1. Details of Module and its structure

Module Detail	
Subject Name	Chemistry
Course Name	Chemistry 01 (Class XI, Semester 01)
Module Name/Title	Structure of Atom: Part 2
Module Id	kech_10202
Pre-requisites	Knowledge of waves
Objectives	After going through this module you will be able to: 1. Know about wave nature and particle electromagnetic radiations 2. Explain photoelectric effect nature of 3. Differentiate between emission and absorption spectrum 4. Understand the cause of line spectrum of hydrogen atom
Keywords	Black Body Radiation, Photoelectric Effect, Atomic Spectra, Hydrogen Atom Spectra

2. Development Team

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1. INTRODUCTION

In the previous module, you have learned that Rutherford proposed a model for the structure of atom on the basis of α scattering experiment. He proposed that the atom consists of a central positively charged nucleus and electrons revolve around it. He was not able to explain why these electrons do not fall into the nucleus. In this module you will study the two important developments which formed the basis of improvement made in Rutherford's model by Neils Bohr.

2. DEVELOPMENTS LEADING TO THE BOHR'S MODEL OF ATOM

Two developments played a major role in the formulation of Bohr's model of atom. These were:

(i) Dual character of the electromagnetic radiation which means that radiations possess both wave like and particle like properties, and

(ii) Experimental results regarding atomic spectra which can be explained only by assuming quantized electronic energy levels in atoms.

2.1 Wave Nature of Electromagnetic Radiation

In the mid-nineteenth century, physicists actively studied absorption and emission of radiation by heated objects. These are called thermal radiations. They tried to find out of what the thermal radiation is made. It is now a well-known fact that thermal radiations consist of electromagnetic waves of various frequencies or wavelengths. It is based on a number of modern concepts, which were unknown in the mid-nineteenth century. First active study of thermal radiation laws occurred in the 1850's and the theory of electromagnetic waves and the emission of such waves by accelerating charged particles was developed n the early 1870's by James Clerk Maxwell, which was experimentally confirmed later by Heinrich Hertz.

James Maxwell (1870) was the first to give a comprehensive explanation about the interaction between the charged bodies and the behaviour of electrical and magnetic fields on macroscopic level. He suggested that when electrically charged particle moves under acceleration, alternating electrical and magnetic fields are produced and transmitted. These fields are transmitted in the forms of waves called **electromagnetic waves** or **electromagnetic radiation**.

Light is the form of radiation known from early days and speculation about its nature dates back to remote ancient times. In earlier days (Newton) light was supposed to be made of particles (corpuscules). It was only in the 19th century when wave nature of light was established.

Maxwell was again the first to reveal that light waves are associated with oscillating electric and magnetic character (Fig 1). Although electromagnetic wave motion is complex in nature, we will consider here only a few simple properties.

(i) The oscillating electric and magnetic fields produced by oscillating charged particles are perpendicular to each other and both are perpendicular to the direction of propagation of the wave. Simplified picture of electromagnetic wave is shown in Fig.1.

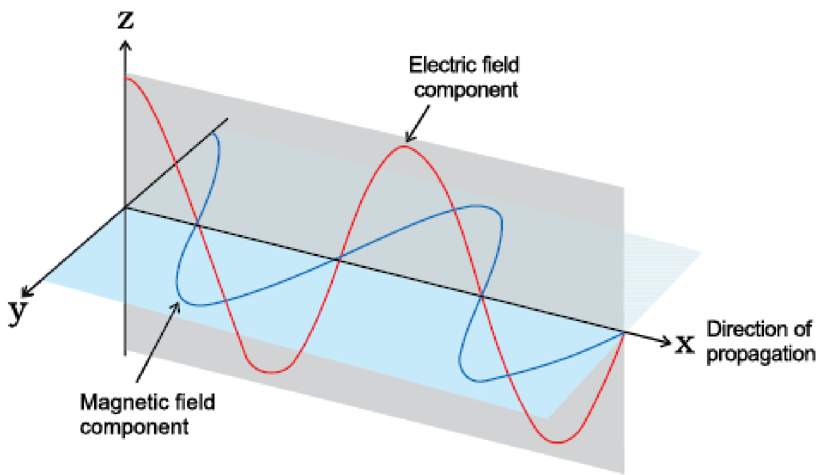

Fig.1 *The electric and magnetic field components of an electromagnetic wave. These components have the same wavelength, frequency, speed and amplitude, but they vibrate in two mutually perpendicular planes.*

(ii) Unlike sound waves or water waves, electromagnetic waves do not require medium and can move in vacuum.

(iii) It is now well established that there are many types of electromagnetic radiations, which differ from one another in wavelength (or frequency). These constitute what is called **electromagnetic spectrum** (Fig.2). Different regions of the spectrum are identified by different names. Some examples are: radio frequency region around 10^6 Hz, used for broadcasting; microwave region around 10^{10} Hz used for radar; infrared region around 10^{13} Hz used for heating; ultraviolet region around 10^{16}Hz a component of sun's radiation. The small portion around 10^{15} Hz, is called **visible light** as it is only this part of electromagnetic spectrum which our eyes can see. Special instruments are required to detect non-visible radiation.

