Chapter-18.

THE ATOMIC SPECTRA

Introduction

The subject of atomic spectra deals with the measurement of the wavelengths and intensities of the electromagnetic radiations emitted or absorbed by atoms. A typical arrangement for observing emission in atomic spectra is shown in Fig.18.1(a). The source S may be an electric arc or spark or an electric discharge passing through a monatomic gas or a heated salt of the material. The light emitted by the source after it passes through a system of lenses and collimating slits, falls on a prism (or diffraction grating). The prism disperses the radiation and different wavelengths are recorded at different points on the photographic plate. The impression on the photographic plate appears as lines (because of the rectangular slit in front of the source) corresponding to different wave lengths. It is because of this appearance that the spectrum obtained is called a line spectrum. Line spectra are typical of atoms and each kind of atom has its own characteristic spectrum. The line spectrum contains a series of lines in the visible region of the spectrum as shown in the fig. 18.1 (a).

We mentioned about the emission spectra of atoms. One may observe absorption spectra of atoms by using the arrangement shown in Fig.18.1(b). In this case, light containing all wavelengths is made to pass through a tube containing the gaseous state of the element under investigation. The light coming out is analysed as before. The spectrum is now consisted of dark lines against a bright back ground.

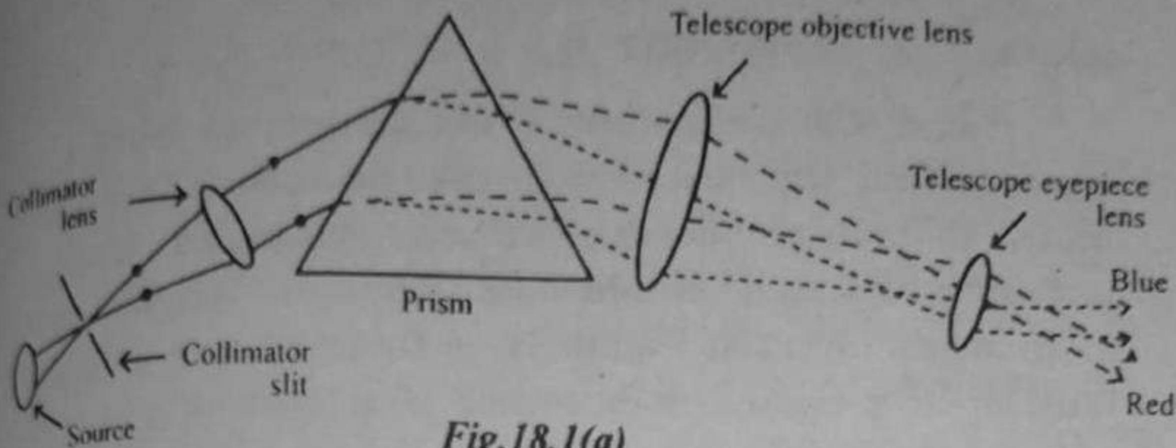

Fig.18.1(a)

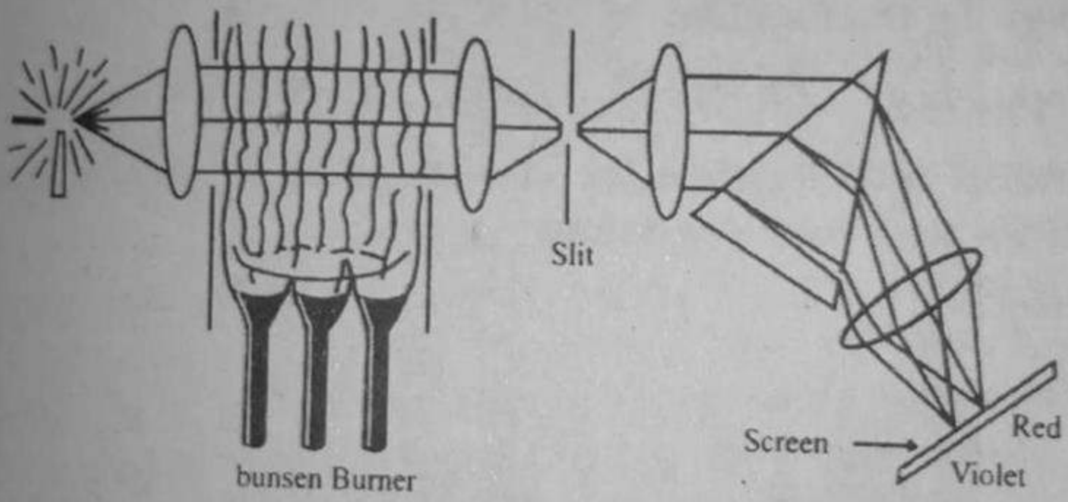

Fig.18.1(b)

18.1. The Spectrum of Hydrogen Atom.

Experimental spectroscopy has been studied since the middle of the nineteenth century, and it was known that atomic spectra of gases consisted of sharp lines of defining frequency. Some atoms had very complicated spectra, though that of hydrogen was simple. The visible spectrum of hydrogen was found to consist of a series of lines shown in Fig. 18.2. Spectroscopists Balmer and Rydberg had been able to fit the frequencies of hydrogen.

Balmer's formula for hydrogen is

Frequency $I = CR\,(1/p^2 - 1/n^2)$ .... (18.1)

where $p = 2$ and $n = p+1,\ p+2,\ p+3\ \ldots\ldots$

and R is a constant. From the relation eq:18.1, it appears that the observed frequencies could be written as the differences between two term values, i.e. R/p^2 is one term and R/n^2 would be other term. The first term in Balmer's formula, Eq.(18.1), is that with $p = 2$. It was natural to suspect the existence of a term for $p = 1$. In 1906 such a term was in fact found by Lyman. He found a series of lines in the hydrogen spectrum in the far ultraviolet, known as the Lyman series, with frequencies given by the formula:

Frequency $CR\ (1/1^2 - 1/n^2)$ .... $n = 2, 3, 4,$------ (18.2).

And in 1908 Paschen found a series in the infrared given by the formula by taking $p = 3$.

