Structure of the Atom

Atoms and molecules are the fundamental building blocks of matter.

An atom is divisible and consists of charged particles.

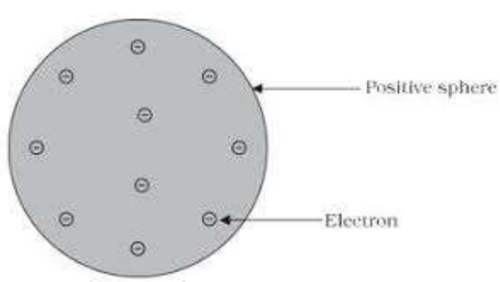
It was known by 1900 that the atom was not a simple, indivisible particle but contained at least one sub-atomic particle-the electron identified by J.J. Thomson. Even before the electron was identified, E. Goldstein in 1886 discovered the presence of new radiations in a gas discharge and called them canal rays. These rays were positively charged radiations which ultimately led to the discovery of another sub-atomic particle. This sub-atomic particle has a charge, equal in magnitude but opposite in sign to that of the electron. Its mass was approximately 2000 times as that of the electron. It was given the name of proton.

- **The Structure of an Atom**

J.J. Thomson was the first one to propose a model for the structure of an atom.

- **Thomson's model of an atom**

Thomson proposed the model of an atom to be similar to that of a Christmas pudding. The electrons, in a sphere of positive charge, were like currants (dry fruits) in a spherical Christmas pudding. We can also think of a watermelon, the positive charge in the atom is spread all over like the red edible part of the watermelon, while the electrons are studded in the positively charged sphere, like the seeds in the watermelon.



Thomson's model of an atom

Thomson proposed that:

(i) An atom consists of a positively charged sphere and the electrons are embedded in it.

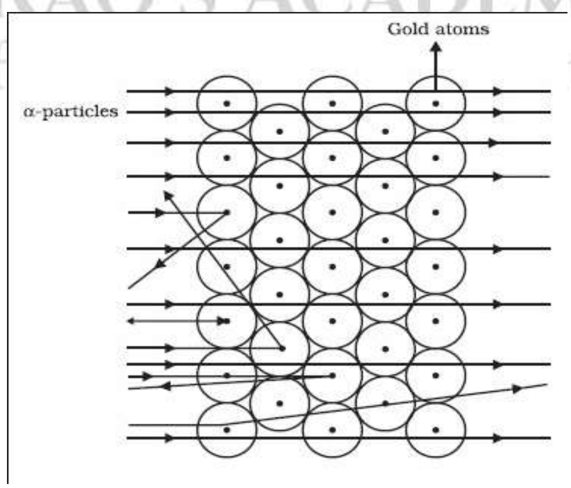
(ii) The negative and positive charges are equal in magnitude. So, the atom as a whole is electrically neutral.

Although Thomson's model explained that atoms are electrically neutral, the results of experiments carried out by other scientists could not be explained by this model, as we will see below.

- **Rutherford's model of an Atom**

Ernest Rutherford was interested in knowing how the electrons are arranged within an atom. Rutherford designed an experiment for this. In this experiment, fast moving alpha (α)-particles were made to fall on a thin gold foil.

- He selected a gold foil because he wanted as thin a layer as possible. This gold foil was about 1000 atoms thick.
- α-particles are doubly-charged helium ions. Since they have a mass of 4 u, the fast-moving α-particles have a considerable amount of energy.
- It was expected that α-particles would be deflected by the sub-atomic particles in the gold atoms. Since the α-particles were much heavier than the protons, he did not expect to see large deflections.



Scattering of α-particles by a gold foil

But, the α-particles scattering experiment gave totally unexpected results. The following observations were made:

(i) Most of the fast moving α-particles passed straight through the gold foil.
(ii) Some of the α-particles were deflected by the foil by small angles.
(iii) Surprisingly one out of every 12000 particles appeared to rebound.

