

Introduction

- i) In 1895, Jean Perin proved that electrons (known as cathode rays that time) are negatively charged.
- ii) Shortly thereafter, J. J. Thomson determined charge (e) to mass (m) ratio of electrons to be of the order of 10^{11} . This meant that m/e of electron is of the order of 10^{-11} and that the mass of electron is very small.
- iii) In 1909, Millikan measured the magnitude of the charge of an electron.
- iv) Studies on X-rays, discovered in 1895, resulted in the discovery of radioactivity.
- v) Rutherford's experiments on radioactivity proved emission of α -particles besides electrons in radioactive radiations. Thus one more particle was discovered.
- vi) In the 19th century, scientists were trying to measure wavelengths of radiations emitted by different gases filled in discharge tubes using the diffraction grating discovered by Henry Rowland. These wavelengths were found to be discrete and dependent on the type of gas filled in the discharge tube.
- vii) Same time Max Planck presented the photon theory and showed that the black body radiation is discrete. Einstein explained photoelectric effect using photon theory of light for which he received the Nobel prize.
- viii) In 1902, J. J. Thomson presented an atomic model according to 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. Hence, 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 electrical neutrality of atom. But, the electrons embedded in the uniform distribution of positive charges should experience a force towards the centre of the atom directly proportional to the distance from the centre. Hence, they should perform either SHM or uniform circular motion. As both these motions are accelerated, electrons should emit continuous radiation according to the electromagnetic theory of Maxwell. This made it difficult to understand the emission of discrete wavelengths from atoms. Also such a model of atom cannot form a stable structure. To overcome these problems, Thomson assumed the charges to remain stationary unless disturbed from outside and thought about different arrangement of electrons in different atoms. He also estimated the size of atoms to be of the order of 10^{-8} cm from the wavelengths of radiations emitted. Despite all these efforts, he could not explain why the radiation consisted of discrete wavelengths.

- ix) In 1906, Rutherford observed that α -particles passing through a slit provided in the chamber and incident on a photographic plate do not give sharp image of the slit. But on evacuating the chamber, the image became sharp. From this, he concluded that the α -particles must be scattered by the air particles in the chamber.

13.1 Geiger-Marsden's experiment and Rutherford's model of atom

The schematic diagram of Geiger-Marsden's experimental arrangement is shown in the figure on the next page.

S is ${}_{83}^{214}\text{Bi}$ source emitting α -particles of energy 5.5 MeV placed inside a thick block of lead. α -particles coming out of the slit in the block were scattered when incident on a thin foil (F) of gold of thickness 2.1×10^{-7} m. Scintillations were observed on a screen E of zinc sulphide mounted on a microscope M. The whole arrangement was enclosed in a cylinder having thick walls. The cylinder was mounted on a thick disc which could be rotated by the arrangement shown by T. Source S and foil F were kept steady on the base but could be rotated around microscope M. Whenever any charged particle strikes at any point on ZnS, a bright spot is formed at that point. By counting such spots, the number of α -particles striking in 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:

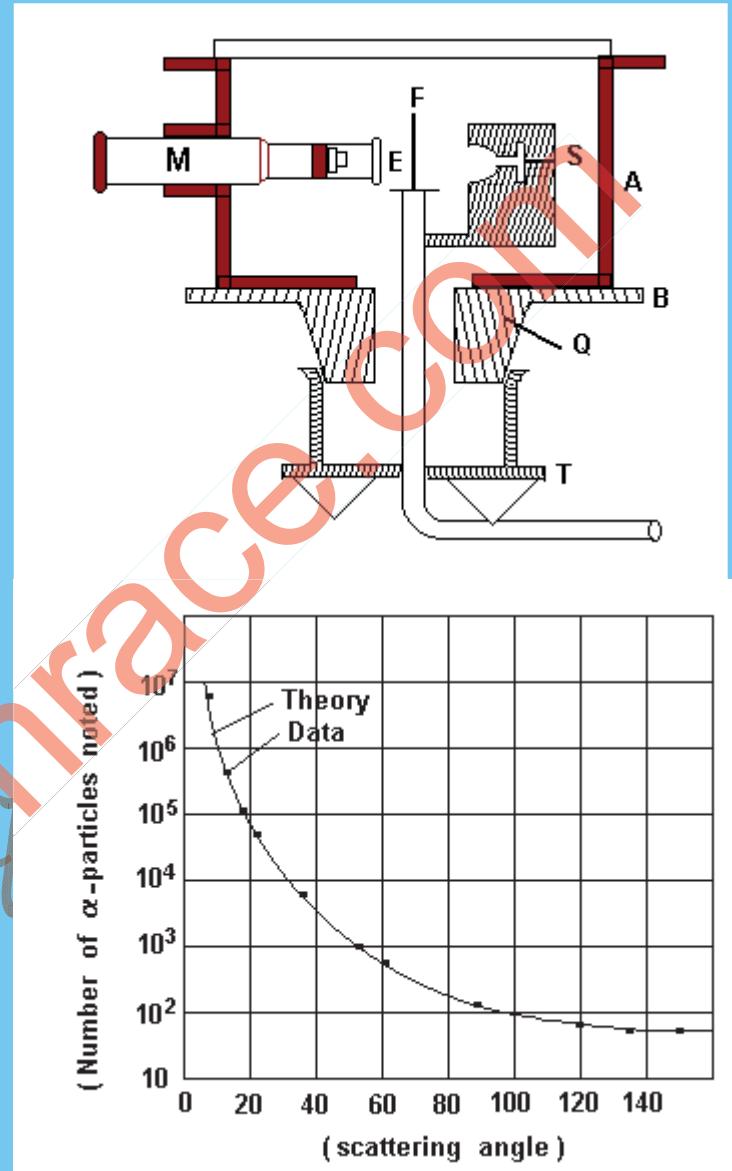
Rutherford suggested that α -particles can undergo only one scattering while passing through the foil as it is very thin. An atom of Gold remains almost stationary during the collision as it is about 50 times heavier than an α -particle. Newton's laws can be used to find the trajectory of α -particles. Rutherford reasoned that the scattering of α -particles at large angles indicate that total positive charge and total mass of the atom must be concentrated in a very small central region of the atom which he called nucleus.

