ATOMS Chapter 12 NCERT PART 2

INSTRUCTIONS

1 This is a short chapter which consists of the following topics as per CBSE syllabus

Alpha-particle scattering experiment; Rutherford's model of atom; Bohr model, energy levels, hydrogen spectrum.

2 It has small weightage of about 3-4 marks in the CBSE examination.

3 To fully prepare this chapter , you can practice previous year question papers of CBSE given along here.

4 Kindly do the assignment in your physics registers.

Main points of the chapter

INTRODUCTION

PLUM PUDDING MODEL – J J THOMPSON

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 and the negatively charged electrons are embedded in it like seeds in a watermelon.

This model was picturesquely called **plum pudding model** of the atom. However subsequent studies on atoms, as described in this chapter, showed that the distribution of the electrons and positive charges are very different from that proposed in this model.

Rutherford's Alpha Scattering Expt.:

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

They directed a beam of 5.5 MeV α-particles emitted from a Bismuth radioactive source at a thin metal foil made of gold.

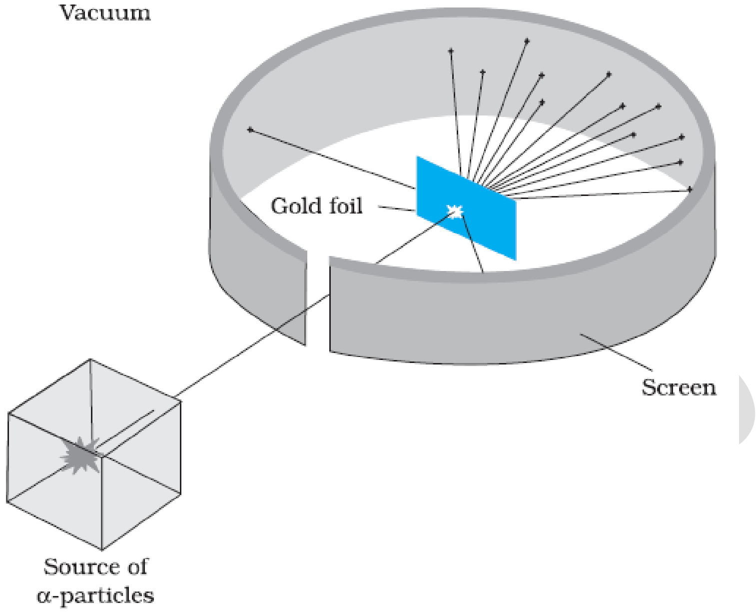
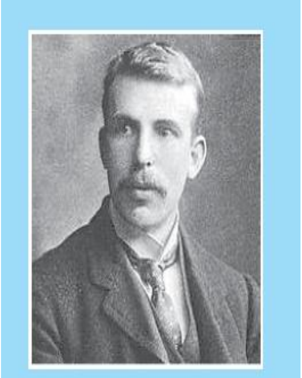


FIGURE 12.1 Geiger-Marsden scattering experiment. The entire apparatus is placed in a vacuum chamber (not shown in this figure).



