Atoms do not exist in a free state. They join together to form chemical structure.

The cohesion of the particles in the chemical structure is due to interactions between the different atoms or ions involved, which lead to the formation of the most stable structure.

These interactions are called ***chemical bonds***.

I-The different types of bonds

There are two types of bonds: strong bonds and weak bonds.

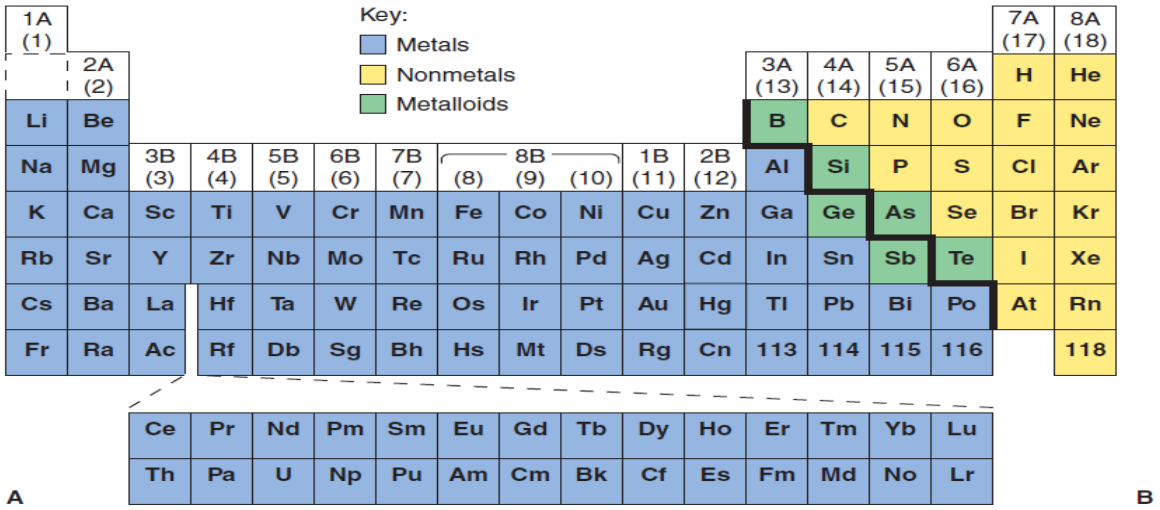

Figure: A comparison of metals and nonmetals.
A, Location within the periodic table.
B, Relative magnitudes of some atomic properties across a period

All chemical bonds involve the valence states of atoms, but the latter are involved in different ways, giving rise to several types of bonds.

I-1-Strong Bonds

I-1-1-Metallic bond

A metallic bond involves a very large number of atoms (several million). These atoms share one or more electrons, known as 'free electrons'.

Metallic bonds are formed between atoms with low electronegativities and few electrons on their outer layer, but they can only form if there is a sufficiently large number of such atoms. This is why it is rarely observed except in certain solids or liquids.

MM
Chapter V Chem

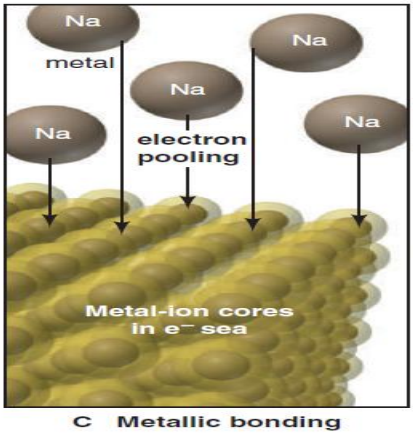

Figure I-1-1: Metallic bond

I-1-2-Covalent bond

a covalent bond is a chemical bond in which each of the bonded atoms shares an electron from one of its valence layers to form a doublet of electrons linking the two atoms.

Only an atom's valence electrons, i.e. electrons in excess of the configuration of the rare gas that precedes the element, can be involved in bond formation.

