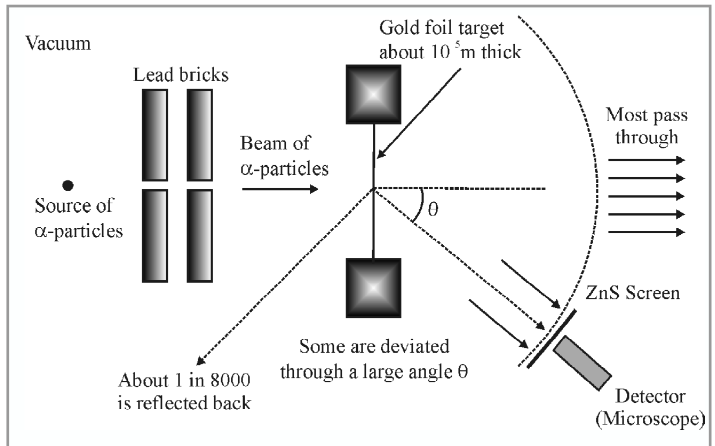
1. ALPHA PARTICLE SCATTERING EXPERIMENT

Thomson's Atom Model

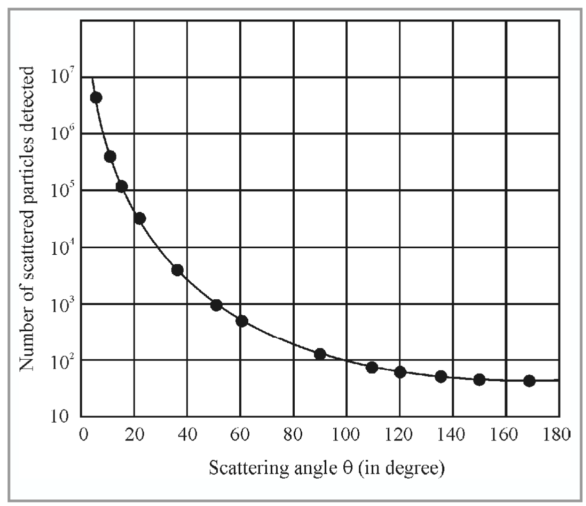
- J. J. Thomson presented an atomic model in which positive charge is distributed uniformly in a small spherical space of atom and electrons are embedded inside it like the seeds of watermelon embedded in its pulp. The model was called watermelon model or plum pudding model.
- The magnitude of positive charge was taken equal to the total negative charge of electrons to explain the electrical neutrality of atom.
- The model failed to explain the emission of discrete wavelengths from atoms. Also such a model of atom cannot form a stable structure.

Geiger–Masden Experiment

Geiger and Masden used a bismuth source, placed inside a thick block of lead in order to get α-particles. The particles having energy of 5.5 MeV coming out of a slit in the block were scattered when incident on a thin foil of gold of thickness 2.1×10^{-7} m.



Scintillations were observed on a screen of zinc sulphide mounted on a (detector) microscope. The whole arrangement was enclosed in a cylinder having thick walls. The cylinder was mounted on a thick disc which could be rotated. The source and foil were kept steady on the base but could be rotated around the microscope. Each α-particle striking the screen created a bright spot in it. By counting the bright spots formed, the number of α-particles striking in a given time interval can be decided.



The graph shows the number of α-particles scattered at different angles in a given time interval. The number of α-particles scattered is about 10^5 at 15° and about 80 at 150°. (Dots indicate experimental values and the continuous line was obtained theoretically by Rutherford).

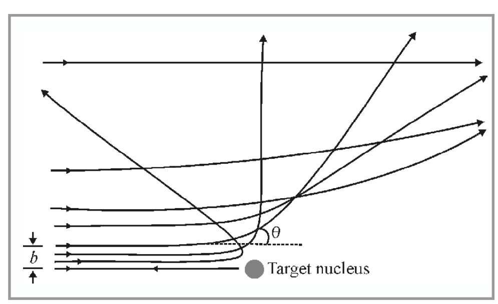
Rutherford's Calculations

- An α-particle can undergo only one scattering while passing through the thin foil.
- An atom of Gold remains almost stationary during the collision as it is about 50 times heavier than the α-particle.
- Most of the α-particles pass through the foil without deviation. Of the scattered particles, most of them do so at small angles. It indicates that the atom contains a lot of empty space.
- A few α-particles are scattered at large angles. Further, a very less number of particles are reflected back. This suggests the total positive charge and total mass of the atom to be concentrated in a very small central region of the atom, which he called the nucleus.

