Chapter 12
Atoms

12.1 Introduction

Thomson Model of Atom- (plum pudding model)

The first model of atom was proposed by J. J. Thomson in 1898.

- According to this model, the positive charge of the atom is uniformly distributed throughout the volume of the atom .
- The negatively charged electrons are embedded in it like seeds in a watermelon.

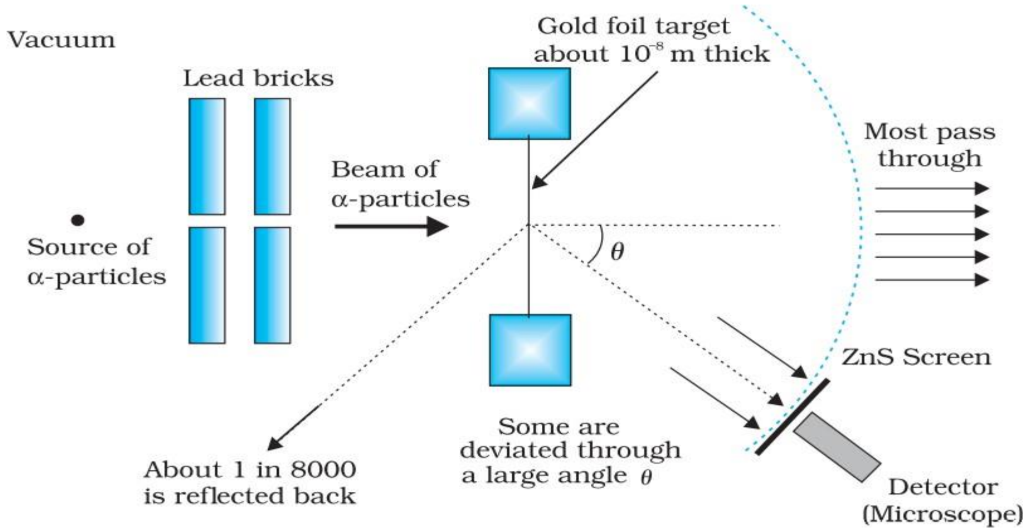
This model is also called plum pudding model of the atom.

Alpha-Particle Scattering and Rutherford's Nuclear Model of Atom

Ernst Rutherford , a former research student of J. J. Thomson, proposed a classic experiment of scattering of these α-particles by atoms to investigate the atomic structure. The explanation of the results led to the birth of Rutherford's planetary model of atom (also called the nuclear model of the atom).

12.2 Alpha-Particle Scattering and Rutherford's Nuclear Model of atom

At the suggestion of Ernst Rutherford, in 1911, H. Geiger and E. Marsden performed scattering experiment.



Alpha-particles emitted by a $^{214}_{83}Bi$ radioactive source were collimated into a narrow beam by passing through lead bricks. The beam was allowed to fall on a thin foil of gold of thickness 2.1×10^{-7} m. The scattered alpha-particles were observed through a rotatable detector consisting of zinc sulphide screen and a microscope.

Observations

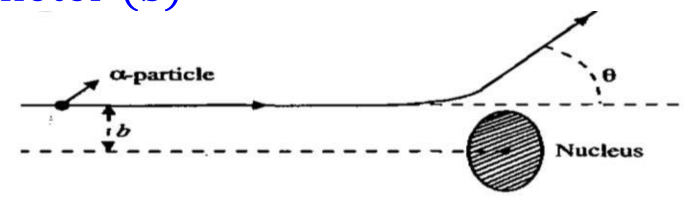
- Many of the α-particles pass through the foil. It means that they do not suffer any collisions.
- Only 0.14% of the incident α-particles scatter by more than 1º.
- About 1 in 8000 of incident α-particles deflect by more than 90º.

Rutherford argued that , greater part of the mass of the atom and its positive charge were concentrated tightly at its centre. When the incoming α-particle make a close encounter with the positive charge ,that would result in a large deflection.

Rutherford's nuclear model of the atom

- Most of an atom is empty space.
- The entire positive charge and most of the mass of the atom are concentrated in the nucleus with the electrons some distance away.
- The electrons would be moving in orbits about the nucleus just as the planets do around the sun.
- The size of the nucleus to be about 10^{-15} m to 10^{-14} m.
- The electrostatic force of attraction, between the revolving electrons and the nucleus provides the centripetal force to keep them in their orbits.

