical conductivity when molten, and when hydrides are electrolyzed, hydrogen gas is released at the anode and the metal is formed at the cathode.

Hydrides are reducing substances; they hydrolyze to form metal hydroxide with the evolution of hydrogen gas.



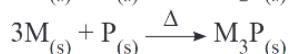
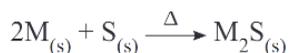
③ Reaction with halogens:

Alkali metals M react violently with halogens X_2 , forming highly stable ionic halides. The reaction is accompanied by an explosion due to the formation of a strong ionic bond accompanied by the release of a large amount of energy.

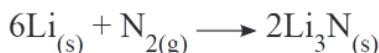


④ Reaction with other nonmetals:

- Hot alkali metals combine directly with sulfur and phosphorus.



- Lithium alone combines with nitrogen (because of its small ionic radius) at room temperature, forming lithium nitride Li_3N , which reacts with water to produce ammonia gas.



⑤ The reaction with water:

The metals of this group are considered the most active metals known and occupy the top of the chemical activity series. Therefore, they displace hydrogen from water, and the reaction is accompanied by the release of a large amount of energy, causing the evolved hydrogen to ignite (Figure 01). The reaction becomes more violent from lithium to cesium.



Figure (10)

Potassium reaction with water

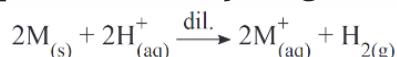


For this reason, sodium fires are not extinguished with water.



⑥ Reaction with acids:

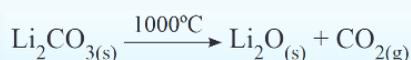
These metals displace hydrogen from diluted acids, and the reaction is more violent than their reaction with water, due to the ease of separating the acid hydrogen compared to the hydrogen in water.



⑦ Thermal decomposition of their oxy salts:

The oxy salts of alkali metals are thermally stable.

- ① Not all alkali metal carbonates decompose by heat, except for lithium carbonate which decomposes at 1000°C



- ⓑ Alkali metal nitrates decompose by heat partially, forming the metal nitrite and oxygen, except for lithium nitrate whose thermal decomposition products are different.



The thermal decomposition of potassium nitrate salt (gunpowder salt) is accompanied by a violent explosion, so it is used in the explosives industry, while sodium nitrate salt (Chilean saltpeter) is not suitable for this industry because it is a deliquescent substance (absorbs water vapor from the atmosphere).

Science, Technology and Society (STS)

- The distinctive physical and chemical properties of alkali metals have been utilized in various technological fields to achieve the goals of sustainable development.

1 Lithium:

- %70 of lithium metal production is used in the manufacture of **lithium-ion batteries** used in smartphones (Figure 11) and laptops and electric cars (Figure 12).



Figure (11)
Falcon 9



Figure (12)
Electric car



Figure (13)
Smartphone battery

Lithium metal is also used with aluminium in the manufacture of **alloys** characterized by low density, so they are used in the manufacture of types of rackets and frames of both sports bicycles and aircraft and fuel tanks of spacecraft such as Falcon 9 (Figure 13).

2 Sodium:

- The distinctive properties of sodium metal and its abundance in the Earth's crust give it many technological applications, including:
 - It is used in the composition of **sodium vapor lamps** which produce yellow light characterized by high efficiency and good fog penetration, so they are used in road lighting (Figure 14).



Figure (14)

Sodium reactor



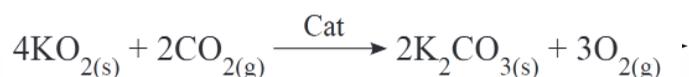
Figure (15)

Use of sodium vapor lamps in lighting roads and streets

- It is used in liquid form **as a coolant** instead of water in a certain type of nuclear reactor known as a **sodium reactor** Natrium reactor (Figure 14) due to its high boiling point—compared to water—and its efficiency in heat transfer.

3 Potassium:

- More than %90 of potassium metal production is used in the production of **agricultural fertilizer**, and the reaction of potassium super oxide with carbon dioxide is utilized as follows:



In replacing carbon dioxide with oxygen in **closed breathing** devices (Figure 61) used by firefighters, and used in submarines and spacecraft that recycle exhaled air, by passing exhaled air containing a high percentage of carbon dioxide over filters containing super oxide compound and a catalyst.



Figure (16)

Closed breathing apparatus

4 Rubidium:

- Although rubidium metal is rare, it is used in some advanced technological applications, such as the use of some of its compounds like rubidium bromide RbBr in the manufacture of types of glass with durable surfaces, scratch resistance, and distinctive refractive properties. This type of glass is used in the manufacture of some precise optical components such as **night vision** devices (Figure 17).



Figure (17)
Night vision device

5 Cesium:

- Cesium metal atoms are used in the manufacture of **cesium atomic clocks** (Figure 18) which are used to measure time with extreme accuracy, unaffected by heat or pressure, and self-adjust automatically.



Figure (18)
Cesium atomic clock

- For this reason, satellite receivers known as GPS carry several cesium atomic clocks, because any small error in time measurement leads to a huge error in location determination.

Extraction of alkali metals from their ores

Alkali metals do not exist in nature in a free state but are found as ionic compounds. To prepare the metal, the lost electron must be returned to its positive ion, and this reduction process is difficult to carry out by ordinary chemical methods, so it is done by **electrolysis** of its molten halides, such as sodium chloride NaCl in the presence of some fluxes, such as calcium chloride, which lowers the melting point of this halide. This method is known as the **Downs process** (Figure 19).

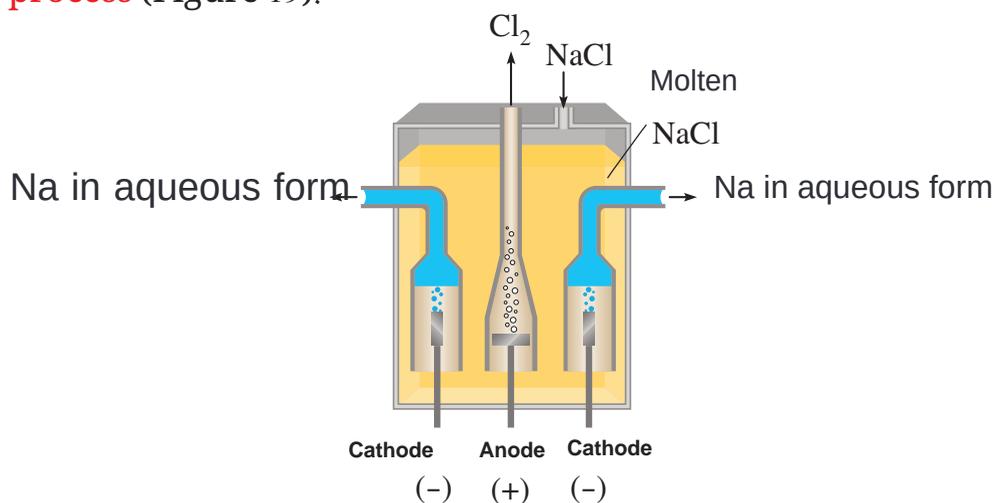
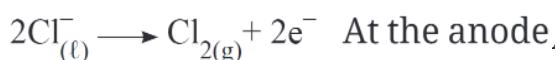


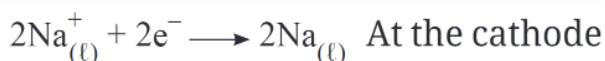
Figure (19)
Downs cell

In it, the following occurs:

- **Oxidation of** chloride anions Cl^- at the anode (positive electrode), turning into chlorine gas Cl_2



- **Reduction of** positive sodium ions Na^+ at the cathode (negative electrode), turning into sodium that floats above molten sodium chloride.



The Downs cell used is designed so that the produced sodium is separated from the formed chlorine so that they do not recombine.