Frequency $I = CR\ (1/3^2 - 1/n^2)$. $n = 4, 5, 6.$ ......----- (18.3)

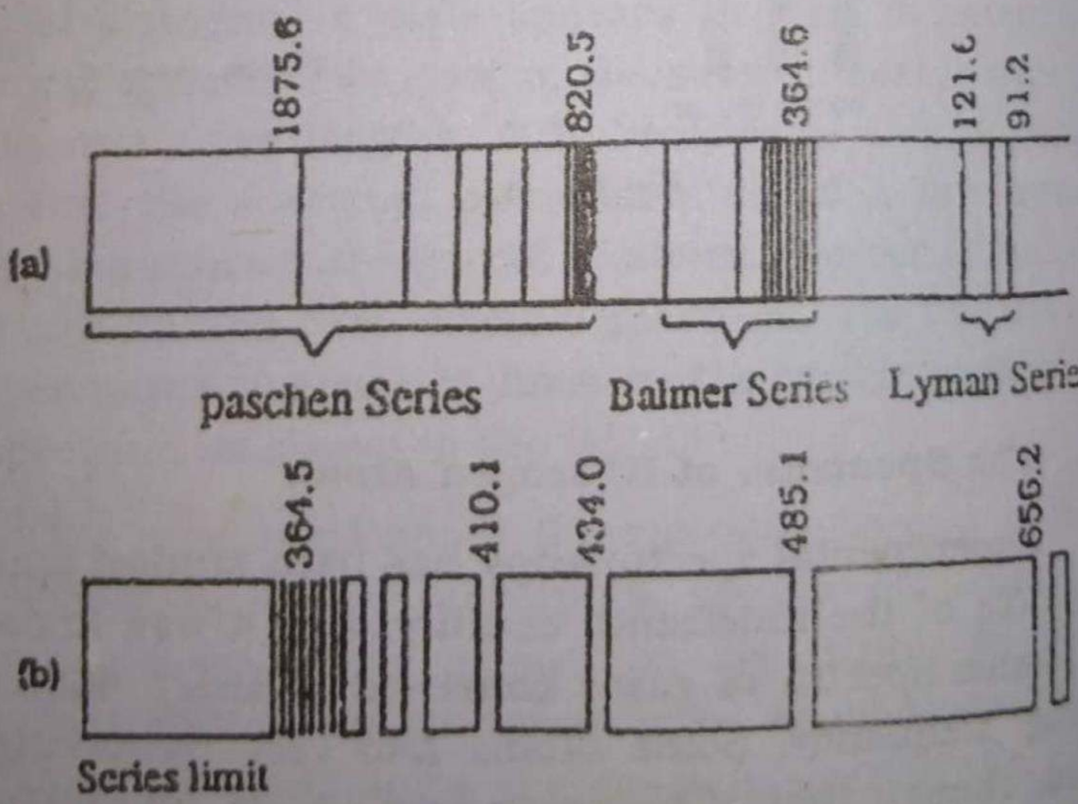

Fig. 18.2

Fig. (18.2) shows these series.

At the end of the first decade of the twentieth century, the significant information at the disposal of Bohr was that the atom radiates fixed frequencies in the form of sharp spectral lines, and have a permanent existence. But this information seemed to be contradicting the picture of the atom proposed by Rutherford (1911) to explain experiments on the scattering of α - particles. According to this model, the atom was assumed to consist of a massive positively charged nucleus surrounded by electrons in orbits of radius of the order of 10^{-10} m. According to classical mechanics, however, an electron moving in the coulomb field of the nucleus would radiate electromagnetic energy at the frequency of its orbital motion. Since no energy is provided to the atom from the outside source so we expect the electrons to lose energy and slowly spiral into the nucleus, emitting radiation of continuously increasing frequency. But real atoms are observed to be stable in their normal state and radiate only certain discrete frequencies when excited. Bohr in 1913 undertook to solve this paradox.

18.2. Bohr's Model for the Hydrogen Atom

In order to develop a quantitative theory for the spectrum of the hydrogen atom, Bohr put forward the following postulates.

i) An electron moves only in those circular orbits for which its orbital angular momentum L is an integral multiple of $h/2\pi$.

ii) The total energy of the electron remains constant as long as it remains in the same orbit.

iii) If the electron jumps from an initial orbit of energy E_i to final orbit of energy E_f ($E_i > E_f$), a photon of frequency is emitted.

$$\upsilon = \frac{E_i - E_f}{h} \qquad \text{------------------ (18.4.)}$$

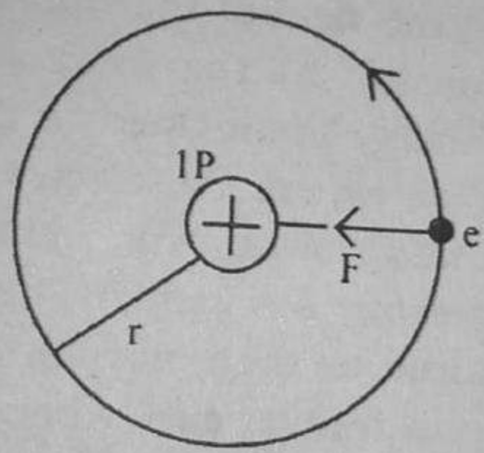

Fig.18.3.

According to Bohr's postulates the hydrogen atom consists of nucleus containing a proton and the electron revolving around the nucleus in definite circular orbits (Fig.18.3).

Since the proton is considered to be stationary being massive and the electron is attracted to it with a force given by Coulomb's law, the attractive force has a magnitude:

$$F = \frac{1}{4\pi\epsilon_o} \frac{e^2}{r^2}$$

where each of the two particles has a charge e and ϵ_0 is the permittivity constant and $1/4\pi\epsilon_0$ is given to be $9.0 \times 10^{+9}$ N m^2 C^{-2}. The electron is revolving in a circular orbit of radius r with the velocity v, the Coulomb attractive force is balanced by the centripetal force $\frac{mv^2}{r}$

Therefore

$$\frac{1}{4\pi\epsilon_0} \frac{e^2}{r^2} = mv^2/r \quad \text{------------------ (18.5)}$$

The electron revolving around the nucleus does not radiate energy as given by classical electromagnetic theory but instead it is assumed that the total energy of the atom remains constant.

The total energy of the atom is the sum of the kinetic energy and the potential energy of the electron. The kinetic energy T of the electron from Eq.(18.5.) is given

$$T = 1/2\ mv^2$$

$$= \frac{e^2}{(4\pi\epsilon_0)2r} \quad \text{(from 18.5)} \quad \text{------------------ (18.6)}$$

The potential energy V of the proton-electron system also depends on r. We take the potential energy to be zero when the electron is infinitely distant from the nucleus. Thus, from electrostatic theory

$$V = -\frac{e^2}{4\pi\epsilon_0 r} \quad \text{------------------ (18.7)}$$

The potential energy is negative as the Coulomb force is attractive. The total energy of the system is, therefore.

