Module -1

ATOMIC STRUCTURE-II & CHEMICAL BONDING

Wavelength (λ) – It is the distance between two neighboring crest or two neighboring troughs of a wave (unit meter)

Frequency (γ) - It is the number of waves that pass a given point in one second (Unit Hertz, Hz)

Velocity (C) - It is the distance travelled by a wave in one second (Unit m/s)

Wave number ($\bar{\nu}$) - It is the number of wave lengths per unit length (Unit m^{-1})

Amplitude (a) - It is the height of crest or depth of trough of a wave.

Electromagnetic Spectrum- It is the arrangement of different regions of electromagnetic radiation in the increasing order of frequency.

Continuous spectra and line spectra – A spectra which contains all the frequency between some range is called continuous spectra. Eg Rainbow (which contains all the frequencies of visible light).

A spectra with discrete lines with some gap in between them is called line spectra. Eg. Emission spectrum of elements (which contains only some lines)

Bohr Model of Atoms – The main postulates of Bohr model are

1. Atom consists of a small positively charged central nucleus and electrons move around the nucleus
2. The circular path of electrons are called orbits. These orbits are designated as K, L, M, N etc with energies E_1, E_2, E_3, E_4 etc
3. As long as the electron remains in a particular shell it neither absorbs nor emit energy. It absorb or emit energy only when it jump from one orbital to another. (Since the energy of each orbit is fixed, Bohr's orbit is called **stationary orbits**)
4. The frequency of radiation absorbed or emitted when the electrons shift from one orbital to another is given by $\nu = \frac{\Delta E}{h}$
5. The angular momentum of electron in a particular orbit is a whole number multiple of $\frac{h}{2\pi}$

(or, Angular momentum, $mvr = \frac{nh}{2\pi}$, Where **n** is a whole number)

Advantages of Bohr's model

1. Bohr model could explain the stability of atom
2. Bohr model could explain the line spectrum of hydrogen atom

Limitation of Bohr's model

1. Bohr's model failed to explain the spectra of multi electron atom
2. It failed to explain the fine structure of line spectrum of hydrogen atom
3. It could not explain the Zeeman effect (Zeeman effect is the splitting of spectral lines in a magnetic field)
4. It could not explain the Stark effect (Stark effect is the splitting of spectral lines in an electric field)
5. It could not explain the three dimensional model of atom. Bohr proposed a flat model of atom
6. It could not explain the chemical bond formation and shape of molecules.
7. Bohr theory is against Heisenberg's uncertainty principle.

Matter wave – The wave associated with a particle is called matter wave or de Broglie wave. The wavelength λ of matter wave is given by $\lambda = \frac{h}{mv}$

Heisenberg's Uncertainty Principle – It is impossible to measure simultaneously the position and momentum of a small particle with absolute accuracy. Mathematically it can be written as $\Delta x \times \Delta p \geq \frac{h}{4\pi}$, ($The\ uncertainty\ in\ momentum, \Delta p = m\Delta v$)

Difference between Orbit and Orbital

Orbit	**Orbital**
Orbit is a well defined circular path around the nucleus	Orbital is the three dimensional space around the nucleus where the probability of finding the electron is maximum
It represents planar motion of electron	Orbital represents three dimensional motion
All orbits are circular shape	Different orbitals have different shape
Orbits do not have directional characteristics	Orbitals have directional characteristics (except 'S' orbital)
The concept of orbit is not in accordance with Heisenberg's uncertainty principle.	The concept of orbital is in accordance with Heisenberg's uncertainty principle.

Quantum Numbers – Quantum numbers are a set of four numbers with the help of which we can get information about an electron in an atom. The four quantum numbers are,

1. Principal quantum number (n)
2. Azimuthal quantum number(l) or angular momentum quantum number
3. Magnetic quantum number (m)
4. Spin quantum number (s)

1. Principal Quantum Number(n) – It represents the main energy level (main shell) in which the electron belongs to. It tells about the size of the orbital, energy of the orbital, and the maximum number of electrons present in the shell ($2n^2$). The possible values of principal quantum number n are 1,2,3,etc. The main shells are also represented as K,L,M,N etc.

2. Azimuthal Quantum Number(*l*) – It indicates the sub shell in which the electron presents. It tells about the shape of the orbital, relative energy of the sub shell, and the angular momentum of electrons. The possible values of Azimuthal Quantum Number(*l*) are 0 to n-1. *l* = 0 is '**S**' orbital, *l* = 1 is '**P**' orbital, *l* = 2 is '**d**' orbital, and *l* = 3 is '**f**' orbital

3. Magnetic Quantum Number(*m*) – It explains the splitting of spectral lines in a magnetic field. It tells about the orientation of each of the sub shell, and the number of orbitals present in a sub shell. The possible values of magnetic quantum number(*m*) are ***-l*** to ***+l*** including zero (total of 2*l +1* values are possible). Each value of m represents an orbital.

4. Spin Quantum Number(*s*) – It explains the the magnetic properties of a substance. It tells about the spin of the electron. The possible values of spin quantum are ***+ 1/2*** *and* ***- 1/2*** which represents clockwise and anticlockwise spinning of electron.

Aufbau Principle – In the ground state of the atoms orbitals are filled with electrons in the order of their increasing energies. The order of increasing energy of orbitals is 1s < 2s < 2p < 3s < 3P < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s

Pauli's exclusion Principle – No two electron in an atom can have the same value for all the four quantum numbers. Pauli's exclusion Principle restricts the maximum number of electrons in an orbital to two.

Hunds Rule of maximum multiplicity – Pairing of electrons in the orbitals of same sub shell does not take place until each orbital of the same sub shell are singly occupied with parallel spin.