Sodium compounds

Sodium compounds are numerous, and only the following will be studied:

- ① **sodium hydroxide (caustic soda).**
- ② **sodium carbonate (washing soda).**

1 Sodium hydroxide (NaOH)

Its preparation in industry:

It is prepared by the electrolysis of concentrated sodium chloride solution (Figure 20), according to the following equation:

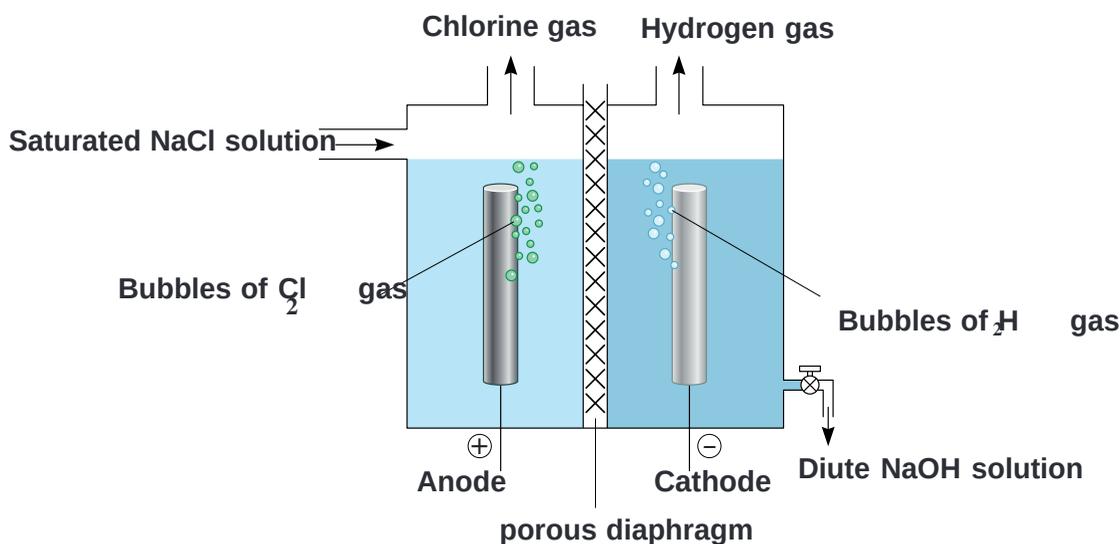
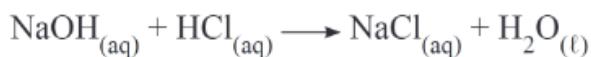


Figure (20)

Electrolysis cell for the production of sodium hydroxide

Physical and chemical properties:

- ① A white solid with a melting point of 318°C
- ② A deliquescent substance (absorbs water vapor from the atmosphere).
- ③ Its solution has a soapy feel and a caustic effect on the skin.
- ④ It dissolves easily in water forming an alkaline solution, with releasing heat energy as a result of this dissolution (exothermic dissolution).
- ⑤ It reacts with acids forming salt and water.



2 sodium carbonate (Na_2CO_3)

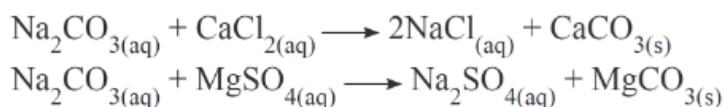
Preparation methods: **A** In the laboratory. **B** In industry.

A Its preparation in the laboratory :

Sodium carbonate is prepared in the laboratory by passing carbon dioxide gas through a hot sodium hydroxide solution, then the solution is left to cool, so that crystals of hydrated sodium carbonate separate out.



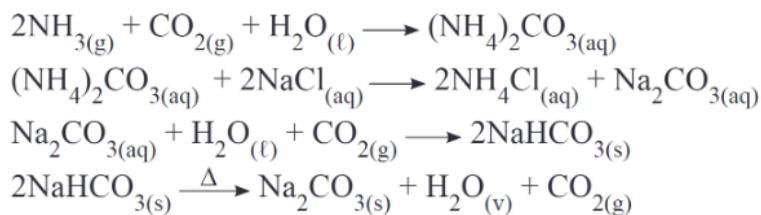
Hydrated sodium carbonate is known as washing soda $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$, because it is used to remove permanent water hardness caused by the presence of Mg^{2+} and Ca^{2+} dissolved ions in water, where the reaction occurs through ion exchange to form calcium and magnesium carbonate, which do not dissolve in water, thus removing the hardness.



B Its preparation in industry (Solvay process):

Sodium carbonate is prepared in industry as follows:

- 1- Ammonia and carbon dioxide gases are passed through a concentrated solution of sodium chloride to form ammonium carbonate.
- 2- Ammonium carbonate reacts with sodium chloride solution, forming sodium carbonate.
- 3- Sodium carbonate, in the presence of carbon dioxide gas and water, is converted to sodium bicarbonate, which precipitates as a result of the solution becoming saturated.
- 4- Sodium bicarbonate is separated and heated to obtain sodium carbonate.



Physical and chemical properties:

- ① Anhydrous sodium carbonate is a white powder that dissolves easily in water.
- ② Its aqueous solution is alkaline.
- ③ It reacts with acids with effervescence, releasing carbon dioxide gas.



Secondly Group (p) elements

- The category includes (p) from the periodic table six groups, of which only the properties of the elements of group (5A) will be studied: The group consists of five elements as shown in Table (3):

Table (3) "For illustration only"

Element	Outer main energy level	Atomic radius (pm)	First ionization potential (kJ/mol)	melting point (°C)	Density (g/cm ³) at rtp
₇ N Nitrogen	2s ² , 2p ³	74	1400	- 210	0.00116
₁₅ P Phosphorus	3s ² , 3p ³	110	1011	44	1.82
₃₃ As Arsenic	4s ² , 4p ³	121	947	817	5.73
₅₁ Sb Antimony	5s ² , 5p ³	141	834	631	6.7
₈₃ Bi Bismuth	6s ² , 6p ³	182	703	271	9.8

Its presence in nature: It is not widespread in nature, except:

- **Nitrogen:** it represents $\frac{4}{5}$ of the volume of atmospheric air approximately.
- **Phosphorus:** It is the most abundant in the earth's crust, where it is found as:
 - Rock calcium phosphate Ca₃(PO₄)₂
 - Apatite CaF₂.Ca₃(PO₄)₂ and it is a double salt of calcium fluoride and phosphate.
- **Arsenic - Antimony - Bismuth:** Their main ores are found as sulphides As₂S₃, Sb₂S₃, Bi₂S₃

General properties of the elements of group 5A (15):

- ① The nonmetallic character predominates in the properties of the elements of this group, and the metallic character increases with the increase in the atomic number of the element.
 - Nitrogen and phosphorus are nonmetals.
 - Arsenic and antimony are metalloids.
 - Bismuth is a metal element, however its electrical conductivity is weak.
- ②
 - a) The nitrogen molecule consists of two atoms (N₂).
 - b) The molecules of phosphorus, arsenic, and antimony in the vapor state (at high temperature) consist of four atoms (P₄, As₄, Sb₄).
 - c) Bismuth forms a metallic crystal, and at high temperatures its vapor molecules consist of two atoms (Bi₂), thus differing from most metals whose molecules are monoatomic in the vapor state.
- ③ The elements of this group are characterized by multiple oxidation numbers in different compounds, ranging between (-3, +5) where they gain three electrons or share up to five electrons.

The following table shows the oxidation numbers of nitrogen in some of its compounds:

Table (4)

Substance	Formula	oxidation number
Ammonia	NH_3	-3
Hydrazine	$\text{NH}_2 - \text{NH}_2 (\text{N}_2\text{H}_4)$	-2
Hydroxylamine	NH_2OH	-1
Nitrogen	N_2	0
Nitrous oxide	N_2O	+1
Nitric oxide	NO	+2
Nitrogen trioxide	N_2O_3	+3
Nitrogen dioxide	NO_2	+4
Nitric acid	HNO_3	+5

It is noted that positive oxidation numbers appear in oxygen compounds, because the electronegativity of oxygen is higher than that of nitrogen.

④ **Its compounds with oxygen:**

All elements in this group form oxides, some of which are acidic, such as: N_2O_5 , P_2O_5 , As_2O_5 , while Sb_2O_3 is an unstable oxide, Bi_2O_5 Basic oxide. The basic character increases (the acidic character decreases) with increasing atomic number.

⑤ **Its compounds with hydrogen:**

Most group elements form compounds with hydrogen in which the oxidation number of the element is (-3) such as: ammonia NH_3 , phosphine PH_3 , and arsine AsH_3 . Since there is still a lone pair of electrons in the valence shell of the central atom in these compounds, it can donate this pair to other atoms or ions and form a coordinate bond.

The basicity of ammonia is stronger than that of phosphine, due to the greater electronegativity of nitrogen compared to phosphorus.

The polarity of the hydride compounds of group (5A) elements decreases with increasing atomic number, and thus both their thermal stability and their solubility in water decrease.

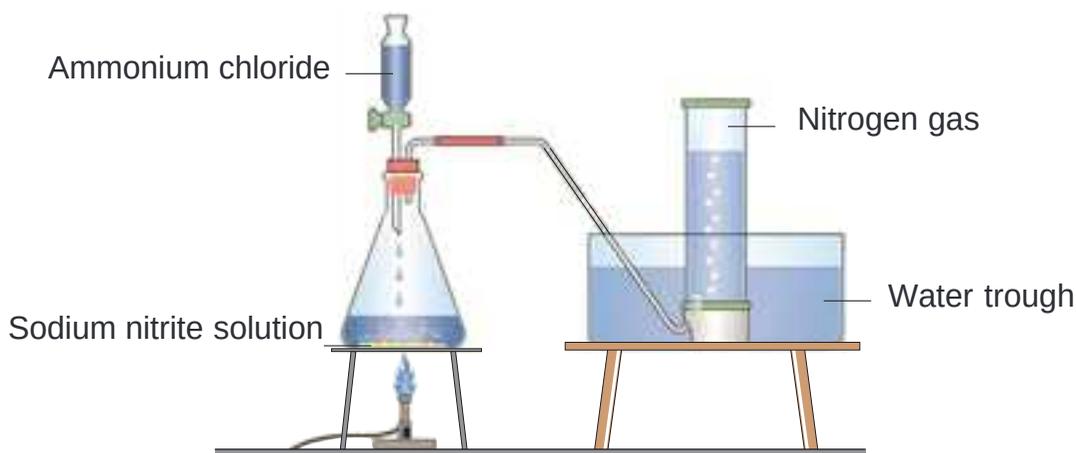
And from the group elements (5A), only the study of the nitrogen element will be covered.