$$E = T + V$$

$$E = -\frac{e^2}{8\pi\epsilon_0 r} \quad \text{------------------ (18.8)}$$

$$= -T$$

In order to decide what particular values of radii of circular orbits are permitted, it was assumed that the angular momentum of the electron must be an integral multiple of $h/2\pi$, where h is the same Planck's constant which relates the energy and frequency of a photon. If L is the orbital angular momentum, then according to this assumption,

$$L = n\ h/2\pi \qquad \text{------(18.9)}$$

where n has values 1,2,3, ... ∝. Since the electron of mass m is moving in a circular orbit of radius r with velocity v, then

$$L = m\,v\,r$$

and from Eq.(18.9) $L = n\hbar$ where $\hbar = h/2\pi$

$$\therefore\ mvr = n\hbar, \quad n = 1,2,3,......$$

solving for v and putting in Eq. (18.5), we get

$$r = 4\pi\epsilon_o \frac{n^2\hbar^2}{me^2} \text{ where } n = 1,2,3,\text{----------(18.10)}$$

which gives the radii of the "non radiating" orbits. For the ground state n=1 and

$$r_1 = \frac{4\pi\epsilon_o \hbar^2}{me^2}$$

Substituting the values, we get

$$r_1 = 0.53 \times 10^{-10}\ m = 0.053\ nm$$

which is called the Bohr radius. From Eq. (18.10)

It is seen that the radii are proportional to the square of the integer number n, called the principal quantum number. Now if r in Eq. (18.8) is replaced by Eq. (18.10) we get,

$$E = E_n = -\frac{m e^4}{32\pi^2 \epsilon_0^2 \hbar^2}(1/n^2) \quad \text{--------}(18.11)$$

The only allowed values of the energy are those given by Eq. (18.11) when n takes the values 1,2,3,...... This shows that the energy is quantized.

Substituting the values of various constants in Eq. (18.11), we get,

$$E_n = -\frac{13.6}{n^2}\ \text{eV where } n = 1,2,3 \text{----------} (18.12)$$

The state of the lowest energy or ground state corresponds to n=1, and its energy is–13.6 eV. The energy of the electron corresponding to n = 2,3,4,...... is given by

$E_2 = -13.6/4 = -3.40$ eV

$E_3 = -13.6/9 = -1.51$ eV

$E_4 = -13.6/16 = -0.85$ eV

. = ------------------

. = ------------------

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These energy levels can now be represented graphically in Fig. (18.4). The quantum numbers are shown at the left and the corresponding energies of hydrogen in electron volts are given at the right. Note that all the states from n = 1 to n = ∞ are bound states, since they have negative energies. The separation of the levels decreases towards the top of the diagram and converges to zero for ∞. Above the line given by n = ∞, the energy states have positive energy, E > 0. The system is then unbound, meaning that the electron is free.

Now according to Bohr's postulate 3, if one electron jumps from an initial state n_i (energy E_i)to another state of lower energy n_f (energy E_f), the frequency of the emitted photon is from Bohr's formula:

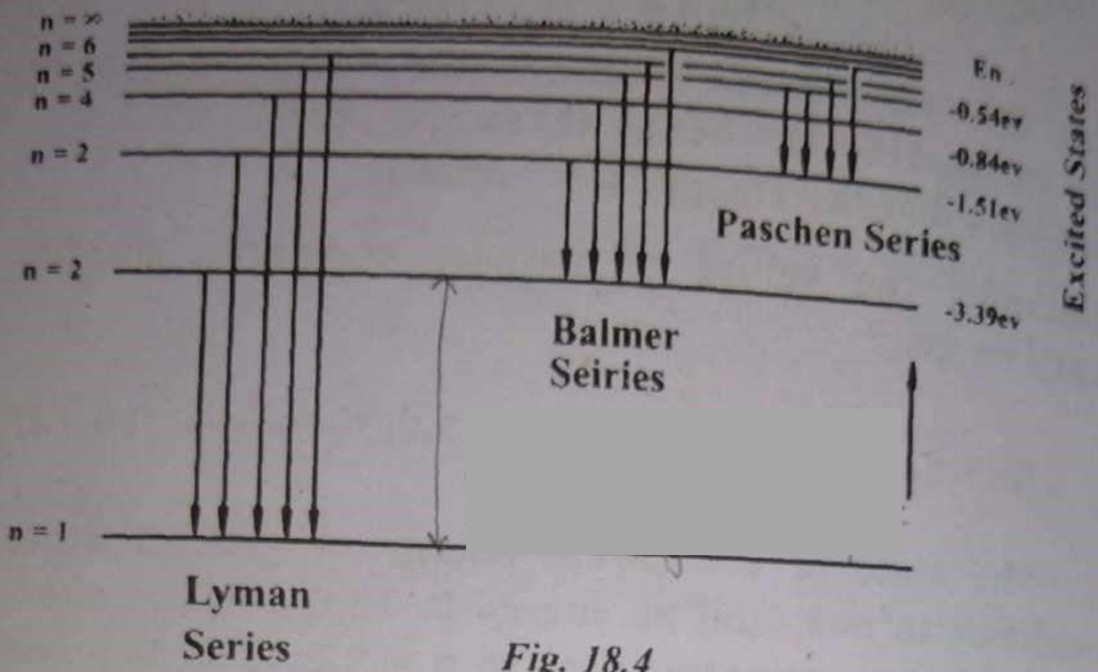


Fig. 18.4

$$\upsilon = \frac{E_i - E_f}{h}$$

$$\frac{E_i - E_f}{2\pi h} \quad (\text{since } \hbar = \frac{h}{2\pi})$$

Substituting for the energies from Eq:(18.8),

we have

$$\upsilon = c/\lambda = \frac{me^4}{64\pi^3 h^3 \epsilon_o^2} \left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

In terms of wave number $\overline{\upsilon}$ or the wavelength λ of the emitted photon is

$$\overline{\upsilon} = \frac{1}{\lambda} = \frac{me^4}{64\pi^3 h^3 \epsilon_o^2 c} \left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right) \quad \text{----(18.13)}$$

Comparing Eq. (18.13) with Eq. (18.1), both have the same form. Therefore from this comparison we find that

$$R_\infty = \frac{me^4}{64\pi^3h^3\epsilon_o^2 c}$$

is the theoretical value of the Rydberg constant.

Eq. (18.13) can now be written as

$$\bar{\upsilon} = 1/\lambda = R\infty\ (1/n^2_f - 1/n^2_i) \text{---------------} (18.14)$$

Hydrogen Series

If we put $n_f = 2$ in Eq: (18.14), we get

$$\bar{\upsilon} = R_\infty\ (1/2^2 - 1/n'^2)$$

where $n' = 3; 4, 5, \ldots$

This relation is identical with the series formula for the Balmer series of the hydrogen atom if $R_\infty = R_H$. The value of $R\infty$ is calculated to be $1.0970 \times 10^7 m^{-1}$ which is in quite good agreement with the experimental value of $R_H = 1.0968 \times 10^7 m^{-1}$. According to Bohr's model, the known series of the hydrogen spectrum arise from transitions in which the electron goes to a certain final quantum state n′. The relations for the different series are:

Lyman series $\bar{\upsilon} = R\infty\ (1/1^2 - 1/n'^2)$ where $n' = 2, 3, 4\ldots$

Balmer series $\bar{\upsilon} = R\infty\ (1/2^2 - 1/n'^2)$ where $n' = 3, 4, 5\ldots$

Paschen series $\bar{\upsilon} = R\infty\ (1/3^2 - 1/n'^2)$ where $n' = 4, 5, 6..$

Brackett series $\bar{\upsilon} = R\infty\ (1/4^2 - 1/n'^2)$ where $n' = 5, 6, 7, 8.$

P'fund series $\bar{\upsilon} = R\infty\ (1/5^2 - 1/n'^2)$ where $n' = 6, 7, 8, \ldots$

As shown in the energy level diagram (Fig.18.4), the region extends beyond $n = \infty$ which corresponds to a state in which the electron is completely removed from the atom and in such a case $E = 0$ and $n = \infty$.