Ernst Rutherford (1871 – 1937) British physicist

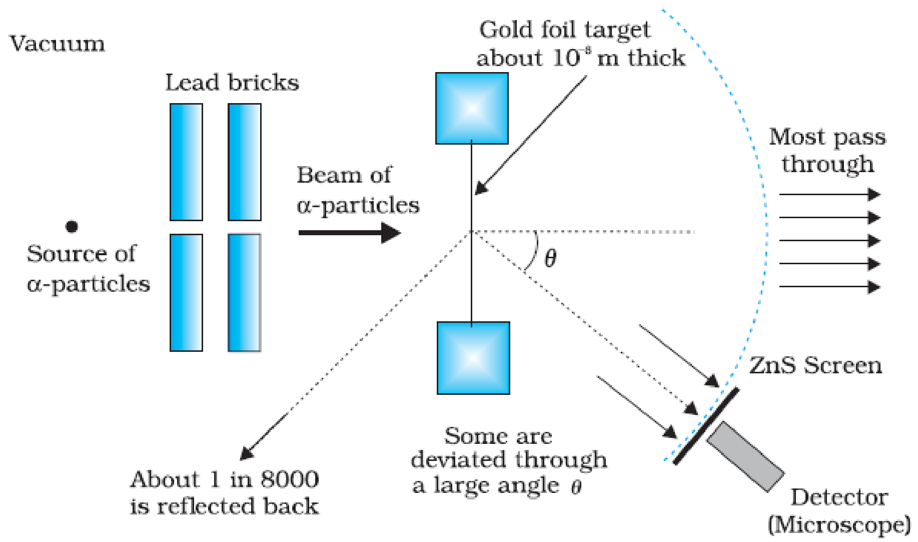


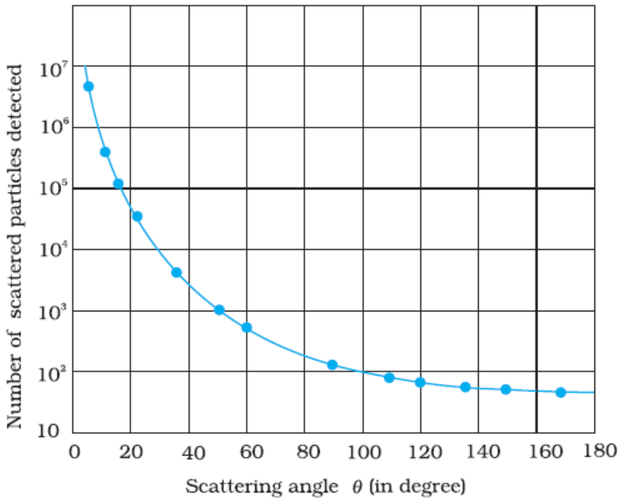
FIGURE 12.2 Schematic arrangement of the Geiger-Marsden experiment.

(Diagram credit NCERT)

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

The scattered alpha-particles on striking the screen produced brief light flashes
or scintillations. These flashes may be viewed through a microscope and the distribution of the number of scattered particles may be studied as a function of angle of scattering.

.The graph showing the number of alpha particles scattered versus the angle of scattering.



(Diagram credit NCERT)

Many of the α-particles pass through the foil.

It means that they do not suffer any collisions. Only about 0.14% of the incident α-particles scatter by more than 1°; and about 1 in 8000 deflect by more than 90°.

Rutherford argued that, to deflect the α-particle backwards, it must experience a large repulsive force. This force will be provided if the greater part of the mass of the atom and its positive charge were concentrated tightly at its centre.

1 In Rutherford's nuclear model of the atom, the entire positive charge and most of the mass of the atom are concentrated in the nucleus with the electrons some distance away.

2 The electrons would be moving in orbits .

3 Rutherford's experiments suggested the size of the nucleus to be about 10^{-15} m to 10^{-14} m

This implies that there is lot of empty space in the atom.

IMPACT PARAMETER

The trajectory taken by an alpha particle is determined by the value of impact parameter (b).

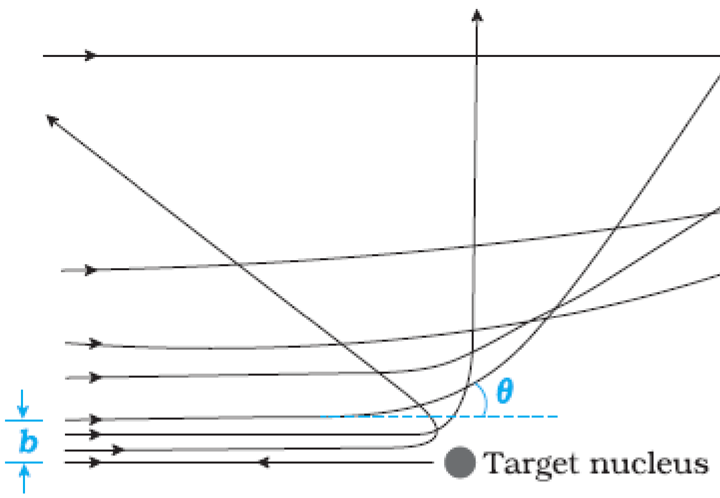


FIGURE 12.4 Trajectory of α-particles in the coulomb field of a target nucleus. The impact parameter, b and scattering angle θ are also depicted.

