

I PUC STRUCTURE OF ATOM

- * The first atomic theory was proposed by John Dalton on the basis of the law of definite proportions. He regarded the atom as the ultimate particle of matter.
- * Dalton's model could not explain that, why should atoms of an element differ in their masses. This led to the discovery of three fundamental particles i.e. electron, proton and neutron.
- * Except H-atom, the atom of all other elements including their isotopes contains all the three fundamental particles.
- * William Crookes (1879) and Julius Plucker (1889) first studied the electric discharge through partially evacuated tubes. A high voltage produces radiations within the tubes. Later on Thomson (1887) noticed the emission of invisible rays which produce fluorescence on the glass and influenced photographic plate.
- * J. J. Thomson discovered electron (the name electron was given by Stoney) and he calculated charge to mass ratio for electron ($e/m = 1.76 \times 10^{11} \text{C kg}^{-1}$). The e/m ratio is also called the specific charge and was found to be independent of the nature of gas and electrode used.

Electrons

- Discovered by J.J Thomson in cathode rays
- e/m value was determined by J.J Thomson
- e value was determined by Mulliken
- It has a mass of $9.109 \times 10^{-31} \text{ kg}$

This is about $\frac{1}{1836}$ of the mass of a H atom

- The charge is $- 1.602 \times 10^{-19} \text{ Coulombs}$.

This quantity of charge is taken as the unit negative charge

Protons

- Discovered by Goldstein in the anode rays
- It has a mass of $1.672 \times 10^{-27} \text{ kg}$

This is equal to 1.007277 amu

- The charge is $+ 1.602 \times 10^{-19} \text{ coulombs}$.

This quantity of charge is equal to that of an electron but opposite in sign

Neutrons

- Discovered by Chadwick (1932) by bombarding beryllium by α - particles.
- Mass = $1.675 \times 10^{-27} \text{ kg}$

This mass is equal to 1.008655 amu

Hence the mass of a neutron is slightly more than the mass of a proton.

- For all practical purposes, the mass of a proton and a neutron are considered as equal and are taken as 1 amu.
- 1 atomic mass unit = $1/12 \times$ mass of one atom of C - 12 isotope = $1.6606 \times 10^{-27} \text{ kg}$

ATOMIC MODELS

Thomson's Model

Atom is made up of positive charge and the negative charges are embedded in the positive charge. It is also called plum – pudding model. It could not satisfactorily explain the properties of atom.

Following are the important feature of the model:

- mass of the atom is evenly spread over the atom.
- explained stability and electrical neutrality of atom.

Rutherford's Nuclear Atom Model

- In order to explain the observations of scattering of α - particles by gold foil, Rutherford proposed the following atom model.
- An atom has a nucleus which is heavy and consists of protons.

An atom is composed of positively charged particles. Majority of the mass of an atom was concentrated in a very small region. This region of the atom was called as the **nucleus** of an atom. It was found out later that the very small and dense nucleus of an atom is composed of neutrons and protons.

- The region outside the nucleus consists of electrons.
- The volume of the nucleus is minute as compared to the volume of the atom.
- Diameter of the nucleus is of the order of 10^{-15} m
- Diameter of the atom is of the order of 10^{-10} m.
- The number of protons in the nucleus is equal to the number of electrons out side the nucleus. So an atom is electrically neutral.
- Electrons are revolving around the nucleus in closed orbits. The centrifugal force developed balances the force of attraction by the nucleus on the electrons and keeps them in their orbits.

Composition of nucleus

- For a nucleus, $A = Z + N$
 A = mass number Z = atomic number (number of protons in the nucleus)
 N = number of neutrons
- Thus A indicates the total number of nucleons, that is, the total number of protons and neutrons present in the nucleus

Rutherford's Model failed to explain

- The electronic configuration of atoms
- The presence of neutrons
- The formation of spectra of elements
- Stability of an electron in an orbit

- a) **Atomic Number (Z):** The atomic number of an element is the number of protons contained in the nucleus of the atom of that element.

Atomic number = Number of protons = Number of electrons

- b) **Nucleons:** Protons and the neutrons are present in a nucleus. So, these fundamental particles are collectively known as nucleons.

- c) **Mass Number (A):** The total number of protons and the neutrons present in the nucleus is called the mass number of the element.