(iv) Different kinds of units are used to represent electromagnetic radiation based on various properties of radiations such as frequency (ν) and wavelength (λ). The SI unit for frequency (ν) is hertz (Hz), after Heinrich Hertz. It is defined as the number of waves that pass a given point in one second.

1 Hz = 1 s^{-1}.

Wavelength has the units of length and the SI units of length is meter (m). Since electromagnetic radiation consists of different kinds of waves having much smaller wavelengths, smaller units of lengths commonly are used. Fig.2 shows various types of electro-magnetic radiations which differ from one another in wavelengths and frequencies.

In vaccum all types of electromagnetic radiations, regardless of wavelength, travel at the same speed, i.e., 3.0×10^8 m s^{-1}(2.997925×10^8m s^{-1}, to be precise). This is called **speed of light** and is given the symbol 'c'. The frequency (ν), wavelength (λ) and velocity of light (c) are related by the equation

$$c = \nu \lambda \tag{1}$$

The other commonly used quantity in spectroscopy, is the **wavenumber** ($\bar{\nu}$). It is defined as the number of wavelengths per unit length. Its units are reciprocal of wavelength unit, i.e., m^{-1}. However, a non SI unit commonly used is cm^{-1}.

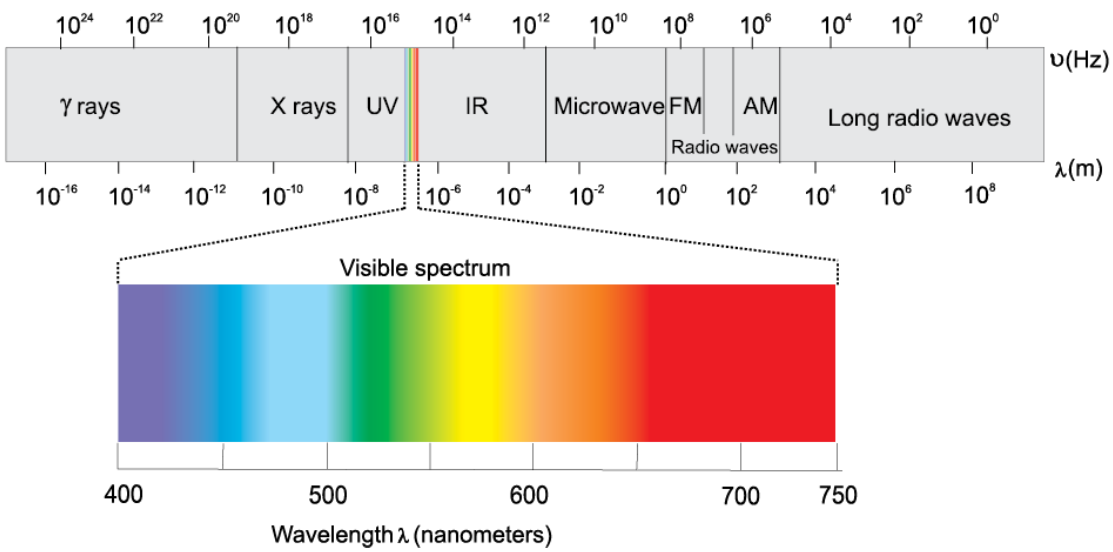

***Fig. 2** (a) The spectrum of electromagnetic radiation. (b) Visible spectrum. The visible region is only a small part of the entire spectrum .*

Problem 1

The Vividh Bharati station of All India Radio, Delhi, broadcasts on a frequency of 1,368 kHz (kilo hertz). Calculate the wavelength of the electromagnetic radiation emitted by transmitter. Which part of the electromagnetic spectrum does it belong to?

Solution

The wavelength, λ, is equal to c/ν , where c is the speed of electromagnetic radiation in vacuum and ν is the frequency. Substituting the given values, we have

$$\lambda = \frac{c}{\nu}$$

$$= \frac{3.00 \times 10^8 \text{m s}^{-1}}{1368 \text{kHz}}$$

$$= \frac{3.00 \times 10^8 \text{m s}^{-1}}{1368 \times 10^3 \text{s}^{-1}}$$

=219.3 m

This is a characteristic radiofrequency wavelength.

Problem 2
The wavelength range of the visible spectrum extends from violet (400 nm) to red (750 nm). Express these wavelengths in frequencies (Hz). (1nm = 10^{-9} m)
Solution
Using equation 1, frequency of violet light
$$\lambda = \frac{c}{\nu} = \frac{3.00 \times 10^{8} \mathrm{m\ s^{-1}}}{400 \times 10^{-9} \mathrm{m}}$$

= 7.50×10^{14} Hz
Frequency of red light
$$\nu = \frac{c}{\lambda} = \frac{3.00 \times 10^{8} \mathrm{m\ s^{-1}}}{750 \times 10^{-9} \mathrm{m}} = 4.00 \times 10^{14} \mathrm{Hz}$$
The range of visible spectrum is from 4.0×10^{14} to 7.5×10^{14} Hz in terms of frequency units.