(Diagram credit NCERT)

The *impact parameter* is the perpendicular distance of the initial velocity vector of the α-particle from the centre of the nucleus,when the alpha particle is far away from the nucleus.

Larger is the value of b, smaller is the deflection. For **alpha** particles, which are bounced back, b is almost 0.

Distance of Closest approach:

It is the closest distance between the nucleus and alpha particle when the alpha particle comes to momentary rest and reverses its direction.

When the alpha particle is far from the nucleus, it possesses K.E. and at the distance of closest approach, the entire KE is converted to electric PE.

KE = PE

$$KE = \frac{KZe \times 2e}{r} \text{ (EPE)}$$

$$r = \frac{2KZe^2}{KE}$$

The distance of closest approach is considerably larger than the sum of the radii of the gold nucleus and the α-particle. Thus, the α-particle reverses its motion without ever actually *touching* the gold nucleus.

Failure of Rutherford's Model:

(i) According to Rutherford's model, electrons are revolving in circular orbits around nucleus but from EM wave theory we know that an accelerated electron must emit an EM wave and loss energy and follow a spiral path and fall into the nucleus. In such an atom, the electrons must emit a continuous spectrum but actually atoms have line spectra.

(ii) The Rutherford's model could not account for the stability of atom

BOHR'S MODEL OF ATOM

Bohr combined classical and early quantum concepts and gave his theory in the form of three postulates:

(i) *An electron in an atom could revolve in certain stable orbits without the emission of radiant energy.*

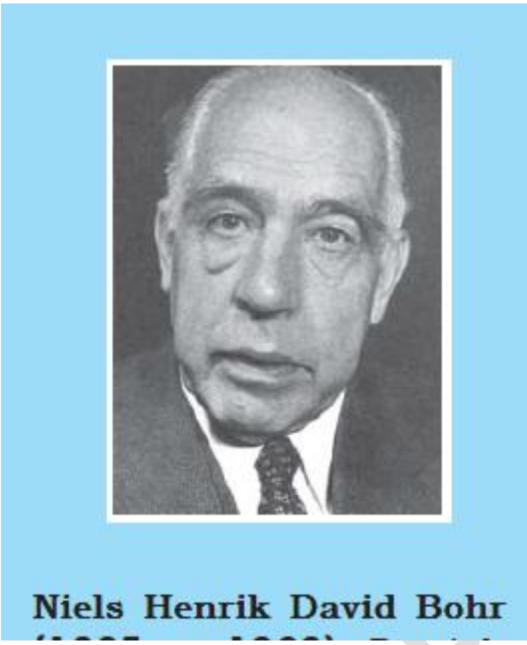
(ii) The electron revolves around the nucleus only in those orbits for which the angular momentum is an integral multiple of $h/2\pi$.

(iii) When an electron undergoes a transition from one energy level to another, 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

$h\nu = E_i - E_f$

where Ei and Ef are the energies of the initial and final states and Ei > Ef .



(picture credit NCERT)

Electron orbits

The Rutherford nuclear model of the atom which involves classical

concepts, pictures the atom as an electrically neutral sphere consisting of a very small, massive and positively charged nucleus at the centre surrounded by the revolving electrons in their respective dynamically stable orbits.