Nitrogen element

Nitrogen gas constitutes about 78% of the volume of atmospheric air, and it is also found in the earth's crust in the form of nitrate and nitrite salts, and enters into the structure of amino acids which represent the basic units forming proteins.

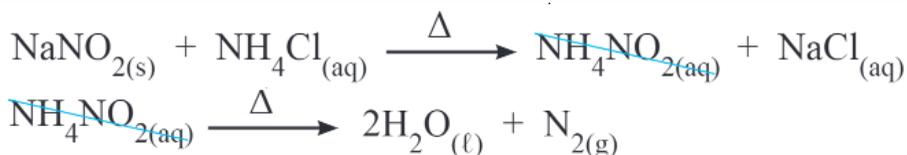
Preparation methods:

- ① **It is prepared in industry** by fractional distillation of liquefied air.
- ② **It is prepared in the laboratory** by adding ammonium chloride solution to sodium nitrite salt with heating, and the gas is collected by downward displacement of water (Figure 21).



(Figure 21)

Apparatus for preparing nitrogen gas in the laboratory



By combining the two equations



Physical properties

Practical activity: Physical properties of nitrogen gas. Prepare several test tubes of nitrogen gas, and discuss the following with your colleagues under the supervision of your teacher:

Table (5)

Practical activity	Observation	Conclusion
(1) The colour and smell of the gas.
(2) The method of collecting the gas when prepared from atmospheric air
(3) The effect of the gas on litmus paper

In light of the previous activity and others, some physical properties of nitrogen can be summarized as follows:

- ① A gas that is colourless, tasteless, and odorless.
- ② Slightly lighter than air because air contains oxygen, which is heavier than nitrogen.
- ③ Sparingly soluble in water.
- ④ Neutral in effect on both red and blue litmus paper.
- ⑤ Its density (at rtp) 1.16 g/L.
- ⑥ Its boiling point(-195.8°C), meaning it can be liquefied at this temperature under normal atmospheric pressure.

Chemical properties

Nitrogen reactions with other elements only occur in the presence of an electric spark or electric arc or intense heating, due to the difficulty of breaking the triple bond between the two nitrogen atoms.

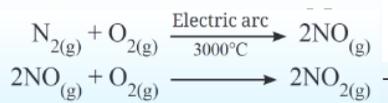
1 With hydrogen

This reaction takes place in a closed vessel in the presence of an **electric spark** which provides the thermal energy needed to initiate the reaction.



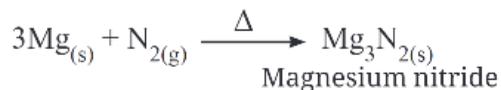
2 With oxygen

This reaction takes place in the presence of an **electric arc** to ensure the continuation of thermal energy, where the temperature reaches 3000°C, forming nitric oxide which quickly oxidizes to nitrogen dioxide.

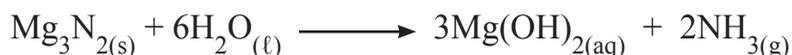


3 With metals at high temperatures

Nitrogen reacts with metals such as magnesium to form the metal nitride.



The formed nitrides decompose easily in water and ammonia gas is released.



Nitrogen compounds

Nitrogen compounds are numerous, and we will limit the study to each of the following:

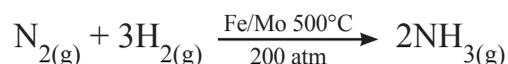
1 Ammonia.

2 Nitric acid.

1 Ammonia NH_3

A Preparation of ammonia in industry:

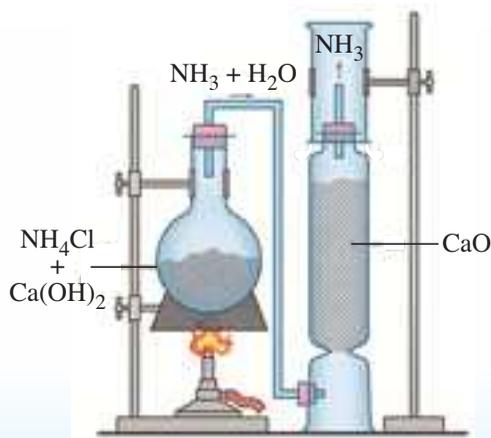
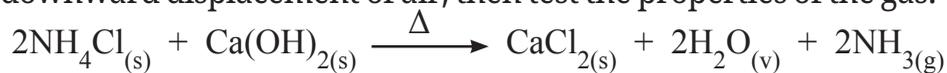
Ammonia gas is prepared in industry from the elements nitrogen and hydrogen in the presence of catalysts which are iron and molybdenum, and under a pressure of 200 atm at a temperature of 500°C . This method is known as the **Haber-Bosch process**.



B Preparation of ammonia in the laboratory:

Practical training:

- 1 Assemble the apparatus shown in the diagram (Figure 22).
- 2 Put ammonium chloride and calcium hydroxide (slaked lime) in the flask, then heat.
- 3 Pass the products of the reaction over a drying agent (quicklime CaO) to remove moisture (drying the gas from water vapor).
- 4 Fill several jars by downward displacement of air, then test the properties of the gas.

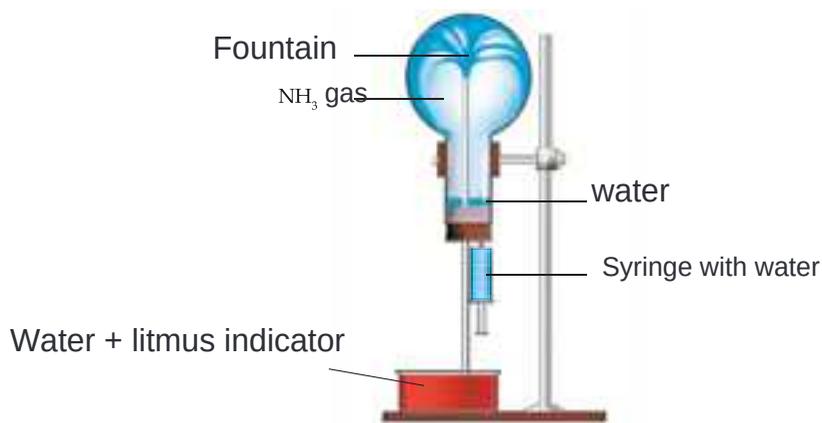


(Figure 22)

Apparatus for preparing ammonia gas in the laboratory

Table (6)

The experiment	Observation	Conclusion
(1) What is the colour of ammonia? And what is its smell?
(2) Bring a red litmus paper moistened with water near the gas... What do you observe?
(3) Bring a burning splint near the mouth of the jar, does the gas ignite? Does the splint go out?
(4) Prepare the fountain apparatus shown in Figure (23) and fill the flask with ammonia gas, then inject the water from the syringe and observe what happens... and why?



(Figure 23)

The fountain experiment to prove that ammonia gas is highly soluble in water and its solution is alkaline

Detection of ammonia gas (ammonia):

The hydrogen chloride gas evolved from concentrated hydrochloric acid HCl is used to detect ammonia gas evolved from concentrated ammonium hydroxide solution (Figure 24), where dense white clouds of ammonium chloride are formed, which is a sublimable solid substance.



(Figure 24)

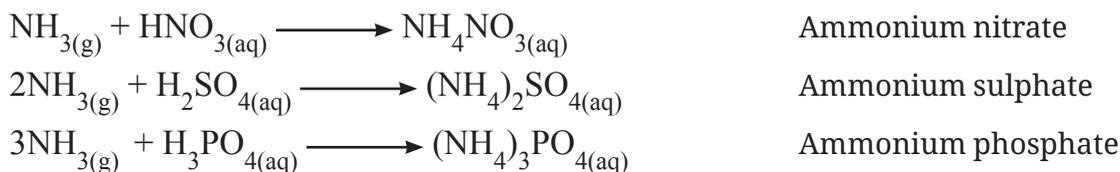
Detection of ammonia gas

Uses of ammonia

Ammonia prepared by the "Haber-Bosch" method is mainly used in the industry of fertilizers:

Nitrogenous fertilizers

- Salts of ammonia are prepared from the reaction of ammonia with suitable acids.



Fertilizer NPK

NPK fertilizer is considered one of the best fertilizers that provide plants with what they need from the elements nitrogen N, phosphorus P, and potassium K. The percentages of these elements are written on fertilizer packages in the form of three numbers representing their percentages as shown in figure (25).



