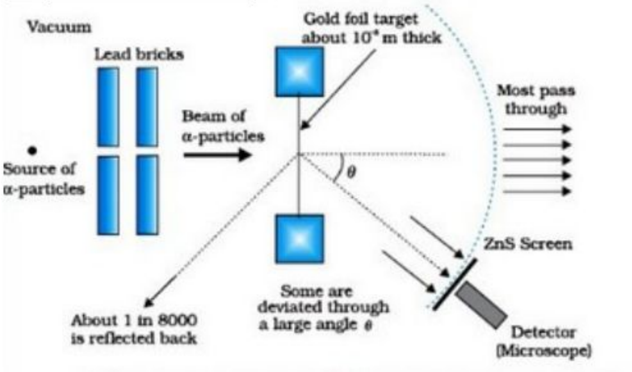
CHAPTER 12
ATOMS

INTRODUCTION

- 1897 – JJ Thomson discovered atoms of different element contains negatively charged particles, the electrons
- 1898 – JJ Thomson proposed first atom model, the plum pudding model
- 1906 – Rutherford proposed Alpha particle scattering experiment

ALPHA PARTICLE EXPERIMENT – RUTHERFORD'S NUCLEAR MODEL OF ATOM

Experimental setup

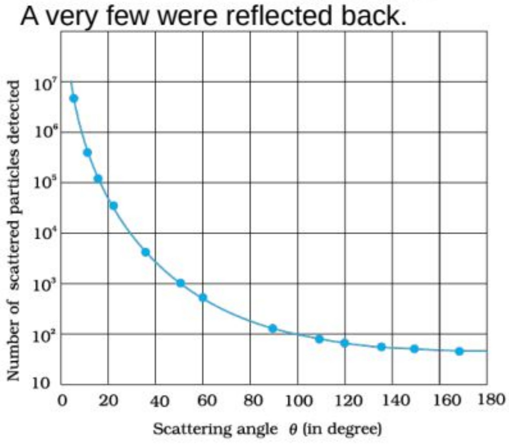


Schematic arrangement of the Geiger-Marsden experiment.

- A narrow beam of α- particles emitted from Bi^{214} (radioactive material) are made to incident on a thin gold foil.
- When the scattered α-particles strike the fluorescent screen (zinc sulphide screen), tiny flashes of light (or scintillations) is produced. This can be observed with the help of a low power microscope

Number of α – particles scattered V/S Scattering Angle

- Most of the α-particles were passed through the gold foil.
- Some were scattered through an angle and some of them in large angles.
- A very few were reflected back.



Conclusions

- Most of the mass and all the positive charges of the atom are concentrated in a very small central core (of diameter about 10^{-14} m) called the nucleus.
- Most of the space inside the atom is empty
- The electrons are revolving around the nucleus. The necessary centripetal force for the revolution of electrons is provided by the electrostatic force of attraction between the electron and the nucleus.

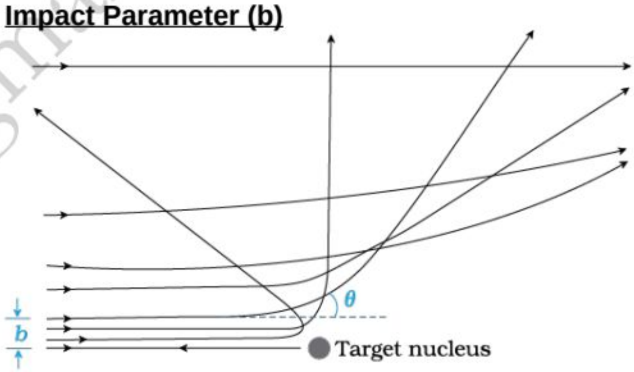
Distance Of Closest Approach (d)

When an alpha particle moves directly towards the nucleus. The velocity and hence the K.E. continues to decrease. And at a particular distance from the nucleus, the α- particle will stop and then start retracing its path. At this distance, the K.E of the α-particle is completely converted into electrostatic potential energy. This distance is called distance of closest approach (d).

ie, At distance of closest Approach

$$KE=\frac{1}{4\pi\epsilon_0}\frac{Ze\,x\,2e}{d}$$

Impact Parameter (b)



The impact parameter is the perpendicular distance of the initial velocity vector of the alpha particle from the centre of the nucleus.

