ATOMIC MODELS-1

Dalton's Atomic Theory

A theory of chemical combination, first stated by John Dalton in 1803. It involves the following postulates:(1) Elements consist of indivisible small particles (atoms).(2) All atoms of the same element are identical; different elements have different types of atom.(3) Atoms can neither be created nor destroyed.(4) 'Compound elements' (i.e. compounds) are formed when atoms of different elements join in simple ratios to form 'compound atoms' (i.e. molecules).Dalton also proposed symbols for atoms of different elements (later replaced by the present notation using letters).

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(3) Atoms can neither be created nor destroyed.

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ATOM

An atom is the smallest unit of ordinary matter that forms a chemical element. Every solid, liquid, gas, and plasma is composed of neutral or ionized atoms. Atoms are extremely small, typically around 100 picometers across. Every atom is composed of a nucleus and one or more electrons bound to the nucleus. The nucleus is made of one or more protons and a number of neutrons.

Electron

The **electron** is a subatomic particle, (denoted by the symbol e^-whose electric charge is negative one elementary charge.[

Charge of an electron (e) = -1.6022×10^{-19} C

The mass of electron (m_e) = $\dfrac{e}{e/m_e}$

$$= \frac{1.6022 \times 10^{-19}\,\text{C}}{1.758820 \times 10^{11}\,\text{C kg}^{-1}}$$

$= 9.1094 \times 10^{-31}$ kg

Proton

The smallest and lightest positive ion was obtained from hydrogen and was called proton. Mass of proton = 1.676×10^{-27} kg

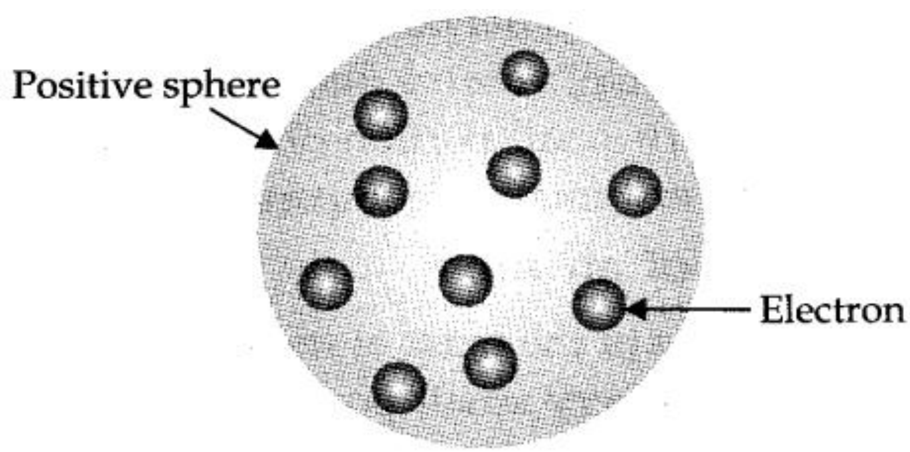
Charge on a proton = (+) 1.602×10^{-19} C

Neutron

It is a neutral particle. It was discovered by Chadwick (1932).

By the bombardment of thin sheets of beryllium with fast moving a-particles he observed • that highly penetrating rays consist of neutral particles which were named neutrons.