Mass number = Number of protons + Number of neutrons

IUPAC notation of an atom (nuclide):

Let X be the chemical symbol of the element. It's atomic number be Z and mass number be A. The element can be represented as ${}_Z X^A$.

- d) **Isotopes:** Atoms of the element with same atomic number but different mass number e.g., ${}_1 H^1$, ${}_1 H^2$, ${}_1 H^3$ are three isotopes of hydrogen.

- e) **Isobars:** Atoms having the same mass number but different atomic numbers, e.g: ${}_{15} P^{32}$ and ${}_{16} S^{32}$, ${}_{54} Xe^{130}$ & ${}_{56} Ba^{130}$ are called isobars.

e.g: ${}_{20} Ca^{40}$, ${}_{21} Sc^{40}$ & ${}_{19} K^{40}$

20 protons + 20 neutrons ${}_{20} Ca^{40}$

21 protons + 19 neutrons ${}_{21} Sc^{40}$

19 protons + 21 neutrons ${}_{19} K^{40}$

- f) **Isotones :** Atoms having the same number of neutrons but different number of protons or mass number .

e.g. ${}_6 C^{14}$, ${}_8 O^{16}$, ${}_7 N^{15}$. ${}_{14} Si^{30}$, ${}_{15} P^{31}$, ${}_{16} S^{32}$

${}_{6}^{14}C$ Protons 6 neutrons 8

${}_{8}^{16}O$ Protons 8 neutrons 8

${}_{7}^{15}N$ Protons 7 neutrons 8

- g) **Isoelectronic :** Atoms, molecules or ions having same number of electrons are isoelectronic.

e.g: Na^+ , Ne , Mg^{+2} , Al^{+3}

h) Isosters : having same number of atoms , same no. of electrons and same physical and chemical properties .

e.g. N_2, CO, CN^- . (number of electrons are 28, number of atoms are 2)

i) Isodiaphers : Atoms having same isotopic number(N-Z)

Ex: ${}_{19}K^{39}$ & ${}_{9}F^{19}$, ${}_{13}Al^{27}$ & ${}_{15}P^{31}$

ELECTROMAGNETIC WAVE THEORY

According to this theory, electromagnetic radiations are made up of electric and magnetic fields oscillating perpendicularly to each other and to the direction of propagation. All electromagnetic radiations travel with the speed of light and do not need any medium for propagation.

IMPORTANT CHARACTERISTICS OF WAVE

A wave propagate in the form of alternate crests and troughs.

(i) **Wave length (λ):** It is the distance between two neighbouring crests or troughs.

(ii) **Frequency (ν):** It is the number of waves passing per second. It is related to wavelength as

$$\nu = \frac{c}{\lambda} \quad \text{or} \quad c = \lambda \times \nu$$

(iii) **Velocity (v):** It is the distance travelled by a wave in one second.

(iv) **Amplitude (A):** It is the maximum height of the crest or depth of the trough.

(v) **Wave number ($\bar{\nu}$):** It is the number of wavelengths per cm.

$$\bar{\nu} = \frac{1}{\lambda}$$

ELECTROMAGNETIC SPECTRUM

The arrangement of electromagnetic radiations in order of increasing wavelengths or decreasing frequencies is called electromagnetic spectrum.

ATOMIC SPECTRUM

(i) **Absorption spectrum:** Some substances absorb the energy when white light is passed from their solution or vapours and this result some dark lines in the continuous spectrum of white light. This is called absorption spectrum.

(ii) **Emission spectrum:** The atom loses energy in the form of radiations by the transition of electrons from higher to lower energy level. This energy corresponds to the line of specific wavelength. These lines constitute the emission spectrum.

Types of emission spectra

(a) **Continuous spectra:** The seven colours in the spectrum of white light are very close and constitute continuous spectra.

(b) **Line spectra:** The spectrum of hydrogen atom contains different lines separated by dark bands. This type of spectra is called line spectra.

Every element gives a characteristic line spectrum different from other elements. Hence, it is like the finger print of the element.

Line spectrum of H – atom: Hydrogen spectrum contains different series of lines are given in the diagram.

