Atoms And Nuclei

1. Discovery of nucleus

- In 1897 J.J. Thomson discovered the electron in the rays emitted from the cathode of discharge tube filled with gas at low temperatures.

- Again 1910 Thomson suggested a model for describing atom , known as 'Thomson's atomic model' which suggests that atom consists of positively charged sphere of radius 10^{-8}cm in which electrons were supposed to be embedded.

- Thomson atomic model failed as it could not give convincing explanation for several phenomenon such as, spectrum of atoms, alpha particle scattering and many more.

- In 1909 Gieger and Marsden employed α-particles (Helium ion) as projectile to bombard thin metallic foil.

- According to Thomson atomic model since all positive charge of atom was neutralized by the negatively charged electrons, there would be rare event for an α-particle to suffer a very large deflection , asexpected force of repulsion would not be very strong.

- Surprisingly experiments of Gieger and Marsden showed large deflections of alpha particles that were many orders of magnitude and more common then expected.
- This result of Gieger and Marsden α-particle scattering experiment was explained by Sir Rutherford in 1911.

- Rutherford proposed a new atomic model in which electrons were located at much greater distance from the positive charge.
- Rutherford proposed that all the positive charge , and nearly all the mass of the atom, was concentrated in an extremely small nucleus.

- The electrons were supposed to be distributed around the nucleus in a sphere of atomic radius nearly equal to 10^{-8}cm.
- In explaining this experiment Rutherford made simple assumptions that both the nucleus and α-particles (Helium ion) were point electrical charges and the repulsive force between them is given by Coulombs inverse square law at all distances of separation.

- These assumptions made by Rutherford were not valid if α-particle approaches the nucleus to a distance comparable with the diameter of the nucleus.
- From this experiment there emerged a picture of internal structure of atoms and it also confirmed the existence of the atomic nucleus.

- Approximate values for size and electrical charge of nucleus were calculated using data of various scattering experiments.

2. Nuclear Composition

- Atomic nuclei are build up of protons and neutrons.
- Nucleus of hydrogen atom contains only single proton.
- Charge on a proton is $+1.6\times10^{-19}$ C and its mass is 1836 times greater then that of electron.

- Neutrons are uncharged particles and mass of a neutron is slightly greater then that of a proton.

- Neutrons and protons are jointly called nucleons.
- Number of protons in nuclei of an element is equal to the number of electrons in neutral atom of that element.

- All nuclei of a given element does not have equal number of neutrons for example99.9 percent of hydrogen nuclei contains only one proton , some contain one proton and one neutron and a very little fraction contains one proton and two neutrons.

- Elements that have same number of protons but differ in number of neutrons in their nucleus are called ISOTOPES.

- Hydrogen isotope deuterium is stable but tritium is radioactive and it decays to changes into an isotope of helium.

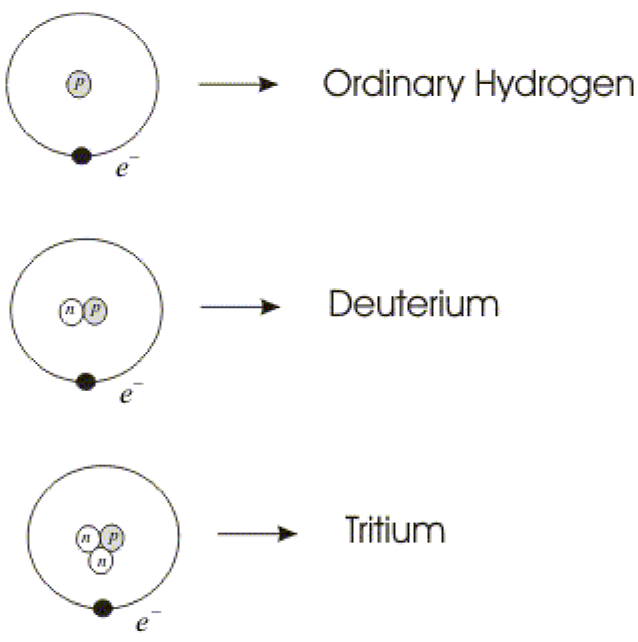

Figure :- Isotopes of Hydrogen

- In heavy water instead of ordinary hydrogen deuterium combines with oxygen.

- Symbol for nuclear species follows the pattern ${}^{A}X_{Z}$ where
  X= Chemical symbol of element
  Z= Atomic number of element or number of protons in the nucleus of that element.
  A= Mass number of nuclide or number of nucleons in the nucleus. A=Z+N where N is the number of neutrons in the nucleus.

- In symbolic form
  (1) hydrogen = ${}^{1}H_{1}$ and Deuterium = ${}^{2}H_{1}$
  (2) Chlorine isotopes are ${}^{35}Cl_{17}$ and ${}^{37}Cl_{17}$

3. Atomic mass

- Atomic masses refer to the masses of neutral atoms , not of bare nuclei i.e., an atomic mass always includes the masses of all its electrons.

- Atomic masses are expressed in mass units (u).
- One atomic mass unit is defined as one twelfth part of the mass of ${}^{12}C_{6}$ atom.
- So the mass of ${}^{12}C_{6}$, the most abundant isotope of carbon is 12u.

- Value of a mass unit is
  $1u=1.66054\times10^{-27}Kg$
- We now calculate the energy equivalent of mass unit. We know that Einsteins Mass-Energy relation is
  $\Delta E=\Delta mc^2$
  here,
  $\Delta m = 1.60\times10^{-27}$ Kg and
  $c = 3\times10^{8}$ m/s
  therefore
  $\Delta E = (1.60\times10^{-27}) \times (3\times10^{8})^2$
  $=1.49\times10^{-10}$ J
  but $1eV = 1.6 \times 10^{-19}$ J
  therefore,

4. Isobars and Isotones

- Nuclei with same A but different Z are known as Isobars for example ${}^{40}K_{19}$ and ${}^{40}Ca_{20}$ share same mass number 40 but differs in one unit of Z.
- Although isobaric atoms share same mass number but they differ slightly in their masses.

- This very slight difference in masses of isobaric atoms is related to difference between energies of two atoms since small mass difference corresponds to considerable amount of difference in energies.

- Nuclei with same number of neutrons but different number of protons are called Isotones for example ${}^{198}Hg_{80}$ and ${}^{198}Au_{79}$