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

$$F = \frac{1}{4\pi\epsilon_0} \frac{(2e)(Ze)}{r^2}, \quad \text{where } 2e = \text{charge of } \alpha\text{-particle}, \quad (e = \text{electronic charge})$$

$$Ze = \text{charge in gold nucleus } (Z = 79 \text{ for gold})$$

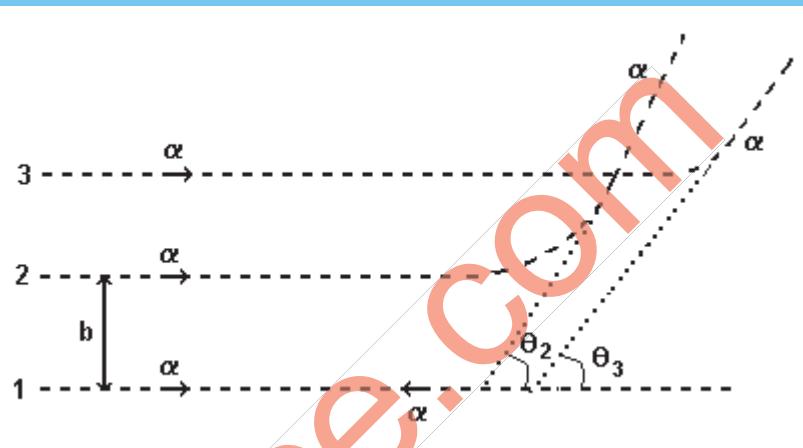
Obviously, this trajectory depends on the initial perpendicular distance of its velocity vector from the nucleus. The minimum perpendicular distance is called the distance of closest approach or impact parameter.



For different impact parameters, the trajectories can be calculated. Such trajectories are shown in the following figure.

It can be seen from the figure that larger the value of the impact parameter, b , smaller is the angle of scattering.

For curve 1, the impact parameter is zero and hence the collision is head on. So the scattering is very large. Calculations in such a case show that the α -particle can go nearest to the nucleus at a distance 10^{-15} m. Hence



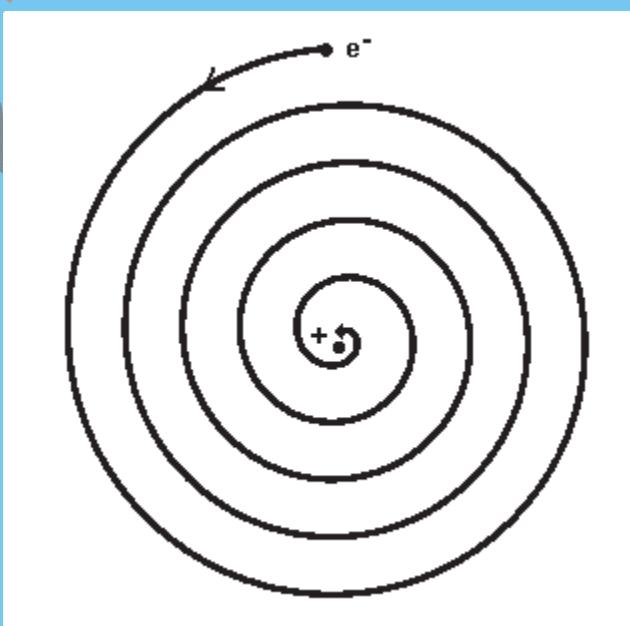
Rutherford concluded that the radius of the nucleus must also be of the order of 10^{-15} m. This is much smaller as compared to the diameter of the atom which is nearly 10^{-10} m.

Rutherford also derived the equation of number of α -particles scattered at different angles. The graph drawn using this equation matched very well with the experimental results of Geiger-Marsden as shown on the previous page.

Rutherford's Atomic Model:

Rutherford proposed that the entire positive charge of an atom resides in a very small region at its centre, where almost all of its mass is concentrated. The negatively charged electrons move around this small central region, called the nucleus in circular orbits.

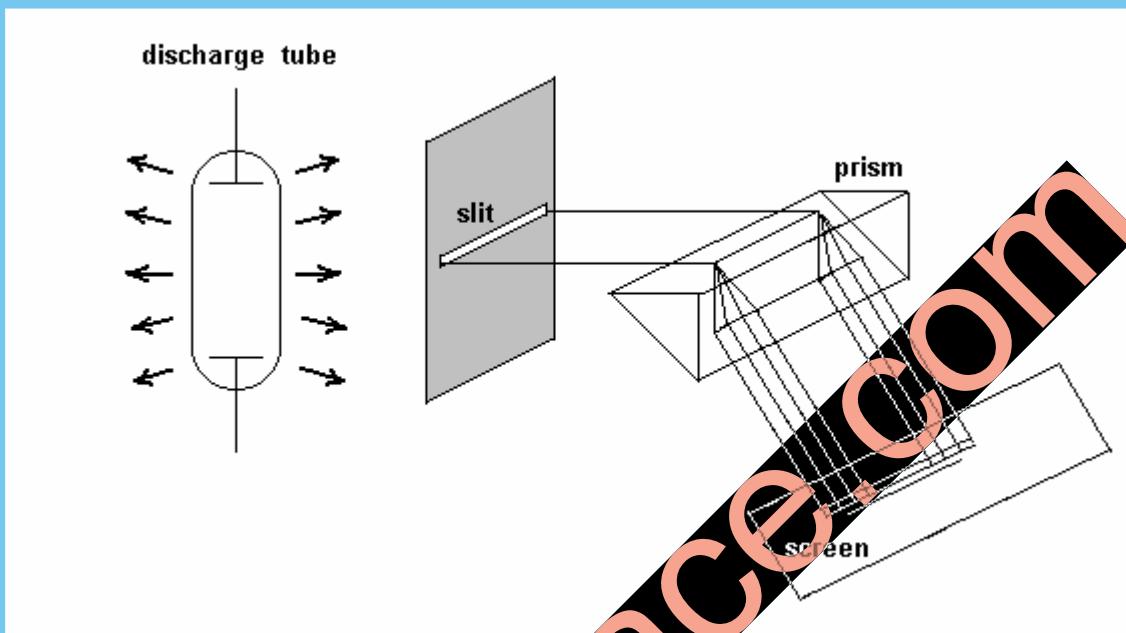
According to classical mechanics, An electron can revolve around the nucleus in any radius depending on its energy. But such a circular motion being accelerated, radiates energy in the form of electromagnetic radiation. As it loses energy, its orbit radius keeps on decreasing resulting in its motion being spiral as shown in the figure and terminating in the nucleus. In this case, the atom cannot remain stable. Thus model failed to explain the stability of the atom.