The repulsive force on α-particle when incident on gold foil due to the nucleus of gold is

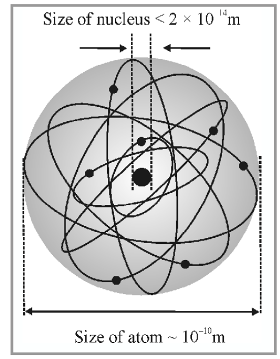
$$F = \frac{1}{4\pi\varepsilon_0}\frac{(Ze)(2e)}{r^2}$$

where 2e = charge of α-particle, Ze = charge in gold nucleus (Z = 79 for gold)



Obviously, this trajectory depends on the initial perpendicular distance of its velocity vector from the nucleus. This minimum perpendicular distance is called the distance of closest approach or impact parameter (*b*).

2. RUTHERFORD'S ATOM MODEL



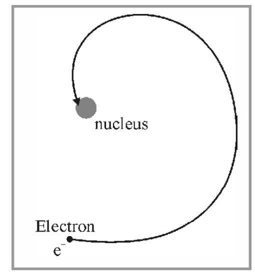
Atom has a central part called nucleus which contains the entire positive charge and almost all the mass of the atom.

Electrons revolve round the nucleus and the required centripetal force is supplied by the electrostatic force of attraction between the electron and the nucleus.

Atom includes wide empty area.

The number of electrons is defined so as to nullify the effect of the positive charge of the nucleus.

Failure of Rutherford's Model



Electron moving around the nucleus is in accelerated motion and so it should emit an electromagnetic wave.

The electron must lose the energy .to become out of balance and spiral inward, and finally collapse on the nucleus (very quickly, within 10^{-8} s). In such a case, the atom will be highly unstable. But this does not happen.

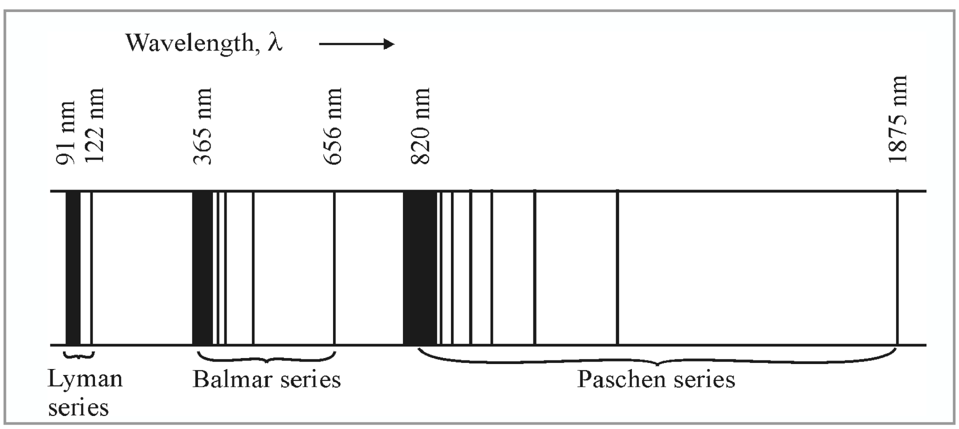
According to this model, the radiations emitted by the electron should be continuous. But, analysis of radiations emitted by hydrogen shows spectral lines instead of continuous band.

3. BOHR MODEL OF HYDROGEN ATOM

Atomic Spectra

Emission spectra: When an atomic gas or vapour is excited at low pressure, it emits radiations in certain specific wavelengths only. A spectrum of this kind is termed as emission line spectrum as spectral analysis shows a few bright lines on a dark background.

The spectrum emitted by atomic hydrogen is shown in the figure.



Each element has a characteristic spectrum of radiation, which it emits. Therefore, study of emission line spectra of a material can be used for the identification of the material.

