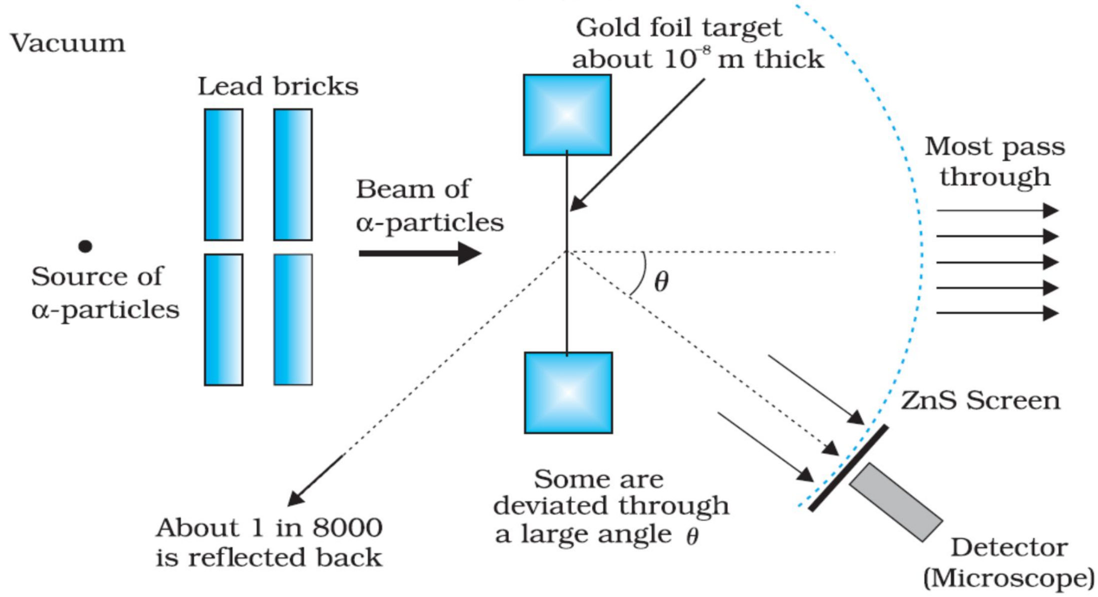
Chapter 12
ATOMS

1. Explain Thomson's model of atom

Ans: According to Thomson's model "***Atom has a spherical shape. It is full of some positive matter and electrons are embedded in this positive matter like plums in pudding. The total positive charge is balanced by the –ve charges of electrons. As a result of this atom is electrically neutral***". This model is sometimes called plum-pudding model.

2. Explain **α – particle scattering experiment** performed by Rutherford and also by H. Geiger and E. Marsden

Ans:



Experimental set up

A narrow beam of α- particles emitted from radon (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.

Variation of no. of α – particles scattered with scattering angle

As the scattering angle increases, the number of α-particles scattered decreases.

Observations

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

Assumptions

i) α-particle and gold nucleus are point charges.
ii) The scattering of α-particles from the gold atoms were due to the coulombic interaction (force) between the positive charge of the α-particle and the positive charge of the gold nucleus.
iii) The nucleus of the gold atom is very massive compared to the α-particle and hence it remains stationary during the experiment (collision)
iv) As the gold foil is very thin, α-particles will not suffer multiple scattering.

Conclusion or Rutherford's model of atom

1. *An atom may be regarded as a sphere of diameter 10^{-10}m.*

2. ***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.***
3. *The nucleus is surrounded by electrons. The electrons are spread over the remaining part of the atom, leaving plenty of empty space in the atom.*
4. *As the atom is electrically neutral, the total positive charge on the nucleus is equal to the total negative charge of the electrons in the atom.*
5. ***Electrons are revolving round the nucleus in circular orbits. The necessary centripetal force for the revolution of electrons is provided by the electrostatic force of attraction between the electron and the nucleus.***