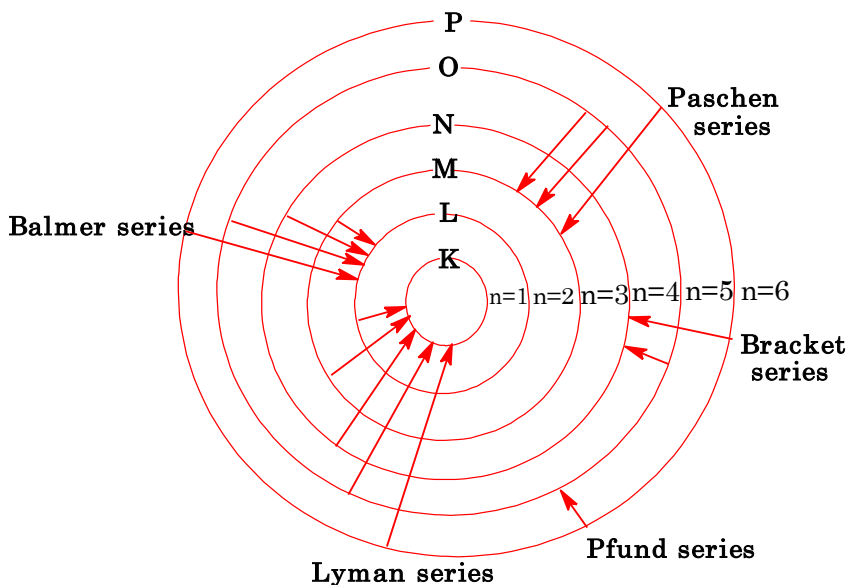


Fig: Different series of H–spectrum

Note:

- (i) The intensity of the spectral lines in a particular series decreases as the value of n of the outer shell increases, e.g. in Lyman series first line ($n_2 = 2, n_1 = 1$) has greater intensity than second line ($n_2 = 3, n_1 = 1$).
- (ii) As the distance from the nucleus increases the energy gap between energy levels decreases.

Hydrogen spectrum

- Hydrogen spectrum recorded in the visible region, consists of 4 prominent lines called Balmer series of lines. The wavelength of each of these lines can be calculated using the Balmer's formula

$$\frac{1}{\lambda} = R \left(\frac{1}{2^2} - \frac{1}{n^2} \right), \text{ where}$$

R = Rydberg's constant = 1.0967800 m^{-1}

n = 3, 4, 5, 6,

On further studies, many more series were also discovered.

The Balmer's formula was modified by Ritz to account for all other series of spectral lines as

$$\frac{1}{\lambda} = R \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$$

n_1 is an integer having a fixed value for each series

n_2 is an integer always greater than n_1

- Lyman Series :** In this series, the spectral lines correspond to the transition of the electron from higher energy state (n) to the lower energy state corresponding to $n_f = 1$.

$$\bar{\nu} = R_H \left[\frac{1}{1^2} - \frac{1}{n^2} \right]; n = 2, 3, 4 \dots \dots \dots \infty$$

These lines are found in the ultraviolet region of the electromagnetic spectrum.

- Balmer Series :** In this series, the spectral lines correspond to the transition of the electron from some higher energy to the lower energy state corresponding to $n_f=2$. Thus,

$$\bar{\nu} = R_H \left[\frac{1}{2^2} - \frac{1}{n^2} \right]; n = 3, 4, 5 \dots \dots \dots \infty$$

These lines are found in the visible region.

- Paschen Series :** Here

$$\bar{\nu} = R_H \left[\frac{1}{3^2} - \frac{1}{n^2} \right]; n = 4, 5, 6 \dots \dots \dots \infty$$

These lines are found in the near infrared region.

- Brackett Series :** Here

$$\bar{\nu} = R_H \left[\frac{1}{4^2} - \frac{1}{n^2} \right]; n = 5, 6, 7 \dots \dots \dots \infty$$

These lines are also found in the near infrared region.

5. **Pfund Series :** Here

$$\bar{\nu} = R_H \left[\frac{1}{5^2} - \frac{1}{n^2} \right]; n = 6, 7, 8 \dots \infty$$

These lines are found in the far infrared region.

The Bohr theory successfully accounts for the spectra for the spectra of hydrogen atom and hydrogen like (Hydrogenic) ions such as He^+ , Li^{2+} , Be^{3+} , B^{4+} , etc.

