Introduction :

Matter is made up of one or different type of elements. Under normal conditions no other element exists as an independent atom in nature, except noble gases. Every system tends to be more stable and bonding is nature's way of lowering the energy of the system to attain stability.

The evolution of various theories of valence and the interpretation of the nature of chemical bonds have closely been related to the developments in the understanding of the structure of atom, the electronic configuration of elements and the periodic table. Different theories and concepts have been put forward from time to time. These are Kössel-Lewis approach, Valence Shell Electron Pair Repulsion (VSEPR) Theory,

I. KÖssel-Lewis Approach to Chemical Bonding

Always electrons in the valence shells of atoms participate in the formation of covalent bonds. In the process of suggesting the process of chemical bonding Lewis provided a very convenient way of representing bonding in simple molecules. This is called Lewis electron-dot structures or simply Lewis structures.

I.1. Lewis symbol

To represent the Lewis symbol of a particular element valence electron is represented as a dot. Some examples of Lewis symbol of various elements are given in the table.

Lewis symbol or electron dot symbols

element	Electronic configation	Valence electrons	Lewis symbol
$_1H$	$1s^1$	1	$\dot{H}$
$_2He$	$1s^2$	2	$\ddot{He}$
$_3Li$	$1s^2\ 2s^1$	1	$\dot{Li}$
$_6C$	Ground state :$1s^2\ 2s^2\ 2p^2$	4	$\cdot C \cdot$
	Excited state : $1s^2\ 2s^1\ 2p^3$	4	$\cdot\dot{\underset{\cdot}{C}}\cdot$
$_8O$	$1s^2\ 2s^2\ 2p^4$	6	$\cdot\ddot{\underset{\cdot\cdot}{O}}\cdot$
$_7N$	$1s^2\ 2s^2\ 2p^3$	5	$\cdot\ddot{\underset{\cdot}{N}}\cdot$

Note : ***Significance of Lewis Symbols :*** The number of dots around the symbol represents the number of valence electrons.

I.2. Octet Rule

tendency of atoms to achieve 8 electrons in their outer most shell is known as Lewis octet Rule (except in case of Hydrogen) for this atom can gain, loose or share the electrons.

Exampels :

H_2 : $_1H : 1s^1$ H.

H-H

$_1H : 1s^1$ H.

F_2 : $_9F : 1s^2/\ 2s^22p^5$ $|\overline{F}—\overline{F}|$ → doublet non liant (é non liants)

$_9F : 1s^2/\ 2s^22p^5$ doublet liant (é liants)

HCl: $_{1}H : 1s^1$

$_{17}Cl: 1s^2 2s^2 2p^6 / 3s^2 3p^5$ } H-Cl

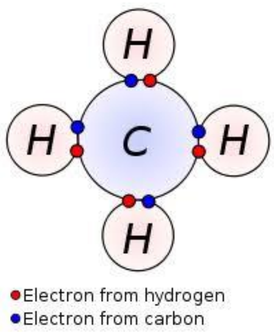
NH_3: $_{1}H : 1s^1$

$_{7}N: 1s^2 2s^2 2p^3$ }

$$\begin{matrix} \overline{} \\ H\text{-}N\text{-}H \\ H \end{matrix}$$

H_2O: $_{1}H : 1s^1$

$_{8}O: 1s^2 2s^2 2p^4$ } $H\text{-}\underline{\overline{O}}\text{-}H$



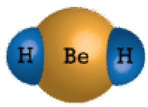
I.3. Exceptions of Octet Rule :

To the octet rule fall into one of three categories: an incomplete octet, odd-electron molecules and an expanded octet.