13.2 Atomic Spectra

On passing an electrical discharge through a tube containing some atomic gas at a low pressure, atoms of the gas get excited and emit radiations consisting of some definite wavelengths characteristic of the nature of the element. The spectrum of such radiations can be obtained and the corresponding wavelengths can be measured with the help of the arrangement shown in the figure on the next page.

It was established that specific groups of the lines of the spectrum can be formed according to their frequencies or wavelengths. In any such group, the wavelengths of the spectral lines can be calculated by a common formula:



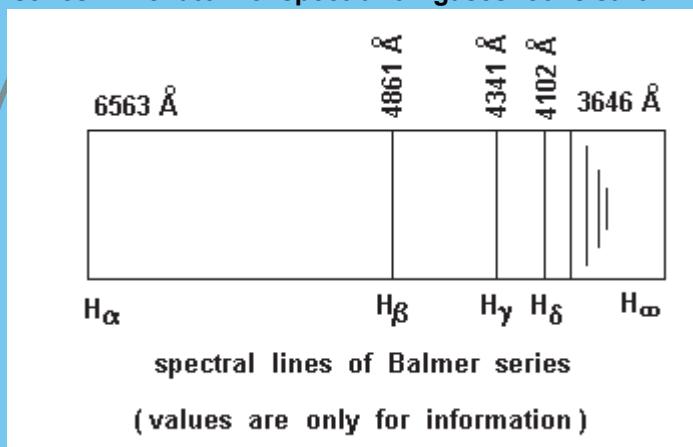
$$\frac{1}{\lambda} = R \left(\frac{1}{m^2} - \frac{1}{n^2} \right) \text{ where } R \text{ is the Rydberg's constant whose value is } 1.097 \times 10^7 \text{ m}^{-1}.$$

The spectral lines in a group form a spectral series. The atomic spectra of gases consist of several such spectral series.

In 1885, first such series was discovered in the visible region by Balmer for hydrogen spectra which is called Balmer series.

The lines of the Balmer series of the hydrogen spectrum with the wavelengths corresponding to each line along with their names are shown in the figure. The wavelengths of the lines in this series are given by the formula

$$\frac{1}{\lambda} = R \left(\frac{1}{2^2} - \frac{1}{n^2} \right) \text{ where } n = 3, 4, 5, \dots$$



The wavelength corresponding to H_α line is obtained by taking $n = 3$, H_β line by taking $n = 4$ and so on. Four other series in the hydrogen spectra in the infrared and the ultraviolet regions of the electromagnetic spectrum are as under.

$$\frac{1}{\lambda} = R \left(\frac{1}{m^2} - \frac{1}{n^2} \right) \text{ where } n > m, \quad n = m+1, \quad m+2, \quad m+3, \quad \dots \text{ Etc.}$$

with $m = 1$ gives Lyman series in the ultraviolet region,

$m = 3$ gives Paschen ")

$m = 4$ " Brackett ") all in the infrared region.

$m = 5$ " Pfund ")

13.3 Energy quantization

The equations of spectral series indicate that there is something discrete in atoms. To understand it, consider the following example.

Suppose an electron is in a one dimensional impenetrable box of length L as shown in the figure. The electron may be anywhere between $x = 0$ to $x = L$.



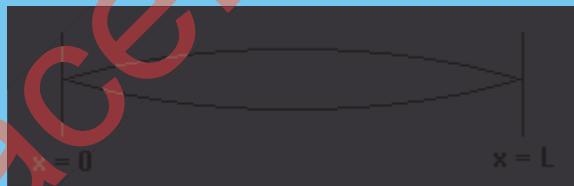
For electron behaving like a wave, it is difficult to find its exact location. The square of the absolute value of the wave function representing the electron represents the probability of the presence of electron in unit length at that point. So the probability of the electron to be at the walls is zero as it is impenetrable. These are the boundary conditions of the electron-wave equation. One such possible wave is shown in the figure.

If λ is the wavelength, $\lambda/2 = L$,

$$\therefore \lambda = 2L \dots \dots \dots \quad (1)$$

But according to de Broglie's hypothesis,

$$\lambda = h/p \dots \dots \dots \quad (2)$$



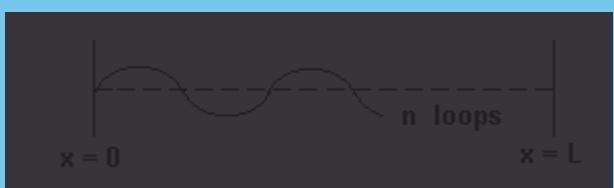
where, h = Planck's constant and p = the momentum of electron.

From equations (1) and (2), $h/p = 2L \quad \blacksquare \quad p = h/2L$.

$$\therefore \text{energy of the electron is } E = \frac{p^2}{2m} = \frac{h^2}{8mL^2}$$

The state of electron represented by the wave as above is called its quantum state one. Here, the probability of electron to be present at the middle of the box is maximum.

Now consider an electron wave of n loops as shown in the adjoining figure. In this case wavelength of the electron is



$$\frac{n\lambda_n}{2} = L \quad \therefore \lambda_n = \frac{2L}{n}$$

$$\blacksquare \quad \frac{h}{p_n} = \frac{2L}{n} \quad \therefore p_n = \frac{nh}{2L}, \quad n = 1, 2, 3, \dots$$

$$\blacksquare \quad \text{in general, energy of the electron, } E_n = \frac{p_n^2}{2m} = \frac{n^2 h^2}{8mL^2}$$

The above equations show that the electron can possess only discrete values of linear momentum such as, $p_1 = \frac{h}{2L}$, $p_2 = \frac{2h}{2L}$, This is known as quantization of the momentum of electron. Also, as the electron can be anywhere between $x = 0$ to $x = L$, uncertainty in its position is $\Delta x = L$. If its momentum, p_n , is considered as uncertainty, Δp , in its momentum, then $\Delta p \sim p_n = \frac{nh}{2L}$

$$\therefore (\Delta x)(\Delta p) = \frac{(L)(nh)}{2L} = \frac{nh}{2}$$

(This equation is similar to Heisenberg's uncertainty principle.)