Example 18.1.

An electron in the hydrogen atom makes a transition from the n = 2 energy state to the ground state (corresponding to n = 1). Find the wavelength and frequency of the emitted photon.

Solution:

We can make use of the equation:

$$1/\lambda = R_\infty (1/n_1^2 - 1/n_2^2)$$

$$1/\lambda = R_\infty (1/1^2 - 1/2^2) = 3/4 R$$

$$\lambda = (4/3)\, 1/R_\infty = \frac{4}{3(1.097 \times 10^7 m^{-1})}$$

$$= 1.215 \times 10^{-7} m$$

The wavelength lies in the ultraviolet region.

Since $c = \upsilon\lambda$ so the frequency of the photon is

$$\upsilon = c/\lambda = \frac{3 \times 10^8 ms^{-1}}{1.215 \times 10^{-7} m} = 2.47 \times 10^{15} H_z$$

Example 18.2.

Calculate the binding energy of the hydrogen atom (the energy binding the electron to the nucleus).

Solution:

The binding energy is numerically equal to the energy of the lowest state. The largest negative value of E in equation.

$$E = -\frac{me^4}{8\epsilon_o^2 h^2} \; 1/n^2 \;\; (n = 1,2,3........)$$

is found for n=1. This gives

$$E = -\frac{(9.11 \times 10^{-31} kg)\;\; (1.60 \times 10^{-19} C)^4}{8(8.5 \times 10^{-12} C^2 - N.m^2)^2 \; (6.63 \times 10^{-34} J.S)^2}$$

$= -2.17 \times 10^{18}$ J

$= -13.6$ eV. (Since 1 e V = 1.6×10^{-19} J)

Example 18.3.

Find the shortest wavelength photon emitted in the Balmer series and determine its energy.

Solution:

The shortest wavelength photon in the Balmer series is emitted when the electron makes a transition from $n = \infty$ to n = 2. Therefore

$$1/\lambda_{min} = R\,(1/2^2 - \infty^2) = R/4$$

$$\lambda_{min}\ 4/R = \frac{4}{1.097 \times 10^{-7} m^{-1}} = 364.6 nm.$$

This wavelength is in the ultraviolet region and corresponds to the series limit. The energy of a photon with this wavelength is:

$$E_{photon} = hc/\lambda_{min} = 3.41 eV$$

This is the maximum-energy of photon in this series, since it involves the largest energy change.

18.3. Excitation and Ionization Potential.

The Bohr model predicts that the total energy of an atomic electron is quantized. For example Eq: (18.11) gives the allowed energy values for the electron in a one electron atom. State for n = 1 is said to be in the ground state whereas states with n = 2,3,4,... are called the excited states. If an atoms is supplied energy so that it reaches one of its allowed values, then the atom is said to have gone to one of its excited states. The excitation potential is defined to be the accelerating potential which moves the electron of the atom from the ground state to the higher state Since the atom has discrete states so definite amounts of energy are required to take the electron to the various excited states of the atom.

There are many ways by which the electron of an atom can be excited to different states. If the atomic gas is heated the electron may be excited by thermal motion collision. If the gas is in an electric discharge, a free electron which has been accelerated by the electric field may hit the atom of the gas and excite it to a higher state. The atoms of the gas when illuminated may absorb energy from a photon and are excited.

If an electron lying in the ground state of the atom is given sufficient energy so that it is raised to the orbit for which $n = \infty$, it will disengage itself from the atom.The atom will then become positively charged. It is then said to be ionized. The ionization potential is therefore, defined to be the accelerating potential which removes an electron completely from an atom.

For hydrogen atom the energy needed to ionize it is 13.6 electron volts and the corresponding ionization potential is 13.6 volts.

Example: 18.4.

(a) What is the longest wavelength of light capable of ionizing a hydrogen atom ?

(b) What energy is needed to ionize a hydrogen atom?

Solution:

(a) The wavelength of light capable of ionizing the hydrogen atom will make the electron to raise it from $n = 1$ to $n = \infty$ So using the relation

$$1/\lambda = R\,(1/n^2{}_1 - 1/n^2{}_2)$$

we have

$$1/\lambda = 1.097 \times 10^7\,(1/1 - 1/\infty)\ m^{-1}$$

$$\lambda = 9.12 \times 10^{-8}\ m$$

(b) The energy required is the energy of the photon in (a). It is

$$Energy = h\upsilon = hc/\lambda = 6.63 \times 10^{-34} Js \left(\frac{3 \times 10^{8} ms^{-1}}{9.12 \times 10^{-8}\ m}\right)$$

$= 2.18 \times 10^{-18}\ J = 13.6\ eV$. (Since $1\ eV = 1.6 \times 10^{-19} J$)

18.4 X-Ray spectra

By X-rays, we usually mean electromagnetic radiation (light) which has a wavelength shorter than that of ultraviolet light though, there is no sharp boundary. The range is usually considered to be 0.1 to 1 nm, which corresponds to quantum energies of 1——— 100 k eV

X-rays are produced if heavier atoms are bombarded by electrons which have been accelerated through thousands of volts. They were first observed by Wilhelm K.Roentgen in 1895, (and are also called Roentgen rays) using an apparatus similar in principle to that shown in Fig.(18.5.). Electrons released from the heated cathode by thermionic emission are accelerated towards the anode by a large potential difference. It is found that at sufficiently high potentials (several thousand volts) a very penetrating radiation is emitted from the surface of the anode. These rays are of the same nature as light or any other electromagnetic wave. The X-rays are detected by photographic plates, film, counting tubes, or more recently, by semiconductor detectors. In 1913, W.H.Bragg discovered the phenomenon of X-ray diffraction by crystals. This technique permitted the precise measurement of the wavelength of X-rays and thus became the basis for a study of X-rays spectrum.

Spectral analysis of x-rays shows that:

i) there is always a continuous, the X-ray bremsstrahlung

ii) and under certain conditions, there is in addition a line spectrum, the characteristic spectrum.