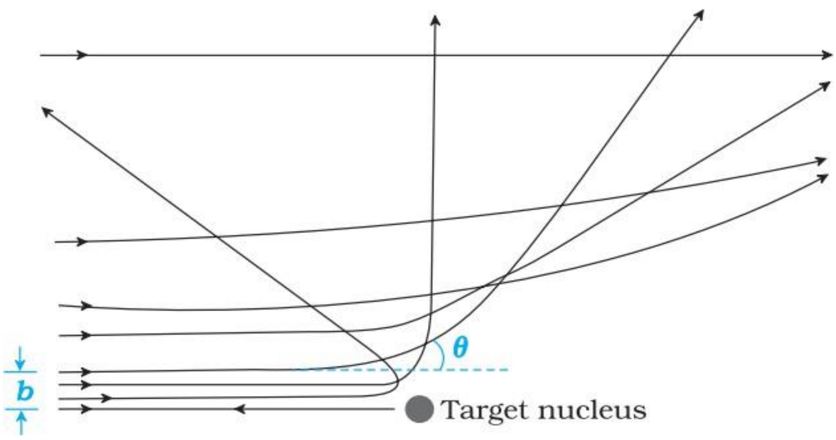
Impact Parameter (b)



Impact parameter is the perpendicular distance of the initial velocity vector of the **α particle** from the centre of the nucleus.

Alpha-particle trajectory

The trajectory traced by an α-particle depends on the impact parameter, b of collision.



- For an α-particle close to the nucleus , impact parameter is small and it suffers large scattering.
  - For head on collision, the impact parameter b=0 and α particle rebounds back ie,angle of scattering $\theta = 180^0$.
  - For large impact parameter, the angle of scattering will be small ($\theta \approx 0^0$) and such α particles go undeviated.

Electron orbits

The electrostatic force of attraction(F_e), between the revolving electrons and the nucleus provides the centripetal force (F_c) to keep them in their orbits.

$$F_c = F_e$$

$$\frac{mv^2}{r} = \frac{1}{4\pi\varepsilon_0}\frac{e^2}{r^2}$$

The kinetic energy (K) of electron

$$K = \frac{1}{2}mv^2 = \frac{e^2}{8\pi\varepsilon_0 r}$$

The potential energy (U) of electron

$$U = \frac{-e^2}{4\pi\varepsilon_0 r}$$

(The negative sign in U signifies that the electrostatic force is in the –r direction.)

Thus the total energy E of the electron in a hydrogen atom is

$$E = K+U$$

$$E = \frac{e^2}{8\pi\varepsilon_0 r} - \frac{e^2}{4\pi\varepsilon_0 r}$$

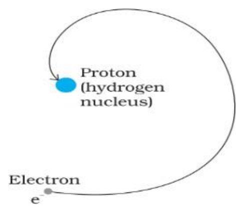
$$E = \frac{-e^2}{8\pi\varepsilon_0 r}$$

The total energy of the electron is negative. This implies the fact that the electron is bound to the nucleus. If E were positive, an electron will not follow a closed orbit around the nucleus.

Limitations of Rutherford Model

Rutherford nuclear model has two main difficulties in explaining the structure of atom:

(a) Rutherford model could not explain stability of matter. The accelerated electrons revolving around the nucleus loses energy and must spiral into the nucleus. This contradicts the stability of matter.



(b) It cannot explain the characteristic line spectra of atoms of different elements.

12.3 Atomic Spectra

Each element has a characteristic spectrum of radiation, which it emits. There are two types of spectra-Emission spectrum and Absorption spectrum.

Emission Spectrum

When an atomic gas or vapour is excited at low pressure, by passing an electric current through it, the emitted radiation has a spectrum which contains certain specific wavelengths only. A spectrum of this kind is termed as emission line spectrum and it consists of bright lines on a dark background. Study of emission line spectra of a material is used for identification of the gas.