Similarly, energy is also quantized as

$$E_1 = \frac{h^2}{8mL^2}, \quad E_2 = \frac{4h^2}{8mL^2}, \quad E_3 = \frac{9h^2}{8mL^2}, \quad \dots \dots$$

This means that the electrons can possess only these values of energy.

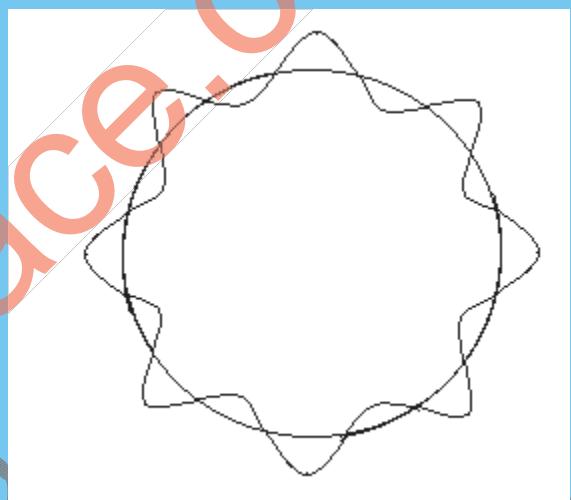
Now consider an electron moving on a circular orbit of radius, r , in a plane. Here also the electron acts as a wave moving in its orbit. If there are n waves on its circumference, (Refer to the figure)

$$2\pi r = n\lambda = \frac{nh}{p} \quad \left(\because \lambda = \frac{h}{p} \right)$$

\therefore angular momentum of the electron,

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

Thus, angular momentum of the electron is also quantized. This fact was presented by Bohr in his hypothesis.



13.4 Bohr Model

In 1913, Neils Bohr gave the following two hypotheses to explain the structure of an atom, its stability and its spectra.

Hypothesis 1:

Of all the orbits permitted by the classical physics, an electron can revolve around the nucleus only in those orbits in which its orbital angular momentum is an integer multiple of $\frac{h}{2\pi}$. The electron can move steadily in such orbits and hence they are called stationary orbits. The electron in a stable orbit does not radiate energy.

Here, h is Planck's constant and its value is 6.625×10^{-34} Js.

Hypothesis 2:

When an electron makes a transition from a stable energy orbit with energy E_i to another stable orbit with a lower energy E_k , it radiates $E_i - E_k$ amount of energy in the form of electromagnetic radiation of frequency f such that $E_i - E_k = hf$. Similarly, when an electron absorbs a quantum of frequency f , it makes a transition which is reverse of the above mentioned transition.

Suppose 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 on the next page.

The centripetal force for the circular motion of the electron is provided by the Coulomb force of attraction between the negatively charged electron and the positively charged nucleus.

$$\therefore \text{Centripetal force} = \frac{mv^2}{r} = \frac{1}{4\pi\epsilon_0} \frac{Ze^2}{r^2} \dots (1)$$

where, ϵ_0 = permittivity of free space

Using the first hypothesis of Bohr, the angular momentum of the electron in an orbit is

$$mv r = \frac{n\hbar}{2\pi} \quad \therefore m^2 v^2 r^2 = \frac{n^2 \hbar^2}{4\pi^2} \dots \dots \dots \dots (2)$$

where, $n = 1, 2, 3, \dots$ is the principal quantum number.

Eliminating v^2 from equations (1) and (2),

$$r = \frac{n^2 \hbar^2 \epsilon_0}{\pi Ze^2 m} \dots (3)$$

Now the total energy of the electron in this orbit is

$E_n = \text{Kinetic energy} + \text{Potential energy}$

$$= \frac{1}{2} mv^2 - \frac{1}{4\pi\epsilon_0} \frac{Ze^2}{r}$$

$$\text{From equation (1), we have } \frac{1}{2} mv^2 = \frac{1}{8\pi\epsilon_0} \frac{Ze^2}{r}$$

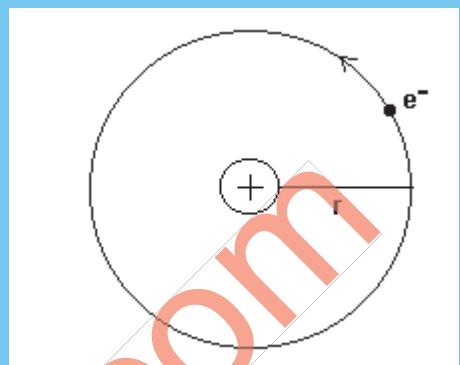
$$\therefore E_n = \frac{1}{8\pi\epsilon_0} \frac{Ze^2}{r} - \frac{1}{4\pi\epsilon_0} \frac{Ze^2}{r} = - \frac{1}{8\pi\epsilon_0} \frac{Ze^2}{r}$$

Substituting the value of r from equation (3) in the above equation,

$$E_n = \frac{1}{8\pi\epsilon_0} \frac{Ze^2 \cdot \pi Z e^2 m}{n^2 h^2 \epsilon_0} = - \frac{Z^2 e^4 m}{8\epsilon_0^2 n^2 h^2}$$

For hydrogen, $Z = 1$,

$$\therefore E_n = - \frac{m e^4}{8\epsilon_0^2 h^2} \frac{1}{n^2} = - \frac{21.76 \times 10^{-19}}{n^2} \text{ J} = - \frac{13.6}{n^2} \text{ eV} \dots \dots \dots (4)$$



Using this equation, the energy of an electron in different orbits can be calculated for the successive integral values of the principal quantum number n for a hydrogen atom. The negative sign is due to the fact that the energy of the stationary electron is taken as zero when placed at an infinite distance from the positively charged nucleus (i.e., a free electron).