Problem 3
Calculate (a) wave number and (b) frequency of yellow radiation having wavelength 5800 Å.
Solution
(a) Calculation of wave number ($\bar{\upsilon}$)
λ = 5800 Å = 5800 x 10^{-8} cm = 5800 x 10^{-10} m
$$\bar{\nu} = \frac{1}{\lambda} = \frac{1}{5800 \times 10^{-10} \mathrm{m}}$$
$$= 1.724 \times 10^{6} \mathrm{m^{-1}}$$
$$= 1.724 \times 10^{4} \mathrm{cm^{-1}}$$
(b) Calculation of the frequency (ν)
$$\nu = \frac{c}{\lambda} = \frac{3.00 \times 10^{8} \mathrm{m\ s^{-1}}}{5800 \times 10^{-10} \mathrm{m}} = 5.172 \times 10^{14} \mathrm{s^{-1}}$$

2.2 Particle Nature of Electromagnetic Radiation: Planck's Quantum Theory

Electromagnetic radiations show the phenomenon of diffraction and interference. Diffraction is the bending of wave around an obstacle. Interference is the combination of two waves of the same or different frequencies to give a wave whose distribution at each point in space is the algebraic or vector sum of disturbances at that point resulting from each interfering wave. Since diffraction and interference are the properties of waves, the electromagnetic radiations were considered to be waves. However, following are some of the observations which could not be explained with the help of the electromagnetic theory of 19th century physics (known as classical physics) which considered the radiations as waves:

(i) the nature of emission of radiation from hot bodies (black -body radiation)
(ii) ejection of electrons from metal surface when radiation strikes it (photoelectric effect)
(iii) variation of heat capacity of solids as a function of temperature
(iv) line spectra of atoms with special reference to hydrogen.

These phenomena indicate that the system can take energy only in discrete amounts. All possible energies cannot be taken up or radiated.

It is noteworthy that the first concrete explanation for the phenomenon of the black body radiation was given by Max Planck in 1900. This explanation resulted in the development of Planck's quantum theory.

Black body radiation

Hot objects emit electromagnetic radiations over a wide range of wavelengths. At high temperatures, an appreciable proportion of radiation is in the visible region of the spectrum. As the temperature is raised, a higher proportion of short wavelength (blue light) is generated. For example, when an iron rod is heated in a furnace, it first turns to dull red and then progressively becomes more and more red as the temperature increases. As this is heated further, the radiation emitted becomes white and then becomes blue as the temperature becomes very high. This means that red radiation is most intense at a particular temperature and the blue radiation is more intense at another temperature. This means intensities of radiations of different wavelengths emitted by hot body depend upon its temperature. By late 1850's it was known that objects made of different material and kept at different temperatures emit different amount of radiation. Also, when the surface of an object is irradiated with light (electromagnetic radiation), a part of radiant energy is generally reflected as such, a part is absorbed and a part of it is transmitted. The reason for incomplete absorption is that ordinary objects are as a rule imperfect absorbers of radiation. An ideal body, which emits and absorbs radiations of all frequencies uniformly, is called a black body and the radiation emitted by such a body is called black body radiation. In practice, no such body exists. Carbon black approximates fairly closely to black body. A good physical approximation to a black body is a cavity with a tiny hole, which has no other opening. Any ray entering the hole will be reflected by the cavity walls and will be eventually absorbed by the walls. A black body is also a perfect radiator of radiant energy. Furthermore, a black body is in thermal equilibrium with its surroundings. It radiates same amount of energy per unit area as it absorbs from its surrounding in any given time. The amount of light emitted (intensity of radiation) from a black body and its spectral distribution depends only on its temperature.At a given temperature, intensity of radiation emitted increases with the increase of wavelength, reaches a maximum value at a given wavelength and then starts decreasing with further increase of wavelength, as shown in Fig.3. Also, as the temperature increases, maxima of the curve shifts to short wavelength. Several attempts were made to predict the intensity of radiation as a function of wavelength.