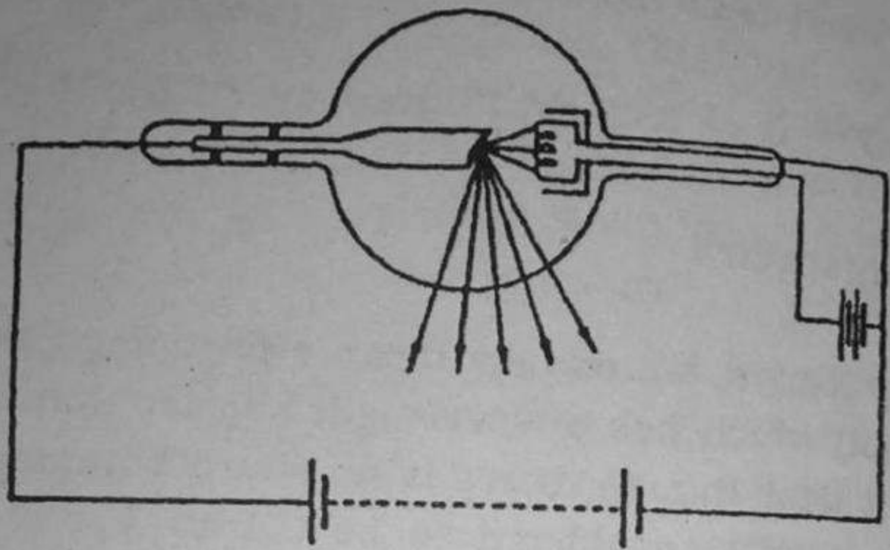


Fig.18.5.

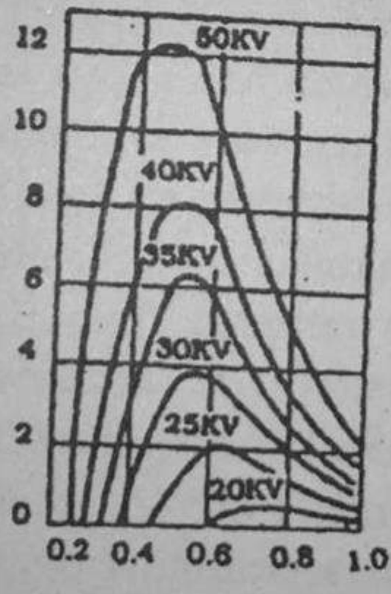


Fig.18.6.

18.5. X-ray Continuous Spectra or X-ray Bremsstrahlung

Ordinarily, a continuous spectrum of frequencies of X-rays is emitted, but the maximum frequency or minimum wavelength was observed to be always directly proportional to the accelerating voltage between the electrodes in the manner indicated in Fig. (18.6.).

Furthermore, this maximum frequency was found to be very nearly independent of the material of which the electrodes were made. These observations can be understood on the basis of quantum hypothesis. The bremsstrahlung spectrum is a result of the fact that when electrons pass close to the atomic nuclei, they are deflected and slowed down (Fig.18.7).

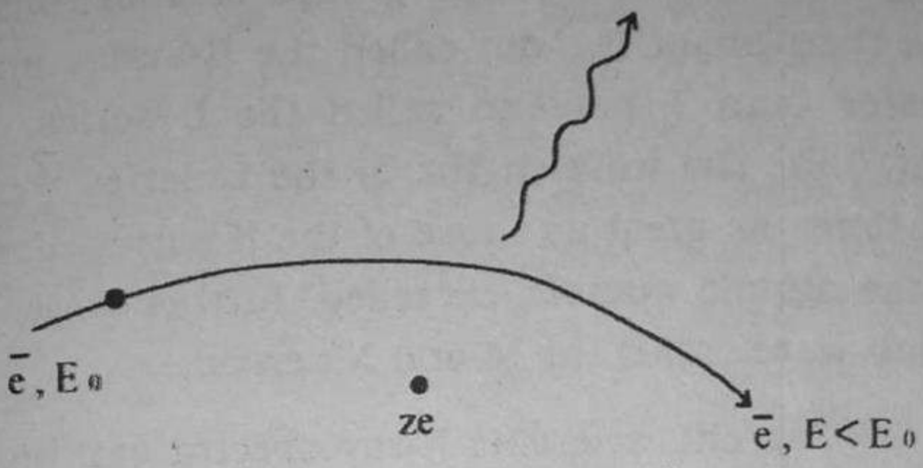


Fig.18.7

A positive or negative accelerated charge will according to classical electrodynamics, emit electromagnetic radiation.This is continuous x-rays bremsstrahlung. In terms of quantum theory, this can be understood as follows: for each braking incident, a quantum of light $h\upsilon = E_0 - E$ is emitted. However, since the beginning and end states are not quantized-the electrons are free, not quantized, The process is represented as

$$\text{Atom} + \bar{e}\,(\text{fast}) \rightarrow \text{Atom} + \bar{e}\,(\text{slow}) + h\upsilon$$

Characteristic X-Ray Spectra:

We have mentioned that the spectrum of X-rays consists of a continuous spectrum upon which is superposed a line spectrum (Fig.18.8). The distribution of energy in the continuous spectrum depends only upon the potential difference across the X-ray tube while the line spectrum is the characteristic of the target. These are called characteristic spectra and were investigated by Moseley in 1913 by making each element in turn the target in an X-ray tube. Thirty nine elements extending from aluminium to gold were examined in this way. The X-rays were analyzed with a Bragg crystal spectrometer. Most elements showed two groups of lines, one generally less than about 0.1 nm called the K-series and another greater than 1 nm and called the L series, similar to Fig.(18.8). The wavelengths of the L series were roughly ten times as great as those of the K series. For elements whose atomic number exceeded, furthçr series appeared which were called the M and N series.

The characteristic X-ray spectra can be explained from the principle of inner shell transitions. The electrons of an atom are ordered according to their arrangement in shells about the nucleus. Each shell has a certain maximum number of electrons. There is a rule that the number of electrons in the nth orbit is equal to $2n^2$. According to this rule, there cannot be more than 2,8,18,32,..... electrons in the orbits or shells for n=1,2,3,4,..... These orbits or shells are called K,L,M,N,.... for n =1,2,3,4,.....respectively.