This means that if positive energy E_n is given to an electron in an orbit with quantum number n , it will become free. In this sense, E_n gives the binding energy of an electron in n th orbit for a hydrogen atom.

The adjoining figure shows the energy level diagram for hydrogen atom for electrons in different orbits.

Electron in an orbit with $n = 1$ has a minimum energy and is said to be in its ground state. The successive energy states with values of $n = 2, 3, 4$, etc. are called the first excited state, the second excited state, etc. respectively.

If an electron makes a transition from higher energy state, E_i (i.e., from $n = n_i$) to lower energy state E_k , then according to the second hypothesis of Bohr,

$$E_i - E_k = h f_{ik}$$

where f_{ik} is the frequency of the radiation emitted when the electron makes this transition. Putting the values of E_i and E_k using equation (4) above in this equation,

$$E_i - E_k = \frac{me^4}{8\epsilon_0^2 h^2} \left(\frac{1}{n_i^2} - \frac{1}{n_k^2} \right) = h f_{ik}$$

$$\therefore f_{ik} = \frac{E_i - E_k}{h} = \frac{me^4}{8\epsilon_0^2 h^2} \left(\frac{1}{n_k^2} - \frac{1}{n_i^2} \right)$$

But wavelength, $\lambda_{ik} = \frac{c}{f_{ik}}$, where c = velocity of light

$$\therefore \frac{1}{\lambda_{ik}} = \frac{me^4}{8\epsilon_0^2 h^3} \left(\frac{1}{n_k^2} - \frac{1}{n_i^2} \right)$$

This equation is similar to the experimental

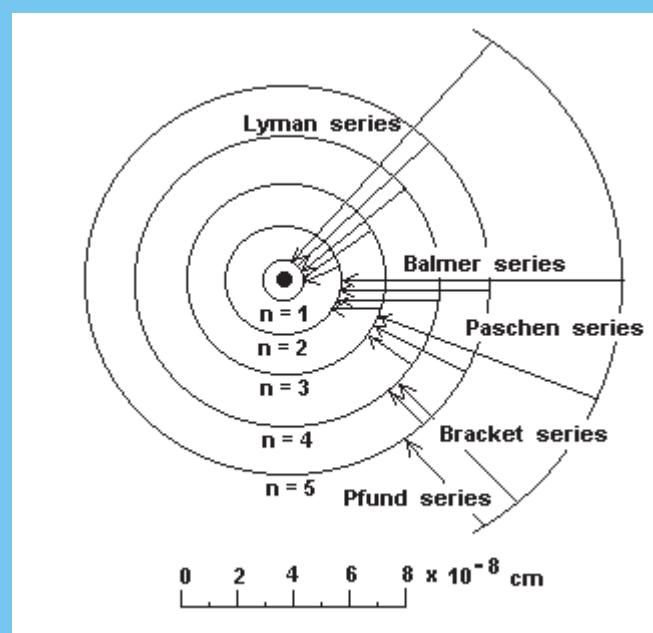
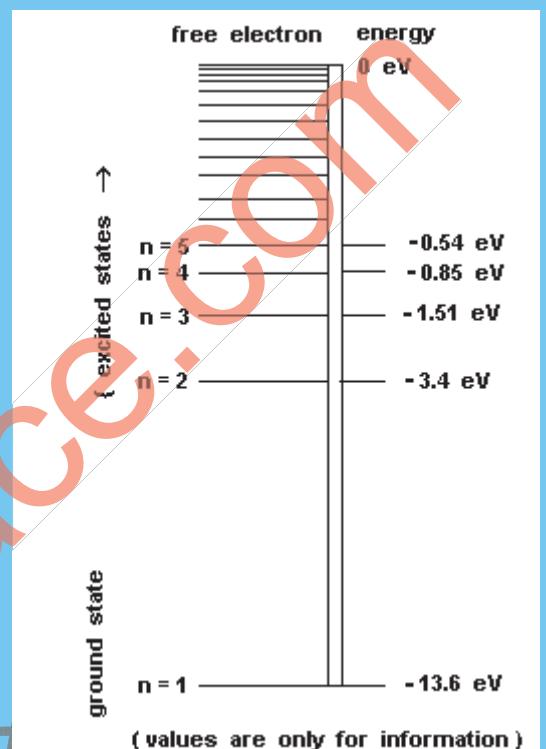
$$\text{result } \frac{1}{\lambda} = R \left(\frac{1}{m^2} - \frac{1}{n^2} \right)$$

The wavelengths of spectral lines of the spectral series can be calculated by taking different values of n_k and n_i . Also the

value of the term $\frac{me^4}{8\epsilon_0^2 ch^3}$ matched very

well with the experimentally obtained value of Rydberg's constant R .

The figure on the right shows transitions of electrons between different levels of energy.



13.5 Limitations of the Bohr model

- (1) On observing hydrogen spectra using a high resolution spectrometer, some more lines are seen which cannot be explained on the basis of the Bohr model.
- (2) Relative intensities of the spectral lines observed in the actual spectra cannot be explained by the Bohr model.
- (3) Orbits of electron need not be circular as assumed in the Bohr model.
- (4) The calculations of energies in these orbits involve an odd combination of classical mechanics and the quantum principles.

Despite all these limitations, Bohr's attempt in applying the principles of quantum mechanics, which were restricted only to the radiation at that time, to the motion of a particle like electron was commendable.

13.6 Emission and Absorption Spectra

In the discharge tube experiments, when the electrons of the gas are excited to higher energy states, they return to their lower energy states of minimum possible energy in about 10^{-8} s and emit radiation according to the equation $E_i - E_k = h f_{ik}$. The spectrum thus obtained is called emission spectrum. For example, Na atoms experience transitions to two excited states from its normal state 3s and come back to original normal state emitting two (yellow) lines of wavelengths 589.0 nm and 589.6 nm.

Now suppose a radiation of continuous wavelengths is incident on an atomic gas filled in a transparent tube. The atoms of the gas which are in their normal characteristic energy states absorb a specific amount of energy needed for them to go from the ground state to another higher quantum state. They do not absorb more energy than needed. Thus out of the continuous radiation incident, the radiations of only suitable wavelengths are absorbed and dark lines appear in the spectrum corresponding to these wavelengths. Such a spectrum is called an absorption spectrum.