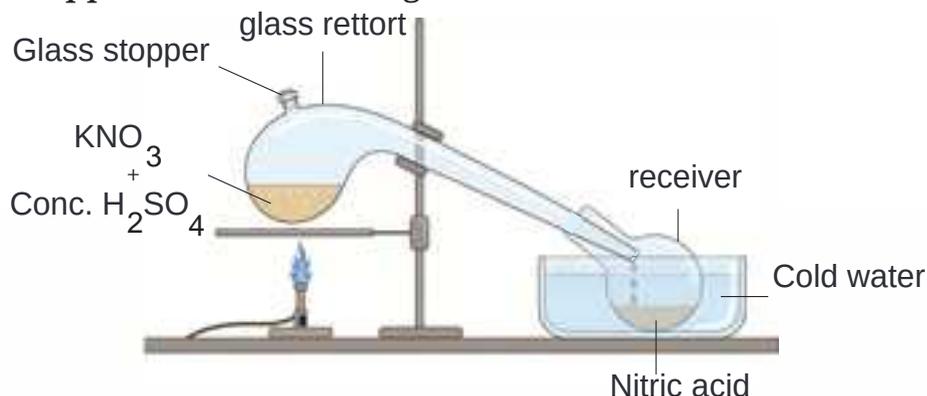
(Figure 25)
Fertilizer NPK

2 Nitric acid HNO₃

Preparation of nitric acid in the laboratory:

Practical training:

- Prepare the apparatus shown in figure (26).



(Figure 26)

- Put potassium nitrate and concentrated sulphuric acid in the glass retort and place the receiver in a basin containing cold water, noting that the stopper should not be made of rubber, so that it does not corrode due to the vapors of the formed acid.

- ③ Heat the contents of the retort provided that the temperature does not exceed 100°C (so that the resulting acid does not decompose) and collect the formed acid in the receiver, then test its properties.



Table (7)

The experiment	Observation	Conclusion
(1) What is the colour of the formed liquid?
(2) Add to the formed liquid blue litmus solution ... What do you observe?
(3) Take a quantity of the liquid in a test tube and add copper filings to it with heating ... What do you observe?
(4) Dilute the acid by adding a quantity of it to water and add iron filings. What is the colour of the evolved gas? And what is the effect of air on it?

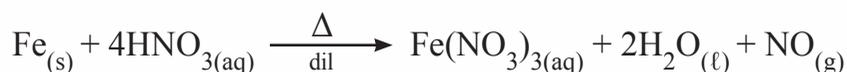
Chemical properties of nitric acid

- ① An oxidizing agent that decomposes at a temperature higher than 100°C forming oxygen gas.



- ② The dilute acid reacts with:

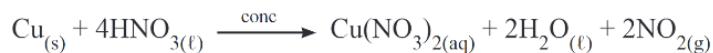
- (1) The metals that precede hydrogen in the chemical activity series, forming the metal nitrate and atomic hydrogen which reduces the acid.



- (2) The metals that follow hydrogen in the chemical activity series, and the reaction occurs on the basis that the acid is an oxidizing agent, where the metal is oxidized, then the oxide reacts with the acid.



- ③ The **concentrated acid** reacts with the metals that follow hydrogen in the chemical activity series, and the oxidation number of nitrogen in the oxide resulting from the reaction of concentrated acid with copper is higher than in the oxide resulting from the reaction of dilute acid with the same metal.



- ④ Some metals that precede hydrogen in the chemical activity series, such as iron, chromium, and aluminum, resist the effect of concentrated acid by forming a non-porous layer of oxide that prevents the continued reaction of the metal with concentrated acid, a phenomenon known as **passivity phenomenon** Passivity.

Science, Technology and Society (STS)

The diversity of the properties of group (5A) elements has been utilized in various technological fields to achieve the goals of sustainable development.

1 Nitrogen:

More than 80% of nitrogen production is consumed in the manufacture of **ammonia gas**, which in turn is used in the manufacture of **fertilizers**.

Nitrogen gas is also used in the industrial sector to provide an inert atmosphere during certain welding processes and in the electronics industry to protect its components from oxidation during production and storage.

Liquid nitrogen (Figure 27) is used in scientific research and medicine as shown below:

- It is used to preserve and store biological samples (Figure 28) such as tissues, blood, sperm, and eggs for long periods.
- It is used in cryotherapy Cryotherapy, such as in the treatment of skin cancer and warts (Figure 29).



Figure (27)
Removal of hand warts using liquid nitrogen



Figure (28)
Liquid nitrogen is used to preserve biological samples



Figure (29)
Liquid nitrogen

2 Phosphorus:

(75 : 85%) of phosphorus production is consumed in the manufacture of **fertilizers** and the remaining part is used

in many technological applications, including:

- Production of white light from the blue light-emitting diode Blue LED (Figure 30).



Figure (30)

Blue and white light-emitting diode

- The oxide P_2O_5 is added to glass to produce a type with a high refractive index, used in lens manufacturing.

Other compounds of it are used as **anti-reflection** coatings on lenses and glasses (Figure 31).



Before



After

Figure (31)

Coating eyeglass lenses with anti-reflection agents

3 Arsenic:

The traditional uses of toxic arsenic in the manufacture of pesticides and wood preservation have decreased because of its health and environmental risks.

At the same time, its uses are increasing in the field of electronics and renewable energy, such as the use of gallium arsenide compound GaAs in the manufacture of mobile phones, flexible solar cell panels, and lasers (Figure 32).



Figure (32)

Flexible solar cell panels

Antimony: 4

Antimony oxide compound $(\text{III})\text{Sb}_2\text{O}_3$ is used as a **flame retardant**, where it is added to some products made of plastic, wood, and fabrics to increase their resistance to ignition (Figure 33).



Figure (33)

Antimony oxide (III) is a flame retardant material

5 Bismuth:

Bismuth is characterized by unique properties that make it suitable for many technological applications, as shown by the following examples:

- The low melting point of bismuth makes it used in the manufacture of alloys with tin that melt at temperatures below 100°C used in fire detection systems and automatic sprinklers (Figure 34).



Figure (34)

Automatic fire sprinkler

- It is used as a substitute for lead in lead-free solder (Figure 35) used in soldering electronic circuit components.



Figure (53)

Lead-free bismuth solder

Second chapter evaluation



1 Choose the correct answer from among the following answers:

- (1) The metallic character in the elements of the first group increases by increasing.....
- (a) The percentage by weight in the Earth's crust. (b) Boiling point.
(c) Atomic number. (d) Melting point.
- (2) The phosphorus molecule in the vapor state contains
- (a) One atom. (b) Two atoms.
(c) Three atoms. (d) Four atoms.
- (3) When magnesium nitride reacts with water, the gas released is
- (a) Ammonia. (b) Hydrogen.
(c) Nitric oxide. (d) Nitrogen dioxide.
- (4) The oxidation number of the first group elements in their compounds is
- (a) (-1). (b) (+1). (c) (-2). (d) (+2).
- (5) Potassium super oxide is used in closed breathing apparatuses to replace the gas
- (a) Carbon dioxide with hydrogen gas.
(b) Carbon dioxide with oxygen gas.
(c) Oxygen with ammonia gas.
(d) Oxygen with carbon dioxide gas.
- (6) When a glass rod moistened with concentrated hydrochloric acid is exposed to ammonia gas, dense white fumes of are formed.
- (a) Ammonium carbonate. (b) Ammonium chloride.
(c) Hydrogen chloride. (d) Ammonium sulphate.
- (7) Which alkali metal combines with nitrogen gas at room temperature?
- (a) Lithium. (b) Sodium.
(c) Potassium. (d) Cesium.

- (8) The Downs cell is used in
- (a) Extraction of nitrogen from atmospheric air.
 - (b) Preparation of nitric acid.
 - (c) Extraction of sodium from sodium chloride.
 - (d) Preparation of sodium hydroxide.
- (9) The Solvay process ends by heating
- (a) Sodium bicarbonate salt.
 - (b) Sodium carbonate salt.
 - (c) Sodium bicarbonate solution.
 - (d) Sodium carbonate solution.
- (10) The fountain experiment is used to prove that ammonia gas is
- (a) Less dense than air.
 - (b) Acidic substance.
 - (c) Basic substance.
 - (d) Does not dissolve in water.
- (11) A fertilizer package contains ammonium sulphate and potassium sulphate, which elements of the NPK fertilizer are present in it?
- (a) N, P only.
 - (b) N, K only.
 - (c) K only.
 - (d) P only.
- (12) Which of the following expresses the continuation of the reaction of chromium metal with nitric acid?

Choices	concentrated acid	dilute acid
(a)	✓	✓
(b)	✗	✓
(c)	✓	✗
(d)	✗	✗

- (13) The metal (X) has the smallest ionic radius. Which of the following expresses the use of (X)?
- (a) Manufacturing atomic clocks.
 - (b) Manufacturing electric car batteries.
 - (c) A coolant material in some nuclear reactors.
 - (d) Manufacturing scratch-resistant glass.
- (14) The metal (A) is prepared by the Downs process. Which of the following expresses the metal (A)?
- (a) Its boiling point is higher than that of water.
 - (b) Reacts with oxygen forming AO_2
 - (c) The radius of $K^+ < A^+$
 - (d) Used in making alloys for sports bicycles.