- The other series of lines arise due to the transition of electrons from various levels as shown in the following table

Series	Transition from level n_2	Transition to Level n_1	Region
Lyman	2, 3, 4, 5, 6	1	UV
Balmer	H_α line 3	2	Visible
	H_β line 4		
	H_γ line 5		
	H_δ line 6		
Paschen	4, 5, 6, 7, 8	3	IR
Brackett	5, 6, 7, 8, 9	4	IR
Pfund	6, 7, 8, 9, 10	5	IR

Number of Spectral Lines

- (i) When the final state is the ground state ($n_1 = 1$) and $n_2 = n$

$$\text{Number of spectral lines} = \frac{n(n-1)}{2}$$

- (ii) When final state is not the ground state ($n_1 \neq 1$)

Number of spectral lines = $\frac{(n_2 - n_1)(n_2 - n_1 + 1)}{2}$ when $n_2 = \infty$, the line produced is called limiting line of the series.

Number of revolutions of electron per second in n^{th} orbit = $\frac{v_n}{2\pi r_n}$

Number of waves in n^{th} orbit = n

$$2\pi r_n = n\lambda$$

When electron jumps from higher to lower level energy equal to difference of energy levels is emitted.

When electron jumps from lower to higher level energy is absorbed.

$$\Delta E = E_2 - E_1$$

where E_1 and E_2 are energies of initial and final states respectively.

Ionization energy:

It is the energy absorbed when electron jumps from ground state to infinity.

$$\text{I.E.} = E_{\infty} - E_1$$

E_{∞} is taken as zero.

Electronic energy in atom is negative because when electron comes close to nucleus from infinity (zero energy) it loses some amount of energy and an e^- is held by attractive force of nucleus

Energy is additive hence, $E_{\text{total}} = E_1 + E_2$

$$\frac{hc}{\lambda} = \frac{hc}{\lambda_1} + \frac{hc}{\lambda_2}$$

$$\frac{1}{\lambda} = \left(\frac{1}{\lambda_1} + \frac{1}{\lambda_2} \right)$$

The velocity of the electron in any orbit is given by $V = \frac{2\pi Ze^2}{nh}$

For a given atom Z is a constant, hence $V \propto \frac{1}{n}$

If the velocity of the electron of the electron is V_1 in the orbit with principal quantum number n_1 and if the velocity is V_2 in another orbit with principal quantum n_2 in the same atom then

$$\frac{V_1}{V_2} = \frac{n_2}{n_1}$$

In the case of the hydrogen atom, the velocity of the electron in any orbit is

$$V = \frac{2.18 \times 10^8}{n} \text{ cm/sec}$$

Velocity of an electron in any orbit in Hydrogen like species

$$V = \frac{2.18 \times 10^8 \times Z}{n} \text{ cm/sec}$$

Energy of an electron:

Let the total energy of the electron be E. It is the sum of kinetic energy and potential energy.

$E = \text{kinetic energy} + \text{potential energy}$

$$= \frac{1}{2} mv^2 - \frac{kZe^2}{r}$$

Putting the value of mv^2 from eq. (i),

$$E = \frac{kZe^2}{2r} - \frac{kZe^2}{r} = - \frac{kZe^2}{2r}$$

Substituting the value of 'r' from $r = \frac{n^2 h^2}{4\pi^2 m k Z e^2}$

$$E = \frac{kZe^2}{2} \times \frac{4\pi^2 m k Z e^2}{n^2 h^2} = - \frac{2\pi^2 Z^2 k^2 m e^4}{n^2 h^2} \dots\dots\dots (ii)$$

For hydrogen atom, $Z = 1$

$$\text{So, } E = - \frac{2\pi^2 k^2 m e^4}{n^2 h^2}$$

Substituting the values of π , k , m , e and h ,

$$E = - \frac{2 \times (3.14)^2 \times (9 \times 10^9)^2 \times (9.1 \times 10^{-31}) \times (1.6 \times 10^{-19})^4}{n^2 \times (6.625 \times 10^{-34})^2}$$

$$= - \frac{21.79 \times 10^{-19}}{n^2} \text{ J per atom}$$

$$E = - \frac{R_H}{n^2} \quad (\text{where, } R_H = 2.18 \times 10^{-18} \text{ J})$$

$$= - \frac{13.6}{n^2} \text{ eV per atom (1J} = 6.2419 \times 10^{18} \text{ eV)}$$

$$= - \frac{313.6}{n^2} \text{ kcal/mol (1 eV} = 23.06 \text{ kcal/mol)} = - \frac{1312}{n^2} \text{ kJ/mol}$$

$$\text{Kinetic energy in } n^{\text{th}} \text{ shell} = \frac{13.6 \times Z^2}{n^2} \text{ eV}$$

$$\text{Potential energy in } n^{\text{th}} \text{ shell} = \frac{-27.2 \times Z^2}{n^2} \text{ eV}$$

$$E_n = \frac{E_1}{n^2} \quad (\text{for hydrogen atom})$$

$$\text{and } E_n = E_1 \times \frac{Z^2}{n^2} \quad (\text{for hydrogen like species})$$

where, E_1 = energy of hydrogen first orbit.