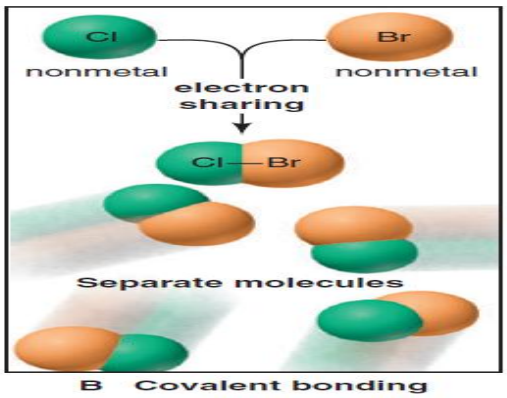

Figure: covalent bond

a- Lewis structure

The Lewis structure consists of the allocation of electrons on or between the atoms of the molecule. Only valence electrons are considered.

The Lewis representation shows the external electronic structure, also known as the valence layer.

Paired electrons are represented by dashes ——

Single electrons are represented by dots .

The number of single electrons gives the valency of the atom

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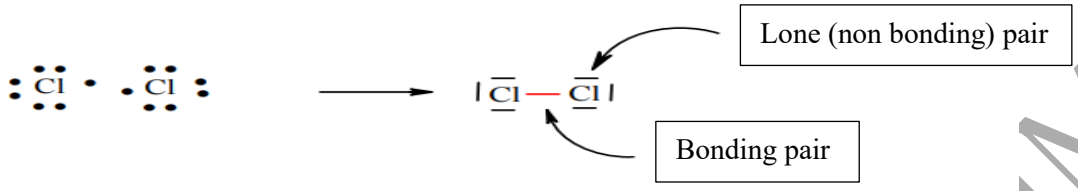
A distinction is made between "bonding doublets (or pairs)" and "non-bonding doublets (or pairs)", also known as "free" doublets.

Example: Formation of the covalent bond of chlorine



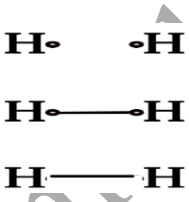
b- The octet Rule

In a molecule, each atom tends to surround itself with eight electrons, except for the hydrogen molecule, to have the configuration of the noble gases ⇒This is what is known as the Octet rule

There are three types of covalent bond:

a-Apolar or non-polar covalent bond: If the difference in electronegativity between these two atoms is zero or less than 0.4, the bond is said to be apolar.

Example: According to the Lewis Representation, H_2 is represented as:



b-Polar covalent bond: A covalent bond is polar if the difference in the electronegativities of the two atoms forming the bond is not zero. The most electronegative atom in a polar bond attracts the electrons in the covalent bond towards itself.

Example:

c-Dative covalent bond: It forms between two atoms where one of the atoms (atom B) (the donor) provides a lone pair from its outer shell, and the other atom (atom A) (the acceptor), which has an electron deficiency (empty box) in its valence shell (rectangle), receives this lone pair.

Example:

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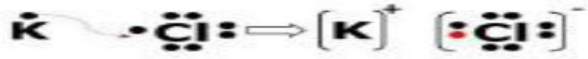
I-1-3-Ionic bond

Ionic bonds are formed between ions of opposite charge. These bonds involve the classic electrostatic force. An ionic bond is the result of one or more ions being transferred from one atom to another. The atoms thus become ions; the one that gives up the electron is the cation and the one that receives the electron is the anion. The transfer is made from the least electronegative atom to the most electronegative atom.

This bond forms between two atoms, A and B, with a significant difference in electronegativity. This difference leads to the transfer of electrons from the less electronegative atom to the more electronegative atom, resulting in the formation of a cation (carrying a positive charge) and an anion (carrying a negative charge).



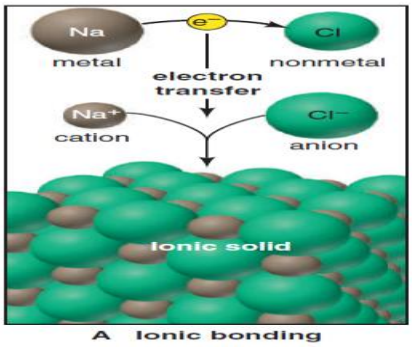


Figure I-1-3: Ionic bond

a- Concept of dipole moment μ

A dipole is a system formed by two equal but opposite charges separated by a distance d.