Thompson Atomic Model

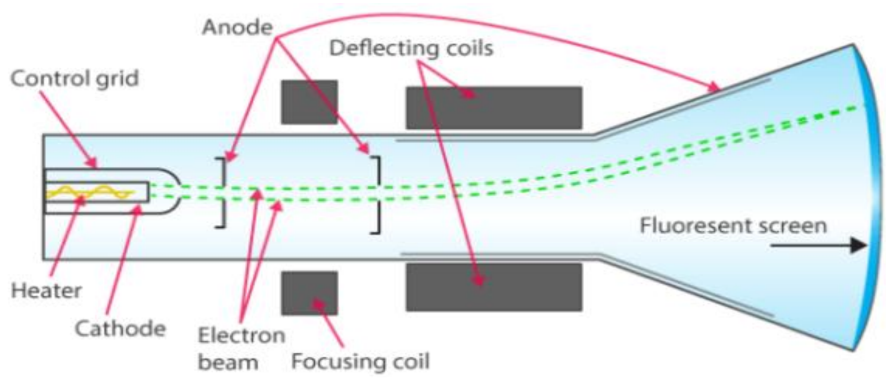


Every atom is uniformly positive charged sphere of radius of the order of 10-10 m, in which entire mass is uniformly distributed and negative charged electrons are embedded randomly. The atom as a whole is neutral.

(i) J. J. Thomson proposed that an atom may be regarded as a sphere of approximate radius 1CT8 cm carrying positive charge due to protons and in which negatively charged electrons are embedded.
(ii) In this model, the atom is visualized as a pudding or cake of positive charge with electrons embedded into it.
(iii) The mass of atom is considered to be evenly spread over the atom according to this model.
Drawback of Thomson Model of Atom
This model was able to explain the overall neutrality of the atom, it could not satisfactorily, explain the results of scattering experiments carried out by Rutherford in 1911.

Goldstein Model

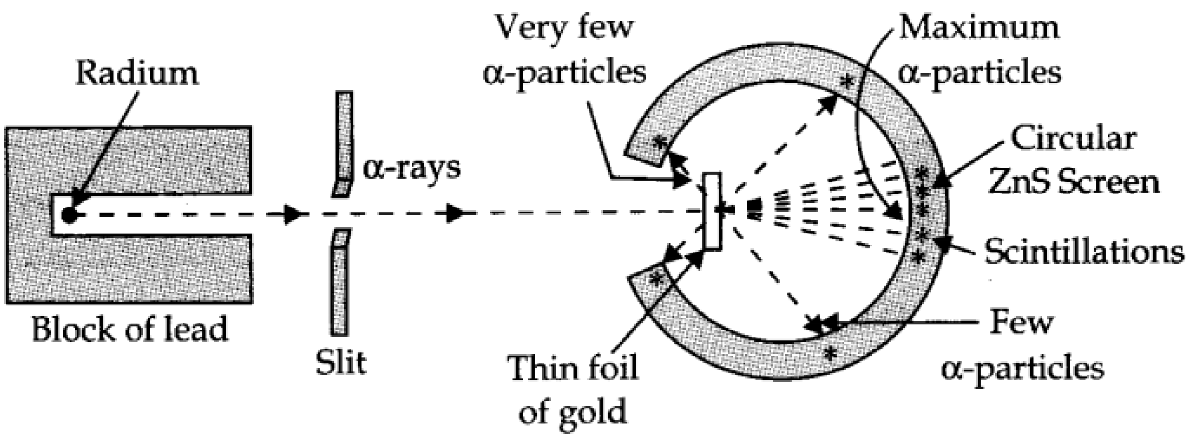
Canal Ray experiment is the experiment performed by German scientist Eugen Goldstein in 1886 that led to the discovery of the proton. The discovery of proton which happened after the discovery of the electron further strengthened the structure of the atom. In the experiment, Goldstein applied high voltage across a discharge tube which had a perforated cathode. A faint luminous ray was seen extending from the holes in the back of the cathode.



When the cathode of a cathode-ray tube was perforated, Goldstein observed rays he called "canal rays," which passed through the holes, or channels, in the cathode to strike the glass walls of the tube at the end near the cathode. Since these canal rays travel in the opposite direction from the cathode rays, they must carry the opposite charge.

Rutherford Atomic Model

Rutherford in 1911, performed some scattering experiments in which he bombarded thin foils of metals like gold, silver, platinum or copper with a beam of fast moving a-particles. The thin gold foil had a circular fluorescent zinc sulphide screen around it. Whenever a-particles struck the screen, a tiny flash of light was produced at that point.