(15) The compound used in manufacturing glass with durable and scratch-resistant surfaces

- (a) KO_2 (b) RbBr (c) Na_2CO_3 (d) KNO_3

(16) The element M is used in manufacturing clocks that are unaffected by heat and pressure, and reacts with pure oxygen under a pressure greater than 1 atm forming

- (a) MO (b) M_2O (c) MO_2 (d) M_2O_2
-

2 Show the electronic configuration of the following elements using the Aufbau principle, then state their possible oxidation numbers in their compounds:

(1) Potassium $_{19}\text{K}$

(2) Cesium $_{55}\text{Cs}$

(3) Nitrogen $_7\text{N}$

3 Give reasons for the following:

(1) Alkali metals are characterized by chemical activity.

(2) The metallic bond between the atoms of group 1 metals is weak.

(3) The use of cesium in the manufacture of photoelectric cells.

(4) The use of alkali super oxide compounds in closed breathing apparatus.

(5) The difficulty of obtaining cesium metal.

(6) water is not used to extinguish sodium fires.

(7) Sodium nitrate is not used in the manufacture of explosives.

(8) The difference in the products of the reaction of alkali metals with oxygen.

(9) The difficulty of extracting alkali metals from their ores by ordinary chemical methods.

(10) The multiplicity of oxidation states of nitrogen.

(11) The basicity of ammonia is stronger than that of phosphine.

(12) The use of concentrated hydrochloric acid in detecting ammonia.

4 Write the chemical equation that shows the reaction of:

- (1) Sodium with oxygen (at 300°C).
 - (2) Alkali super oxide compounds with carbon dioxide.
 - (3) The thermal decomposition of alkali nitrates.
 - (4) The electrolysis of concentrated sodium chloride solution.
 - (5) The electrolysis of molten sodium chloride.
 - (6) The thermal decomposition of ammonium nitrite.
-

5 Explain the effect of the following substances on sodium metal:

- (1) Hydrochloric acid.
 - (2) Hydrogen.
 - (3) Water.
-

6 Explain the similarity between:

- (1) Salts **and** hydrides.
- (2) Gunpowder salt **and** Chilean saltpeter.

Chapter Three

Nuclear Chemistry





Objectives

At the end of the student's study of the chapter on nuclear chemistry, he should be able to:

- 1- Identify the historical sequence of the discovery of the activity of the nuclei of radioactive elements.
- 2- Identify the concept of isotopes.
- 3- Distinguish between alpha, beta, and gamma radiations.
- 4- Calculate the conversions of mass to energy.
- 5- Define nuclear binding energy and calculate nuclear binding energy.
- 6- Illustrates the relation between nuclear binding energy and the stability of the nucleus.
- 7- Distinguish between the two types of activity (natural and industrial).
- 8- Define half-life period and how to calculate it.
- 9- Define the smoke detector and how it works.
- 10- Compare between nuclear fission and nuclear fusion.
- 11- Calculate the amount of produced energy from a nuclear reaction.
- 12- Compare between chemical reactions and nuclear reactions.
- 13- Identify applications to science, technology, and society.

- In 1896, **Becquerel** discovered—through coincidence—that one of the uranium compounds emits invisible radiations that lead to the formation of images or shadows on sensitive photographic plates, and he proved that this radiation is emitted by all uranium compounds, and he called this phenomenon **activity**.



(Figure 1)
Henri Becquerel

And **Chadwick** discovered the existence of **neutrons** with neutral charge inside the nucleus. The sum of the numbers of protons and neutrons was called **mass number (A)** or what is known as **nucleons** (Figure 2).

Number of neutrons (N)
= **mass number (A) – atomic number (Z)**

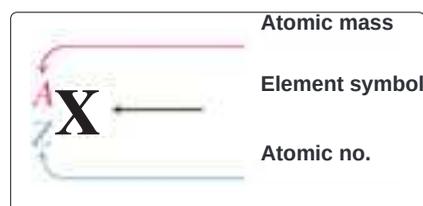


Figure (2)

mass number and
atomic number

- In 1942, **Fermi** managed to initiate the first artificial chain nuclear reaction, and studies and research followed after that.

Isotopes Isotopes

- Scientists discovered that there are atoms of the same element that differ in their mass numbers but have the same atomic number, which they agreed to call **isotopes**.

The table (1) illustrates the isotopes of hydrogen (Figures 5 - 3):

Table (1)

Isotope	Protium ${}^1_1\text{H}$	Deuterium ${}^2_1\text{H}$	Tritium ${}^3_1\text{H}$
shape	 Figure (3)	 Figure (4)	 Figure (5)
atomic number	1	1	1
mass number	1	2	3
Number of neutrons	0	1	2

- The isotopes of the same element agree in chemical properties due to having the same atomic number (**and thus** the number of electrons in the outer energy level) and differ in physical properties due to the difference in mass number.

Types of isotopes

• There are two types of isotopes, **which are:**

- ① **Stable isotopes** do not emit any radiation, such as the oxygen isotope 16 – and the sodium isotope 23 –
- ② **Radioactive isotopes** have unstable nuclei and emit radiations (alpha (α), beta (β^-), gamma (γ)). The effect of the electric field on them differs (Figure 6).

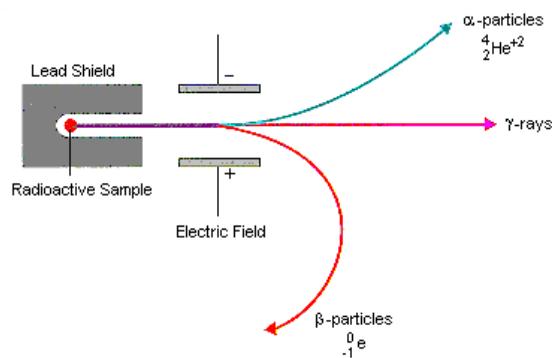


Figure (6)

The effect of the electric field on alpha, beta, and gamma radiations

The table (2) illustrates the aspects of comparison between them:

Table (2)

Aspects of comparison	Alpha radiation α	Beta radiation β^-	Gamma radiation γ
Nature	Helium atom nucleus ${}^4_2\text{He}$	High-speed electron ${}^0_1\text{e}$	Photon High energy
Charge	Positively charged (+2)	Negatively charged (1-)	Electrically neutral
Affected by the magnetic field	Affected	Affected	Not affected
Affected by the field electric	Deflects slightly Towards the negative pole Due to its large relative mass	Deflects greatly towards the positive pole due to its small mass	Not affected

Calculations of the conversion of mass to energy

- In nuclear reactions, mass is converted into energy. The amount of produced energy E resulting from the conversion of mass can be calculated in the unit of :
 - The joule (J) if the converted mass is in kg by applying **Einstein's** equation, as follows:

$$E \text{ (J)} = m \text{ (kg)} \times c^2 \text{ (} 3 \times 10^8 \text{ m/s)}^2$$

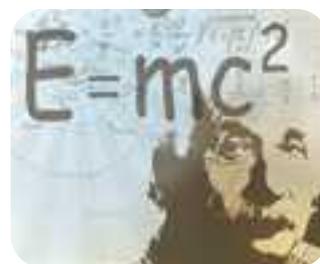


Figure (7)
Einstein's equation

- mega electron volt (MeV), if the converted mass is in atomic mass unit (u) **from the relation:**

$$E \text{ (MeV)} = m \text{ (u)} \times 931$$

It is noted that: $1 \text{ MeV} = 1.6 \times 10^{-13} \text{ J}$

Nuclear binding energy Binding Energy

- All precise measurements of the masses of different nuclei have proven that:

- Proton mass 1.0073 u, neutron mass 1.0087 u and electron mass 0.000548 u
- The actual mass of the nucleus (mass of the bound nucleons) **is less than** the theoretical mass (mass of the separated nucleons) (Figure 8).

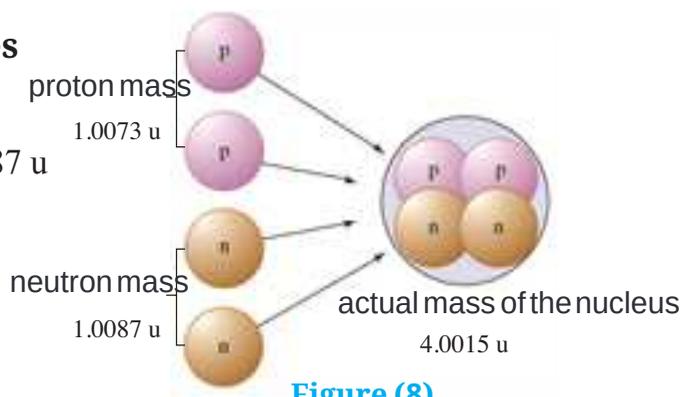


Figure (8)
The actual mass is less than the theoretical mass

 **Application:** Helium atom nucleus

The actual mass of the nucleus ${}^4_2\text{He} = 4.0015 \text{ u}$

The theoretical mass of the nucleus of the helium atom ${}^4_2\text{He}$
 $= (2 \times 1.0073 \text{ u}) + (2 \times 1.0087 \text{ u}) = 4.032 \text{ u}$

The amount of mass defect = theoretical mass – actual mass = 0.0305 u
 $= 4.0015 \text{ u} - 4.032 \text{ u}$

And since energy and mass are two forms of the same thing according to **Einstein's theory**, mass can be converted into energy and vice versa, and the amount of mass defect is equivalent to the amount of energy required to overcome the electrostatic repulsion between the protons inside the nucleus, and this energy is called **nuclear binding energy (BE)**.