APPLYING RUTHERFORD MODEL TO HYDROGEN ATOM

For a dynamically stable orbit in a hydrogen atom

$$F_e=F_c$$

$$=> \quad \frac{1}{4\pi\epsilon_0}\frac{e^2}{r^2}=\frac{mv^2}{r}$$

Electron Radius

$$r=\frac{e^2}{4\pi\epsilon_0 mv^2}$$

Energy of an Orbiting Electron

The kinetic energy (K) and electrostatic potential energy (U) of the electron in hydrogen atom are

$$K=\frac{1}{2}mv^2=\frac{e^2}{8\pi\epsilon_0 r} \quad \text{and} \quad U=\frac{-e^2}{4\pi\epsilon_0 r}$$

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

$$E = K + U = \frac{e^2}{8\pi\epsilon_0 r} - \frac{e^2}{4\pi\epsilon_0 r}$$

$$E = \frac{-e^2}{8\pi\epsilon_0 r}$$

Limitations Of Rutherford Model

An object which moves in a circle is being constantly accelerated . According to classical em theory, an accelerating charged particle emits radiation in the form of electromagnetic waves. The energy of an accelerating electron should therefore, continuously decrease. The electron would spirally fall into the nucleus. It doesn't happen.

ATOMIC SPECTRA

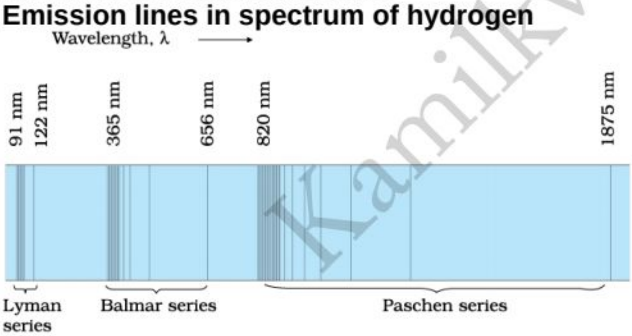
Emission Spectrum

If current is given to hydrogen gas taken in a discharge tube, radiations of certain frequencies were emitted. This is known as emission spectrum. Emission spectrum contains bright lines in a dark background.

Absorption Spectrum

If light is passed though hydrogen gas taken in a glass tube, the atoms will absorb certain frequencies for the excitation to higher states. Thus if we analyse the light coming out we can see dark lines in a bright background. The dark lines correspond to the frequencies absorbed by the hydrogen atoms. This type of spectrum is known as absorption spectrum.

Emission lines in spectrum of hydrogen



BOHR MODEL OF THE HYDROGEN ATOM

Bohr applied the quantum theory of radiation developed by Max Planck and Einstein to the Rutherford's model.

Following are the 3 postulates added by Niels Bohr.

- Electrons are revolving around the nucleus only in certain stable orbits called stationary orbits. Electrons do not radiate while they are in stable orbits.
- Electrons revolve only in those orbits (stationary orbits) in which their angular momentum is an integral multiple of $\frac{h}{2\pi}$
- Electrons might make transitions from one outer orbit to a lower orbit. Then they emit the energy equivalent to the energy gap, in the form of radiations.

$$h\nu = E_2 - E_1$$

Orbital Radius

According to Bohr the radius of nth possible orbit found to be

$$r_n = \frac{n^2 h^2 4\pi\epsilon_0}{4\pi^2 m e^2}$$

Energy of the Electron

The total energy of the orbiting electron according to Bohr model is

$$E_n = \frac{-me^4}{8n^2\epsilon_0^2 h^2}$$

$$E_n = \frac{2.18 x 10^{-18}}{n^2} J \quad \text{(or)} \quad E_n = \frac{-13.6}{n^2} eV$$

Different Energy Levels of Hydrogen Atom (Bohr model)

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

For n=1 (K shell) $E_1 = \frac{-13.6}{1^2} eV = -13.6 eV$

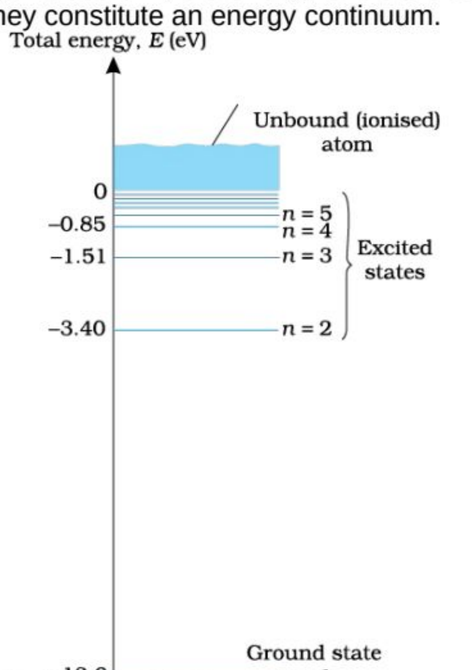
For n=2 (L shell) $E_2 = \frac{-13.6}{2^2} eV = -3.4 eV$