(i) Most of the a-particles passed through the foil without undergoing any deflection,

(ii) A few a-particles underwent deflection through small angles.

(iii) Very few mere deflected back i.e., through an angle of nearly 180°.

From these observations, Rutherford drew the following conclusions:

(i) Since most of the a-particles passed through the foil without undergoing any deflection, there must be sufficient empty space within the atom.

(ii) A small fraction of a-particles was deflected by small angles. The positive charge has to be concentrated in a very small volume that repelled and deflected a few positively charged a-particles. This very small portion of the atom was callednucleus.

(iii) The volume of nucleus is very small as compared to total volume of atom.

• **Rutherford's Nuclear Model of an Atom**

(i) The positive charge and most of the mass of the atom was densely concentrated in an extremely small region. This very small portion of the atom was called nucleus by Rutherford.

(ii) The nucleus is surrounded by electrons that move around the nucleus with a very high speed in circular paths called orbits.

(iii) Electrons and nucleus are held together by electrostatic forces of attraction.

Planck's quantum theory:

According to this theory:

1. Energy emitted or absorbed is not continuous, but is in the form of packets called quanta .In terms of light it is called as photon.
2. Each photon carries an energy which is directly proportional to the frequency of wavelength i.e. E depends upon v (nu).

1. Or E=hv (where v is frequency)
2. Value of h =6.634 x 10^{-34}Jsec
3. Energy associated with no of packets is given by:
   E=nhv (where n is an integral multiple)

This formula can also be written as:

E = (nhc)/ (λ)

(Because we know frequency=speed of light/wavelength)

V = (c/ λ)

EXPLANATION OF BLACK BODY AND PHOTOELECTRIC EFFECT ON THE BASIS OF QUANTUM THEORY.

For the ejection, minimum frequency is required called threshold frequency .Let's say light that falls, has energy equal to hv .Out of this h , a amount of h is used as binding energy and rest is given to electron as kinetic energy.

ATOMIC MODELS-2

Bohr's Atomic Model

Niels Bohr in 1913, proposed a new model of atom on the basis of Planck's Quantum Theory. The main points of this model are as follows:
(i) In an atom, the electrons revolve around the nucleus in certain definite circular paths called orbits.
(ii) Each orbit is associated with definite energy and therefore these are known as energy
levels or energy shells. These are numbered as 1, 2, 3, 4……….. or K, L, M, N………..
(iii) Only those energy orbits are permitted for the electron in which angular momentum of the electron is a whole number multiple of h/2π
Angular momentum of electron (mvr) = nh/2π (n = 1, 2, 3, 4 etc).
m = mass of the electron.
v = tangential velocity of the revolving electron.
r = radius of the orbit.
h = Planck's constant.
n is an integer.
(iv) As long as electron is present in a particular orbit, it neither absorbs nor loses energy and its energy, therefore, remains constant.
(v) When energy is supplied to an electron, it absorbs energy only in fixed amounts as quanta and jumps to higher energy state away from the nucleus known as excited state. The excited state is unstable, the electron may jump back to the lower energy state and in doing so, it emits the same amount of energy. ($\Delta E = E_2 - E_1$).

• **Achievements of Bohr's Theory**

1. Bohr's theory has explained the stability of an atom.
2. Bohr's theory has helped in calculating the energy of electron in hydrogen atom and one electron species. The mathematical expression for the energy in the nth orbit is,

$$E_n = -\frac{2\pi^2 m_e e^4 Z^2}{n^2 R^2}$$

By substituting the values of,

m_e (mass of electron)
e (charge of electron)
R (Rydberg constant)
Z (Atomic number)

The value comes out to be,

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

$$= -\frac{1312 \times Z^2}{n^2} \text{ kJ mol}^{-1}$$

For hydrogen electron, $Z = 1$

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

$$= -\frac{1312}{n^2} \text{ kJ mol}^{-1}$$

3. Bohr's theory has explained the atomic spectrum of hydrogen atom.

• **Limitations of Bohr's Model**

(i) The theory could not explain the atomic spectra of the atoms containing more than one electron or multielectron atoms.