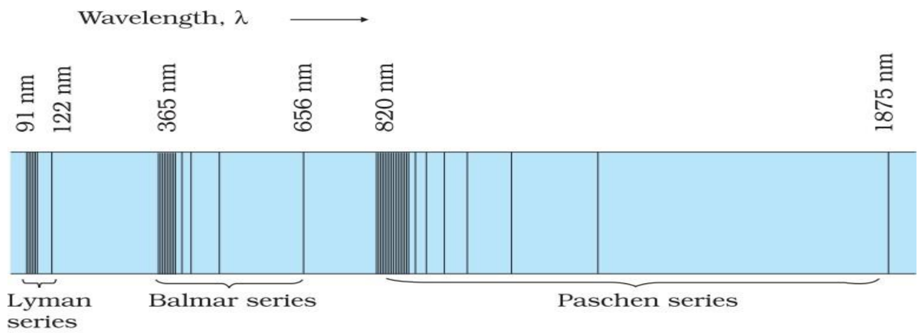


FIGURE 12.5 Emission lines in the spectrum of hydrogen.

Absorption Spectrum

When white light passes through a gas and we analyse the transmitted light using a spectrometer we find some dark lines in the spectrum. These dark lines correspond precisely to those wavelengths which were found in the emission line spectrum of the gas. This is called the absorption spectrum of the material of the gas.

12.4 Bohr Model of Hydrogen Atom

Niels Bohr made certain modifications in Rutherford's model using the ideas of quantum hypothesis. Bohr combined classical and early quantum concepts and gave his theory in the form of three postulates.

1) Bohr's first postulate states that an electron in an atom revolves in certain stable orbits without the emission of radiant energy.

2) Second postulate states that the electron revolves around the nucleus only in those orbits for which the angular momentum is an integral multiple of h/2π where h is the Planck's constant

$$L = mvr = \frac{nh}{2\pi} \quad \text{,where } n = 1,2,3\ldots\ldots\ldots$$

n is called principal quantum number

3) Third postulate states that when an electron make a transition from higher energy level to lower energy level a photon is emitted having energy equal to the energy difference between the initial and final states. The frequency of the emitted photon is then given by

$$h\nu = E_i - E_f$$

Energy of Hydrogen Atom

Total energy of n^{th} energy level

$$E_n = \frac{-e^2}{8\pi\varepsilon_0 r_n}$$

The radius of hydrogen atom

$$r_n = \frac{n^2 h^2 \varepsilon_0}{\pi m e^2}$$

Substituting for r_n from eqn(3)

$$E_n = \frac{-e^2}{8\pi\varepsilon_0 \left(\frac{n^2 h^2 \varepsilon_0}{\pi m e^2}\right)}$$

$$E_n = \frac{-me^4}{8n^2 \varepsilon_0{}^2 h^2}$$

$$E_n = \frac{-13.6}{n^2}\text{ eV}$$

$$E_n \propto \frac{1}{n^2}$$

The negative sign of the total energy of an electron moving in an orbit means that the electron is bound with the nucleus.

Energy levels

The energy of an atom is the least (largest negative value) when its electron is revolving in an orbit closest to the nucleus for $n = 1$. The energy is progressively larger in the outer orbits.

Ground State

The lowest energy state of an atom is called the Ground State, with the electron revolving in the orbit of smallest radius, the Bohr radius, a_0.

For ground state n=1

$$E_1 = \frac{-13.6}{1^2}\text{ eV} = -13.6\text{ eV}$$

At room temperature most of the Hydrogen atoms are in ground state.

Excited States

When Hydrogen atom receives energy by the process such as collisions, the atoms may acquire sufficient energy to raise the electrons to higher energy states. Then atom is said to be in an excited state.