The electrostatic force of attraction, Fe between the revolving electrons and the nucleus provides the requisite centripetal force (Fc) to keep them in their orbits. Thus, for a dynamically stable orbit in a hydrogen atom

Fe = Fc

$$K\frac{e^2}{r^2} = \frac{mv^2}{r}$$

Hence Kinetic Energy = $\frac{1}{2}mv^2 = K\frac{e^2}{2r} = \frac{e^2}{8\pi\epsilon_o r}$

Potential energy=$-K\frac{e^2}{r} = \frac{-e^2}{4\pi\epsilon_o r}$

Total energy = KE +PE=$\frac{-e^2}{8\pi\epsilon_o r}$

Expressions for radius, velocity and Energy of an electron in a bohrs orbit

Since according to the first postulate

Fe = Fc

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

Or

$$\frac{e^2}{4\pi\epsilon_o} = mv^2 r ----(1)$$

Also as per the second postulate

$$mvr = \frac{nh}{2\pi} ----(2)$$

Dividing them we get

$$v = \frac{e^2 2\pi}{4\pi\epsilon_o nh}$$

Or

$$v = \frac{e^2}{2\epsilon_o nh} - - - - - - (3)$$

This shows that velocity of an electron in orbit is inversely proportional to n

Expression for Radius

Substituting the value of v from (3) in (2)

We get

$$m\left(\frac{e^2}{2\epsilon_o nh}\right)r = \frac{nh}{2\pi}$$

Or

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

$$r = \frac{n^2h^2\epsilon_o}{\pi me^2} = \frac{n^2(h^2\epsilon_o)}{(\pi me^2)}$$

This shows that radius of orbit is directly proportional to n^2

The term in the bracket is a constant and is the radius of the first Bohr's radius,

a0 = 5.29×10^{-11} m

The radius of any orbit is given by

rn = n^2 a0

The radii of orbits are in the ratio of 1: 4 : 9 : 16……

Energy

Since total energy = $\frac{-e^2}{8\pi\epsilon_o r}$

Substituting the value of r from above

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

$$E = \frac{-\pi m e^4}{8\pi n^2 h^2 {\epsilon_o}^2}$$

This shows that Energy of electron is inversly proportional to n^2

Solving the values

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

The negative sign of the total energy of an electron moving in an orbit means that the electron is bound with the nucleus. Energy will thus be required to remove the electron from the hydrogen atom. , this energy is also known as ionization energy of the electron.

THE LINE SPECTRA OF THE HYDROGEN ATOM

According to Bohr's third postulate, when electron makes a transition from higher energy state (ni) to a lower energy state (nf) where ni < nf, the energy is emitted as a photon with energy

$h\nu = E_2 - E_1$

Subtituting value of En from equation (6), we get

$As\ h\nu = \frac{hc}{\lambda} = E_2 - E_1$

$$h\nu = \frac{hc}{\lambda} = \frac{-\pi m e^4}{8\pi n_2^2 h^2 {\epsilon_o}^2} - \left(-\frac{-\pi m e^4}{8\pi n_1^2 h^2 {\epsilon_o}^2}\right)_2$$

Or

$$\frac{hc}{\lambda} = \frac{-\pi m e^4}{8\pi h^2 {\epsilon_o}^2}\left(\frac{1}{n_2^2} - \frac{1}{n_1^2}\right)$$

$$\frac{hc}{\lambda} = \frac{\pi m e^4}{8\pi h^2 {\epsilon_o}^2}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

Or

$$\frac{1}{\lambda} = \frac{\pi m e^4}{c\,8\pi h^3 {\epsilon_o}^2}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

The term $\frac{\pi m e^4}{c\,8\pi h^3 {\epsilon_o}^2}$ *is called Rydberg Constant (R) and its value for H atom is* $1.03 \times 10^7/m$

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. The spectrum is called ***emission spectrum.***