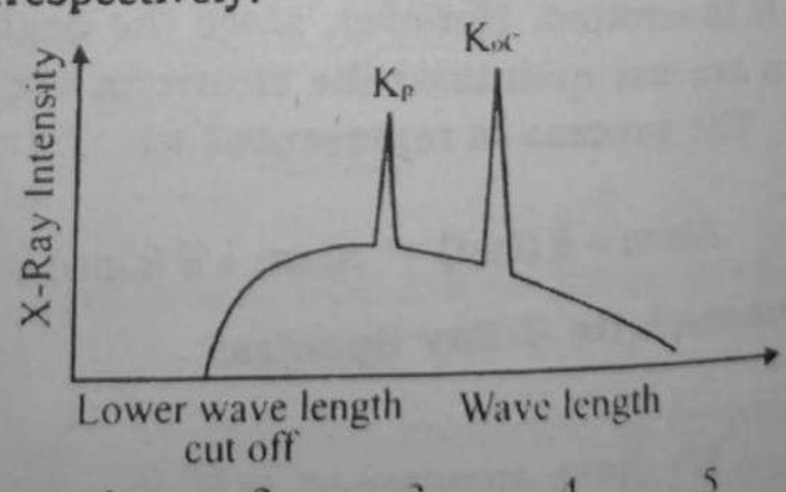


Fig.18.8

N	1	2	3	4	5
Spectroscopy Notation	K	L	M	N	O

When highly energetic incident electrons knock an electron from the K shell, there occurs a vacancy in that shell. This vacant space in the K shell is filled by an electron from the L shell giving up energy in the form of an X-ray. This radiation, which is characteristic of the target material, is labeled the K_α linc. The electron in the M shell that fills a vacancy in the K shell gives up energy as an X-ray called the $K\beta$ linc. These transitions from the shells L,M,N and so on to the K shell give rise to a series of lines K_α, K_β, K_γ, and so on called the K series. When incident electrons dislodge electrons from the L shell, these are filled by electrons from the remaining M,N,O, shells etc. These transitions give rise the L series, the first line of which is L_α The nomenclature of these transitions is illustrated in Fig.18.9.and 18.10. Upon closer observation each of the characteristic X-ray line is found to be composed of a number of closely spaced lines called the X-rays fine structure, the explanation of which is not taken up here, since this topic is beyond the level of this book.

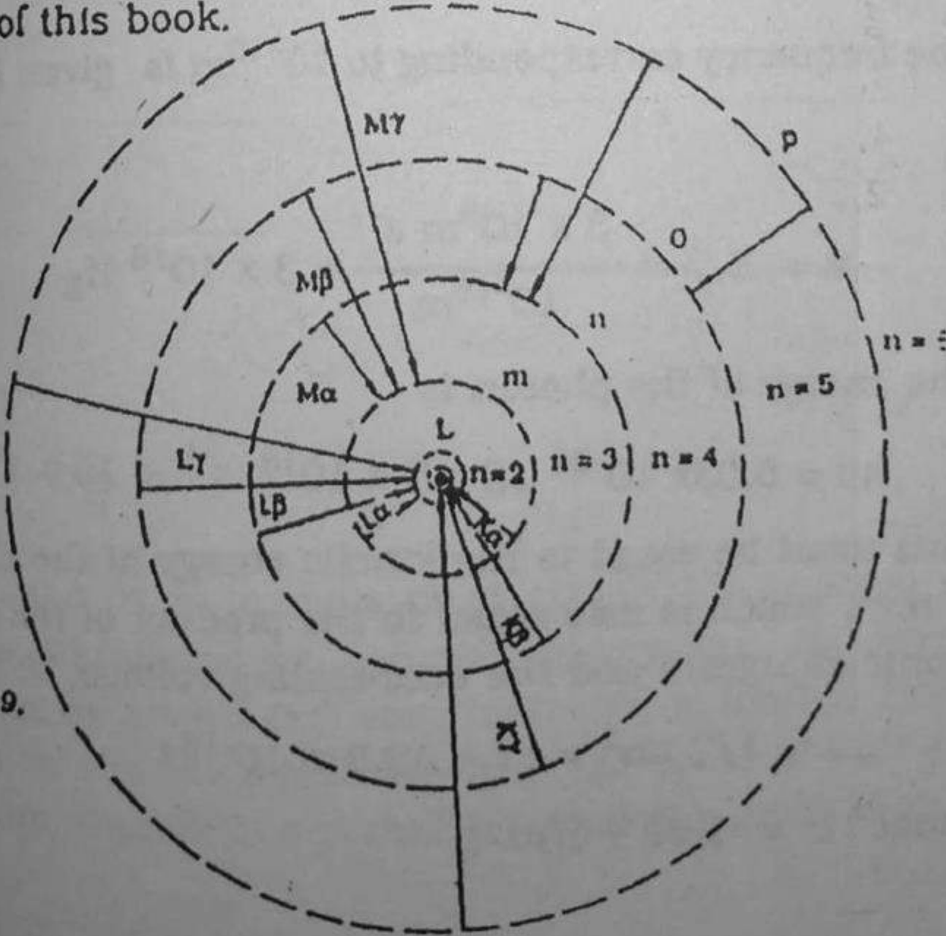


Fig.18.9.

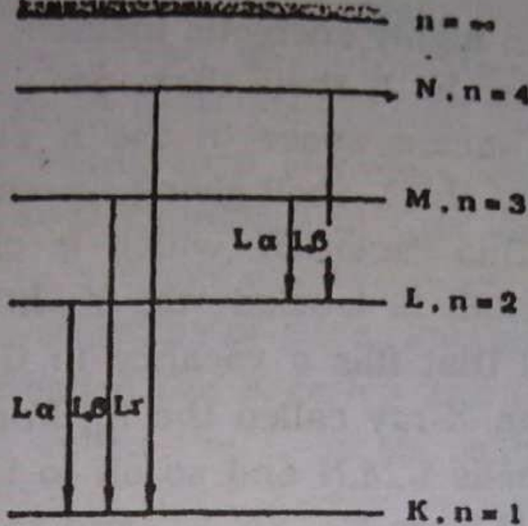


Fig. 18.10.

Example: 18.5

Calculate the potential difference through which an electron must be accelerated in order that the short wave limit of the continuous X-ray - spectrum shall be exactly 1×10^{-10}m.

Solution:

The frequency corresponding to 10^{-10}m is given by

$$\nu = E/\lambda = \frac{3 \times 10^{8} \text{m s}^{-1}}{10^{-10}\text{m}} = 3 \times 10^{18} \text{ H}_z$$

The energy of the photon is

$$h\nu = 6.63 \times 10^{-34} \text{ JS} \times 3 \times 10^{18} \text{ s}^{-1} = 19.9 \times 10^{-16} \text{J}$$

This must be equal to the kinetic energy of the electron $\frac{1}{2}mv^2$, which is also equal to the product of the electronic charges e and the accelerating voltage, V:

$$\text{i.e. } 1/2\, mv^2 = eV = 19.9 \times 10^{-16}\text{J}$$

Since $e = 1.60 \times 10^{-19}$C

so $V = 19.90\times 10^{-16} J/1.60\times 10^{-19} C = 12.400\ V.$

18.6. Introduction to Lasers

Lasers are one of the most important discoveries made in the second half of the twentieth century. The laser is a device for producing very intense, highly directional, coherent and monochromatic light beams. The name laser stands for Light Amplification by Stimulated, Emission of Radiation. Different types of lasers: solid state, gas, semiconductor, liquid-are used for producing light at frequencies from the far infrared to the ultraviolet regions. In order to understand the basic principle of lasers, we must first discuss the three radiation processes.