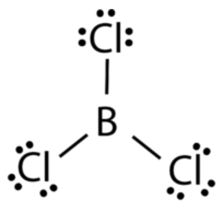
a. Incomplete Octet :

In some compounds, the number of electrons surrounding the central atom in a stable molecule is fewer than eight. example Beryllium only has two valence electrons, it does not typically attain an octet through

$$H\cdot \quad \cdot Be\cdot \quad H\cdot \longrightarrow H:Be:H$$

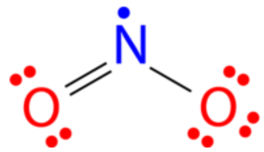


Boron and aluminum, with three valence electrons, also tend to form covalent compounds with an incomplete octet. The central boron atom in boron trichloride (BCl_3) has six valence electrons



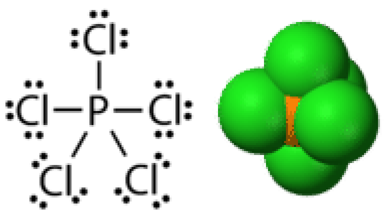
b. Odd-Electron Molecules :

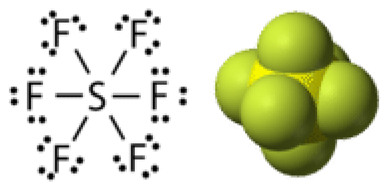
There are a number of molecules whose total number of valence electrons is an odd number. It is not possible for all of the atoms in such a molecule to satisfy the octet rule. An example is nitrogen dioxide (NO_2).



c. Expanded Octets

Atoms of the second period cannot have more than eight valence electrons around the central atom. However, atoms of the third period and beyond are capable of exceeding the octet rule by having more than eight electrons around the central atom. Starting with the third period, the d sublevel becomes available, so it is possible to use these orbitals in bonding, resulting in an expanded octet. example: Phosphorus and sulfur





II. Chemical Bond :

Three different types of bond may be formed depending on the electropositive or electronegative character of the atoms involved.

- Electropositive element + Electronegative element = Ionic bond (electrovalent bond)
- Electronegative element + Electronegative element = Covalent bond
- Electropositive + Electropositive element = Metallic bond.

The difference in electronegativity's of two elements can be used to predict the nature of the chemical bond.

• If the difference in electronegativity's is between:

1.7 to 4.0: Ionic

0.3 to 1.7: Polar Covalent

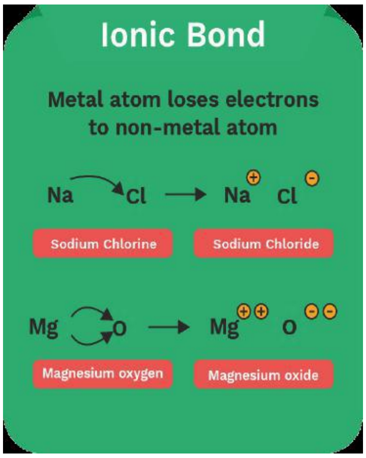
0.0 to 0.3: Non-Polar Covalent

Ionic Bond or Electrovalent Bond

y It is formed by complete transfer of e–s from one to another atom, so as to complete their octet by acquiring 8 electrons or duplet in case of H_2, Li_2 etc. and hence acquire the nearest stable noble gas configuration.

Example : NaCl, $MgCl_2$ etc.

For NaCl : Difference in electronegativity (ΔEN) is 2.1 (Na = 0.9, Cl = 3.0), so this is an ionic bond



II.2. **Metallic bond :**

Metallic bonding arises from the electrostatic force of attraction between conduction electrons and positively charged metal ions (Kernels)

II.3. Covalent Bonds

Types of covalent bonds:

- Non-polar covalent bond.
- polar covalent bond.