- The nuclear binding energy in the nucleus of the helium atom (MeV)
= mass defect (u) \times 931 = 5030.0 \times 139 = MeV 28.3955

This amount represents the released energy when the nucleus of the helium atom is formed or the absorbed energy when the nucleus of the helium atom is split.

And the nuclear binding energy per nucleon is considered $\left(\frac{BE}{A}\right)$ Also a suitable measure as well for the degree of nuclear stability, **and it is calculated from the relation:**

$$\text{Nuclear binding energy per nucleon } \left(\frac{BE}{A}\right) = \frac{\text{Total nuclear binding energy}}{\text{Number of nucleons (mass number)}}$$

- Nuclear binding energy per nucleon in the nucleus of the helium atom ${}^4_2\text{He}$
= $\frac{28.3955}{4}$ = MeV 7.099

The stability of the nucleus increases with the increase of nuclear binding energy per nucleon $\left(\frac{BE}{A}\right)$.

Example

Calculate the nuclear binding energy - in mega electron volt - for the nucleus of the oxygen atom ${}^{16}_8\text{O}$

Then **calculate nuclear binding energy per nucleon**. [Knowing that: the atomic mass of the oxygen isotope is 15.995 u and the mass of the proton is 1.0073 u and the mass of the neutron is 1.0087 u.]

Solution

The theoretical mass of the nucleus of the oxygen atom

$$= (8 \times 1.0087) + (8 \times 1.0073) = 16.128 \text{ u}$$

The amount of mass defect (u) $0.133 \text{ u} = 15.995 - 16.128 =$

$$\text{Nuclear binding energy} = 123.823 \text{ MeV} = 931 \times 0.133$$

$$\text{Nuclear binding energy per nucleon} = 7.739 \text{ MeV} = \frac{123.823}{16}$$

Nuclear changes and the energy associated with them

- The changes that occur in the nucleus of an atom of an element, which lead to its transformation into the nucleus of another atom, are known as **nuclear reactions**, and the element that undergoes the change is called **the parent element and the resulting element is called** the daughter element.
- The algebraic sum of the atomic numbers and mass numbers must be equal on both sides of the nuclear equation.
- There are two types of changes that occur inside the nuclei of elements: the first is known as **natural radioactivity** and the other is known as **artificial radioactivity**.

Firstly natural radioactive activity

- The nuclei of atoms of radioactive elements tend to emit radiations, such as: alpha, beta, and gamma from their nuclei spontaneously to transform into nuclei of other more stable elements.

1 Alpha emission α

- The alpha particle α (**the nucleus of a helium atom**) contains 2 protons, 2 neutrons, and is denoted by the symbol ${}^4_2\text{He}$, and for this reason it is considered **positively charged** (2+).

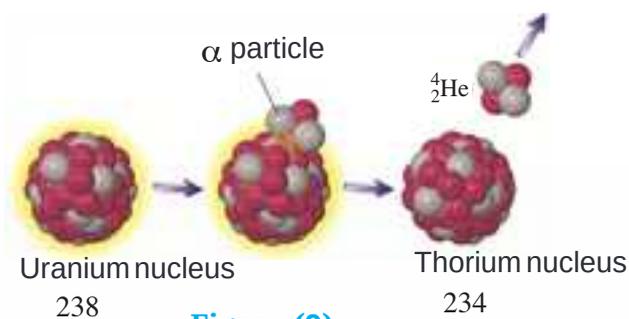
The emission of an alpha particle α from the nucleus of a radioactive atom leads to the formation of **a new element** whose atomic number is **less** by 2 and whose mass number is **less** by 4

Example:

(1) The transformation of uranium isotope 238 - into the isotope thorium 234 - (Figure 9):



Alpha particle Thorium 234 – Uranium 238 –
(the daughter element) (the parent element)



Emission of an alpha particle from an unstable nucleus

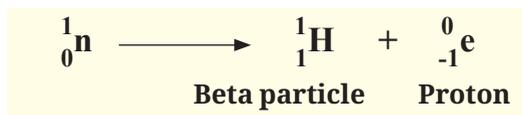
(2) The transformation of actinium isotope 227 - into francium isotope 223 -



Alpha particle Francium – 223 Actinium – 227

2 Negatron emission Negatron Emission

- The **negatron** (negative beta particle β^-), also known as **nuclear electron** ${}_{-1}^0e$, originates from the nucleus of the radioactive atom due to the transformation of a neutron into a proton.



The **emission** of a negatron from the nucleus of a radioactive atom leads to the formation of **a new element** whose atomic number is **greater** by 1 with **no change** in the mass number, **such as**:

- * **The transformation of the carbon isotope 14 - into the nitrogen isotope 14 - (Figure 10).**

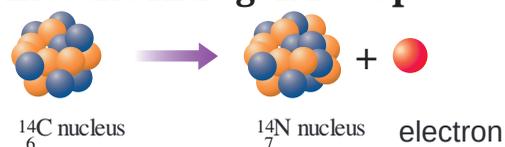
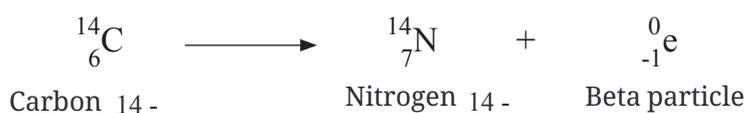
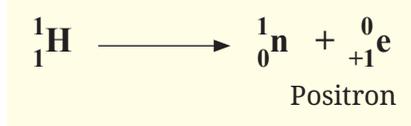


Figure (10)

Emission of a nuclear electron from an unstable nucleus

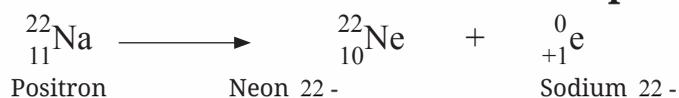
3 Positron emission

- The **positron** (positive beta particle β^+) is similar to the negative beta particle in all aspects, except that it carries a positive charge ${}_{+1}^0e$
- The positron is produced by the transformation of a proton into a neutron.



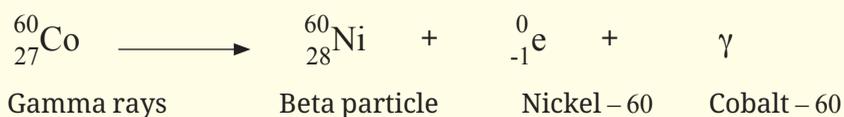
The **emission** of a positron leads to the formation of **a new element** whose atomic number is **less** by 1 with **no change** in the mass number, **such as**:

- * **The transformation of sodium isotope 22 -into neon isotope 22 -**



4 Gamma emission γ

- Gamma ray emission often accompanies the emission of alpha or beta particles.



- **Gamma rays** γ are photons (electromagnetic rays) of high energy and without mass or charge (Figure 11), and thus their emission does not cause a change in the mass number or atomic number of the radioactive element, but it leads to the stabilization of the nucleus of the same element by releasing the excess energy from the nucleus.

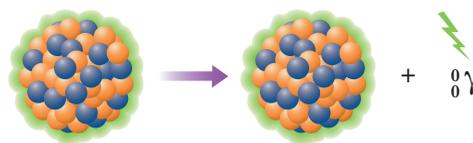


Figure (11)

Emission of gamma rays from the nucleus of a radioactive element

Example: Gamma emission from the nucleus of barium 137 –

Half-life time Half life time

- Scientists deduced from the study of radioactivity that the activity of the radioactive substance decreases over time (Figure 12), and the time required for half of the nuclei of the radioactive element to decay is called **the half-life time** $\frac{1}{2} t$, and it repeats at successive and equal time intervals, and this time varies from one radioactive element to another.

The number of decays D is calculated from the relation:

$$D = \frac{\text{Total decay time } t}{\text{Half-life time } t \frac{1}{2}}$$

- The half-life time of the radioactive element may be millions of years or it may be several seconds, as illustrated in table (3).

Example

The manganese isotope - 56 decays by losing beta particles, with a half-life time of 2.6 h

Calculate the time required for the activity of the sample to become $\frac{1}{4}$ what it was.