Absorption spectra: When white light is passed through a gas and the transmitted light is analysed using a spectrometer, a few dark lines are observed in the spectrum. These dark lines correspond to the wavelengths absorbed by the gas and the set of these dark lines is called the absorption spectrum of the gas.

Spectral Series

Emission spectrum of a material consists of the wavelengths emitted by it. There is no regular pattern for the entire spectrum. But, the spacing between lines within certain sets of the spectrum may change in a regular way. Each of these sets can be called a spectral series.

The atomic hydrogen emits a line spectrum consisting of various series. The first series observed was Balmer series in the visible region.

The wavelengths in the Balmer series follow a simple formula, which is given below.

$$\frac{1}{\lambda} = R\left(\frac{1}{2^2} - \frac{1}{n^2}\right), \quad n = 3, 4, 5, 6, ...$$ {Balmer series is observed in visible region}

λ is the wavelength and R is a constant called the Rydberg constant with value $1.097 \times 10^7\ m^{-1}$

Other series in hydrogen spectrum and their respective formulae are given below:

Lyman series (UV):

$$\frac{1}{\lambda} = R\left(\frac{1}{1^2} - \frac{1}{n^2}\right), \quad n = 2, 3, 4, ...$$

Paschen series (IR):

$$\frac{1}{\lambda} = R\left(\frac{1}{3^2} - \frac{1}{n^2}\right), \quad n = 4, 5, 6, ...$$

Brackett series (IR):

$$\frac{1}{\lambda} = R\left(\frac{1}{4^2} - \frac{1}{n^2}\right), \quad n = 5, 6, 7, ...$$

Pfund series (IR):

$$\frac{1}{\lambda} = R\left(\frac{1}{5^2} - \frac{1}{n^2}\right), \quad n = 6, 7, 8, ...$$

Rutherford's atom model could not explain why radiations are emitted in certain wavelengths only.

Bohr Model

Niels Bohr made modifications to Rutherford's atom model. The postulates of Bohr model are:

- An electron in an atom could revolve in certain stable orbits without the emission of radiant energy.
- Bohr quantisation condition: Electron revolves around the nucleus only in those orbits for which the angular momentum is an integral multiple of $\frac{h}{2\pi}$, where h is the Planck's constant. The angular momentum (L) of the orbiting electron is quantised.

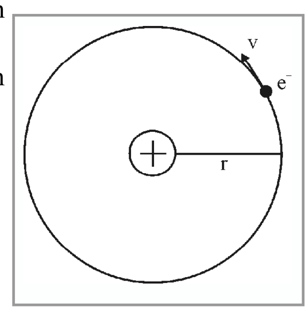
  $$L = \frac{nh}{2\pi}$$
- Bohr frequency condition: An electron could make a transition from one of its specified non-radiating orbits to another. When transition takes place to another orbit of lower energy, a photon is emitted having energy equal to the energy difference between the initial and final states. If ν is the frequency of the emitted photon, then:

  $$h\nu = E_f - E_i$$

  Energy is to be supplied to transfer electron to a higher energy orbit.

Theory

Let an electron having mass m and charge e revolves around a nucleus having charge Ze in a circular orbit of radius r with a linear speed v as shown in the figure.



The centripetal force is provided by the Coulomb force of attraction between the electron and nucleus.

$$\frac{mv^2}{r} = \frac{1}{4\pi\varepsilon_0}\frac{Ze^2}{r^2} \quad ...(i)$$

The angular momentum of the electron in an orbit

$$L = mvr = \frac{nh}{2\pi}$$ where, n = 1, 2, 3, ... is the principal quantum number.

Then, $m^2v^2r^2 = \frac{n^2h^2}{4\pi^2}$ ...(*ii*)

From (*i*) and (*ii*), the radius of the orbit,

$$r = \frac{n^2h^2\varepsilon_0}{\pi Ze^2 m} \quad \text{...}(iii)$$

The speed of electron in the orbit : $v = \frac{Ze^2}{2\varepsilon_0 nh}$ ...(*iv*)

The energy of electron in the orbit is given by:

Kinetic energy, $K = \frac{1}{8\pi\varepsilon_0}\frac{Ze^2}{r}$

Potential energy, $U = -\frac{1}{4\pi\varepsilon_0}\frac{Ze^2}{r}$

Total energy (in the nth orbit),

$$E_n = K + U = K = -\frac{1}{8\pi\varepsilon_0}\frac{Ze^2}{r} \quad \text{...}(v)$$