Rutherford concluded from the α-particles scattering experiment that-

(i) Most of the space inside the atom is empty because most of the α-particles passed through the gold foil without getting deflected.
(ii) Very few particles were deflected from their path, indicating that the positive charge of the atom occupies very little space.
(iii) A very small fraction of α-particles were deflected by the 180°, indicating that all the positive charge and mass of the gold atom were concentrated in a very small volume within the atom.

On the basis of the experiment, Rutherford put forward the nuclear model of an atom.

(i) There is a positivity charged centre in an atom called the nucleus. Nearly all the mass of an atom resides in the nucleus.
(ii) The electrons revolve around the nucleus in circular paths.
(iii) The size of the nucleus is very small as compared to the size of the atom.

Drawbacks of Rutherford's model of the atom

The revolution of the electron in a circular orbit is not expected to be stable. Any particle in a circular orbit would undergo acceleration. During acceleration, charged particles would radiate energy. Thus, the revolving electron would lose energy and finally fall into the nucleus. If this were so, the atom should be highly unstable and hence matter would not exist in the form that we know. We know that atoms are quite stable.

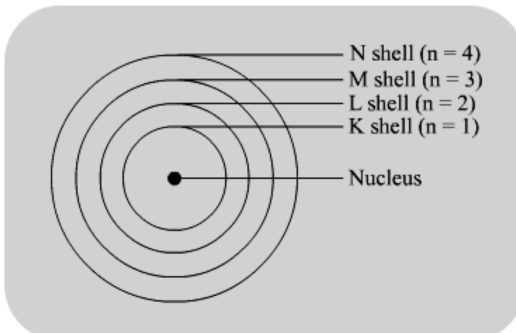
- **Bohr's Model of Atom**

Neils Bohr put forward the following postulates about the model of an atom:

(i) Only certain special orbits known as discrete orbits of electrons, are allowed inside the atom.

(ii) While revolving in discrete orbits the electrons do not radiate energy.

These orbits or shells are called energy levels. Energy levels in an atom are shown in –



A few energy levels in an atom

These orbits or shells are represented by the letters K,L,M,N,… or the numbers, n=1,2,3,4,….

- **Neutrons**

In 1932, J. Chadwick discovered another sub-atomic particle which had no charge and a mass nearly equal to that of a proton. It was eventually named as neutron.

- **Atomic Number and Mass Number**

Atomic Number

It is the number of protons of an atom, which determines its atomic number. For hydrogen, Z = 1, because in hydrogen atom, only one proton is present in the nucleus. Similarly, for carbon, Z = 6.

Mass Number

We can conclude that mass of an atom is practically due to protons and neutrons alone. These are present in the nucleus of an atom. Hence protons and neutrons of an atom called nucleons. For example, mass of carbon is 12 u because it has 6 protons and 6 neutrons, 6 u + 6 u = 12 u. Similarly, the mass of aluminium is 27 u (13 protons + 14 neutrons). The mass number is defined as the sum of the total number of protons and neutrons present in the nucleus of an atom.

Mass Number

Symbol of element

Atomic Number

For example, nitrogen is written as $^{14}N_7$.

Isotopes

A number of atoms of some elements have been identified, which have the same atomic number but different mass numbers. For example, protium ($^{1}H_1$), deuterium ($^{2}H_1$ or D) and tritium ($^{3}H_1$ or T). The atomic number of each one is 1, but the mass number is 1, 2 and 3, respectively. Other such examples are (i) Carbon, $^{12}C_6$ and $^{14}C_6$ (ii) chlorine, $^{35}Cl_{17}$ and $^{37}Cl_{17}$ etc.

Isotopes are defined as the atoms of the same element, having the same atomic number but different mass numbers.

The chemical properties of isotopes are similar but their physical properties are different.

Applications

(i) An isotope of uranium is used as a fuel in nuclear reactors.
(ii) An isotope of cobalt is used in the treatment of cancer.
(iii) An isotope of iodine is used in the treatment of goiter.

Isobars

Calcium, atomic number 20, and argon, atomic number 18. The number of electrons in these atoms is different, but the mass number of both these elements is 40. That is, the total number of nucleons is the same in the atoms of this pair of elements. Atoms of different elements with different atomic numbers, which have the same mass number, are known as isobars.