(ii) Bohr7s theory failed to explain the fine structure of the spectral lines.

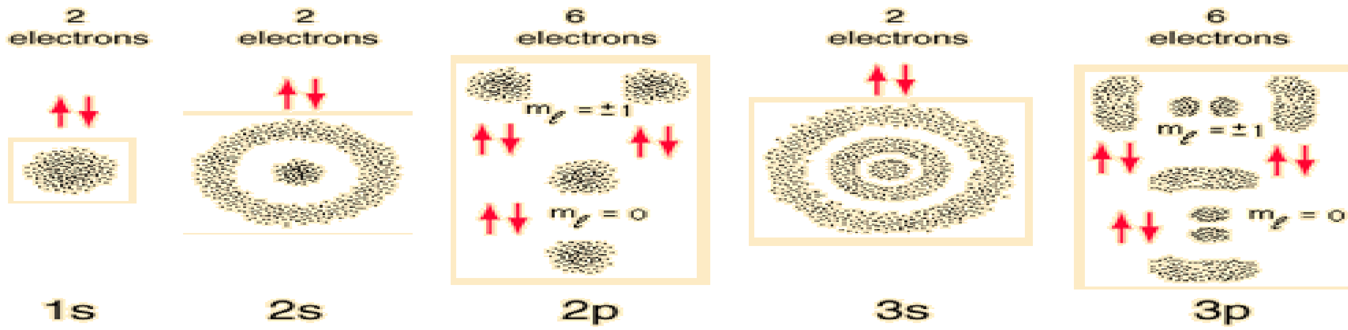
(iii) Bohr's theory could not offer any satisfactory explanation of Zeeman effect and Stark effect.

(iv) Bohr's theory failed to explain the ability of atoms to form molecule formed by chemical bonds.

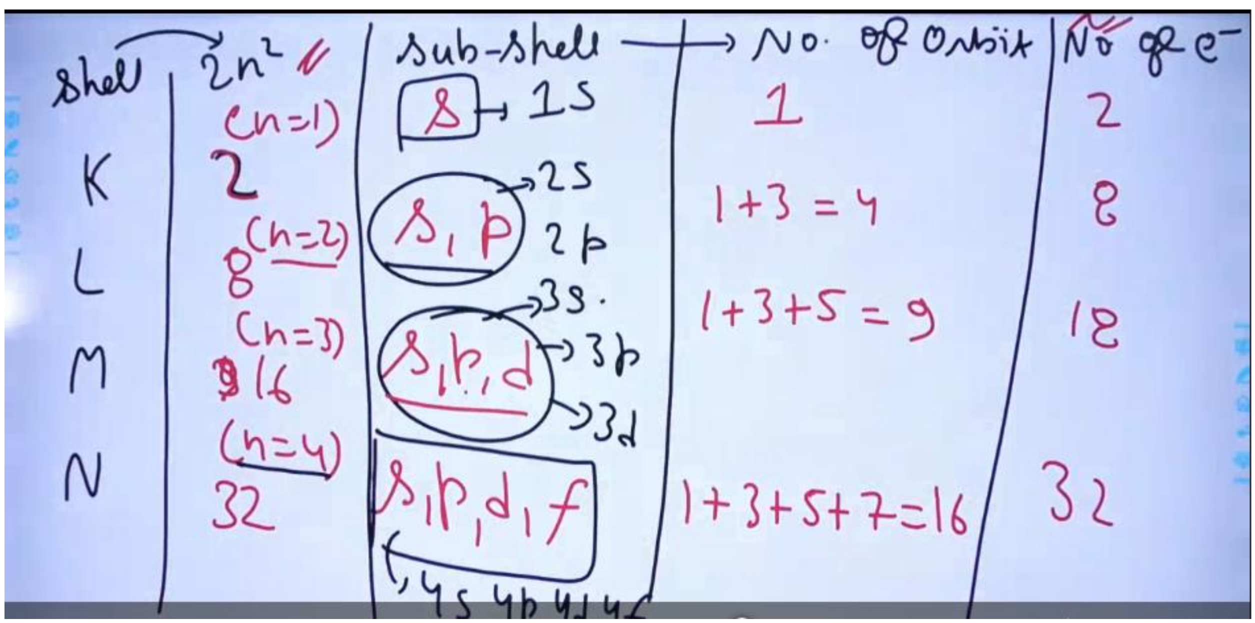
(v) It was not in accordance with the Heisenberg's uncertainty principle.

Orbitals

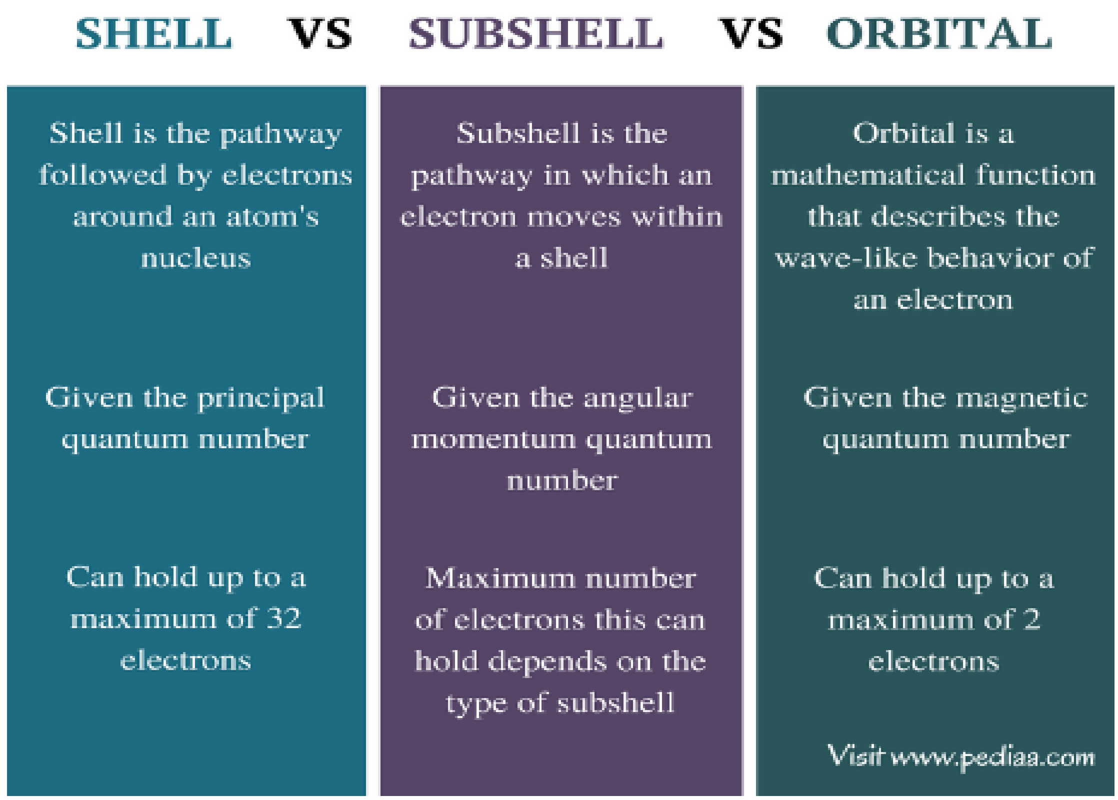
An orbital is a three dimensional description of the most likely location of an electron around an atom. Below is a diagram that shows the probability of finding an electron around the nucleus of a hydrogen atom. Notice that the 1s orbital has the highest probability. This is why the hydrogen atom has an electron configuration of $1s^1$. Orbitals are combined when bonds form between atoms in a molecule.There are four types of orbitals that you should be familiar with s, p, d and f (sharp, principle, diffuse and fundamental). Within each shell of an atom there are some combinations of orbitals. In the n=1 shell you only find s orbitals, in the n=2 shell, you have s and p orbitals, in the n=3 shell, you have s, p and d orbitals and in the n=4 up shells you find all four types of orbitals.



Difference between Bohr's Law and Orbitals



Difference between Shell, Subshell & Orbital



SHELL	SUBSHELL	ORBITAL
Shell is the pathway followed by electrons around an atom's nucleus	Subshell is the pathway in which an electron moves within a shell	Orbital is a mathematical function that describes the wave-like behavior of an electron
Given the principal quantum number	Given the angular momentum quantum number	Given the magnetic quantum number
Can hold up to a maximum of 32 electrons	Maximum number of electrons this can hold depends on the type of subshell	Can hold up to a maximum of 2 electrons

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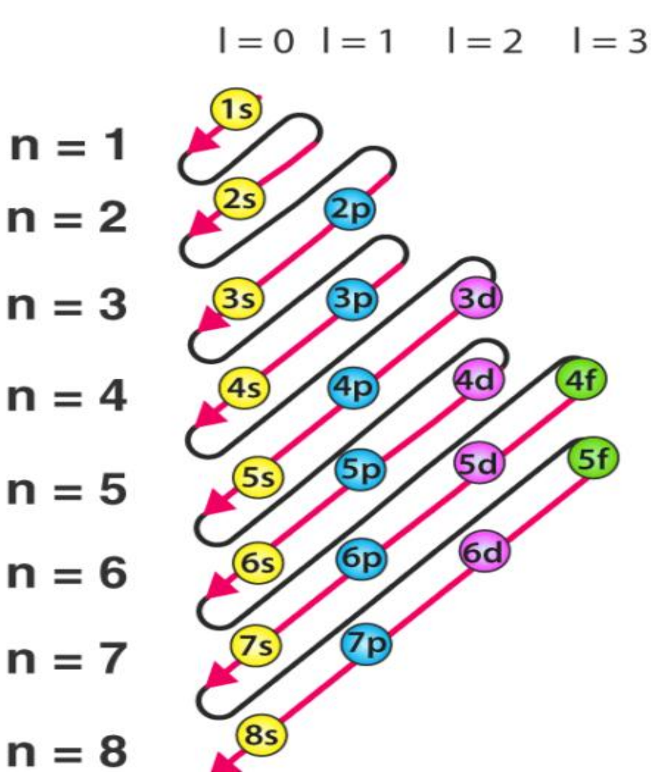
Aufbau's Principle & Electronic Configuration

The Aufbau principle dictates the manner in which electrons are filled in the atomic orbitals of an atom in its ground state. It states that electrons are filled into atomic orbitals in the increasing order of orbital energy level. According to the Aufbau principle, the available atomic orbitals with the lowest energy levels are occupied before those with higher energy levels.



Salient Features of the Aufbau Principle

- According to the Aufbau principle, electrons first occupy those orbitals whose energy is the lowest. This implies that the electrons enter the orbitals having higher energies only when orbitals with lower energies have been completely filled.
- The order in which the energy of orbitals increases can be determined with the help of the (n+l) rule, where the sum of the principal and azimuthal quantum numbers determines the energy level of the orbital.
- Lower (n+l) values correspond to lower orbital energies. If two orbitals share equal (n+l) values, the orbital with the lower n value is said to have lower energy associated with it.
- The order in which the orbitals are filled with electrons is: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p, and so on.