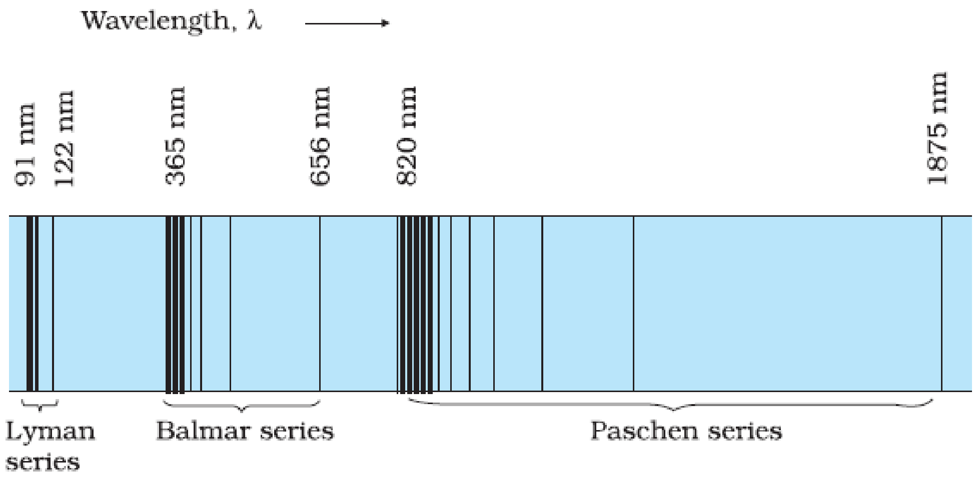
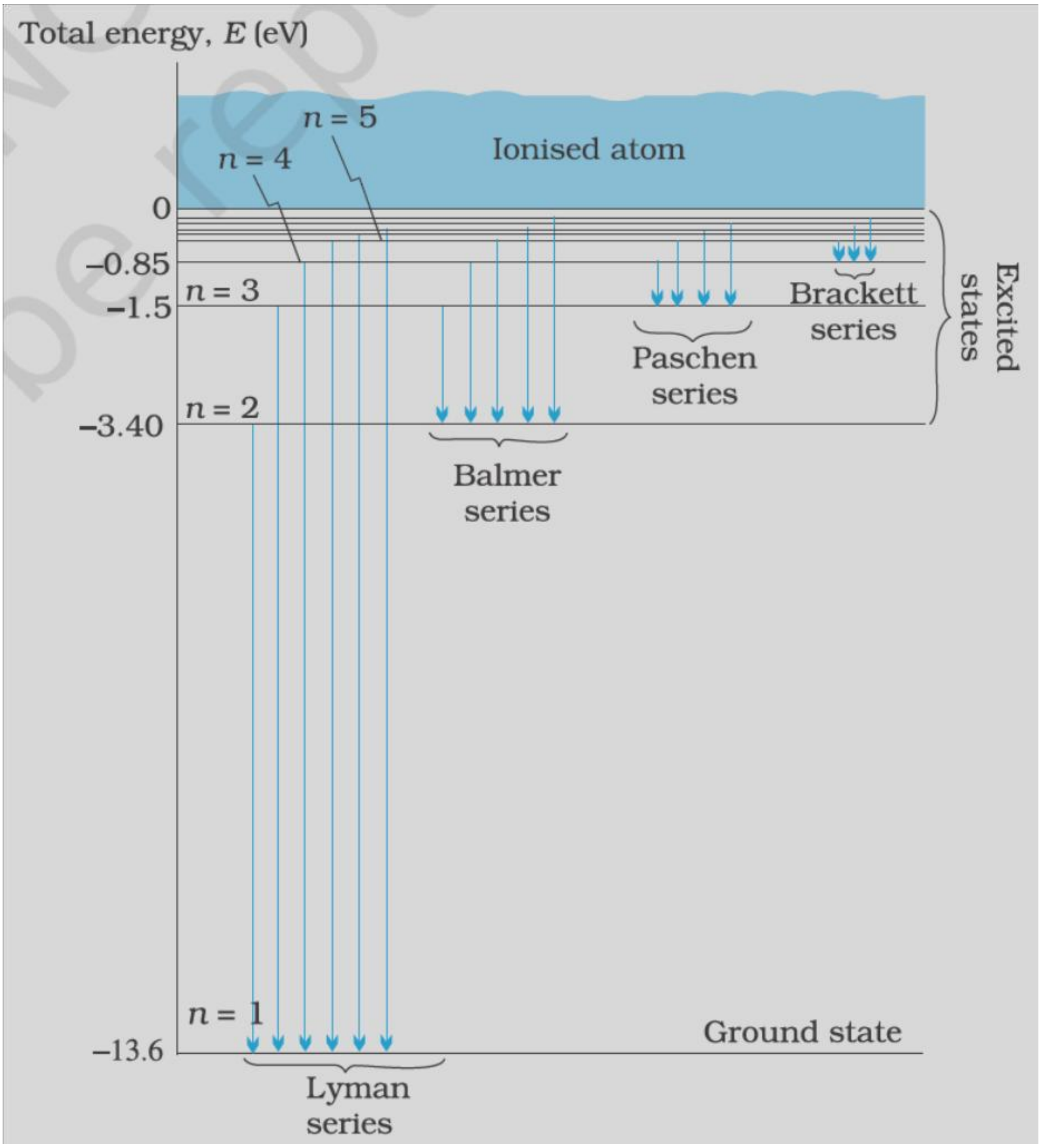