It is considered an electric dipole if the difference in electronegativity between the two atoms creates polarization, it results in a negatively charged pole rich in electrons and a positively charged pole lacking electrons at a distance (d) from each other

This dipole is characterized by its dipole moment or electric moment (μAB). This quantity (μAB) is represented by a vector oriented from the negative charge (more electronegative element) to the positive charge (less electronegative element).

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The notations +δ and - δ represent partial charges, the more electronegative atom that carries the partial negative charge.

A dipole is characterised by its electric dipole moment μ such that:

$$\left\|\vec{\mu}\right\| = q.d = \mu$$

Where:

d is the length of the A-B bond, δ is the partial charge.

The effective partial charges are: +q=+δ·e and −q=−δ·e, where:

δ is the partial charge,

e is the elementary charge.

$$\boldsymbol{\mu = |\delta| \cdot e\ d}$$

μ is expressed in Coulomb. Metre (C.m). This quantity is also often expressed in Debye :

1 Debye = 1 D= 3.33. 10^{-30} C.m

By convention, we can know the nature of the bond by simply calculating the difference in electronegativity of the atoms forming the bond.

Δχ : difference in electronegativity $\Delta\chi = |\chi_2 - \chi_1|$

- $\Delta\chi \leq 0.4$ ⇨ non polar covalente bond non polar
- $0.4 < \Delta\chi < 1.7$ ⇨ polar covalente bond
- $\Delta\chi \geq 1.7$ ⇨ ionique bond

b-Partial ionic character of a bond

the percentage of ionic character of the bond is calculated from the following formula:

$$\textbf{\% ionique} = \frac{\mu_{exp}}{\mu_{theo}} . \mathbf{100} = \frac{\delta.d}{e.d} . \mathbf{100}$$

Where:

d: length of the bond

e: electron charge

δ: partial charge

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I-2- Weak Bonds

Weak bonds have dissociation energies ranging from about 1 to 40 $kJ \cdot mol^{-1}$, in any case, less than 50 $kJ \cdot mol^{-1}$. These bonds result from cohesive forces between atoms incapable of forming valence bonds or between molecules where the possibilities of strong bonds are already saturated. There are two types of weak bonds: Hydrogen bond and van der Waals bonds.

II- Geometry of molecules: GILLEPSIE or VSEPR theory (Valence Shell Electron Pair Repulsion)

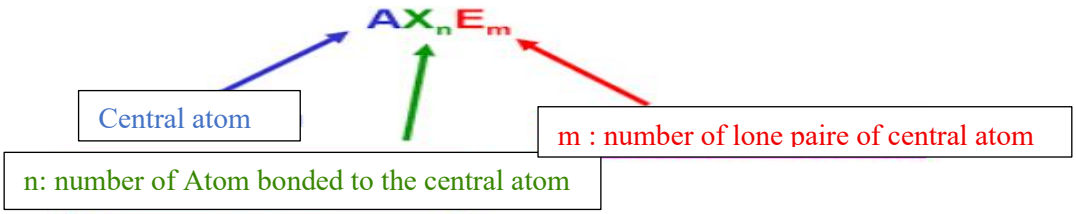
Gillepsie's theory or the VSEPR method is a fairly recent method (1960). The aim of this theory is to predict the geometry of molecules.

The geometry of a polyatomic molecule is the arrangement of its atoms in space.

The doublets of the outer electrons (binding or non-binding) of a single atom repel each other: the geometry adapted by a molecule is one in which the electron doublets of each atom are as far apart from each other as possible.