Solution

$$1 \xrightarrow[(1)]{\frac{t_1}{2}} \frac{1}{2} \xrightarrow[(2)]{\frac{t_1}{2}} \frac{1}{4}$$

$$t = D \times t_1 = 2 \times 2.6 = 5.2 \text{ h}$$

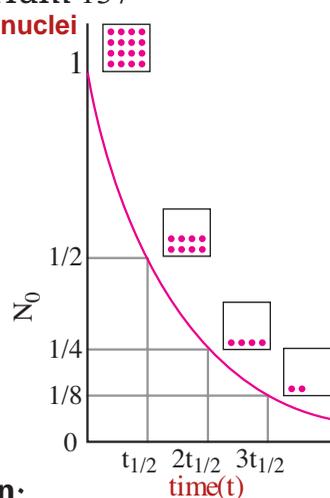


Figure (12)

the relation between the number of unstable nuclei and their decay time

Table (3) “For illustration only”

Radioactive isotope	Half-life time
Uranium - 238	4.51×10^9 years
Radon - 222	3.82 days
Thorium - 231	25.5 hours
Bismuth - 215	8 minutes
Polonium - 214	1.6×10^{-4} second

Science, Technology and Society (STS)

Smoke detector

The smoke detector is a device used to issue an alarm bell in case of fire (Figure 13), and its working principle depends on the activity of the americium isotope -241, which is characterized by its long half-life time.

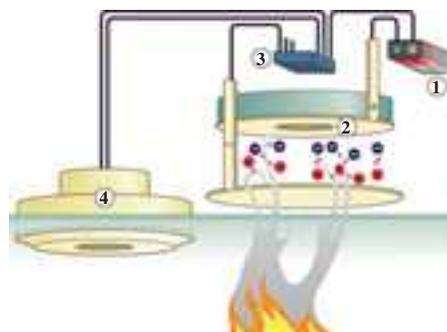


Figure (13)

Smoke detector

Structure:

- ① A source of electric current.
- ② A sensor (probe) containing the americium isotope -241
- ③ An electronic chip.
- ④ Alarm bell.

How it works:

- (1) The sensor is connected to a source of electric current, at the same time the americium isotope -241 emits alpha radiation that ionizes air molecules.
- (2) Under normal conditions (absence of smoke), air ions conduct electric current between the electrodes, so the electric current flows in the electronic chip.
- (3) When a fire occurs, smoke particles rise to the sensor (probe), disrupting the ionization process and thus stopping the air from conducting electric current. The electronic chip responds by conducting electric current to the alarm bell, which emits a sound to alert to the occurrence of a fire.

Secondly artificial radioactive activity

Artificial activity is classified into three types of reactions, which are:

- ① Transmutation reactions.
- ② Nuclear fission reactions.
- ③ Nuclear fusion reactions.

1 Transmutation reactions

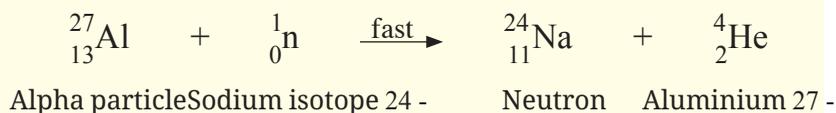
- **Transmutation** reactions are defined as nuclear reactions in which the nucleus of an element (known as the target) is bombarded by a particle with suitable kinetic energy (known as the projectile), transforming it into a new nucleus with different physical and chemical properties. Table (4) illustrates the projectiles used in such reactions, and the **neutron** projectile is considered the best of these projectiles because it is electrically neutral and does not lose part of its energy while reaching the nucleus, as it is not affected by the electrons surrounding the nucleus.

Table (4)

the particle	symbol
Alpha	${}^4_2\text{He}$
proton	${}^1_1\text{H}$
Neutron	${}^1_0\text{n}$
Deuteron	${}^2_1\text{H}$

Examples of transmutation reactions include

- Transformation of the aluminium isotope 27 - into the sodium isotope 24 - using a **neutron projectile**.



2 Nuclear fission reactions

- Nuclear fission reactions are defined as reactions in which the nucleus of a radioactive (heavy) atom is bombarded by a light nuclear projectile with low kinetic energy, splitting into two nuclei of new, more stable elements, and their combined masses are less than the mass of the radioactive element. A number of neutrons and enormous energy are also produced (Figure 14).

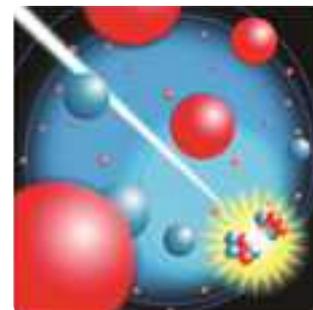
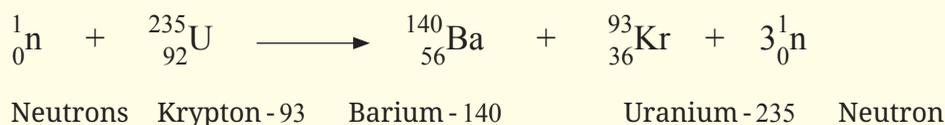


Figure (14)
Nuclear fission

As illustrated by the fission of uranium isotope - 235:



The neutrons produced from the fission reaction act as projectiles for similar fission reactions, and thus the fission reaction continues once it starts, and for this reason it is called **the chain fission reaction** (Figure 15).

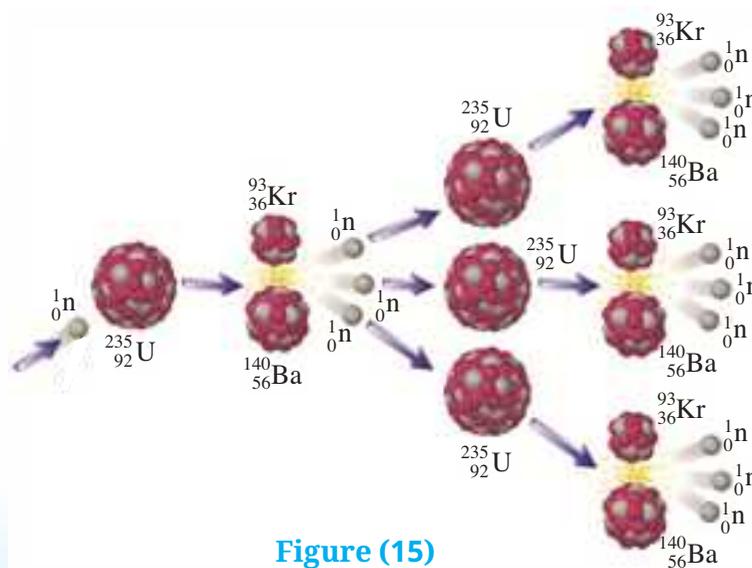


Figure (15)

Chain fission reaction

- Fission nuclear reactors are among the important peaceful applications of chain fission reactions, and the main reaction in them is the fission of uranium isotope - 235 (nuclear fuel) by bombarding it with slow neutrons. The Dabaa reactor project (Figure 16) is considered the first nuclear power plant in Egypt.



Figure (16)

Model of the Dabaa reactor

Uranium enrichment

- **Natural uranium contains:**

- Uranium isotope - 238 at about % 99.3, which is not fissionable.
- Uranium isotope - 235 at about % 0.7, which is fissionable.

Therefore, the process of increasing the proportion of uranium isotope 235 - in natural uranium is carried out, and this process is known as **uranium enrichment**. The enrichment percentage varies according to the type of targeted nuclear reaction.

In power generation reactors, enrichment is low (% 3 - 5), while it reaches more than % 90 in the manufacture of internationally prohibited nuclear weapons.

Fission reactions produce a huge amount of thermal energy; the produced energy from the fission of 1 g of uranium - 235 is **equivalent to** the thermal energy produced from burning 3 tons of coal.

3 Nuclear fusion reactions

- **Nuclear fusion** reactions are defined as the process of merging two light nuclei to form the nucleus of another element whose mass is less than the sum of the masses of the fused nuclei, where the mass deficit is converted into energy, according to Einstein's law. Nuclear fusion reactions (Figure 17) are considered the opposite of nuclear fission reactions.

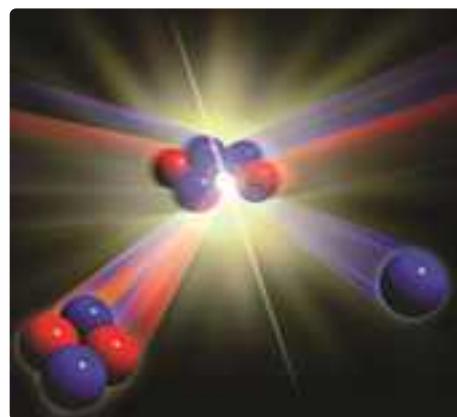


Figure (17)

Nuclear fusion

Examples of fusion reactions that occur in fusion nuclear reactors include the reaction represented in Figure (18).