Electronic configuration of ,

Nitrogen (N - 7) – $1s^2\ 2s^2\ 2p^3$

Magnesium (Mg - 12) – $1s^2\ 2s^2\ 2p^6\ 3s^2$

Potassium (K - 19) – $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

Argon (Ar - 18) – $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

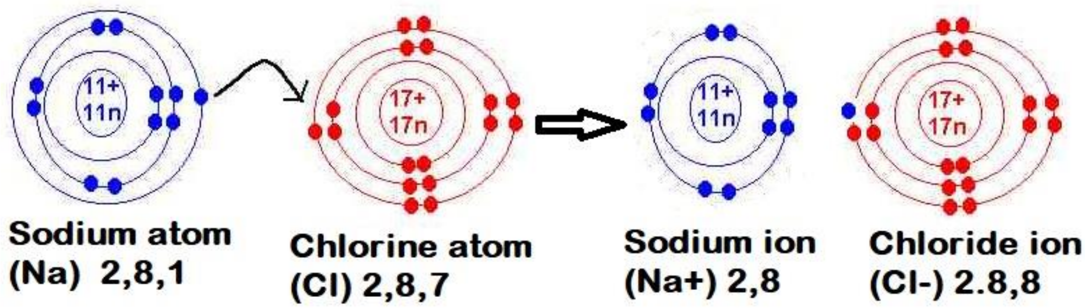
Chemical Bonding

The attractive force that holds the constituent atoms or ions together in a molecule is called **chemical bond**.

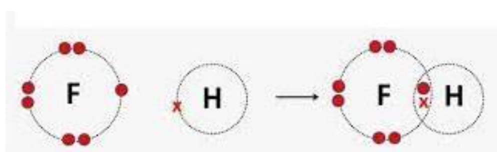
Octet Rule - Octet rule says that, atoms of different element combine with each other to attain eight electronic configuration (Octet) in their outermost shell.

Electronegativity – It is the ability of an atom in a molecule to attract the shared pair of electrons towards itself. Fluorine is the most electronegative element . Metals are generally electropositive and nonmetals are generally electronegative in nature.

Ionic Bond – It is the bond formed by transfer of one or more valence electrons from one atom to another. The atom which lose the electrons becomes positive in charge(cations) and the atom which gain electrons become negative in charge(anion). The attractive force between the oppositely charged ions gives the ionic bond. Ionic bond is formed between metals and nonmetals. Eg. Formation of Sodium Chloride



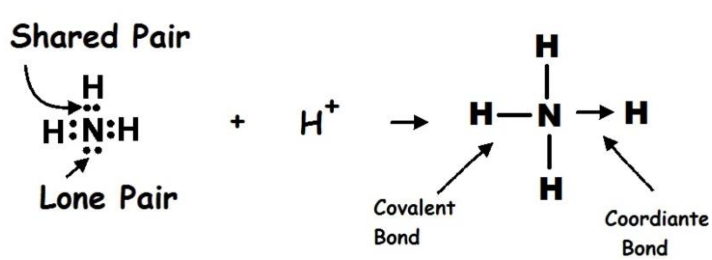
Covalent Bond – chemical bond formed by the sharing of one or more pairs of valence electrons between two atoms is called covalent bond. Eg. Formation of Hydrogen molecule, Hydrogen fluoride molecule



Hydrogen Fluoride

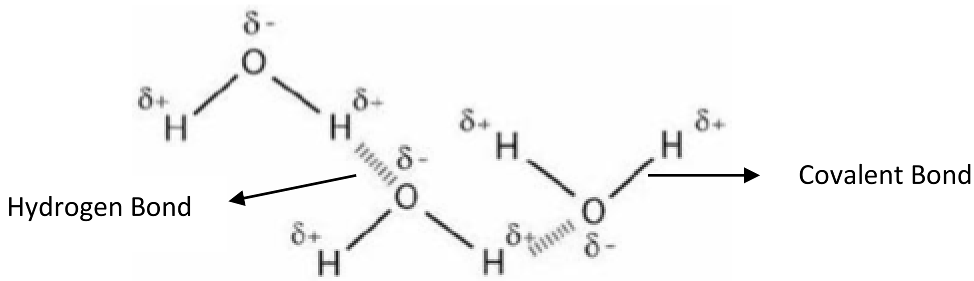
Bond Pair and Lone Pair of electrons – A shared pair of electron between two atoms in a bond is called bond pair. A pair of valence electron not involved in bonding is called Lone pair of electron.

Coordinate Bond or Dative Bond – chemical bond formed by the sharing of a pair of electrons contributed by one of the atoms alone is called Coordinate Bond or Dative Bond. The atom which donates the shared pair of electron is called donor and the atom that accepts electrons for sharing is called acceptor. Eg. Formation of Ammonium ion (NH_4^+), Formation of Ammonium ion (H_3O^+), Formation of SO_2 etc.



Or,	$\mathbf{NH_3}$	**+**	$\mathbf{H^+}$	**→**	$\mathbf{NH_4^+}$

Hydrogen Bond – The attractive force between hydrogen of one molecule with electronegative element of another molecule is called hydrogen bond. Eg. Hydrogen bonding in H_2O and HF



Hydrogen bonding in Water (H_2O)

Anomalous behaviour of water – Water is a liquid at room temperature due to strong intermolecular hydrogen bonding. When water is heated from 0°C to 4°C its volume decreases and hence density increases. After 4°C volume increases and density is decreased. This means **water has a maximum density at 4°C**. When water is heated from 0°C more and more hydrogen bond break and the water molecules come closer and hence the volume decreased. But after 4°C increase in volume with rise in temperature is the dominant factor and the density is decreased