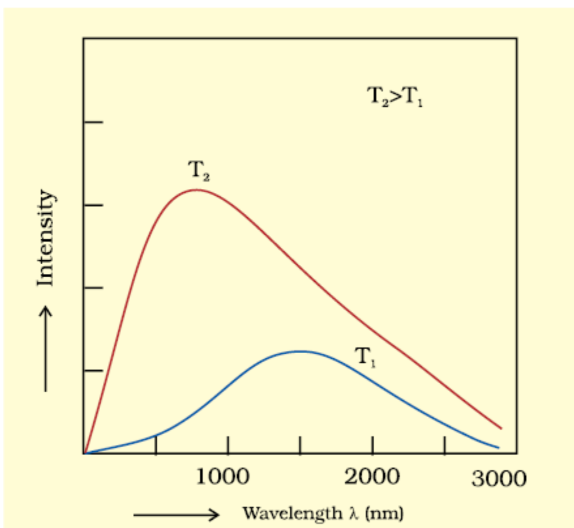

Fig. 3 Wavelength-intensity relationship

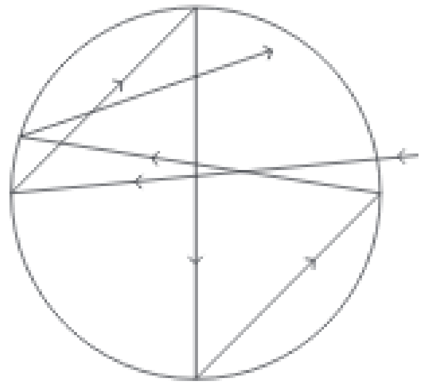

Fig.4 Black body

But the results of the above experiment could not be explained satisfactorily on the basis of the wave theory of light. Max Planck arrived at a satisfactory relationship by making an assumption that absorption and emmission of radiation arises from oscillator i.e., atoms in the wall of black body (Fig. 4). Their frequency of oscillation is changed by interaction with oscilators of electromagnetic radiation. Planck assumed that radiation could be sub-divided into discrete chunks of energy. He suggested that atoms and molecules could emit or absorb energy only in discrete quantities and not in a continuous manner. He gave the name quantum to the smallest quantity of energy that can be emitted or absorbed in the form of electromagnetic radiation. The energy (E) of a quantum of radiation is proportional to its frequency (ν) and is expressed by equation:

$$E = h\nu \qquad (2)$$

The proportionality constant, '*h*' is known as Planck's constant and has the value 6.626×10^{-34} J s.
With this theory, Planck was able to explain the distribution of intensity in the radiation from black body as a function of frequency or wavelength at different temperatures.

Quantisation has been compared to standing on a staircase. A person can stand on any step of a staircase, but it is not possible for him/her to stand in between the two steps. The energy can take any one of the values from the following set, but cannot take on any values between them.

E = 0, hυ, 2hυ, 3hυ....nhυ.....

Photoelectric Effect

In 1887, H. Hertz performed a very interesting experiment in which electrons (or electric current) were ejected when certain metals (for example potassium, rubidium, caesium etc.) were exposed to a beam of light as shown in Fig.5.

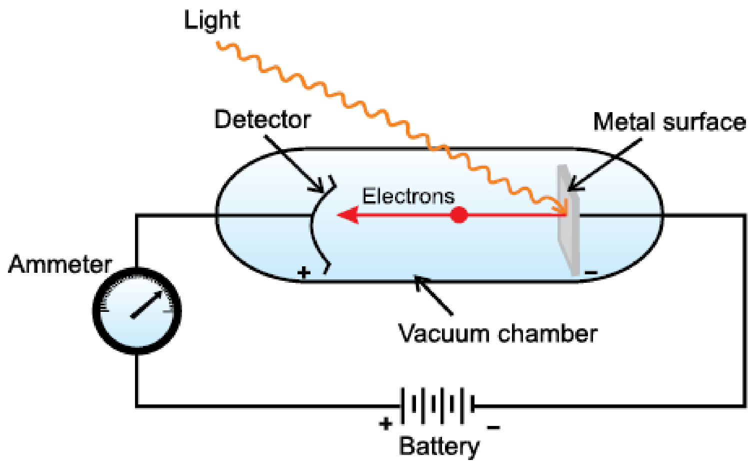

Fig.5 Equipment for studying the photoelectric effect. Light of a particular frequency strikes a clean metal surface inside a vacuum chamber. Electrons are ejected from the metal and are counted by a detector that measures their kinetic energy.

The phenomenon is called **Photoelectric effect**. The results observed in this experiment were:

(i) The electrons are ejected from the metal surface as soon as the beam of light strikes the surface, i.e., there is no time lag between the striking of light beam and the ejection of electrons from the metal surface.

(ii) The number of electrons ejected is proportional to the intensity or brightness of light.

(iii) For each metal, there is a characteristic minimum frequency of light,ν_0 (also known as **threshold frequency**) below which photoelectric effect is not observed. At a frequency $\nu > \nu_0$, the ejected electrons come out with some kinetic energy. The kinetic energies of these electrons increase with the increase of frequency of the light used.

All the above results could not be explained on the basis of laws of classical physics. According to latter, the energy content of the beam of light depends upon the brightness of the light. In other words, number of electrons ejected and kinetic energy associated with them should depend on the brightness (intensity) of light. It has been observed that though the number of electrons ejected does depend upon the brightness of light, the kinetic energy of the ejected electrons does not. For example, red light [$\nu = (4.3 \text{ to } 4.6) \times 10^{14}$ Hz] of any brightness (intensity) may shine on a piece of potassium metal for hours but no electrons are ejected. However, as soon as even a very weak yellow light ($\nu = 5.1–5.2 \times 10^{14}$ Hz) shines on the potassium metal, the photoelectric effect is observed. The threshold frequency (ν_0) for potassium metal is 5.0×10^{14} Hz.