The radiation emitted by the lower layer of photosphere, which is at a higher temperature, is continuous when this radiation passes through the outer layer of photosphere which is at a lower temperature. Radiation of some wavelengths are absorbed and hence dark lines corresponding to these wavelengths called Fraunhofer lines are observed.

[Photosphere is the visible surface of the Sun which is its about 100 km thick outer layer.]

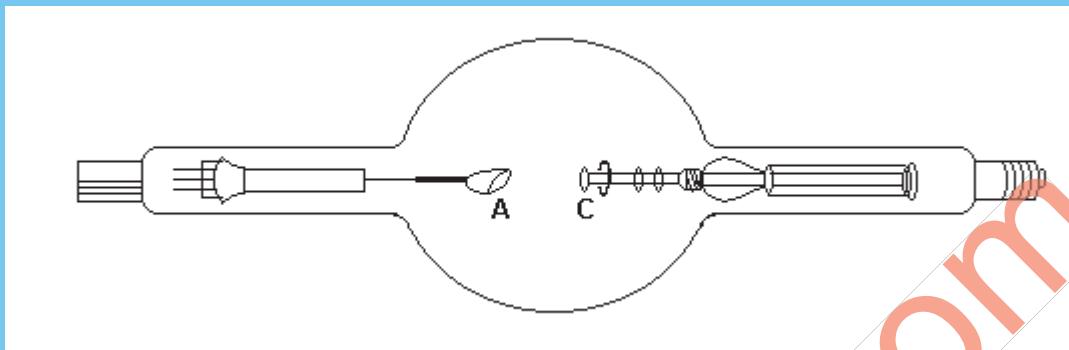
13.7 Many Electron Atoms Study notes of this topic are omitted here as it is not included in the syllabus for the purpose of examination.

13.8 X-rays

Rontgen discovered X-rays in 1895. The wavelengths of X-rays range from .001 to 1 nm. Coolidge designed a special tube for emission of X-rays, the figure of which is shown on the next page.

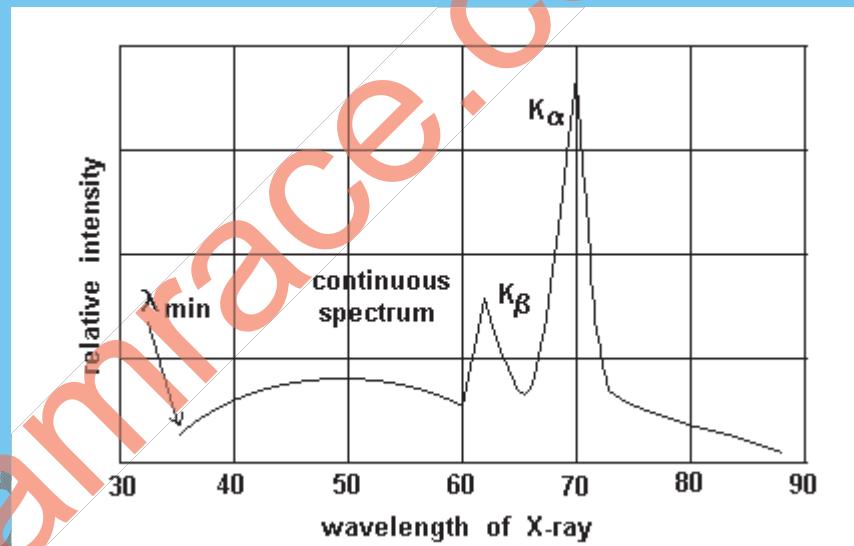
Here C is the cathode. When current is passed through the filament it gets heated and heats the cathode which emits electrons. These electrons are accelerated under a p.d. of 20 to 40 KV and reach the anode. As a result, X-rays are emitted from the surface of the anode.

Normally, the anode is made from a transition element (e.g., Mo).

X-ray spectrum:

The figure shows the graph of wavelengths of X-rays emitted from the Mo target by 3 KeV electrons against the relative intensity. Such a graph is called the X-ray spectrum corresponding to the given element and energy of electrons.

- (1) The graph, starting from some minimum wavelength (λ_{\min}) is continuous.
- (2) The relative intensity is very large for some definite wavelength.
- (3) λ_{\min} is a definite wavelength.



The continuous curve in the graph is called continuous spectrum. The peaks obtained for certain wavelengths indicate line spectrum. This is the characteristic curve of given element.

Explanation of X-ray Spectrum:

Highly energetic electrons collide with atoms of the anode and lose some energy during each collision. Thus they keep losing energy during multiple collisions and emit X-rays of different frequencies which form a continuous spectrum of continuous frequencies (wavelengths).

Explanation of λ_{\min}

When any electron makes head on collision with an atom of the anode, its total kinetic energy gets completely converted into X-rays of maximum frequency (minimum wavelength).

$$\therefore \frac{1}{2} m(v_{\max})^2 = eV = hf = h \frac{c}{\lambda_{\min}} \quad \therefore \lambda_{\min} = \frac{hc}{eV}$$

where, h = Planck's constant = 6.62×10^{-34} J s,

c = velocity of light = 3.0×10^8 m s⁻¹

e = charge of electron = 1.6×10^{-19} C, V = potential difference (V)

Explanation of characteristic X-ray spectrum

Some incident electrons penetrate deep inside the atom and knock off the inner orbit electrons. The electrons from outer shells undergo transition and fill up these vacancies emitting radiations of definite frequencies. The radiation is called K_{α} X-rays when an electron undergoes transition from $n = 2$ (L shell) to $n = 1$ (K shell) and is called K_{β} if the transition of electron is from $n = 3$ to $n = 1$. Thus many lines of X-ray spectrum are obtained and the spectrum so formed is called characteristic spectrum.

Such spectra depend on the type of element of the anode (target) as energies of electrons in K, L, M ... shells in the atoms of different types are also different. Hence, the wavelengths of K_{α} , K_{β} , L_{α} , ... radiations are also different for different elements. That is why such curves are called characteristic spectrum of the given element.