For n=3 (M shell) $E_3 = \frac{-13.6}{3^2} eV = -1.5 eV$

For n=4 (N shell) $E_3 = \frac{-13.6}{4^2} eV = -0.85 eV$

For n=∞ $E_\infty = \frac{-13.6 eV}{\infty} = 0$

For large values of n, the energy levels are so close they constitute an energy continuum.



Excitation energy

Excitation energy is the energy required to excite an electron from its ground state to an excited state.

Thus First excitation energy of hydrogen atom required to excite the electron from n = 1 to n = 2 orbit of hydrogen atom is

$$-3.4\,eV-(-13.6\,eV)=10.2\,eV$$

The second excitation energy of hydrogen atom required to excite the electron from n = 1 to n = 3 orbit of hydrogen atom is

$$-1.51\,eV-(-13.6\,eV)=12.09\,eV$$

Ionization energy

Ionisation is the process of knocking an electron out of the atom. Ionisation energy is the energy required to knock an electron completely out of the atom. (ie from the ground state to n = ∞)

ie, $E_{\infty}-E_1=13.6\,eV$

Drawbacks of Bohr's theory

i) Bohr's theory is applicable to only single electron atoms like hydrogen.

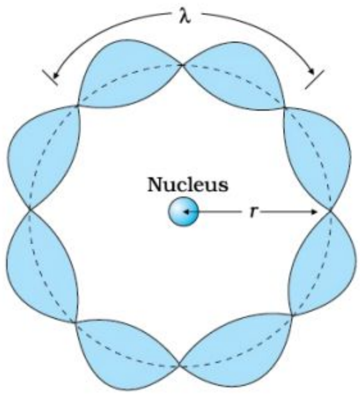
ii) This theory gives no idea about relative intensities of spectral lines.

iii) The fine structure of certain spectral lines of hydrogen, as observed by high resolving power instruments, could not be explained by Bohr's theory.

DE BROGLIE'S EXPLANATION OF BOHR'S SECOND POSTULATE OF QUANTISATION

According to Bohr's 2nd postulate the angular momentum of electron orbiting around the nucleus is quantized. Why should angular momentum can have only those values that are integral multiples of $\frac{h}{2\pi}$?.

This was a puzzle for many years. The French physicist Louis de Broglie explained this puzzle in 1923, ten years after Bohr proposed his model.



De Broglie stated that electron has wave nature with wavelength $\lambda=\frac{h}{mv}$.

For an electron moving in n th circular orbit of radius 'r_n', the wavelength of electron is such that

$2\pi r_n=n\lambda$ where $\lambda=\frac{h}{mv}$

=> $2\pi r_n=\frac{nh}{mv}$ => $mv\,r_n=\frac{nh}{2\pi}$

where, n= 1, 2, 3.......

PREVIOUS QUESTIONS

1.	Energy of electron in the n th orbit of hydrogen atom is $En=\frac{-13.6}{n^2}eV$ What is the energy required to make electron free from first orbit of hydrogen atom ?	1
2.	The minimum energy required to free an electron from the ground state of hydrogen atom is .... (a) + 13.6 J (b) + 13.6 eV (c) – 13.6 J (d) – 13.6 eV	1
3.	Explain Rutherford's alpha particle scattering experiment.	3
4.	The expression for de Broglie wavelength associated with a particle is	1
5.	Write the postulates of Bohr's atom model.	2
5.a)	In alpha particle scattering experiment, what is the impact parameter for scattering angle of 180° ?	1
b)	The ground state energy of hydrogen atom is -13.6 eV. What is the ionisation energy of the electron in this state ?	1
6.	If radius of first electron orbit of hydrogen is r_o , radius of second electron orbit of hydrogen is -------------	1
7.	What is meant by ionisation energy?	2