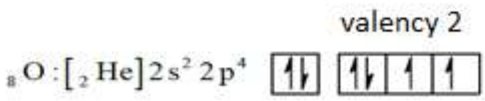
Chapter Six: Chemical Bonds

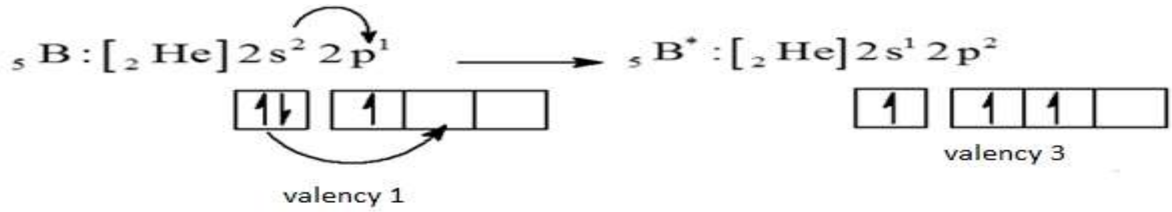
1-Chemical Valency

Valency is the number of electrons gained, lost, or shared by an atom with another atom to form a chemical bond. It is not necessary for the valency of an atom to be equal to the number of valence electrons in all cases. For example, oxygen has six valence electrons, but its valency is 2.



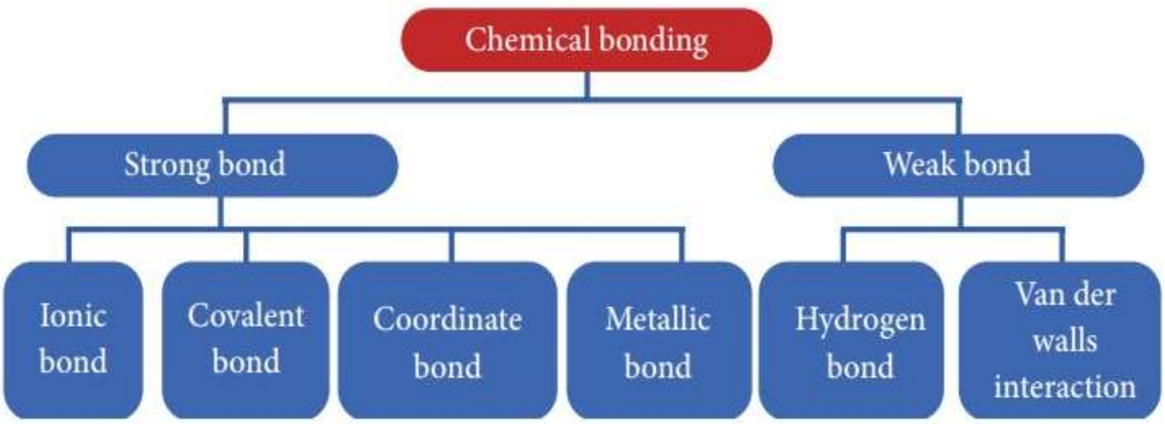
Additionally, the valency of an element can be increased or decreased by exciting it, where electrons can change energy levels. An atom's valency can only be increased if it possesses both paired electrons and vacant quantum energy levels in the outermost shell simultaneously.

Example: Boron atom $_5B$ has a valency of 1, but when excited, it becomes 3.



2- Chemical Bonds

Atoms of elements bond with each other to achieve stability or attain a configuration similar to the nearest inert gas. Chemical bonds are classified based on how valence electrons contribute to the formation of the bond, which can be strong if formed between atoms in the same molecule or weak if formed between different molecules.



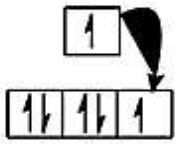
2-1- Ionic Bond (Electrovalent Bond)

The ionic bond is present in compounds resulting from the reaction between metallic and non-metallic elements, where electrons transfer from the outermost shell of the metallic atom to the non-metallic atom. The ionic bond is encountered in compounds resulting from the union of elements with extremely disparate electronegativities, such as halogens, with other elements possessing very high electronegativities, like those in the first column of the periodic table. Where the electronegativity difference exceeds the value of 1.7. The ionic bond is an electrostatic attraction, meaning it has no material existence and no specific direction. It arises from the attraction between positive ions and negative ions.

EXP :

$_{11}Na:[_{2}He]2s^2 2p^6 3s^1$

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



$_{11}Na^+:[_{2}He]2s^2 2p^6 \equiv {}_{10}Ne$

$_{17}Cl^-:[_{10}Ne]3s^2 3p^6 \equiv {}_{18}Ar$

2-2- Covalent Bond (Coordinate - Covalent Bond)

A bond forms between two non-metal atoms when the electronegativity difference is less than 1.7. Atoms tend to share electrons from their valence shells in an attempt to achieve the stable electron configuration of a noble gas. If two atoms share one electron each, the bond is called a single bond, for example, H-Cl. If they share two electrons, it is a double bond, as in O=O, and if they share three electrons, it is a triple bond, as in N≡N.