Einstein (1905) was able to explain the photoelectric effect using Planck's quantum theory of electromagnetic radiation as a starting point. Accordingly Einstein considered the beam of light as a collection of energy particles called photons. The energy of the photon is given by the relation

$E = h\nu$

Here, ν is the frequency of light. Shining a beam of light on to a metal surface can be viewed as shooting a beam of particles, the photons. When a photon of sufficient energy strikes an electron in the atom of the metal, it transfers its energy instantaneously to the electron during the collision and the electron is ejected without any time lag or delay. Greater the energy possessed by the photon, greater will be transfer of energy to the electron and greater the kinetic energy of the ejected electron. In other words, kinetic energy of the ejected electron is proportional to the frequency of the electromagnetic radiation. Since the striking photon has energy equal to ν and the minimum energy required to eject the electron is $h\nu_0$ (also called work function, W_0; Table 1), then the difference in energy ($h\nu - h\nu_0$) is transferred as the kinetic energy of the photoelectron. Following the conservation of energy principle, the kinetic energy of the ejected electron is given by the equation.

$$h\nu = h\nu_0 + \frac{1}{2}m_e v^2 \qquad (3)$$

where m_e is the mass of the electron and v is the velocity associated with the ejected electron. Lastly, a more intense beam of light consists of larger number of photons, consequently the number of electrons ejected is also larger as compared to that in an experiment in which a beam of weaker intensity of light is employed.

2.3 Dual Behaviour of Electromagnetic Radiation

The particle nature of light posed a dilemma for scientists. On the one hand, it could explain the black body radiation and photoelectric effect satisfactorily but on the other hand, it was not consistent with the known wave behaviour of light which could account for the phenomena of interference and diffraction.

Table 1. Values of Work Function (W_0) for a Few Metals

Metal	Li	Na	K	Mg	Cu	Ag
W_0 /eV	2.42	2.3	2.25	3.7	4.8	4.3

The only way to resolve the dilemma was to accept the idea that light possesses both particle and wave-like properties, i.e., light has dual behaviour. Depending on the experiment, we find that light behaves either as a wave or as a stream of particles. Whenever radiation interacts with matter, it displays particle like properties in contrast to the wavelike properties (interference and diffraction), which it exhibits when it propagates. This concept was totally alien to the way the scientists thought about matter and radiation and it took them a long time to become convinced of its validity. It turns out, as you shall see later, that some microscopic particles like electrons also exhibit this wave-particle duality.

Problem 4

Calculate energy of one mole of photons of radiation whose frequency is 5×10^{14} Hz.

Solution

Energy (E) of one photon is given by the expression

$E = h\nu$

$h = 6.626 \times 10^{-34}$ J s

$\nu = 5 \times 10^{14}$ s^{-1} (given)

$E = (6.626 \times 10^{-34}$ J s$) \times (5 \times 10^{14}$ s$^{-1}) = 3.313 \times 10^{-19}$ J

Energy of one mole of photons = $(3.313 \times 10^{-19}$ J$) \times (6.022 \times 10^{23}$ mol$^{-1}) = 199.51$ kJ mol^{-1}

Problem 5

A 100 watt bulb emits monochromatic light of wavelength 400 nm. Calculate the number of photons emitted per second by the bulb.

Solution

Power of the bulb = 100 watt = 100 J s^{-1}

Energy of one photon $E = h\nu = hc/\lambda$

$$= \frac{6.626 \times 10^{-34} \text{ J s} \times 3 \times 10^{8} \text{ m s}^{-1}}{400 \times 10^{-9} \text{m}}$$

$$= 4.969 \times 10^{-19} \text{ J}$$

Number of photons emitted

$$= \frac{100 \text{ J s}^{-1}}{4.968 \times 10^{-19} \text{ J}} = 2.012 \times 10^{20} \text{ s}^{-1}$$

Problem 6

When electromagnetic radiation of wavelength 300 nm falls on the surface of sodium, electrons are emitted with a kinetic energy of 1.68 $\times 10^{5}$ J mol^{-1}. What is the minimum energy needed to remove an electron from sodium? What is the maximum wavelength that will cause a photoelectron to be emitted?