Suppose that a beam of photons of energy $\upsilon h=E_2-E_1$ is incident on a sample in which atoms are in the ground state. If the photons interact with an atom in the ground state, the atom absorbs the photon and reaches the excited state E_2 (Fig.18.11) as the atom was induced

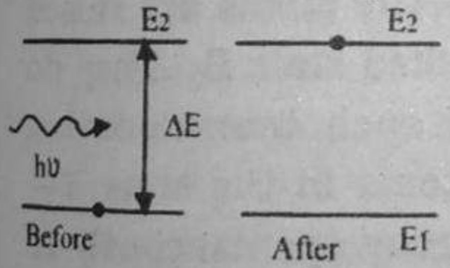


Fig. 18.11

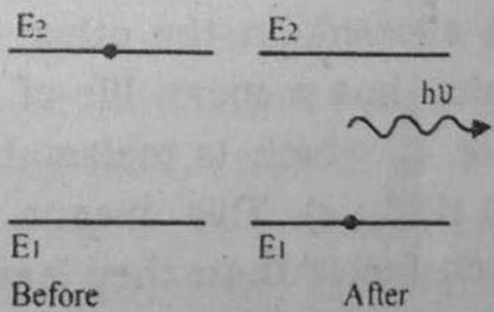


Fig. 18.12

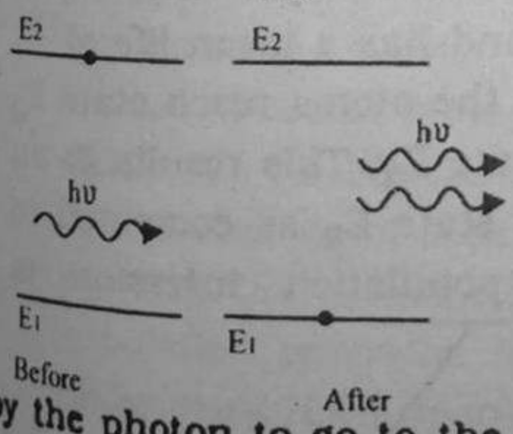


Fig. 18.13

by the photon to go to the excited state. This process is known as stimulated or induced absorption. An atom can remain in an excited state only for a limited time called life time of the state-usually of the order of 10^{-8} S. There exists, therefore, a probability that the atom in the excited state E_2 will fall back to the lower state E_1. Since this transition is spontaneous (not of external origin), then the radiation emitted is called spontaneous emission (Fig.18.12). The emitted radiation is of random character and is incoherent. The third process is called stimulated emission. In this process, an atom in an excited state E_2 (Fig.18.13) absorbs a photon of energy $h\nu = E_2 - E_1$. This incident photon will increase the probability that the atom will go back to its ground state and thereby emits a second photon of energy $h\nu$. The two identical photons emitted in this process-the incident photon and the emitted photon are shown in the Fig. The emitted photon will be exactly in phase with the incident photon.

18.7. The Laser Principle

Principle of a laser is explained by considering that atoms of a material have a number of energy levels and atleast one of which is metastable, the state having much longer life time than 10^{-8} s. We consider a three level system as shown in Fig.18.14.The atoms are raised from the ground state E_1 to the excited state E_3 .They do not fall back to state E_1 because such transitions are not allowed.On the other hand, atoms in the state E_3 (which has a mean life of 10^{-8}s) decay spontaneously to state E_2 which is metastable (and has a mean life of 3×10^{-3} s). This means that the atoms reach state E_2 much faster than they leave state E_2. This results in an increased number of atoms in state E_2 as compared to the number in E_1. Thus population inversion is

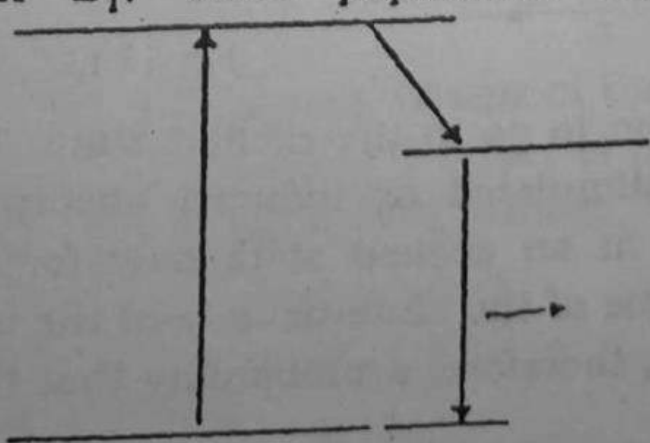


Fig.18.14

achieved. After population inversion is obtained, the state E_2 is exposed to beam of photons of energy $h\upsilon = E_2 - E_1$, which causes induced emission. In order to sustain this process, some method is employed to maintain the population inversion in the states. This is achieved by confining the emitted radiations in an assembly. The ends of which are fitted with mirrors. One end mirror is totally reflecting while the other is made partially reflecting for the laser beam to be taken out.

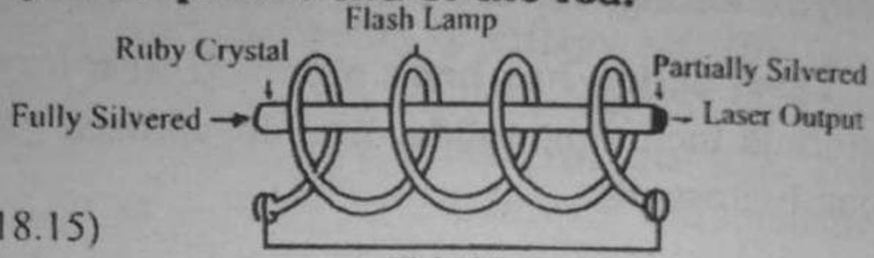
Laser action has been obtained in a large variety of materials including solids, liquids, ionized gases, dyes, semiconductors etc.

We give a brief account of a typical laser called ruby laser.