FIGURE 12.5 Emission lines in the spectrum of hydrogen.

(Diagram credit NCERT)

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. 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. The spectrum is called* ***absorption spectrum.***



(Diagram credit NCERT)

SPECTRAL SERIES OF HYDROGEN

Lyman series

$$\frac{\boldsymbol{1}}{\boldsymbol{\lambda}} = R\left(\frac{1}{\boldsymbol{n_1^2}} - \frac{1}{\boldsymbol{n_2^2}}\right)$$

For this series value of n_1 =1 and value of n_2 can be 2,3,4 and so on till infinity

It lies in UV region.

BALMER Series:

$$\frac{\boldsymbol{1}}{\boldsymbol{\lambda}} = R\left(\frac{1}{\boldsymbol{n_1^2}} - \frac{1}{\boldsymbol{n_2^2}}\right)$$

For this series value of n_1 =2 and value of n_2 can be 3,4 ,5 and so on till infinity

It lies in visible region of spectrum.

The line with the longest wavelength, 656.3 nm in the red is called ***Hα****; the next line with wavelength 486.1 nm in the blue-green is called* ***Hβ****, the third line 434.1 nm in the violet is called* ***Hγ****; and so on. As the wavelength decreases, the lines appear closer together and are weaker in intensity*

PASCHEN Series:

$$\frac{\boldsymbol{1}}{\boldsymbol{\lambda}} = R\left(\frac{1}{\boldsymbol{n_1^2}} - \frac{1}{\boldsymbol{n_2^2}}\right)$$

For this series value of n_1 =3 and value of n_2 can be 4,5,6 and so on till infinity

Lies in ***infra red region***

BRACKETT Series:

$$\frac{\boldsymbol{1}}{\boldsymbol{\lambda}} = R\left(\frac{1}{\boldsymbol{n_1^2}} - \frac{1}{\boldsymbol{n_2^2}}\right)$$

For this series value of n_1 =4 and value of n_2 can be 5,6,7 and so on till infinity

Lies in ***infra red region***

PFUND series

$$\frac{\mathbf{1}}{\boldsymbol{\lambda}} = R\left(\frac{1}{\boldsymbol{n_1^2}} - \frac{1}{\boldsymbol{n_2^2}}\right)$$

For this series value of n_1 =5 and value of n_2 can be 6,7,8 and so on till infinity

Lies in ***infra red region***

Failure of Bohr's Model:

(i) Bohr's model is only applicable to hydrogen or hydrogen like atoms and fails even in atoms with two electrons. As it neglects interaction between electrons.

(ii) Bohr's model could not explain difference in relative intensities of the transition lines. Some of the spectral lines are brighter means there is higher probability of these transitions.

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

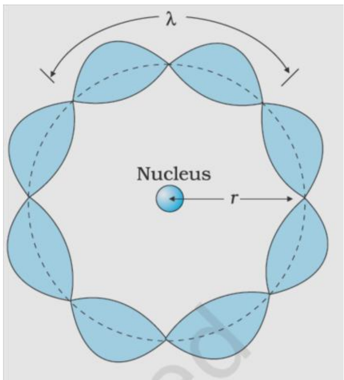
Fromtheory of stationary waves, for standing waves to be formed, the length of the string must be equal to some multiple of wavelength, for an orbit of radius rn,

$2\pi\, r_n = n\lambda$

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

Or $2\pi\, mvr_n = nh$

$$l = \frac{nh}{2\pi}$$



(Diagram credit NCERT)

But this is the quantization condition imposed by Bohr.