Solution

The energy (E) of a 300 nm photon is given by

$h\nu = hc/\lambda$

$$= \frac{6.626 \times 10^{-34} \text{ J s} \times 3 \times 10^{8} \text{ m s}^{-1}}{300 \times 10^{-9} \text{m}}$$

= 6.626 × 10^{-19} J

The energy of one mole of photons = 6.626 $\times 10^{-19}$ J × 6.022 $\times 10^{23}$ mol^{-1}

= 3.99 × 10^{5} J mol^{-1}

The minimum energy needed to remove one mole of electrons from sodium

= (3.99 –1.68) 10^{5} J mol^{-1} = 2.31 × 10^{5} J mol^{-1}

The minimum energy for one electron

$$= \frac{2.31 \times 10^{5} \text{ J mol}^{-1}}{6.022 \times 10^{23} \text{electrons mol}^{-1}} = 3.84 \times 10^{-19} \text{ J}$$

This corresponds to the wavelength

$$\lambda = \frac{hc}{E} = \frac{6.626 \times 10^{-34} \text{ J s} \times 3 \times 10^{8} \text{ m s}^{-1}}{3.84 \times 10^{-19} \text{ J}}$$

= 517 nm

(This corresponds to green light)

Problem 7

The threshold frequency ν_0 for a metal is 7.0 $\times 10^{14}$ s^{-1}. Calculate the kinetic energy of an electron emitted when radiation of frequency ν =1.0 $\times 10^{15} s^{-1}$ hits the metal.

Solution

According to Einstein's equation

Kinetic energy = ½ $m_e v^2 = h(\nu - \nu_0)$
$= (6.626 \times 10^{-34}$ J s$)(1.0 \times 10^{15}$ s$^{-1} - 7.0 \times 10^{14}$ s$^{-1})$
$= (6.626 \times 10^{-34}$ J s$)(10.0 \times 10^{14}$ s$^{-1} - 7.0 \times 10^{14}$ s$^{-1})$
$= (6.626 \times 10^{-34}$ J s$) \times (3.0 \times 10^{14}$ s$^{-1})$
$= 1.988 \times 10^{-19}$ J

2.4 Atomic spectra

The speed of light depends upon the nature of the medium through which it passes. As a result, the beam of light is deviated or refracted from its original path as it passes from one medium to another. It is observed that when a ray of white light is passed through a prism, the wave with shorter wavelength bends more than the one with a longer wavelength. Since ordinary white light consists of waves with all the wavelengths in the visible range, a ray of white light is spread out into a series of coloured bands called **spectrum**. The light of red colour which has longest wavelength is deviated the least while the violet light, which has shortest wavelength is deviated the most. The spectrum of white light, that we can see, ranges from violet at 7.50×10^{14} Hz to red at 4×10^{14} Hz. Such a spectrum is called **continuous spectrum.** It is called continuous because violet merges into blue, blue into green and so on. A similar spectrum is produced when a rainbow forms in the sky. Remember that visible light is just a small portion of the electromagnetic radiation (Fig.2). The scientists studied the spectrum of various sources of white light like sun, star and flame. The spectrum from Sun showed some dark regions.

When electromagnetic radiation interacts with matter, the atoms or molecules of constituting matter may absorb energy and reach to a higher energy state. With higher energy, these are in an unstable state. For returning to their normal (more stable, lower energy states) energy state, the atoms and molecules emit radiations in various regions of the electromagnetic spectrum. In nineteenth century, spectrum of light emitted by various elements were studied. These were observed as a series of coloured lines with dark spaces in between. The pattern of lines and dark spaces obtained from elements is called atomic spectra. The origin of atomic spectra could not be explained by assuming continuous electronic energy.

3. EMISSION AND ABSORPTION SPECTRA

The spectrum of radiation emitted by a substance that has absorbed energy is called an **emission spectrum**. Atoms, molecules or ions that have absorbed radiation are said to be **“excited”**. To produce an emission spectrum, energy is supplied to a sample by heating it or irradiating it and the wavelength (or frequency) of the radiation emitted when the sample gives up the absorbed energy, is recorded.

An absorption spectrum is like the photographic negative of an emission spectrum. A continuum of radiation is passed through a sample which absorbs radiation of certain wavelengths. The missing wavelength which corresponds to the radiation absorbed by the matter, leave dark spaces in the bright continuous spectrum.

The study of emission or absorption spectra is referred to as **spectroscopy**. The spectrum of the visible light, as discussed above, was continuous as all wavelengths (red to violet) of the visible light are represented in the spectra. The emission spectra of atoms in the gas phase, on the other hand, do not show a continuous spread of wavelength from red to violet, rather they emit light only at specific wavelengths with dark spaces between them. Such spectra are called **line spectra** or **atomic spectra** because the emitted radiation is identified by the appearance of bright lines in the spectra (Fig. 6).

Line emission spectra are of great interest in the study of electronic structure. Each element has a unique line emission spectrum. The characteristic lines in atomic spectra can be used in chemical analysis to identify unknown atoms in the same way as finger prints are used to identify people. The exact matching of lines of the emission spectrum of the atoms of a known element with the lines from an unknown sample quickly establishes the identity of the latter, German chemist, Robert Bunsen (1811-1899) was one of the first investigators to use line spectra to identify elements.