Above three equations show that
K.E = – (T. E)
P. E = 2 T.E
The negative sign indicates that the electron is bound to the nucleus.
From (*iii*) and (*iv*),

$$E_n = -\frac{Z^2e^4m}{8\varepsilon_0^2n^2h^2} \quad \text{...}(vi)$$

Hydrogen Atom
For hydrogen atom, Z = 1
The radius of the nth orbit from (*iii*),

$$r_n = \frac{n^2h^2\varepsilon_0}{\pi e^2 m} \quad \text{...}(vii)$$

Radius of the first (innermost) orbit is called the Bohr radius. It is denoted by a_{-0} and can be obtained by putting n = 1 and substituting for symbols in (*vii*). Bohr radius is calculated to be 0.53 Å
Energy in the nth orbit from (*vi*),

$$E_n = -\frac{me^4}{8\varepsilon_0^2h^2n^2} = -\frac{13.6}{n^2}\ eV \quad \text{...}(viii)$$

Using equation (*viii*), the energy of an electron in different orbits can be calculated for *n* = 1, 2, 3, …

4. LINE SPECTRA OF HYDROGEN ATOM

Energy Levels: The total energy of the electron in the nth orbit of hydrogen atom is given by:

$$E_n = -\frac{13.6}{n^2}\ \text{eV} \quad \ldots (i)$$

The negative sign indicates that the electron is bound to the nucleus by a force of attraction. Additional energy is to be supplied from outside to release an electron.

Maximum energy of an electron is taken as zero (for $n = \infty$)

The various energy levels for the electron are given below:

$n = 1,\quad E_1 = -\frac{13.6}{1^2} = -13.6\,\text{eV}$, ground state

$n = 2,\quad E_2 = -\frac{13.6}{2^2} = -3.4\,\text{eV}$, first excited state

$n = 3,\quad E_3 = -\frac{13.6}{3^2} = -1.51\text{eV}$, second excited state

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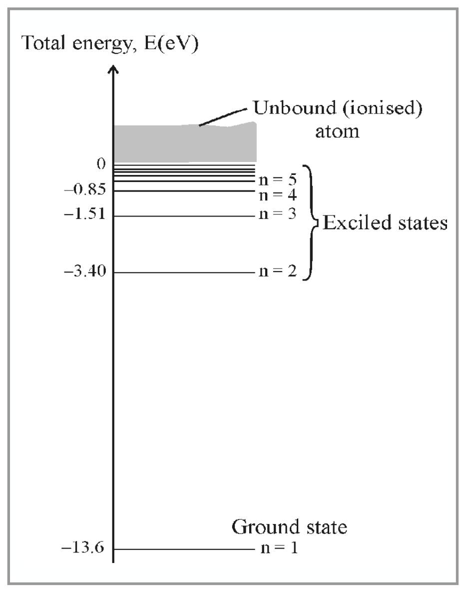
$n = \infty$, $E_\infty = 0$ the atom is said to be ionised.

The energy level diagram is shown below:



Line Spectra of hydrogen atom

The line spectra of hydrogen atom arise from the transition of electron from higher energy levels to the lower energy levels. The excess of energy is emitted as a photon.

Transition from a higher energy level n_i to a lower energy level n_f causes the emission of a photon of frequency ν_{fi} such that:

$$h\nu_{fi} = E_{n_f} - E_{n_i} = \frac{me^4}{8\varepsilon_0^2 h^2}\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

$$\nu_{fi} = \frac{me^4}{8\varepsilon_0^2 h^3}\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right) \ldots (ii)$$