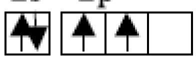
- **Lewis model**

The Lewis model can be used to represent molecules of covalent compounds and illustrate how bonds are formed through Lewis structures. The Lewis method involves writing the element symbol and surrounding it with dots, representing the valence electrons possessed by the atom. These dots are distributed in a way that does not exceed two in each of the four directions around the atom. A single dot represents a lone electron, and a pair of electrons is depicted as two adjacent dots or a small straight line, symbolizing a shared bond. If the shared electron pair is between two atoms, it is referred to as a bonding pair, while if it is not, it is termed a non-bonding pair.

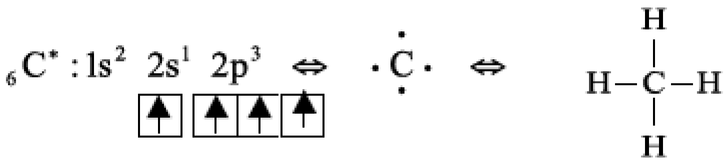
The central atom is the one with the most free electrons, and in cases where two atoms have the same number of free electrons, the central atom is the one with the lower electronegativity.

EXP:

2- CH_4 $\quad {}_6C:1s^2\ 2s^2\ 2p^2 \quad \Leftrightarrow \quad \cdot\overline{C}\cdot$ et $4H\cdot$

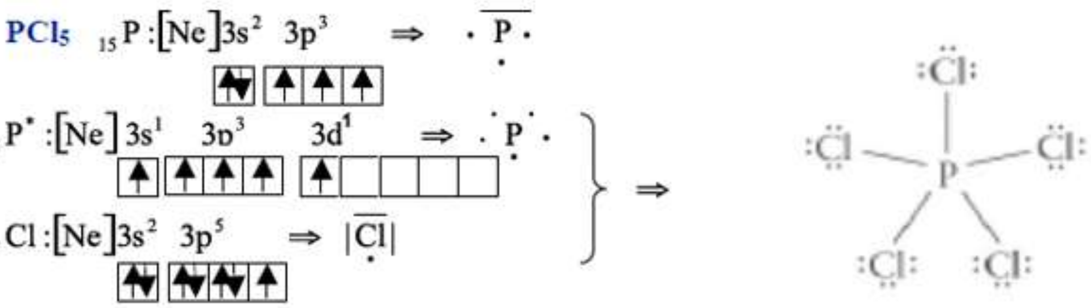


Carbon transitions from its ground state to an excited state (C*) to have 4 unpaired electrons and form 4 bonding pairs with hydrogen (H)



- **Octet Rule:** When atoms within a molecule are surrounded by eight electrons, we say that they satisfy the Octet Rule. In doing so, they tend to acquire the electronic structure of noble gases ($ns^2 np^6$).
- **Duet Rule:** When atoms within a molecule are surrounded by only two electrons, we say that they satisfy the Duet Rule, aiming to acquire the electronic structure of helium.

However, there are exceptions: Some elements can accept more than 8 electrons, especially those belonging to the third period, such as S and P. Also, some second-period elements like B and Be.

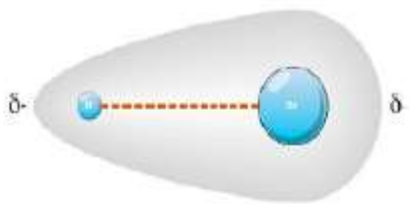


La règle de l'octet n'est pas vérifiée

- **Polarity of Bonds**

If the atoms participating in a covalent bond are identical, as in H2 and Cl2, the shared electron pair is equally distributed between the atoms, and the electron cloud is symmetrically arranged. The charge carried by each atom is zero, and in this case, we say the molecule is nonpolar.
However, if the atoms are not identical, as in HCl, HBr, CH_4, or any case where the atoms have different electronegativities, the electrons are not equally shared between the atoms. The electron cloud will be distorted towards the atom with higher electronegativity. In this case, the molecule is termed polar. For example, in hydrogen bromide (HBr), bromine carries a partial negative charge, δ^-, and hydrogen carries a partial positive charge, δ^+.
When: $\delta = 0$, the bond is purely covalent.
$\delta = 1$, the bond is purely ionic (completely polar).
$0 < \delta < 1$, the bond is polar covalent (partially polar).
The type of bond can be predicted based on the electronegativity difference between two atoms:

- If the difference is less than 0.3, the bond is nonpolar covalent.



- If it's greater than 1.7, the bond is ionic.
- Any value in between indicates a polar covalent bond.

➢ Dipole Moment

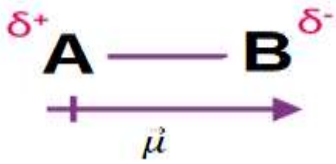
The polarity of a molecule is measured by its dipole moment, which is equal to:

$$\vec{\mu} = \delta.\vec{d}$$

Where:

- δ is the charge carried by one end of the atoms (partial charge).
- d is the bond length.
- μ is the dipole moment (the vector representing the strength and direction of the dipole).

δ < e because B attracts the electron towards itself and does not remove it completely (in the case of an ionic bond $\vec{\mu} = e.\vec{d}$).