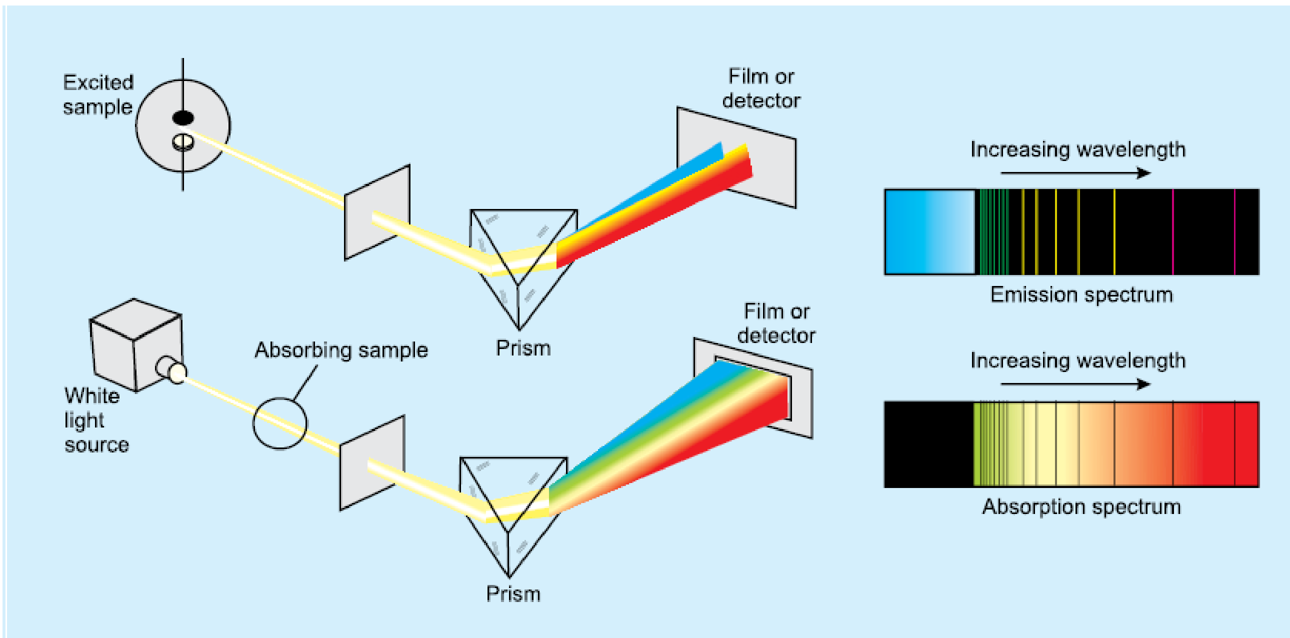

Fig. 6(a) Atomic emission. *The light emitted by a sample of excited hydrogen atoms (or any other element) can be passed through a prism and separated into certain discrete wavelengths. Thus an emission spectrum, which is a photographic recording of the separated wavelengths is called as line spectrum. Any sample of reasonable size contains an enormous number of atoms. Although a single atom can be in only one excited state at a time, the collection of atoms contains all possible excited states. The light emitted as these atoms fall to lower energy states is responsible for the spectrum.* ***(b) Atomic absorption****. When white light is passed through unexcited atomic hydrogen and then through a slit and prism, the transmitted light is lacking in intensity at the same wavelengths as are emitted in (a) The recorded absorption spectrum is also a line spectrum and the photographic negative of the emission spectrum.*

Elements like rubidium (Rb), caesium (Cs) thallium (Tl), indium (In), gallium (Ga) and scandium (Sc) were discovered when their minerals were analysed by spectroscopic methods. The element helium (He) was discovered in the sun by spectroscopic method.

4. LINE SPECTRUM OF HYDROGEN

When an electric discharge is passed through gaseous hydrogen, the H_2 molecules dissociate and the energetically excited hydrogen atoms produced emit electromagnetic radiation of discrete frequencies. The hydrogen spectrum consists of several series of lines named after their discoverers. Balmer showed in 1885 on the basis of experimental observations that if spectral lines are expressed in terms of wave number ($\bar{\upsilon}$) then the visible lines of the hydrogen spectrum obey the following formula:

$$\bar{\nu} = 109{,}677\left(\frac{1}{2^2} - \frac{1}{n^2}\right)\ \text{cm}^{-1} \qquad (4)$$

where n is an integer equal to or greater than 3 (i.e., n = 3,4,5,....)

The series of lines described by this formula are called the **Balmer series**. The Balmer series of lines are the only lines in the hydrogen spectrum which appear in the visible region of the electromagnetic spectrum. The Swedish spectroscopist, Johannes Rydberg, noted that all series of lines in the hydrogen spectrum could be described by the following expression :

$$\bar{\nu} = 109{,}677 \times \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)\ \text{cm}^{-1} \qquad (5)$$

where n_1=1,2........

$n_2 = n_1 + 1,\ n_1 + 2$......

The value 109,677 cm^{-1} is called the **Rydberg constant** for hydrogen. The first five series of lines that correspond to n_1 = 1, 2, 3, 4, 5 are known as Lyman, Balmer, Paschen, Bracket and Pfund series, respectively, Table 2 shows these series of transitions in the hydrogen spectrum. Fig 7 shows the Lyman, Balmer and Paschen series of transitions for hydrogen atom.