Magnitude of repulsive force between α-particle and gold nucleus

Charge of gold nucleus – **Ze**

Z ⟶ atomic number of **Au** (i.e., 79)

e = 1.6 × 10^{-19}C

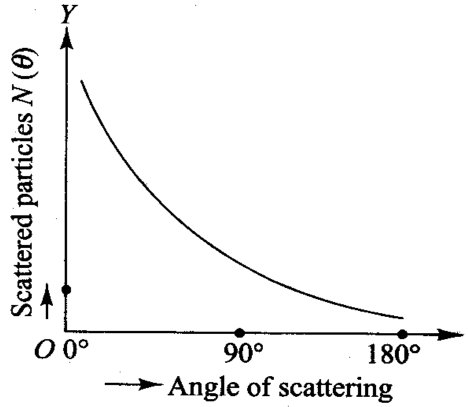
Charge of α-particle = **2e**

Let 'd' be the distance between α-particle and Gold nucleus then,

$$\mathbf{F} = \frac{1}{4\pi\varepsilon_0}\frac{Ze\times 2e}{d^2}$$

3. Draw a graph between the number of alpha particles scattered and scattering angle

Ans:



3. What is meant by distance of closest approach?

Ans: 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 if called distance of closest approach.

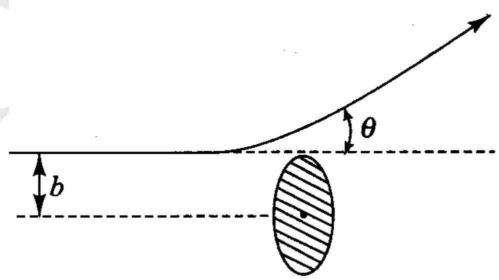
$$\frac{1}{2}mv^2 = \frac{1}{4\pi\varepsilon_0}\frac{Ze\times 2e}{d}$$

Distance of closest approach

$$\mathbf{d} = \frac{1}{4\pi\varepsilon_0}\frac{Ze\times 2e}{\left(\frac{1}{2}mv^2\right)}$$ (Distance of closest approach is of the order of 10^{-14}m.)

4. Define Impact Parameter (**b**)

Ans:



The impact parameter is defined as the perpendicular distance of the velocity of the α-particle from the centre of the nucleus when it is far away from the nucleus.

Impact parameter,

$$b = \frac{Ze^2\cot\left(\theta/2\right)}{4\pi\varepsilon_0\left(1/2\, mv^2\right)}$$

where θ is the scattering angle.

If the impact parameter is zero, the scattering angle is 180^0.

5. Explain the success of Rutherford's model

Ans:

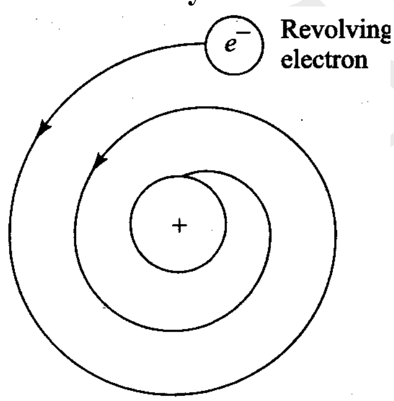
i) Large angle scattering of alpha particles through thin foils could be explained.

ii) The classification of elements in the periodic table on the basis of their atomic number, instead of atomic weight, was justified.

6. What are the **limitations of Rutherford's model** of atom?

Ans:

i) According to the classical electromagnetic theory, a charged particle in circular motion (accelerated motion) should radiate energy in the form of electromagnetic radiations. If the charged particle completes ' ν ' revolutions in one second, it should emit a radiation of frequency ν . As a result of continuous emission of radiations, the energy of electrons should continuously decrease. The electron should follow a spiral path and finally fall into the nucleus. So the atom must not be stable. But the atom is a very stable structure. Thus **Rutherford's model fails to account for the stability of the atom**.



ii) Rutherford's model does not give any particular value for the radius of electronic orbit. This should mean that an electron can emit radiations of all possible frequencies. Thus atom should give a continuous spectrum. But the experiments give a line spectrum. Thus **Rutherford's model does not explain the line spectra of atoms**.

7. Explain **Bohr's Model** of Hydrogen atom

Ans: The drawbacks of Rutherford's model were removed by Niels Bohr in 1913. He applied the quantum theory of radiation, as developed by Max Planck and Einstein to the Rutherford's model.

Following are the 3 postulates added by Niels Bohr.

i) ***Electrons in an atom can revolve in certain stable orbits without radiating energy****. Since the total energy of the electrons remains constant when they revolve in non-radiating orbits, these orbits are called stationary orbits.*

ii) ***Electrons can revolve only in those orbits in which their angular momentum is an integral multiple of*** $\frac{h}{2\pi}$.