Strengths of Covalent Bonds

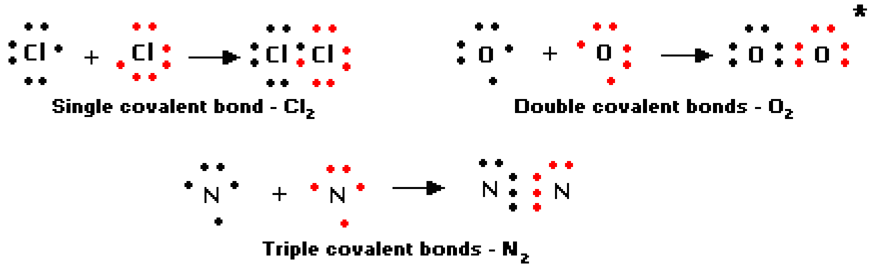
- ✓ We know that multiple bonds are shorter than single bonds.
- ✓ We know that multiple bonds are stronger than single bonds.
- ✓ As the number of bonds between atoms increases, the atoms are held closer and more tightly together.

II.3.1. Non-polar covalent bond

• It is a covalent bond in which electrons are shared equally between two atoms.

• Non-polar covalent bonds are formed, when the sharing atoms have the same electronegativity so the electrons are shared equally. Electron cloud is not displaced

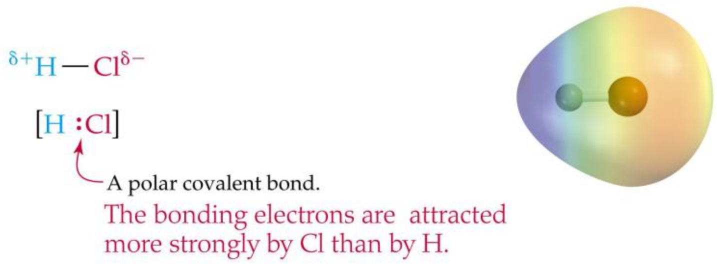
Example : Covalent bonds that are formed between identical atoms, as : Cl_2, O_2 and N_2.



II.3.2. Polar covalent bond

A type of covalent bond between two different atoms in which electrons are shared unequally. The shared electrons tend to be pulled more toward one atom than the other (because of electronegativity) and results on partial charges on the atom (i.e one end of the molecule has a slightly negative charge and the other a slightly positive charge.

Example : bond between the two atoms H and Cl



Note: Atoms of oxygen, nitrogen and phosphorus have a particularly strong tendency to pull electrons toward themselves when they bond with other atoms.

- ➢ Dipole moment

Differences in electron negativity cause electron density to be centered around one side of a molecule. This causes the molecule to become polarized with a partial negative and partial positive charge

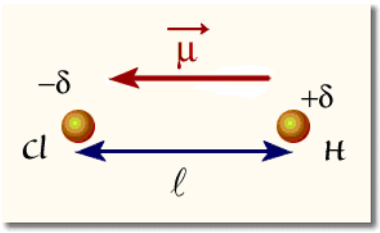
As a result of polarisation, the molecule possesses the **dipole moment** (depicted below) which can be defined as the product of the magnitude of the charge and the distance between the centres of positive and negative charge. It is usually designated by a Greek letter 'μ'. Mathematically, it is expressed as follows :

Dipole moment (μ) = charge (Q) × distance of separation (d)

Note : Dipole moment is usually expressed in Debye units (D). The conversion factor is : 1 D = 3.33564×10^{-30} C m ; where C is coulomb and m is meter.

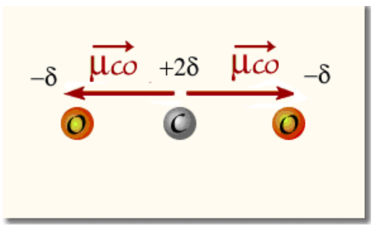
Example : The dipole moment of hydrogen chloride, with bond distance 127 pm, is 1.03D. The percent ionic character of its bond is about:

Solution



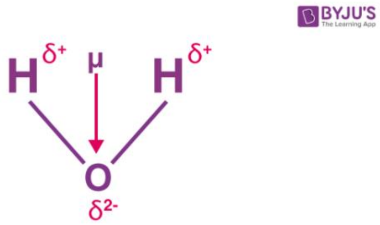
Theoritical dipole moment $=q\times d=1.6\times10^{-19}\times127\times10^{-12}=203.2\times10^{-31}$ Coloumb meter =6.1D (as $1D=3.33\times10^{-30}$C.m)

Thus, in the case of the CO_2 molecule, it is found that, due to its linear and symmetrical geometry, the two bonding moments C=O compensate each other. The resulting dipole moment is therefore zero.