Now from Bohr's atomic model,

$$\text{Energy, } E_n = - \frac{m Z^2 e^4}{8 \epsilon_0^2 h^2 n^2} .$$

In a multi-electron atom, an electron cannot see the complete charge of the nucleus due to other electrons screening the charge of the nucleus and can see only $(Z - 1)e$ charge on the nucleus. Hence taking $Z = Z - 1$ in the above equation,

$$E_n = - \frac{m(Z - 1)^2 e^4}{8 \epsilon_0^2 h^2 n^2} = - \frac{13.6(Z - 1)^2}{n^2} \text{ eV} \quad \left[\because \frac{m e^4}{8 \epsilon_0^2 h^2} = 13.6 \text{ eV} \right]$$

To calculate the frequency of K_{α} radiation of a target of atomic number Z,

$$E_2 - E_1 = hf = 13.6(Z - 1)^2 \left[\frac{1}{1^2} - \frac{1}{2^2} \right] \times 1.6 \times 10^{-19} \text{ J} \quad (\because 1 \text{ eV} = 1.6 \times 10^{-19} \text{ J})$$

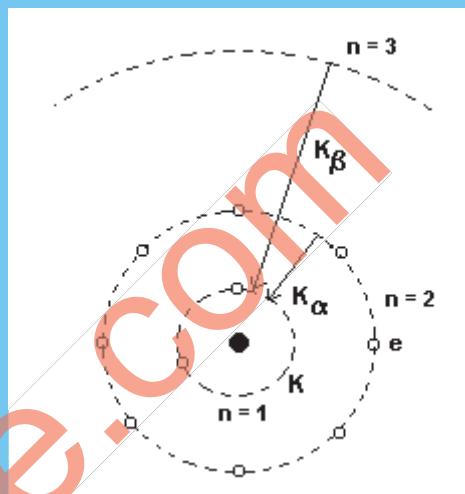
$$\sqrt{f} = \left(\frac{13.6 \times 1.6 \times 10^{-19} \times 3}{6.62 \times 10^{-34} \times 4} \right)^{\frac{1}{2}} (Z - 1)$$

$$\sqrt{f} = C Z - C \quad \text{where, } C = 4.965 \times 10^7 \text{ unit}$$

This equation represents a line, i.e., the graph of \sqrt{f} versus Z is a straight line.

In 1913, Moseley used a specially designed X-ray tube with a series of targets to obtain their characteristic spectra. The \sqrt{f} versus Z graph (next page) shows K_{α} lines of 21 different elements.

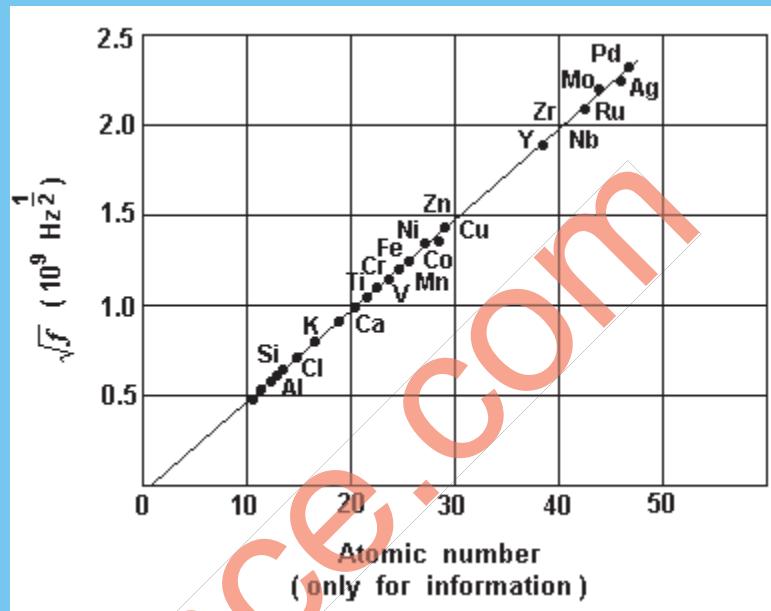
The following points emphasize the scientific importance of Moseley's work.



(1) During Moseley's time, different elements were arranged in the periodic table based on their atomic masses. This seemed inappropriate with respect to their chemical properties. Moseley suggested arranging elements according to their atomic numbers which was found to be proper with respect to the chemical properties of the elements.

(2) At that time, some places in the periodic table were missing. Atomic numbers Z of such missing elements could be decided from Moseley's work and such missing positions could be filled with appropriate elements. The chemical properties of Lanthanides (or rare earth elements) were found to be very similar. Moseley's work was useful in deciding their positions with certainty.

- (3) Positions of elements coming after uranium could be fixed in the periodic table after obtaining their atomic spectra.
- (4) K_{α} X-radiation associated with $n = 1$ shell helped obtain charge of the nucleus.
- (5) Ordinary emission or absorption spectra associated with the transition of the valence electrons were not useful in obtaining the charge of the nucleus.

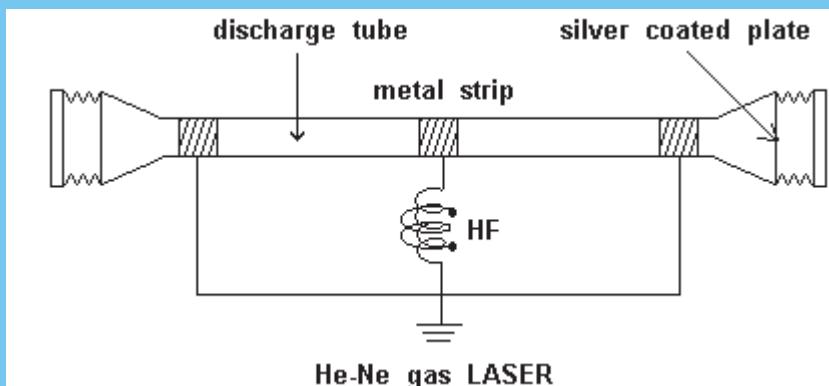


13.9 LASER

LASER means "Light Amplification by Stimulated Emission of Radiation" which represents the process occurring in the Laser device. A schematic diagram of He-Ne gas LASER is shown below.