$\mathbf{L} = \frac{nh}{2\pi}$

or $\mathbf{mvr} = \frac{nh}{2\pi}$, n = 1, 2, 3, ……

n is the principal quantum number

iii) Electrons might make transitions from one orbit to a lower orbit. Then they emit the energy equivalent to the energy gap, in the form of radiations.

$\mathbf{E_2 - E_1 = h\upsilon}$

8. Using Bohr's Theory derive expressions for (i) radii of orbits (ii) velocity of electrons in orbits (iii) energy of electrons in orbits of Hydrogen atom

Ans:

(i) **<u>Radii of Bohr's Stationary orbits</u>**:-

The centripetal force for the revolution of electrons round the nucleus is provided by the electrostatic force of attraction between the nucleus and the electron.

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

$$\Rightarrow mv^2 = \frac{1}{4\pi\varepsilon_0}\frac{e^2}{r} \quad \text{........ (1)}$$

But $mvr = \frac{nh}{2\pi}$ (Bohr's postulate)

$$\Rightarrow v = \frac{nh}{2\pi mr}$$

.............(2)

From equation (1),

$$m\left(\frac{nh}{2\pi mr}\right)^2 = \frac{1}{4\pi\varepsilon_0}\frac{e^2}{r}$$

$$m\frac{n^2h^2}{4\pi^2m^2r^2} = \frac{1}{4\pi\varepsilon_0}\frac{e^2}{r}$$

$$\Rightarrow \frac{n^2h^2}{\pi mr} = \frac{e^2}{\varepsilon_0}$$

$$\Rightarrow r = \frac{\varepsilon_0 n^2h^2}{\pi me^2}$$

In general,

$$\boxed{r_n = \left(\frac{\varepsilon_0 h^2}{\pi me^2}\right)n^2}$$

$$\mathbf{r_n \propto n^2}$$

∴ The radii of the stationary orbits are in the ratio:

$1^2 : 2^2 : 3^2 : 4^2 :$ ....

Or 1 : 4 : 9 : 16, etc

The stationary orbits are not equally spaced.

For n = 1 we get the smallest radius. This represents the normal state of the hydrogen atom.

$$r_1 = \frac{\varepsilon_0 h^2}{\pi me^2}\times 1^2$$

$$= \frac{8.854\times10^{-12}\times6.626\times10^{-34}}{3.14\times9.1\times10^{-31}\times(1.6\times10^{-19})^2}$$

$= 0.53 \times 10^{-10}$m $= 0.53$A^0.

This is the radius of the lowest orbit and is known as **Bohr radius.**

$$\boxed{\mathbf{r_n = r_1.n^2}}$$

ii) **Velocity of electron**

$$mvr = \frac{nh}{2\pi} \Rightarrow v = \frac{nh}{2\pi mr} \quad \text{..... (1)}$$

Also we have $r = \left(\frac{\varepsilon_0 h^2}{\pi me^2}\right)n^2$

$$\therefore (2) \rightarrow v = \frac{nh}{2\pi m\left(\frac{\varepsilon_0 h^2}{\pi me^2}\right)n^2}$$

$$v = \frac{e^2}{2\varepsilon_0 nh}$$

In general $\boxed{v_n = \frac{e^2}{2\varepsilon_0 nh}}$

iii) **Energy of electron**

An electron revolving around the nucleus possesses both **KE** and **PE**.

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

$$\text{KE, } \mathbf{E_k} = \frac{1}{2}mv^2 = \frac{1}{2}\times\frac{1}{4\pi\varepsilon_0}\frac{e^2}{r}$$

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

$$\boxed{E_k = \frac{e^2}{8\pi\varepsilon_0 r}}$$

$$\text{PE, } \mathbf{E_p} = \frac{-1}{4\pi\varepsilon_0}\frac{e^2}{r}$$

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

∴ Total energy

$E = E_k + E_p$

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

$$= \frac{e^2}{4\pi\varepsilon_0 r}\left[\frac{1}{2}-1\right]$$

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

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

But $r = \frac{\varepsilon_0 h^2 n^2}{\pi me^2}$

$$\therefore E = \frac{-e^2}{8\pi\varepsilon_0 \times \frac{\varepsilon_0 h^2 n^2}{\pi m e^2}}$$

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

Substituting the values,

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

KE = -E

PE = 2E

As n increases, E_n becomes less negative ie, energy increases.