For n^{th} orbit of hydrogen atom ($Z = 1$), the energy will be

$$E_n = - \frac{2.178 \times 10^{-18}}{n^2} \text{ J}$$

The energy of the first of nearest energy state for hydrogen atom is

$$E_1 = - 2.178 \times 10^{-18} \text{ J}$$

The radius of the orbit is given by $r = \frac{n^2 h^2}{4\pi^2 mze^2}$

Where m is mass of the electron

e = charge of the electron

For an atom of a given element Z is a constant. Hence the radius of the orbit in an atom is proportional to the square of the principal quantum number. $r \propto n^2$, radius of the Bohr orbit is 0.529 \AA (0.53 \AA)

Thus, radius of 1st orbit = $0.529 \times 10^{-8} \times 1^2 = 0.529 \times 10^{-8} \text{ cm} = 0.529 \times 10^{-10} \text{ m}$

Radius of 2nd orbit = $0.529 \times 10^{-8} \times 2^2 = 2.11 \times 10^{-8} \text{ cm} = 2.11 \times 10^{-10} \text{ m}$

Radius of 3rd orbit = $0.529 \times 10^{-8} \times 3^2 = 4.76 \times 10^{-8} \text{ cm} = 4.76 \times 10^{-10} \text{ m}$ and so on.

Radius of n^{th} orbit $r_n = r_1 \times n^2$ for hydrogen atom

and $r_n = 0.529 \times \frac{n^2}{Z} \text{ \AA}$ (for hydrogen like species)

if r_1 and r_2 are the radii of two orbits whose principal quantum numbers are n_1 and n_2 , then

$$\frac{r_1}{r_2} = \frac{n_1^2}{n_2^2}$$

Bohr's Theory

Assumptions

- The electrons revolve round the nucleus in fixed closed orbits known as stationary states without emitting radiation.
- Each orbit is at a definite distance from the nucleus.
- The energy of the electron depends on the orbit it occupies.
- Farther the orbit from the nucleus, the greater is the energy associated with it.
- With the absorption of energy by an atom, the electrons in it go to higher energy levels called stationary states
- The atom radiates electromagnetic radiations when its electrons undergo transition from a higher to a lower orbit.
- The frequency of the radiation emitted depends on the energy gap between the two levels

Limitations of Bohr's Theory

- Bohr's theory is applicable only to hydrogen atom and other uni - electron systems like He^+ , Li^{2+} , Be^{3+} etc.
- Bohr's theory is not applicable for multi - electron systems
- Bohr's theory gives a flat model for an atom

PLANCK'S QUANTUM THEORY

According to this a hot body emits radiation energy not continuously but in small packets called quanta. Energy of each quantum is given by $E = h\nu$ and for 'n' quanta $E_{\text{total}} = nh\nu$

Where ν is the frequency of light and h is Planck's constant having value 6.624×10^{-34} Js.

$$\text{Energy of a quanta} = h\nu = \frac{hc}{\lambda} = \frac{hc}{\lambda}$$

1 Einstein = energy of 1 mole quanta

EINSTEIN'S EQUATION

It is $E = mc^2$, where 'E' is the energy of photon, 'm' is the mass and 'c' is the velocity of photon.

Separation Energy: If an electron is already present in the excited state, then the energy required to remove that electron is called separation energy. It is shown as

$$E_{\text{separation}} = E_{\infty} - E_{\text{excited}}$$

- * When solid are heated they emit radiations over a wide range of wavelength. The ideal body, which emits and absorbs radiations of all frequencies is called a black body and the radiation emitted by this body are called black body radiations. The exact frequency distribution of emitted radiation from a black body depends upon its temperature.