Table 2. The Spectral Lines for Atomic Hydrogen

Series	n_1	n_2	Spectral Region
Lyman	1	2,3....	Ultraviolet
Balmer	2	3,4....	Visible
Paschen	3	4,5....	Infrared
Brackett	4	5,6....	Infrared
Pfund	5	6,7....	Infrared

Of all the elements, hydrogen atom has the simplest line spectrum. Line spectrum becomes more and more complex for heavier atom. There are however certain features which are common to all line spectra, i.e., (i) line spectrum of element is unique and (ii) there is regularity in the line spectrum of each element. The questions which arise are: What are the reasons for these similarities? Is it something to do with the electronic structure of atoms? These are the questions need to be answered.

We shall find later that the answers to these questions provide the key in understanding electronic structure of these elements.

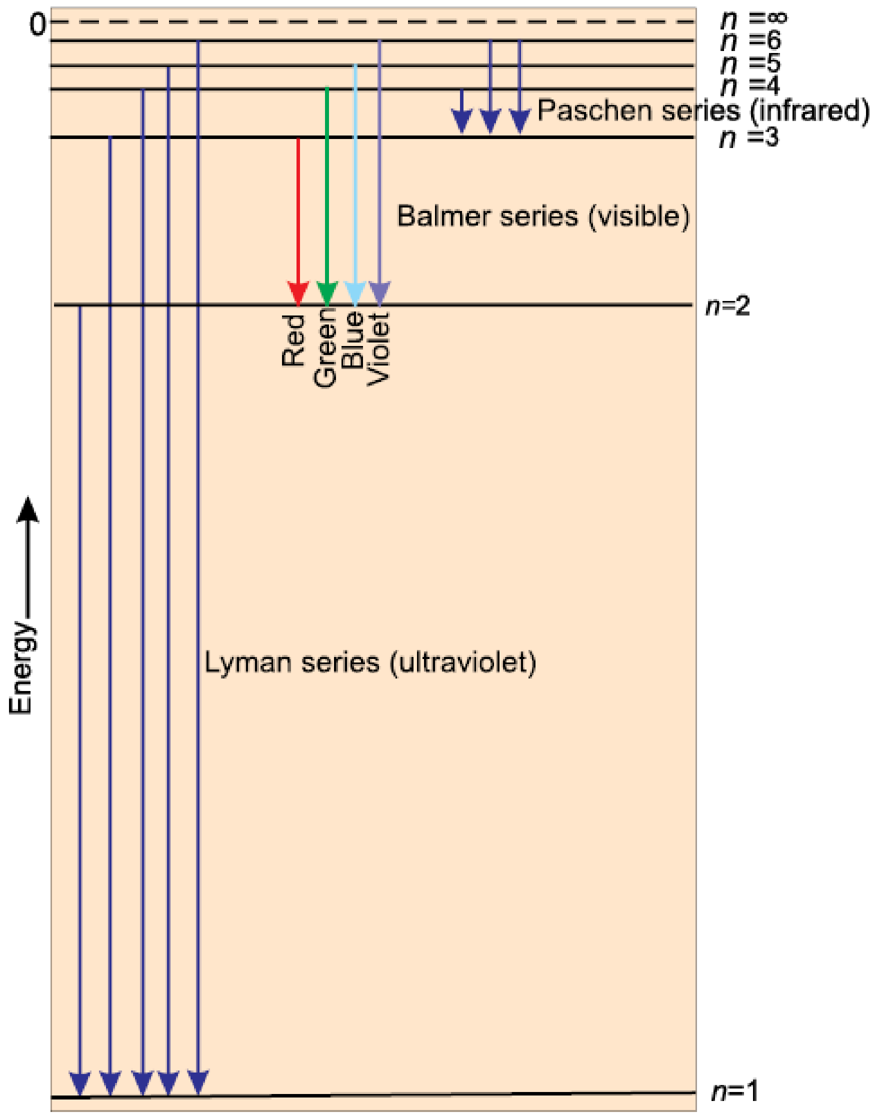

Fig. 7. Transitions of the electron in the hydrogen atom (The diagram shows the Lyman, Balmer and Paschen series of transitions)

5. SUMMARY

- There are many types of electromagnetic radiation which differ from one another in wavelength and constitute the electromagnetic spectrum.
- The phenomenon of diffraction and interference can be explained by the wave nature of the electromagnetic radiations.
- The ideal body which emits and absorbs radiations of all frequencies is called a black body

- Quantum is the smallest quantity of energy that can be emitted or absorbed in the form of electromagnetic radiations
- The electrons are ejected from the metal surface when a beam of light of a particular frequency strikes the surface of a metal. It is called photoelectric effect.
- Electromagnetic spectra may be emission or absorption spectrum on the basis of energy emitted or absorbed.