Construction:

In a He Ne gas LASER, about 1 m long glass discharge tube is used. The tube is filled with He at a partial pressure of 1 mm of Hg and Ne at a partial pressure of 0.1 mm of Hg. At both ends of the tube, two polished silver coated plates are fixed parallel to each other such that no gas leakage occurs. A high p.d. of high frequency is applied with the help of three metal strips provided on the tube lying outside it. Tesla coils can be used for this purpose.



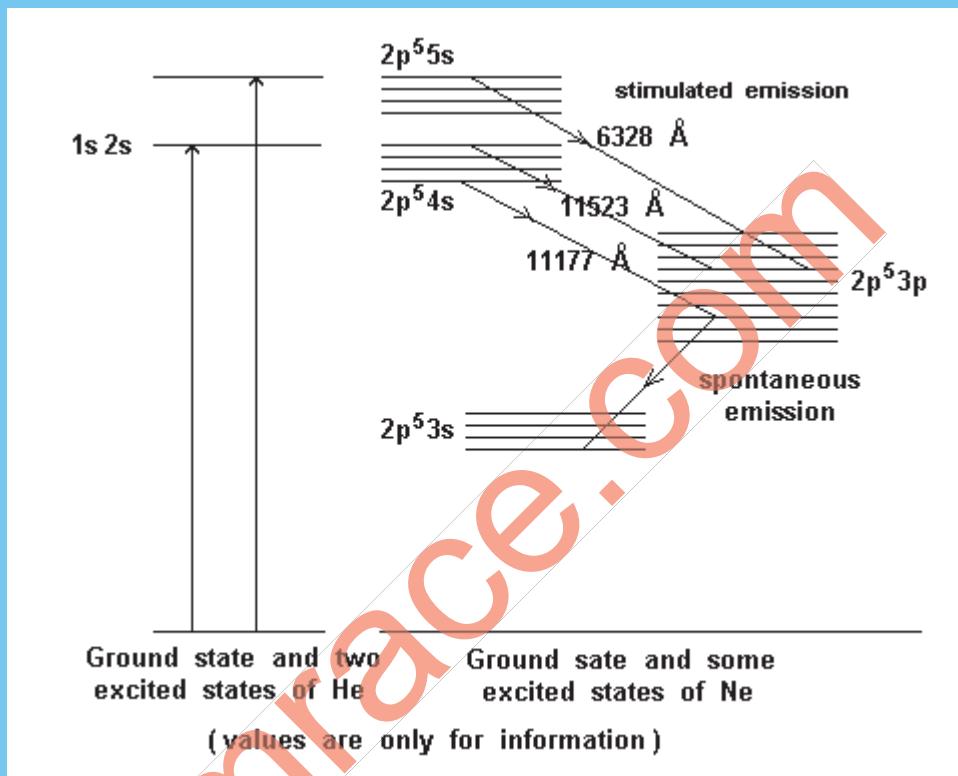
Working:

The ground state and two excited states of He are shown in the figure. In the normal state, the number of atoms in the ground state, N_0 , having energy, E_0 , will be much more than the number of atoms, N_x , having energy, E_x , in the excited state.

$$N_x = N_0 e^{(E_x - E_0)/kT}$$

where, k = Boltzmann constant
and T = temperature

When one electron of Ne from its $2p^6$ orbit goes to $3s$, $3p$, $4s$ or $5s$, the number of excited states obtained are 4, 10, 4 and 4 respectively as shown in the figure.



During discharge in the tube, when the electrons collide with He atoms, one electron of He goes to $2s$ and acquires one of the two excited states. Thus, He atoms are in two excited states as shown in the figure. These two states are meta stable states of He where electrons can stay for a long time. The process of exciting electrons from the ground state to the meta stable state is called optical pumping.

From an ordinary excited state, the electrons return to their states of minimum energy in about 10^{-8} s and emit radiations. Such transitions are called spontaneous transitions and radiations are called spontaneous emission. These are shown in the figure. When the excited He atoms collide with Ne atoms, they knock out Ne atoms from their ground state to the uppermost states $2p^5 5s$ and $2p^5 4s$. The remaining energy appears in the form of kinetic energy of atoms. After the collision, He atoms return to the ground state whereas the population of Ne atoms in aforesaid states goes on increasing. This phenomenon is called population inversion.

In this situation, if photons of appropriate frequency are incident on the excited Ne atoms, two photons of frequency same as that of the incident photon are emitted. This phenomenon is called stimulated emission. Initially, the stimulated emission of radiation occurs due to the transition from $2p^5 5s$ and $2p^5 4s$ to $2p^5 3p$ and then from $2p^5 3p$ to $2p^5 3s$. The wavelengths emitted are shown in the figure.

If two photons obtained from one photon are confined to the system by reflectors instead of letting them escape, they will induce emission of even more number of photons increasing their number considerably. These photons can then be taken out of the system suitable to obtain a very narrow, highly parallel and intense beam of light called LASER beam. The waves become parallel and are in phase due to multiple reflections. The one foot diameter of the LASER beam on Earth becomes no more than a mile on reaching the moon.

Properties of LASER light:

- (1) Highly monochromatic.
- (2) Highly coherent.
- (3) Highly parallel.
- (4) Can be focused sharply.

Uses of LASER light:

- (1) Used to bore holes
- (2) Used in long distance surveying.
- (3) Small lasers (\approx) used in optical fibres.
- (4) Large lasers used in nuclear fusion.
- (5) Used in the medical field for retina detachment surgery, blood vessel cut, etc.

MASER

MASER means “Microwave Amplification by Stimulated Emission of Radiation”. In such devices, microwaves are amplified. MASER using NH_3 molecules is described here.

The molecules of ammonia are of two types: (i) N atom above the plane formed by the three hydrogen atoms and (ii) N atom below the plane formed by the three hydrogen atoms. This phenomenon is known as the inversion of the ammonia molecule.

Each oscillatory and rotatory layers of the ammonia molecule divide into two types of layers. Because of these layers, a characteristic behaviour of the molecules is observed in an electric field.

When a beam of molecules of the two above mentioned types is passed through a special type of electric field, the molecules of down position get thrown out of the beam. The molecules of the second type are allowed to enter a cavity where their two layers are tuned with their corresponding frequencies. As a result, an amplified radiation is obtained due to transition from the upper to the lower layer.