PHOTOELECTRIC EFFECT

When light of certain frequency (\geq threshold frequency, ν_0) is incident on a metal surface, electrons are ejected from the surface. If the incident radiation have frequency $\nu > (\nu_0)$, the difference of energy ($h\nu - h\nu_0$) is converted to kinetic energy of photoelectrons.

$\frac{1}{2}mv^2 = h\nu - h\nu_0$ where $h\nu_0$ is the minimum energy required for the emission of photoelectrons and is called work function of metal.

No electron is ejected if the energy of incident light is less than the work function of a given metal. The number of electrons ejected depends upon the intensity and velocity and kinetic energy of photoelectrons depend upon the frequency of incident radiations.

COMPTON EFFECT

When monochromatic X – rays are allowed to fall on some light element, X – ray interact with electrons and the scattered X – rays have less frequency than incident X – rays. This is called Compton's effect.

de Broglie Equation:

$\lambda = \frac{h}{mv}$ where λ is the wavelength of a particle, h is Planck's constant, m is mass and v is velocity, de Broglie equation is applicable to only sub atomic particles.

Heisenberg's uncertainty principle

It is impossible to determine the exact position and momentum of a electron simultaneously with accuracy

$\Delta x \cdot \Delta p \geq \frac{h}{2\pi}$ where Δx is the error in the determination of position, Δp is the error in the determination of momentum.

WAVE OR QUANTUM MECHANICAL MODEL OF ATOM

This model is proposed by Schrodinger (1920) and is based on dual nature of electron and Heisenberg's uncertainty principle. He derived an equation which describes the wave motion of electron in three dimensional space. This is known as Schrodinger wave equation

$$\frac{\partial^2 \psi}{\partial x^2} + \frac{\partial^2 \psi}{\partial y^2} + \frac{\partial^2 \psi}{\partial z^2} + \frac{8\pi^2 m}{h^2} (E - V) \psi = 0 \Rightarrow \nabla^2 \psi + \frac{8\pi^2 m}{h^2} (E - V) \psi = 0$$

where, ∇^2 is $\frac{\partial^2}{\partial x^2} + \frac{\partial^2}{\partial y^2} + \frac{\partial^2}{\partial z^2}$

ψ – Amplitude of electronic wave and is called wave function

E – Total energy of the system

V – Potential energy of electron

Square of the wave function ψ^2 gives the probability of finding electrons within a small three dimensional space.

The acceptable solutions of the above equation for energy E are called *Eigen* values and corresponding wave functions ψ are called *Eigen* functions.

Atomic Orbital

It is the three dimensional space around the nucleus within which the probability of finding electrons is maximum.

Quantum Mechanical Model of an Atom

A set of 4 quantum numbers define the energy of an electron in an atom.

Principal Quantum Number (n)

- This quantum number
 - a) defines the main energy level to which the electron belongs
 - b) determines the size of the atomic orbital.
- Principal quantum number may have any positive integral value except 0
- In the ground state n values range from 1 to 7.

Azimuthal Quantum Number (ℓ)

- It specifies the shape of the atomic orbital.
- Values range from 0 to $(n - 1)$ for a given value of n .
- Each value of ℓ represents a sublevel.
- Sublevels with ℓ values 0, 1, 2, and 3 are designated as s , p , d and f respectively
- The energies of the sublevels of a given major energy level increases in the order s , p , d and f .
- Orbital angular momentum $(\mu_\ell) = \sqrt{\ell(\ell+1)} \frac{h}{2\pi}$

Magnetic Quantum Number (m)

- Magnetic quantum number represents the orientation of an atomic orbital
- The permitted values of m are dependent on the ℓ values. For a given value of ℓ , the values of m are $-\ell$ to $+\ell$ including 0.
- So any sublevel has $(2\ell + 1)$ orientations.
- Each orientation is called an orbital.
- s , p , d and f sublevels have 1, 3, 5 and 7 orbitals respectively
- Orbital is a three dimensional region in space around the nucleus where the probability of locating an electron is maximum.

Spin Quantum Number (s)

- This quantum number describes the spin of the electron about its own axis
- The spin can be either clockwise or anticlockwise.
- It has values $+\frac{1}{2}$ and $-\frac{1}{2}$
- Maximum number of electrons that can be accommodated in an orbital are 2
- Two electrons with opposite spins are called paired up electrons. s , p , d and f sublevels can accommodate a maximum of 2, 6, 10 and 14 electrons respectively.