9. Derive **Rydberg Formula**

Ans: According to Bohr's frequency condition,

$$h\nu = E_{n_2} - E_{n_1}$$

$$\Rightarrow h\nu = \frac{-me^4}{8\varepsilon_0^2 h^2 n_2^2} - \frac{-me^4}{8\varepsilon_0^2 h^2 n_1^2}$$

$$= \frac{me^4}{8\varepsilon_0^2 h^2}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

$$\nu = \frac{me^4}{8\varepsilon_0^2 h^3}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

$$\frac{c}{\lambda} = \frac{me^4}{8\varepsilon_0^2 h^3}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

$$\frac{1}{\lambda} = \frac{me^4}{8\varepsilon_0^2 h^3 c}\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]$$

$\frac{1}{\lambda}$ is the wave number.

It is denoted by $\bar{\nu}$.

The term $\frac{me^4}{8\varepsilon_0^2 h^3 c}$ is known as Rydberg constant. It is denoted by R.

$$R = \frac{me^4}{8\varepsilon_0^2 h^3 c}.$$

Its value is $\mathbf{1.097 \times 10^7 m^{-1}}$.

$$\therefore \boxed{\frac{1}{\lambda} = R\left[\frac{1}{n_1^2} - \frac{1}{n_2^2}\right]}$$

This equation is called Rydberg formula.

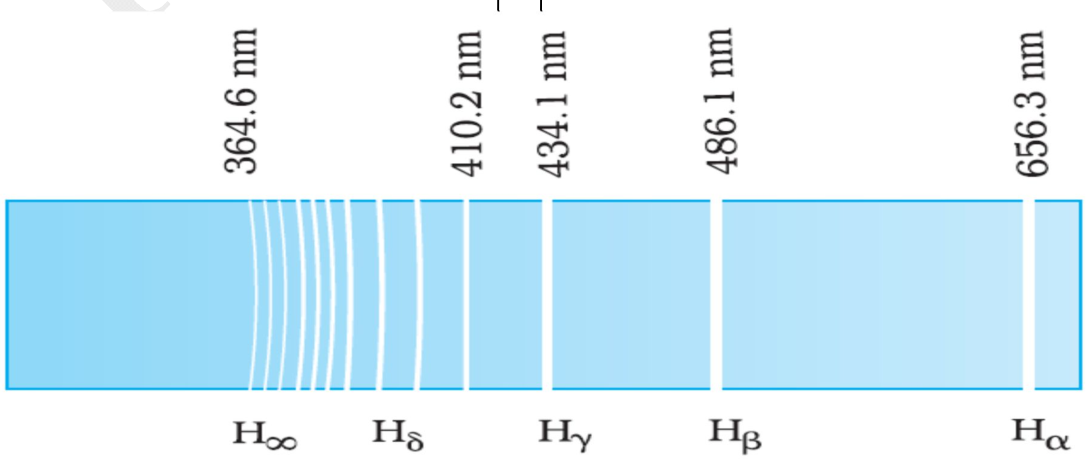
10. What is **Emission Spectrum**?

Ans: 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.

11. What is **Absorption Spectrum?**

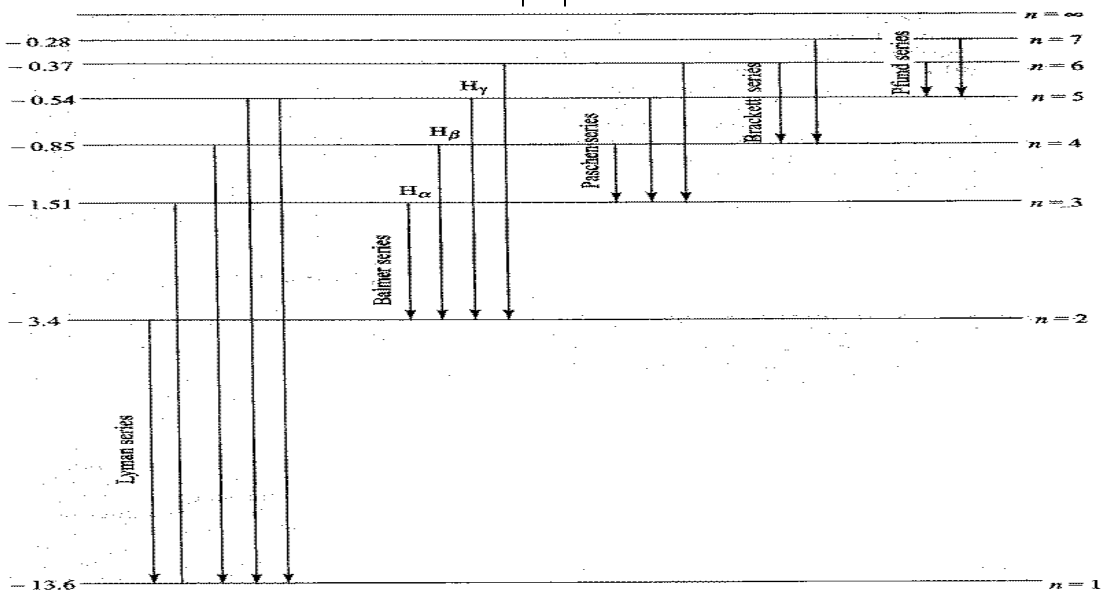
Ans: If light is passed though hydrogen gas taken in a glass tube, the atoms will absorb certain frequencies for the excitation to higher states. Thus if we analyse the light coming out we can see dark lines in a bright background. The dark lines correspond to the frequencies absorbed by the hydrogen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