In case of polyatomic molecules the dipole moment not only depend upon the individual dipole moments of bonds known as bond dipoles but also on the spatial arrangement of various bonds in the molecule. In such case, the dipole moment of a molecule is the vector sum of the dipole moments of various bonds. For example in

H_2O molecule, which has a bent structure, the two O–H (**μ (OH) = 1.5 D)** bonds are oriented at an angle of 104.50.



$\mu (H_2O) = 2\mu (OH) \times \cos(104.5/2) = 1.5 D \times 0.612 = 1.836 D$

$\mu (H_2O)$ $= 1.85 D = 1.836 \times 3.33564 \times 10^{-30}$ C m $= 6.12 \times 10^{-30}$ C m

The Valence Shell Electro n Pair Repulsion (VSEPR) Theory

Lewis concept is unable to explain the shapes of molecules. This theory provides a simple procedure to predict the shapes of covalent molecules. Sidgwick

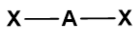
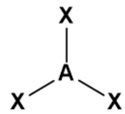
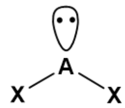
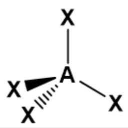
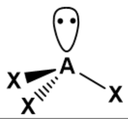
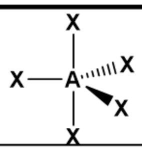
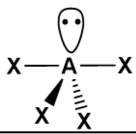
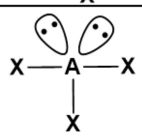
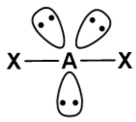
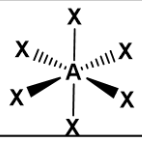
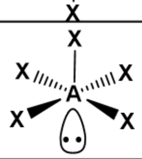
III.1. VSEPR Theory

VSEPR stands for valence shell electron pair repulsion. Valence shell electrons are on the outer part of an atom. These electrons can participate in bonding (bonding pairs) or act as non-bonding electrons (lone pairs). These bonding and lone pairs electrons repel each other.

- **AXE Method**

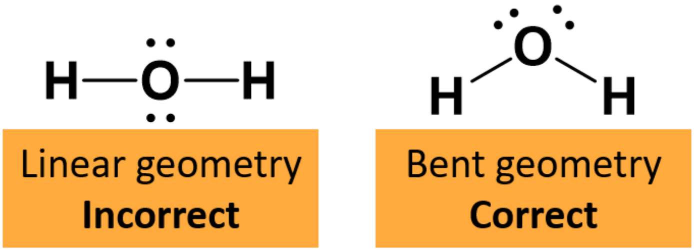
The AXE method is a way to characterize molecules and determine geometry. *A* stands for the central atom. The central atom in the atom which other atoms are bonded to, some molecules will have multiple central atoms. *X* stands for a ligands or atoms bond to central a central atom. *E* stands for the lone pairs on the central atom. For our water example the AXE method abbreviation is AE2X2. All molecules with this same designation will have a bent geometry.

In the chart below, the molecular geometry is given for some types of molecules. Notice that the molecular geometry does not include the lone pairs. If lone pairs are taken into account it is called lone pair geometry. The chart is broken into sections denoted by the solid lines, the electron geometry is bolded for that section.

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°	

➢ Water Molecule and VSEPR Theory

Let's consider a water molecule or H2O! Try drawing a water molecule with bonds represented by black line starting from the oxygen and connecting to hydrogens. A hint is that there are two lone pairs or two non-bonding electron pairs.



These are two ways someone might try to draw water molecule, but the correct way is with the bent geometry on the right.