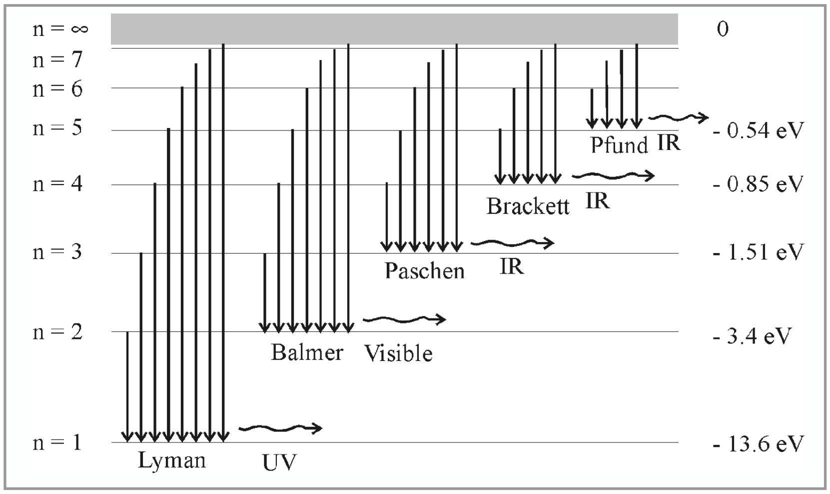
Equation (*ii*) is referred to as Rydberg formula for the spectrum of hydrogen atom.

But $\quad \nu_{fi} = \frac{c}{\lambda_{fi}}$

Then, $\quad \frac{1}{\lambda_{fi}} = \frac{me^4}{8\varepsilon_0^2 h^3 c}\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right) = R\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$

$R = \frac{me^4}{8\varepsilon_0^2 h^3 c}$ is called the Rydberg constant.

Various spectral series emitted by hydrogen atom are explained as below:



Lyman series: arise due to the transition from any higher energy level to the ground state (n = 1). These are UV radiations.

$$\frac{1}{\lambda_{f1}} = R\left(\frac{1}{1^2} - \frac{1}{n_i^2}\right), \qquad n_i = 2, 3, 4, \ldots$$

Balmer series: arise due to the transition from any higher energy level to the first excited state (n = 2). This is visible light.

$$\frac{1}{\lambda_{f2}} = R\left(\frac{1}{2^2} - \frac{1}{n_i^2}\right), \qquad n_i = 3, 4, 5 \ldots$$

Paschen series: arise due to the transition from any higher energy level to the second excited state (n = 3). These are IR radiations.

$$\frac{1}{\lambda_{f3}} = R\left(\frac{1}{3^2} - \frac{1}{n_i^2}\right), \qquad n_i = 4, 5, 6, \ldots$$

Bracket series: arise due to the transition from any higher energy level to the third excited state (n = 4). These are also IR radiations.

$$\frac{1}{\lambda_{f4}} = R\left(\frac{1}{4^2} - \frac{1}{n_i^2}\right), \qquad n_i = 5, 6, 7, \ldots$$

Pfund series: arise due to the transition from any higher energy level to the ground state (n = 5). These are also IR radiations.

$$\frac{1}{\lambda_{f5}} = R\left(\frac{1}{5^2} - \frac{1}{n_i^2}\right), \qquad n_i = 6, 7, 8, \ldots$$

Hydrogen gas can give rise to absorption spectra also. If photons with a continuous range of frequencies pass through rarefied hydrogen gas and the transmitted photons are analysed with a spectrometer, then a series of dark spectral absorption lines appear in the spectrum. These dark lines indicate the frequencies that have been absorbed. The absorbed energy is used to transfer electrons from the lower energy levels to the higher energy levels.

Note: A material absorbs the same wavelengths which they can emit.

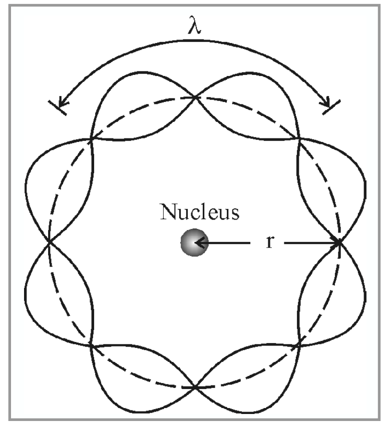
Limitations of Bohr Model

- ♦ Orbits of electron need not be circular as assumed in the Bohr model.
- ♦ The calculation of energies in the orbits involve complex combination of classical
- ♦ Mechanics and the quantum principles.
- ♦ Additional lines are seen with many of the spectral lines of hydrogen spectra, which cannot be explained on the basis of the Bohr model.
- ♦ Bohr model fails to explain the relative intensities of the spectral lines observed in the actual spectra.