Emission Lines in Balmer Series



12. Explain Spectral lines of hydrogen atom?

Ans:



i. Lyman series

The spectral line of this series corresponds to the transition of an electron from some higher energy state to the innermost orbit (n = 1).

For Lyman series, $n_1 = 1$

and $n_2 = 2, 3, 4, \ldots$

$$\bar{\upsilon} = R\left[\frac{1}{1^2}-\frac{1}{n_2^2}\right]$$

Lyman series lies in the **ultra violet** region.

ii) Balmer series

The spectral lines of this series correspond to the transition of an electron from some higher state to an orbit having n = 2.

For Balmer series $n_1 = 2$,

$n_2 = 3, 4, 5, \ldots$

$$\bar{\upsilon} = R\left[\frac{1}{2^2}-\frac{1}{n_2^2}\right]$$

The first four lines of the Balmer series lie in the **visible region** of the spectrum and the rest of the series goes to the **u.v. region**.

iii) Paschen series

For Paschen series

$n_1 = 3$, $n_2 = 4, 5, 6, \ldots$

$$\bar{\upsilon} = R\left[\frac{1}{3^2}-\frac{1}{n_2^2}\right]$$

Paschen series lies in the **infrared region.**

iv) Bracket series

For Bracket series, $n_1 = 4$, $n_2 = 5$,

v) 6, 7, …………… $\bar{\upsilon} = R$

$$\left[\frac{1}{4^2}-\frac{1}{n_2^2}\right]$$

Bracket series lies in the **infrared region.**

v) Pfund series

For Pfund series

$n_1 = 5$ and $n_2 = 6, 7, 8, \ldots$

$$\bar{\upsilon} = R\left[\frac{1}{5^2}-\frac{1}{n_2^2}\right]$$

This series is in the **far infrared region** of the spectrum.

13. Calculate the different Energy levels of hydrogen

Ans: We have

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

This equation gives the binding energy of the electron in the n^{th} orbit of hydrogen atom. The negative sign shows that the electron is bound to the nucleus.

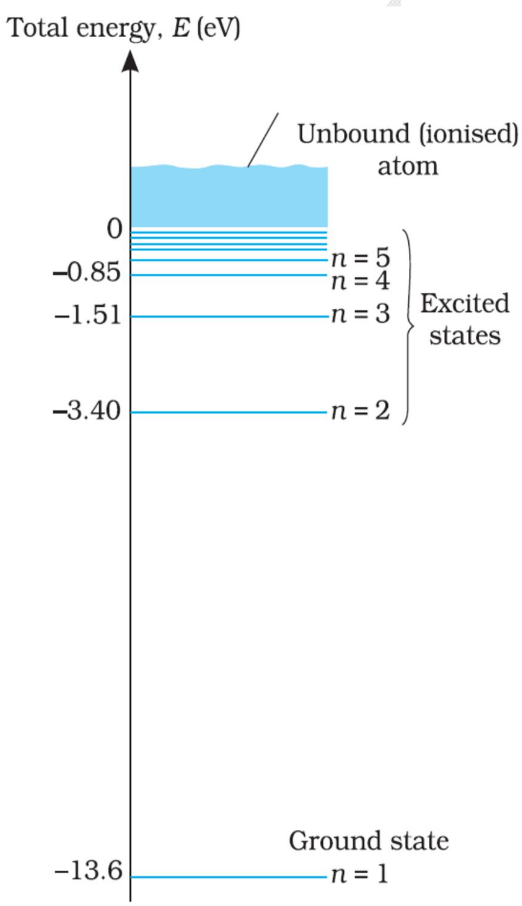
For n = 1 (K shell)

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

This is the ground state energy of hydrogen atom. This gives the ionisation energy of electron from n = 1 to n = ∞ , which is 13.6eV.