Thus de Broglie hypothesis provided an explanation for Bohr's second postulate for the quantization of angular momentum of the orbiting electron. The quantised electron orbits and energy states are due to the wave nature of the electron and only resonant standing waves can persist.

Video links

RUTHERFORD EXPERIMENT

https://www.youtube.com/watch?v=u7xrDhpY-oA

Bohr's Theory

https://www.youtube.com/watch?v=au2HCVn9IJI

Emission Spectrum

https://www.youtube.com/watch?v=lJh2Ra1eygA

Assignment on the topic

ATOMS(questions from previous year CBSE papers)

1 Name the series of hydrogen spectrum which has least wavelength.

2 The ground state energy of Hydrogen atom is -13.6 eV. What is the KE of an electron in the 3rd excited state?

3 Calculate the radius of the third Bohr orbit of hydrogen atom and energy of electron in that orbit.

4 What is impact parameter? What is the value of impact parameter for a head on collision?

5 The value of ground state energy of hydrogen atom is: -13.6 eV; (i) What does the negative sign signify? & (ii) How much energy is required to take an electron in this atom from the ground state to the first excited state?

6 Calculate the frequency of the photon which can excite an electron to -3.4 eV from -13.6 eV.

7 The wavelength of the first member of Balmer series in the hydrogen spectrum is 6563Å.Calculate the wavelength of the first member of Lyman series in the same spectrum.

8 Find the ratio of maximum wavelength of Lyman series in hydrogen spectrum to the maximum wavelength in Paschen Series?

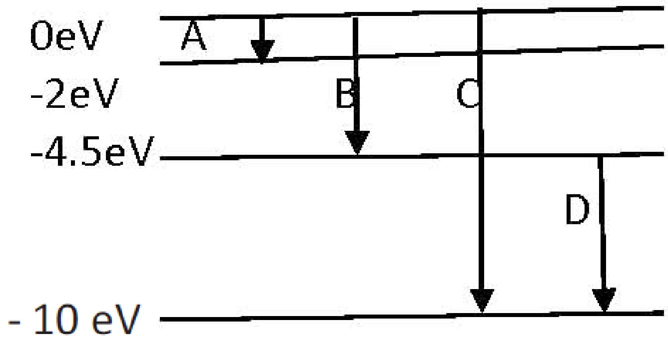
9 Calculate the ratio of wavelength of photon emitted due to transition of electrons of hydrogen atom from

i) Second permitted level to first level

ii) Highest permitted level to second level

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

11 The energy levels of an atom are as shown below. a) Which of them will result in the transition of a photon of wavelength 275 nm? b) Which transition corresponds to the emission of radiation maximum wavelength?



Theory Based questions

1 Define distance of the closest approach. An alpha particle of kinetic energy ‘K’ is bombarded on a thin gold foil. The distance of the closet approach is ‘r’. What will be the distance of closest approach of double the kinetic energy

2 State Bohr’s postulates. Using these postulates, drive an expression for total energy of an electron in the nth orbit of an atom. What does negative of this energy signify?

3.Derive an expression for the radius of stationary orbit. Prove that the various stationary orbits are not equally spaced.

4.Derive mathematical expressions for: (i) kinetic energy, & (ii) potential energy of an electron revolving in an orbit of radius ‘r’; how does the potential energy change with increase in principal quantum number (n) for the electron and why?

5.In Rutherford’s scattering experiment, mention two important conclusions which can be drawn by studying the scattering of α particles by an atom. Draw the schematic arrangement of Geiger and Marsden experiment showing the scattering of α particle by a thin foil of gold. How does one get the information regarding the size of the nucleus in this experiment?

6Sketch the energy level diagram for hydrogen atom. Mark the transitions corresponding to Lyman and Balmer series.