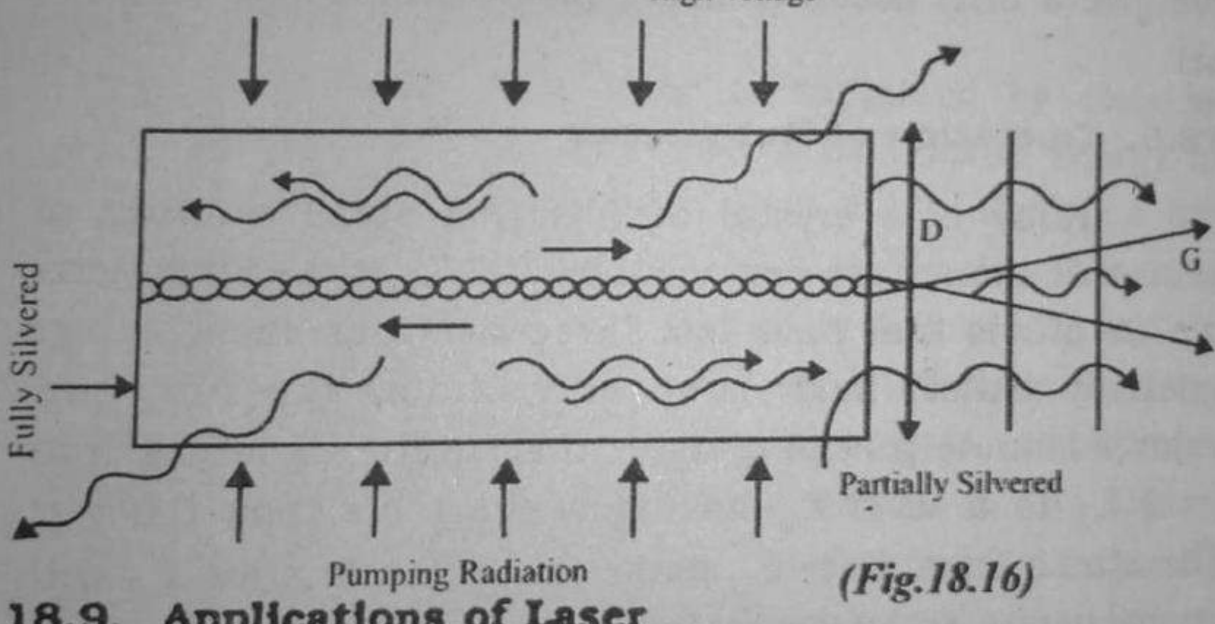
18.8. Operation of Ruby Laser

Ruby is a crystal of Al_2O_3, a small number of whose Al atoms are replaced by Cr^{+++} ions. Such ions are Cr atoms that have lost three electrons each. A high intensity helical flash lamp surrounding the ruby provides adequate pumping light to raise the Cr atoms from level E_1 to a level E_3 having a short life time (10^{-8} s). The atoms from state E_3 make transition to state E_2 with spontaneous emission making the number of atoms larger than those in state E_3, since E_2 is a metastable state which has its life time of the order 10^{-3}. In this process the number n_1 of atoms from state E_1 are going faster to E_3 than the number n_2 of atoms leaving the state E_2. Population inversion has been created. A few Cr atoms make transitions spontaneously from level E_2 to level E_1 and these emitted photons of $\lambda = 694.3$ nm stimulate further transitions. Stimulated emission will dominate stimulated absorption (because $n_2 > n_1$) and we obtain an intense coherent monochromatic red beam of light.

In practice, the ruby laser is a cylindrical rod with parallel, optically flat reflecting ends, one of which is only partly reflecting as shown in Figure.18.15'. These emitted photons which travel exactly in the direction of the axis are reflected several times, and they are capable of estimulating emission repeatedly. Those photons not in the direction of axis leave through the sides. Thus the number of photons is built up rapidly and leave through the partially transparent end of the rod.



(Fig.18.15)



(Fig.18.16)

18.9. Applications of Laser

There are a number of applications of Laser technology. Some applications are as follows:-

1. Three dimensional images of objects obtained by using lasers in a process called Holography.
2. As surgical tool for "welding" detached retina.
3. To perform precision surveying and length measurements.
4. As a potential energy source for inducing nuclear fusion reactions.

5. For telephone communications along optical fibers.

6. For precision cutting of metals and other materials. These and other applications are possible only because of the unique characteristics of Laser light.

QUESTIONS.

18.1. The Bohr theory of hydrogen atom is based upon several assumptions. Do any of these assumptions contradict classical physics?

18.2. Why does the hydrogen gas produced in the Laboratory not glow and emit radiations?

18.3. Why are the energy levels of the hydrogen atom less than zero.

18.4. If hydrogen gas is bombarded by electrons of energy 13.6 eV would you expect to observe all the lines of hydrogen spectrum.

18.5. Hydrogen gas at room temperature absorbs light of wavelengths equal to the lines in the Lyman Series but not those of the Balmer Series. Explain.

18.6. How X-rays are different from the visible radiations.

18.7. What property of X-rays makes them so useful in seeing otherwise invisible internal structures?

18.8. Explain the difference between laser light and light from an incandescent bulb.

18.9. Name some applications of Lasers.

PROBLEMS

18.1. Calculate the following (a) the orbit radius (b) the angular momentum (c) the linear momentum (d) the kinetic energy (e) the potential energy (f) the total energy for the Bohrs hydrogen atom in ground state.

Ans.[(a) 5.3×10^{-11}m (b) 1.1×10^{-34}Js
(c) 2.1×10^{-24}kgm/s (d) 13.6eV (e) −27.2eV
(f) 13.58 eV]

18.2. What is the wavelength of the radiation that is emitted when a hydrogen atom undergoes a transition from the state n=3 to n=1

Ans.(103 nm.)

18.3. Light of wavelength 486.3 nm is emitted by a hydrogen atom in Balmer series. what transitions of the hydrogen atom is responsible for this radiation.

Ans.(n = 4.)

18.4. In the hydrogen atoms an electron experiences a transition from a state whose binding energy is 0.54 eV to another state whose excitation energy is 10.2 eV (a) What are the quantum numbers for these states? (b) Compute the wavelength of the emitted photon. (c) To what series does this line belong.

Ans (a) 5,2 (b) 434 nm (c) Balmer Series.

18.5. Photon of 12.1 eV absorbed by a hydrogen atom, originally in the ground state, raises the atom to an excited state. What is the quantum number of this state:

Ans: (3.)

18.6. (a) Find the wavelength of the first three lines of the Lyman series of hydrogen.

Ans: [121.5 nm, 102.5 nm, 97.2nm]

18.7 In an experiment, the excitation potentials of hydrogen are found at 10.21 V and 12.10 V. three different spectral lines are emitted. Find their wavelengths.

Ans: [102.55 nm, 121.54 nm, 655.92 nm.]

18.8. What minimum energy is needed in an X-ray tube in order to produce X-rays with a wavelength of 0.1×10^{-10} m.

Ans: [1.99×10^{-16} J]

18.9. A certain atom emits spectrum lines at 300, 400 and 1200 nm. Assuming that three energy levels are involved in the corresponding transitions, calculate the quantum of energy emitted at each wave length:

Ans: [4.14 eV, 3.1eV, 1.03 eV]

18.10. Calculate the energy of a photon whose frequency is

(a) (i) 4×10^{14} Hz (ii) 20 GHz

(iii) 30 MHz. Express your answer in eV.

Ans: [(i) 1.65 eV (ii) 8.28×10^{-5} eV

(iii) 1.24×10^{-7} eV.]

(b) Describe the corresponding wavelengths for the photons described in (a)

Ans:[(i) 750 nm, (ii) 0.015m (iii) 10m.]