5. DE BROGLIE'S EXPLANATION OF BOHR'S QUANTISATION

Going with the wave nature of electrons, the condition of non-radiation of energy can be linked to the formation of stationary waves. For the vibrations of a wire loop, the stationary waves would be formed if each wave joins smoothly with the next, *i.e.* the number of wavelengths on the loop must be an integer.

Similarly, an electron can circle around an atomic nucleus indefinitely without radiating energy if the circumference of its orbit is an integral multiple of the wavelength of the electron.



$2\pi r = n\lambda$

$= \frac{nh}{p}$ (as $\lambda = \frac{h}{p}$, the de Broglie wavelength)

Or, $pr = n\frac{h}{2\pi}$

$mvr = \frac{nh}{2\pi}$ (as $p = mv$)

Previous Year Questions

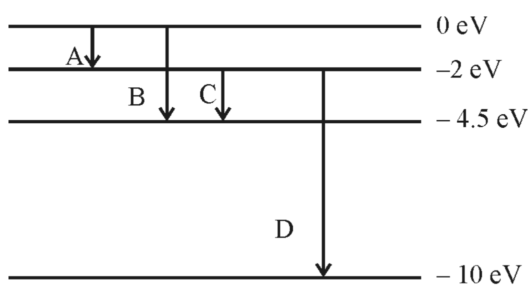
2012

1. State de-Broglie hypothesis.

2. (a) Using Bohr's second postulate of quantization of orbital angular momentum show that the circumference of the electron in the n^{th} orbital state in hydrogen atom is n times the de Broglie wavelength associated with it. (b) The electron in hydorgen atom is initially in the third excited state. What is the maximum number of spectral lines which can be emitted when it finally moves to the ground state?

2011

1. The energy levels o a hypothetical atom are shown below. Which of the shown transitions will result in the emission of a photon of wavelength 275 nm? Which of these transitions correspond to emission of radiation of (i) maximum and (ii) minimum wavelength ?

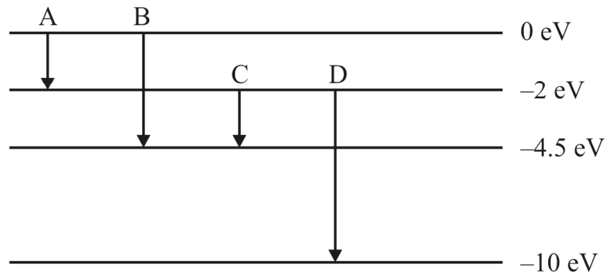


2010

1. Write the expression for Bohr's radius in hydrogen atom.

2009

1. (*a*) The energy levels of an atom are as shown below. Which of them will result in the tranistion of a photon of wavelength 275 nm.



(*b*) Which transition corresponds to emission of radiation of maximum wavelength. **[Ans. (a) B, (b) A]**

2008

1. The ground state energy of hydrogen atom is – 13.6 eV.

(*i*) What is the kinetic energy of an electron in the third excited state ? **[Ans. 1.5 eV]**

(*ii*) If the electron jumps to the ground state from the third excited state, calculate the wavelength of the spectral line emitted. **[Ans. 102 nm]**

2. In Bohr's theory of hydrogen atom, calculate the energy of the photon emitted during a transition of the electron from the first excited state to its ground state. Write in which region of the electromagnetic spectrum this transition lies. Given Rydberg constant $R = 1.03 \times 10^7\ m^{-1}$. **[Ans. 121 nm, Lyman series]**

3. The energy of the electron in hydrogen atom is known to be expressible in the form $E_n = -\frac{13.6}{n^2} eV$ $(n = 1, 2, 3)$.

Use this expression to show that the

(*i*) Electron in the hydrogen atom cannot have an energy of – 2 eV.

(*ii*) spacing between in lines (consecutive energy levels) within the given set of the observed hydrogen spectrum decreases as n increases.

2007

1. The ground state energy hydrogen atom is – 13.6 eV. If an electron makes a transition from an energy level –1.5 eV to –3.4 eV, calculate the wevelength of the spectral line emitted. To which series of hydrogen spectrum, does this wavelength belong ? **[Ans. 654 nm , Balmer series]**