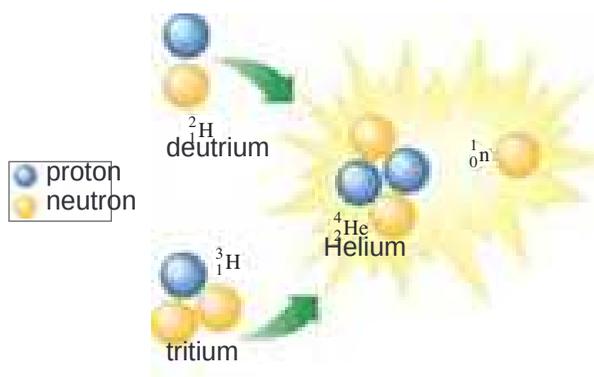
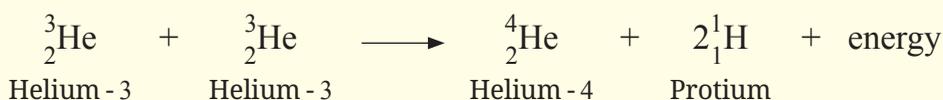


Figure (18)

Fusion of the deuterium nucleus with the tritium nucleus



- And among the nuclear fusion reactions that occur inside the Sun's star:



- **The differences between chemical reactions and nuclear reactions can be summarized in the table (5):**

Table (5)

Chemical reactions	Nuclear reactions
• Occur through the electrons of the outer energy level.	• Occur through the nucleons of the nucleus.
• Follow the law of conservation of matter.	• Follow the law of conservation of energy.
• Are usually carried out in uncomplicated conditions.	• Are carried out in complex conditions and with special safety precautions.
• Are accompanied by the release or absorption of a limited amount of energy.	• Are accompanied by the release of huge amounts of energy.

Science, Technology and Society (STS)

- Nuclear radiation is considered a fundamental pillar in a vast number of peaceful technological applications that drive progress in many fields, such as:

A The field of medicine

- Radioactive isotopes are used in imaging, diagnosis, and treatment, **and the most important of these are:**
 - * **Technetium isotope 99 -** : It is one of the most widely used radioactive isotopes in modern medicine, especially in imaging bones and lungs.
 - * **Thallium isotope 201-** : Used to assess the activity of the heart.

B The field of industry

- **Nuclear radiation emitted from radioactive isotopes is used in:**
 - * Detecting leakage locations in natural gas and petroleum pipelines.
 - * Nuclear batteries that are characterized by long life and high performance for up to 80 years.
 - * Some industrial reactions, such as using gamma rays as an alternative to high heat, pressure, and catalysts in certain chemical reactions to convert non-ionized molecules into ions, which facilitates the occurrence of these reactions.

C The field of agriculture

- **Radioactive isotopes are used in** eliminating harmful insects by exposing them to gamma rays emitted from cobalt-60 isotope. They are also used in breeding and developing wheat strains resistant to drought and salinity, which can be cultivated in desert environments.

Some radiation is also used to kill bacteria that spoil food, without changing the taste or quality (Figure 19).



Figure (19)

Two samples of strawberries were left in the air for 3 days (the left sample was exposed to gamma rays)

d The field of scientific research

- Isotopes are used in many scientific studies, **such as:**
 - * Tracking the path of phosphorus in plants by mixing normal phosphorus with radioactive phosphorus, and determining the time it reaches the plant leaves by using a Geiger counter (Figure 20) used to detect nuclear radiation.
 - * Determining the source of oxygen produced from the photosynthesis process using the oxygen isotope - 18, where it was found that its source is water and not CO_2 gas
- It is used to estimate the ages of pharaohs' mummies using carbon isotope 14 - (Figure 21).



Figure (20)
Geiger counter



Figure (21)
**Mummy of King
Ramses III**

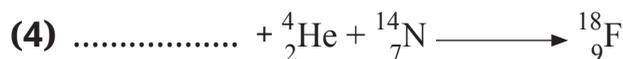
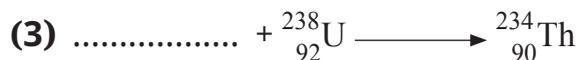
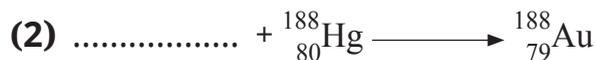
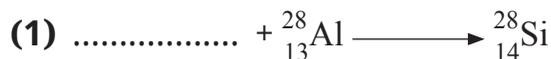
Evaluation of Chapter Three



1 Choose the correct answer from the given answers:

- (1) The number of nucleons is represented by each of
- (a) Protons and electrons. (b) Protons and neutrons.
(c) Alpha particles and beta particle.
(d) Positrons and electrons.
- (2) Isotopes of the same element agree in all of the following, except
- (a) Number of protons. (b) Number of electrons.
(c) Physical properties. (d) Chemical properties.
- (3) All of the following are affected by the electric field, except
- (a) Alpha rays. (b) Positron rays. (c) Beta rays. (d) Gamma rays.
- (4) 0.00234 u of a substance is converted to energy of
- (a) 2.179 MeV (b) 5.146 MeV (c) 9.302 MeV (d) 13.541 MeV
- (5) The emission of an alpha particle from the nucleus of a radium atom - 226 leads to the formation of the isotope
- (a) Radon - 220 (b) Radon - 222
(c) Thorium - 232 (d) Thorium - 228
- (6) The positron is produced from the transformation of
- (a) Proton to electron. (b) Neutron to proton.
(c) Electron to neutron. (d) Proton to neutron.
- (7) A sample of a radioactive element contains 4.8×10^{12} atom and its half-life is 2 yr. What is the number of decayed atoms after 8 yr?
- (a) 2.4×10^{12} atom (b) 4.2×10^{12} atom
(c) 3.6×10^{12} atom (d) 4.5×10^{12} atom
- (8) The operation of the smoke detector depends on the isotope
- (a) Americium - 241 (b) Uranium - 235
(c) Nitrogen - 14 (d) Magnesium - 24

3 Complete the following nuclear equations with what is appropriate:

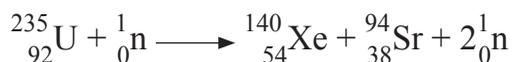


4 Compare between each of the following:

(1) Alpha radiation **and** beta radiation **and** gamma radiation.

(2) Chemical reactions **and** nuclear reactions.

5 calculate the amount of energy produced from the following reaction Estimated in (MeV):



And knowing the data shown in the opposite table.

mass (u)	Substance
235.044	${}_{92}^{235}\text{U}$
1.0087	${}_0^1\text{n}$
139.905	${}_{54}^{140}\text{Xe}$
93.906	${}_{38}^{94}\text{Sr}$

6 calculate the number of alpha particles emitted from the nucleus of the thorium isotope ${}_{90}^{228}\text{Th}$ in order to transform into the polonium isotope ${}_{84}^{216}\text{Po}$

7 If the half-life of the cobalt isotope 60 - is 5.3 yr , **calculate the remaining mass** of a sample of it with a mass of mg 1 after 15.9 yr

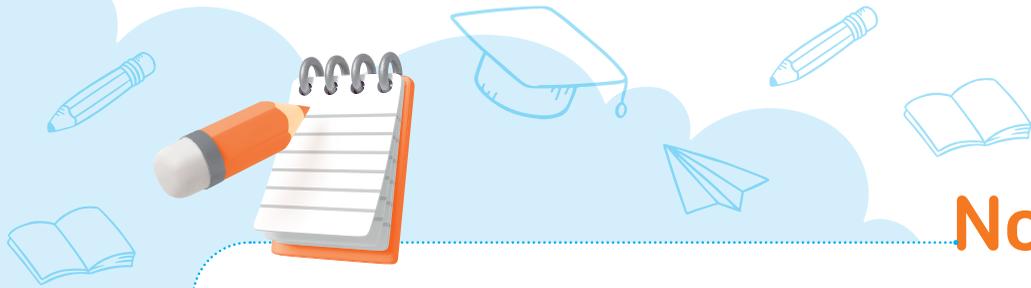
8 Explain the working principle of the smoke detector.

9 Explain the importance of each of the following isotopes:

(1) Thallium isotope 201 - in the field of medicine.

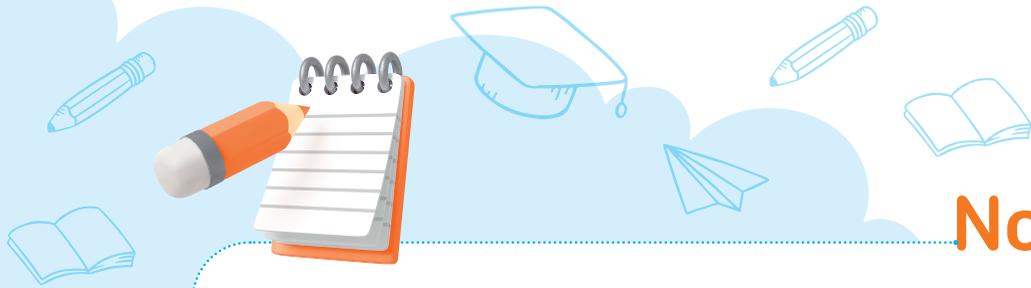
(2) Cobalt isotope 60 - in the field of agriculture.

(3) Oxygen isotope 18 - in the field of research.



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Notes

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