For n= 2, $E_2 = \frac{-13.6}{4} eV = -3.4 eV$

For n = 3, $E_3 = \frac{-13.6}{9} eV = -1.5 eV$



For n = 4, $E_4 = \frac{-13.6}{16} eV = -0.85 eV$

For n = 5, $E_5 = -0.54 eV$

For n = ∞, $E_\infty = 0$ eV

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

14. Define **Excitation energy**

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

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

First excitation energy of H atom

= -3.4 – (13.6) = 10.2eV

Second excitation energy of hydrogen atom is the energy required to excite the electron from n = 1 to n= 3 orbit of hydrogen atom is called second excitation energy H atom.

Second excitation energy of H atom

= -1.51-(-13.6) = 12.09eV.

15. Define **excitation potential**.

Ans: ***Excitation potential of an excited state is the potential difference through which electron in an atom has to be accelerated so as to excite it from its ground state to the given excited state.***

The **first excitation potential of H atom** is 10.2V. and the **second excitation potential of H atom** is 12.09V.

16. Define **ionization energy**

Ans: 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 = ∞)

Ionisation energy of H atom

= E_∞ - E_1 = 0 - (-13.6) = 13.6eV

17. What are the **drawbacks of Bohr's theory**?

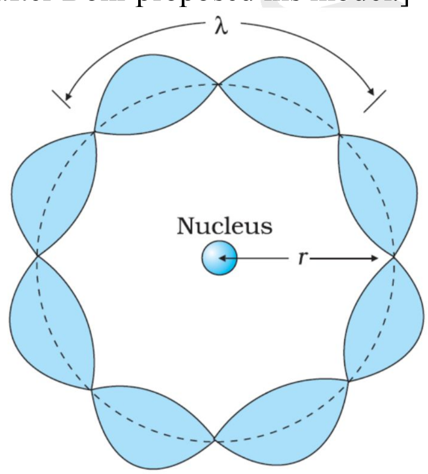
i) Bohr's theory could explain the spectra of only single electron atoms like hydrogen. For atoms having more number of electrons, Bohr's theory was found inadequate.

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.

18. Give **De Broglie's explanation of Bohr's second postulate of quantization.**

Ans: 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 $h/2\pi$?. 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

$$\mathbf{2\pi r_n = n\lambda}, \quad n = 1, 2, 3, \ldots..$$

$2\pi r_n = n\lambda$

But $\lambda = \frac{h}{mv}$

$\Rightarrow 2\pi r_n = \frac{nh}{mv}$

$\Rightarrow mvr_n = \frac{nh}{2\pi}$, $n = 1, 2, 3, \ldots\ldots$

Where $\mathbf{mvr_n}$ is the angular momentum (**L**)of the revolving electron. Hence the proof.

Problems

1. What is the shortest wave length present in the paschen series of spectral lines?

2. The ground state energy of hydrogen atom is – 13.6 eV. What are the kinetic and potential energies of the electron in this state?

3. A hydrogen atom initially in the ground level absorbs a photon, which excites it to the n = 4 level. Determine the wavelength and frequency of photon.

4. The radius of the innermost electron orbit of a hydrogen atom is 5.3×10^{-11} m. What are the radii of the n = 2 and n = 3 orbits?

5. Calculate the radius of third Bohr orbit of hydrogen atom and energy of electrons in that orbit.

6. Calculate the radius of third orbit of hydrogen atom and the energy of electrons in that orbit.
($h=6.625\times10^{-34}$ JS,
$\varepsilon_0=8.85\times10^{-12}C^2/Nm^2$,
$e=1.6\times10^{-19C}$, $m_e=9.1\times10^{-31}Kg$)

7. Find out the shortest wavelength of spectral line emitted in Balmer series.