The unit of the dipole moment is the debye (D), where:

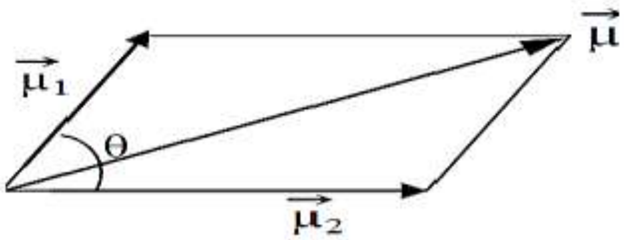
$$1\ \text{D}=3.33\times10^{-30}\ \text{C·m}$$

In the case of a molecule composed of more than two atoms, to calculate the total dipole moment, you sum the individual dipole moments vectorially for all the bonds in the molecule. This requires knowledge of the molecular geometry. Using one of the two relationships:

$$\vec{\mu}_T = \sum \vec{\mu}_J$$

$$\mu = \sqrt{\mu_1^2 + \mu_2^2 + \mu_1\mu_2\cos\theta}$$

$$\mu = 2\mu_1 \cos\frac{\mu_1\mu_2}{2}$$



EXP

- ## **Ionic character of the covalent bond (P.I)**

The ionic character of the covalent bond can be evaluated using the ratio of the experimentally measured dipole moment (μ) to the dipole moment calculated assuming an ionic bond ($\mu_i = |e|.d$). The formula to calculate this ratio is:

$$(C.I)P.I\% = \frac{\mu}{\mu_i} \times 100$$

According to general rules:

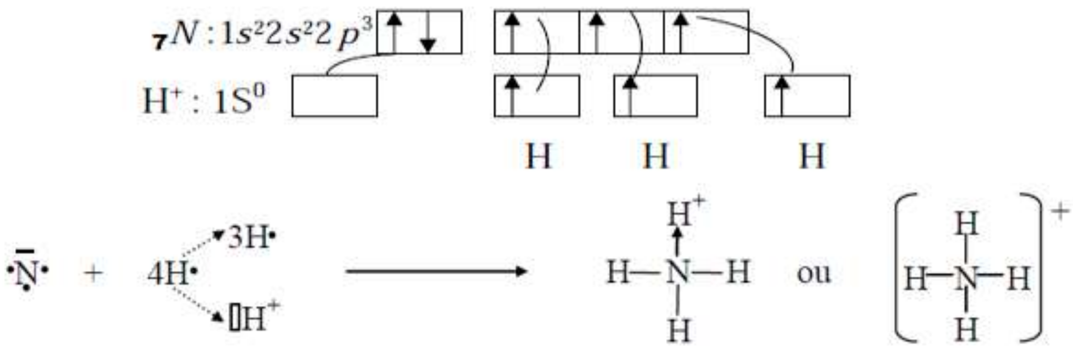
- If *P.I* is between 0 and 5%, the bond is considered nonpolar covalent.
- If it is between 5 and 50%, the bond is considered polar covalent.
- If *P.I* is exactly 50%, the character is 50% covalent.
- If it is between 50 and 100%, the ionic character predominates.

EXP:

What is the ionic character of the HF molecule, knowing that the measured dipole moment has μ = 1.99D and the bond length $d_{H-F} = 0.92\text{Å}$.

2-3-Coordinate (donor) bonding

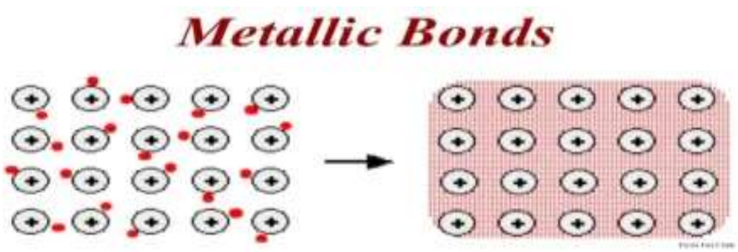
Coordinate (donor) bonding is a type of covalent bond that forms as a result of one atom contributing a pair of non-shared electrons to another in a bond. The atom that donates the electron pair is called the donor atom, while the receiving atom is called the acceptor atom, creating an empty orbital. These bonds can occur between atoms to form molecules or between an atom in a molecule and an ion, or between an atom in one molecule and an atom in another molecule.



2-4- Mettallic Bonds

Metals exhibit strong bonding with each other, and the explanation for this is that metal atoms readily lose some of their valence electrons, leaving behind positively charged particles (positive ions). These electrons form a cloud surrounding the positive ions.

These electrons are not attracted to just one or two ions but instead are attracted to the vast number of positive ions present. As a result, metal atoms are held together by a very strong force.



2-5- Hydrogen Bonding

Hydrogen bonding occurs between a hydrogen atom bound to an atom with high electronegativity and another atom with high electronegativity that is also bound in a different molecule. This second atom typically possesses a lone pair of electrons, as seen in elements such as fluorine (F), oxygen (O), nitrogen (N), and chlorine (Cl). A hydrogen bond is formed, represented as a hydrogen bridge.

$$A^{\delta-}{-}H^{\delta+}\text{-----}B^{\delta-}$$

EXP: