

## Discovery of electrons

By J.J Thomson Charge to Mass ratio  
of electrons =  $1.758820 \times 10^{11} \text{ C kg}^{-1}$   
Charge on electron =  $1.6022 \times 10^{-19} \text{ C}$   
Mass of Electron =  $9.1094 \times 10^{-31} \text{ Kg}$

## Discovery of Neutron

By James Rutherford  
Charge on Neutron = 0  
Mass of Neutron =  $1.675 \times 10^{-27} \text{ kg}$

## Rutherford's Nuclear Model of Atom

Postulates:

- Positive charge and most of the mass of atoms was densely concentrated in extremely small region i.e. nucleus.
- Nucleus is surrounded by electrons that move around the nucleus with high speed in circular path called orbits.
- Electrons and nucleus are held together by electrostatic forces of attraction.

Drawbacks:

- It cannot explain the stability of an atom
- It does not say anything about the electronic structure of atoms

## Isotopes

Atoms of different elements having same atomic number but different mass number. (Isotopes of Hydrogen)  
Protium  ${}^1\text{H}$ , Deuterium  ${}^2\text{D}$ , and Tritium  ${}^3\text{D}$

## Isobars

Atoms of different elements with different atomic number but same mass number. (Ca and Ar)

## Quantum Mechanical Model of Atom

Postulates:

- Electron in H atoms can move around the nucleus in a circular path of fixed radius and energy called as orbits. these orbits are arranged concentrically around the nucleus.
- Each of these orbits has a definite energy known as energy levels or stationary states.
- When an electron jumps from a lower energy level to higher one, some energy is absorbed.

$$v = \frac{E_2 - E_1}{h} \text{ Bohr's frequency rule.}$$

Angular momentum of electrons:  $m_e v_e r = n \frac{h}{2\pi}$   
 $n = 1, 2, 3, \dots$

Limitations:

- Unable to account for finer details of H atom. Spectrum observed by sophisticated spectroscopic techniques.
- Could not explain the ability of atoms to form molecules by chemical bonds.

## Schrodinger

Fundamental Equation was developed by Schrodinger as  $\hat{A}\psi = E\psi$  where  $\hat{A}$  = Hamiltonian

## Quantum Number

(i) Principal Quantum number (n):  $n = 1, 2, 3, 4, \dots$  Shell = K, L, M, N, ...

(ii) Azimuthal Quantum number: For given value of n,  $l = 0$  to  $n-1$

(iii) Magnetic Quantum number (m): For Subshell with 'l' value  $m = -l$  to  $+l$

(iv) Spin Quantum Number ( $m_s$ ):  $+1/2$  or  $-1/2$

## Electronic Configuration of Atom

(i)  $s^2 p^3 d^1$ ..... notation

(ii) Orbital diagram



## Energy of Orbitals

Lower the value of (n+l) for an orbital, lower is its energy.

## Discovery of Protons

By Ernest Rutherford  
Charge on Proton =  $1.6022 \times 10^{-19} \text{ C}$   
Mass of Proton =  $1.672 \times 10^{-27} \text{ kg}$

## Thomsons Model of Atom

Atoms possess a spherical shape in which the positive charge is uniformly distributed.



## Atomic Number

Number of protons in (Z) +  
Number of neutrons (n)

## Mass Number

Number of protons in nucleus  
of an atom or Number of  
electrons in a neutral atom

## Bohr's Nuclear Model of Atom

- Electrons revolve around the nucleus in certain definite orbits called stationary states having fixed energies.
- Electrons revolve only in those orbits for which the angular momentum is an integral multiple of  $\frac{h}{2\pi}$ .
- Radii of stationary states,  $r_n = \frac{n^2 a_0}{Z}$
- Energies of different stationary states,  $E_n = -\frac{1312}{n^2} Z^2 \text{ kJ mol}^{-1}$
- Velocity of electrons in  $n^{\text{th}}$  orbit,  $v_n = 2.188 \times 10^8 \times \frac{Z}{n} \text{ cm s}^{-1}$
- No. of spectral lines =  $\frac{(n_2 - n_1)(n_2 - n_1 + 1)}{2}$

## Photoelectric effect

It was found by H. Hertz

It is the phenomenon of ejection of electrons from the surface of a metal when light of suitable frequency strikes on it.

## Atomic Spectra

- Spectrum of radiation emitted by a substance Emission Spectra: that has absorbed energy.
- It is like photographic negative of an Absorption Spectra: emission spectra.
- Emission Spectra which do not show a Line / Atomic Spectra: continuous spread of wavelength from red to violet, rather they emit light only at specific wavelength with dark space between them.

$$V = 109677 \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right) \text{ cm}^{-1} \text{ where}$$

$$n_1 = 1, 2, \dots, n_2 = n_1 + 1, n_1 + 2, \dots$$

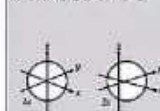
Series	$n_1$	$n_2$	Spectral Region
Lyman	1	2, 3, ...	Ultraviolet
Balmer	2	3, 4, ...	Visible
Paschen	3	4, 5, ...	Infrared
Brackett	4	5, 6, ...	Infrared
Pfund	5	6, 7, ...	Infrared

## Filling of orbitals in atoms

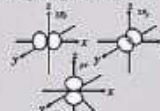
- Aufbau Principle: In the ground state of atoms, the orbitals are filled according to increasing energies.
- Pauli Exclusion Principle: No 2 electrons in an atom can have same set of four quantum numbers.
- Hund's Rule: Pairing of electrons in the orbitals belonging to same subshell does not take place until each electron belonging to the subshell is singly occupied.

## Shape of Orbitals

s Orbitals: Spherically symmetric. Size increases with increase in size in  $1s, 2s, 3s, 4s, 5s, 6s, 7s$



p Orbitals: Each p orbital consists of two sections called lobes on either side of plane passing through the nucleus.



d Orbitals: "Clover leaf" distribution. - two angular nodes.