Spin angular momentum (μ_s) = $\sqrt{s(s+1)} \frac{h}{2\pi}$ where $s = \frac{n}{2}$ and 'n' is number of unpaired electrons.

Permissible Values of Quantum Numbers for Atomic Orbitals

n	l	m	Subshell Notation	Number of orbitals in the subshell
1	0	0	1s	1
2	0	0	2s	1
2	1	-1,0,+1	2p	3
3	0	0	3s	1
3	1	-1, 0, +1	3p	3
3	2	-2,-1,0,-1,+2	3d	5
4	0	0	4s	1
4	1	-1,0,+1	4p	3
4	2	-2,-1,0+1,+2	4d	5
4	3	-3, -2, -1, 0, +1, +2, +3	4f	7

Differences between orbit and orbital

ORBIT	ORBITAL
Circular path around the nucleus where electron moves	3 dimensional path where probability of finding electron is maximum
Planar concept	Three dimensional concept
Circular in shape	Shape depends on orbitals
No directional characteristics	Except 's' orbital, remaining are Directional
Orbit can accommodate a maximum of $2n^2$ Electrons	Orbital can accommodate maximum of 2 Electrons

Spherical nodes (or radial nodes) and nodal planes (or angular node):

The spherical surface where the probability of finding an electron is zero is called a spherical node. Number of spherical nodes in a orbital = $n-l-1$

Where 'n' principal quantum number, 'l' is the azimuthal quantum number.

Example:

For 1s, $n=1$ and $l=0$

No. of spherical nodes = $1-0-1=0$

For 2s, $n=2$ and $l=0$

No. of spherical nodes = $2-0-1=1$

The planes in which the probability of finding an electron is zero is called nodal plane
Number of nodal planes for an orbital = l

Example:

For s- orbital ($l=0$) has no nodal plane

p- orbital ($l=1$) has one nodal plane

Total nodes = radial nodes + angular nodes = $(n - 1)$

For hydrogen wave function, number of nodes can be calculated as:

(i) Number of radial nodes = $(n - l - 1)$

(ii) Number of angular nodes = l

(iii) Total number of nodes = $(n - 1)$

(iv) Number of nodes planes = l

The region of maximum electron density is called **antinode**

Rules involved with Electronic Configuration of Atoms

Pauli's Exclusion Principle

No two electrons in the same atom can have the same values for the four quantum numbers.
This rule limits the maximum occupation of an orbital to 2 electrons with opposite spins

Aufbau Principle

In the ground state of an atom, the electrons enter into the various available orbitals in the order of increasing energies

The order of increasing energy is found to be

$1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s$ and so on.

$(n + l)$ Rule

An orbital with a lower sum of $(n + l)$ is of lesser energy and hence gets filled first.

If two orbitals have the same values of $(n + l)$, then the orbital with the lower value of n is of lower energy and is filled first.

Illustration of $(n + l)$ rule

Type of orbital	Value of n	Value of l	Value of $(n + l)$	Relative energy
1s	1	0	$1+0=1$	Lowest energy
2s	2	0	$2+0=2$	Higher energy than 1s orbital

2p	2	1	2+1=3	2p orbital (n=2) have lower energy than 3s orbital (n=3)
3s	3	0	3+0=3	

Hund's Rule of Maximum Multiplicity:

Electron pairing does not take place in the *p*, *d* and *f* orbitals until each orbital of the given set is singly occupied

This means orbitals are filled in such a way as to give maximum number of unpaired electrons.

Stability of half filled and completely filled orbital

It is found that electronic configuration consisting of half filled or completely filled orbitals is stable.

Hence the electronic configuration of Cr ($Z = 24$) is $[\text{Ar}] 4s^1 3d^5$ and not $[\text{Ar}] 4s^2 3d^4$ because of the extra stability of half filled orbitals

Electronic configuration of Cu ($Z = 29$) is $[\text{Ar}] 4s^1 d^{10}$ and not $[\text{Ar}] 4s^2 3d^9$ because of extra stability of completely filled 3d orbital's

Electronic Configuration of elements in their compounds

To write the electronic configuration of elements in their compounds, the oxidation state of the element is first determined. From this, the number of electrons belonging to that atom is calculated. On the basis of number of electrons present now, the electronic configuration is written

Example:

Electronic configuration of nitrogen in HNO_3 (oxidation state of nitrogen = + 5) is $1s^2$