The **VSEPR** method can be used to predict the geometry of simple molecules or ions. It applies to molecules or ions of type : $\mathbf{AX_nE_m}$

A: designates the central atom which is bonded to n atoms X and has m pairs of free electrons E (or m non-bonding doublets E)



II-1-Principle of the VSEPR method

The VSEPR method is based on a number of assumptions, particularly concerning the nature of the bonds between atoms:

- Atoms in a molecule are linked by pairs of electrons.
- Two atoms can be linked by more than one pair of electrons. These are known as multiple bonds.
- Some atoms may also have pairs of electrons that are not involved in a bond. These are known as non-bonding doublets.
- The electrons making up these bonding or non-bonding doublets exert repulsive electrical forces on each other. The doublets are therefore arranged around each atom in such a way as to minimise the values of these forces.

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- Non-bonding doublets take up more space than bonding doublets.
- Multiple bonds take up more space than single bonds.

The sum (n+m), called the total number of electrons pairs, provides information about the basic geometry of the molecule

In table below we present different Possible Shapes Using the VSEPR Theory

Table II-1: Different Possible Shapes Using the VSEPR Theory

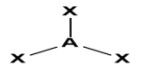
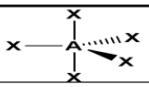
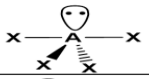
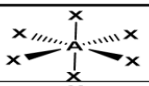
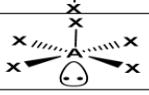
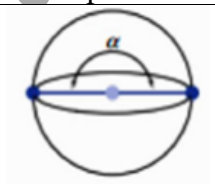
AXE Formula	Molecular Geometry	Bond Angle	Molecule Shape
AX_2E_0	**Linear**	180°	
AX_3E_0	**Trigonal planar**	120°	
AX_2E_1	Bent	119°	
AX_4E_0	**Tetrahedral**	109.5°	
AX_3E_1	Trigonal pyramidal	107.3°	
AX_2E_2	Bent	104.5°	
AX_5E_0	**Trigonal bipyramidal**	90°, 120°, 190°	
AX_4E_1	See-saw	86.5°,102°, 187°	
AX_3E_2	T-shape	87.5°, 185°	
AX_2E_3	Linear	180°	
AX_6E_0	**Octahedral**	90°	
AX_5E_1	Square pyramidal	84.8°, 180°	

Table II-2: Examples for different Possible Shapes Using the VSEPR Theory

Repulsion	AXnEm	n+m	n	Geometry	Example
 α=180°	AX_2	2	2	linear	$BeCl_2$, CO_2, HCN

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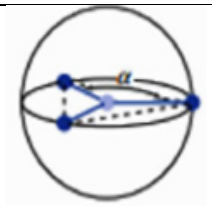
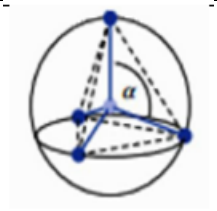
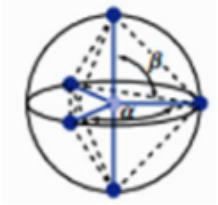
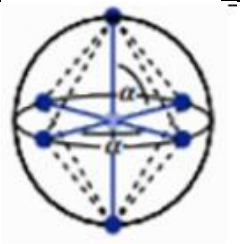
 α=120°	AX_3	3	3	Trigonal plan	BF_3, $AlCl_3$, NO_3^-
	AX_2E_1		2	in V	SO_2, $SnCl_2$,
 α=109,47°	AX_4	4	4	Tetraedric	CH_4, $SiCl_4$, NH_4^+
	AX_3E_1		3	Pyramid	NH_3, H_3O^+, PCl_3
	AX_2E_2		2	in V	H_2O, H_2S
 α=120°, β=90°	AX_5	5	5	Bipyramide	PCl_5
	AX_4E_1		4	Bipyramide	$TeCl_4$
	AX_3E_2		3	in T	IC_{l3}
	AX_2E_3		2	linear	XeF_2
 α=90°	AX_6	6	6	Octaedric	SF_6
	AX_5E_1	6	5	Pyramid	BrF_5
	AX_4E_2		4	scare	XeF_4

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