ATOMIC MODELS-3

Heisenberg's Uncertainty Principle

It states that, "It is impossible to determine simultaneously, the exact position and exact momentum (or velocity) of an electron".

Mathematically, it can be given as,

$$\Delta x \times \Delta P_x \geq \frac{h}{4\pi}$$

or $$\Delta x \times \Delta (mv_x) \geq \frac{h}{4\pi}$$

or $$\Delta x \times \Delta v_x \geq \frac{h}{4\pi m}$$

where Δx is the uncertainty in position and ΔP_x (or ΔV_x) is the uncertainty in momentum (or velocity) of the particle and h is Planck's constant.

• **Significance of Uncertainty Principle**

(i) It rules out existence of definite paths or trajectories of electrons and other similar particles.

(ii) The effect of Heisenberg's uncertainty principle is significant only for microscopic objects and is negligible for macroscopic objects.

General Rules

Pauli Exclusion Principle

According to this principle, no two electrons in an atom can have the same set of four quantum numbers.

Pauli exclusion principle can also be stated as: Only two electrons may exist in the same orbital and these electrons must have opposite spins.

• **Hund's Rule of Maximum Multiplicity**

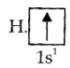
It states that: pairing of electrons in the orbitals belonging to the same subshell (p, d or f) does not take place until each orbital belonging to that subshell has got one electron each i.e., it is singly occupied.

Rules for filling orbitals

The distribution of electrons into orbitals of an atom is called its electronic configuration. The electronic configuration of different atoms can be represented in two ways.
For example:

(i) $s^a\ p^b\ d^c$ ..... notation
(ii) Orbital diagram

example: The electronic configuration of hydrogen $-1s^1$.



Quantum numbers

They are set of 4 numbers, which give complete information about the address of electron.

There are 4 types of quantum numbers:

- Principal quantum number.
- Azimuthal quantum number.
- Magnetic quantum number.
- Spin quantum number.

Principal quantum number:

- It is represented as 'n'.
- It was given by Bohr.
- It represents the orbit where electron is going to be present.

Uses:

1. It gives number of electron in orbit by formula 2 n^2.
2. It gives angular momentum of electron.
3. It gives energy of electron.
4. It gives radius of orbit.

1. **Azimuthal quantum number:**
   - It gives information about sub shell of an atom.
     - It is represented as 'l'.
     - It was introduced by Somerfield.
     - It always has value (n-1).

Example: if n=1 ,l=0

If n=2,l=0,1

If n=3,l=0,1,2

1. **Magnetic quantum number**:
   - It describes the behavior of electron in magnetic field.
   - It is represented as ‘m’.
   - It was given by land.
   - Its value is equal to –l,0, +l

   For example: if n=1, l=0, m=0 that is only one orbital

   If n=2, l=0,1, m=-1,0, +1 that is three orbitals

1. **Spin quantum number:**
   - It gives the info about spinning of electron about its axis i.e. clockwise or anticlockwise
   - It is denoted by‘s’.
   - Its value is either =+1/2, -1/2

ISOTOPES, ISOBARS, ISOELECTRONICAND ISOTONES:

- **ISOTOPES:** Are those elements which have same atomic number ,but different mass number.

  Example : H^1, H^2, H^3 (HYDROGEN)

  O^{16}, O^{17}, O^{18} (OXYGEN)

- **ISOBARS:** Are those elements which have same mass number, but different atomic number.

  Example : ${}_{11}Na^{24}$ ${}_{12}Mg^{24}$

  ${}_{18}Ar^{40}$ ${}_{20}Ca^{40}$

- **ISOTONES:** Are those elements which have same number of neutrons.
- Example:- $C_6N_7O_8$ (all have 8 electrons)

- **ISOELECTRONIC**: the species containing same number of electrons.
- Example:- O^{7-}, F^{-}, Mg^{2+},Al^{3+} (all have 10 electrons)