For first excited state (second energy level)

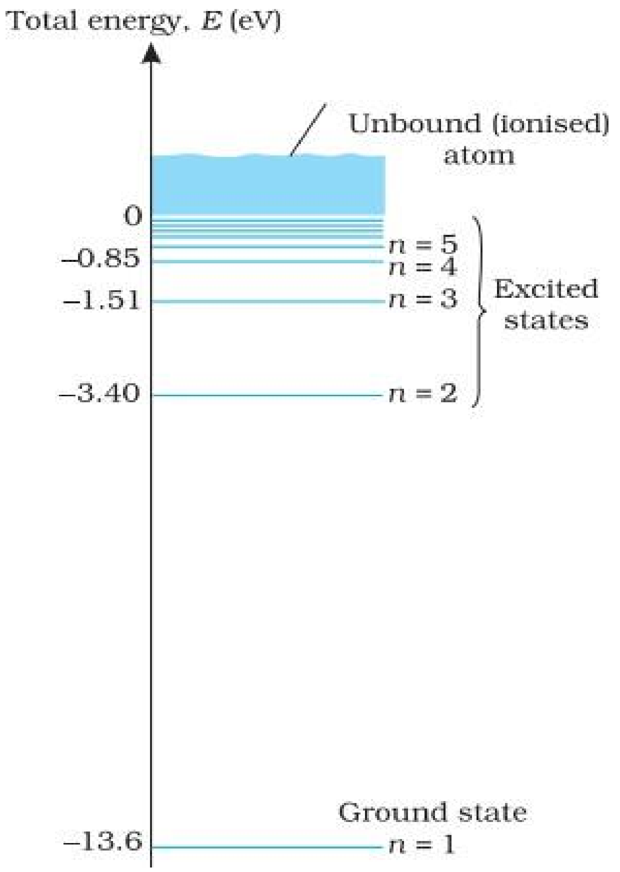
$$n = 2, \qquad E_2 = \frac{-13.6}{2^2}\ \text{eV} = -3.4\ \text{eV}$$

For second excited state (third energy level)

$$n = 3, \qquad E_3 = \frac{-13.6}{3^2}\ \text{eV} = -1.51\ \text{eV}$$

And so on..

The energy level diagram for the hydrogen atom

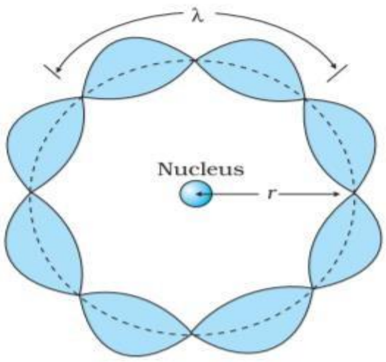


12.5 The Line spectra of Hydrogen Atom

The various lines in the atomic spectra are produced when electrons jump from higher energy state to a lower energy state and photons are emitted. These spectral lines are called emission lines. But when an atom absorbs a photon that has precisely the same energy needed by the electron in a lower energy state to make transitions to a higher energy state, the process is called absorption. Thus if photons with a continuous range of frequencies pass through a rarefied gas and then are analysed with a spectrometer, a series of dark spectral absorption lines appear in the continuous spectrum. The dark lines indicate the frequencies that have been absorbed by the atoms of the gas.

12.6 De Broglie's Explanation of Bohr's second postulate of Quantisation

De Broglie argued that electron in its circular orbit behaves as a particle wave. The particle wave can produce standing wave under resonant condition.



For $\boldsymbol{n}^{th}$orbit of radius r_n , the resonant condition is

$$2\,\pi\, r_n = n\,\lambda \text{----------- (1)} \quad \text{where } n=1,2,3.....$$

But by de Broglie hypothesis , for matter waves

$$\lambda = \frac{\mathbf{h}}{\mathbf{mv}} \text{-------------- (2)}$$

Substituing eqn (2) in eqn (1),

$$2\,\pi\, r_n = n\,\frac{\mathbf{h}}{\mathbf{mv}}$$

$$mv\,\mathbf{r_n} = \frac{\mathbf{nh}}{\mathbf{2\,\pi}} \quad \text{where } n=1,2,3......$$

This Bohr's second postulate of Quantisation.

Limitations of Bohr Atom Model

(i) The Bohr model is applicable to hydrogenic atoms. It cannot be extended two or more electron atoms. Difficulty lies in the fact that each electron interacts not only with the positively charged nucleus but also with all other electrons.

(ii) While the Bohr's model correctly predicts the frequencies of the light emitted by hydrogenic atoms, the model is unable to explain the intensity variations of the frequencies in the spectrum.