### Failure of Rutherford's Model

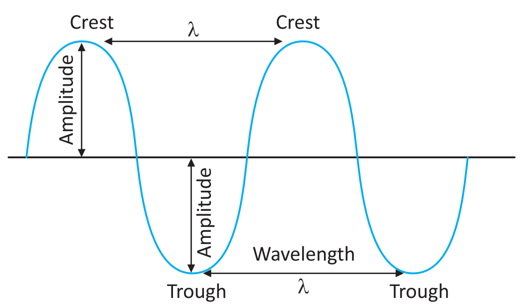
According to *Classical Theory of Electromagnetism*, whenever a charge is subjected to acceleration around an opposite charge, it emits radiation continuously. Hence the electron in Rutherford's atom will loose energy and will not be able to stay in a circular path around the nucleus and should ultimately go into a spiral motion. Such an electron will fall into the nucleus. This, of course, does not happen for electrons in an atom and the discrepancy could not be explained at that time.

**Note :** Later Niel Bohr, a student of Rutherford analysed atomic spectra of Hydrogen atom in terms of *Quantum Theory of Radiation* and applied the results of *Photoelectric Effect* to it and developed a model of atom which was widely accepted at that time.

## NATURE OF CHARACTERISTICS OF RADIANT ENERGY SECTION - 2B

Newton was first person to comment on the nature of light in terms of *Corpuscular Theory of Light*. According to this theory, light is a stream of particles commonly known as *corpuscles of light*. He was able to explain *reflection* and *refraction*, the most common phenomenon of light. But the other phenomenon like *diffraction* and *interference* could not be explained on the basis of this theory.

Maxwell, in 1865 proposed that radiant energy (light) has wave characteristics. Light according to him is *Electromagnetic Wave* arising due to the disturbance created by electric and magnetic fields oscillating perpendicular to each other in space. Like all other mechanical waves, it is characterised by velocity, $c$ ; frequency, $\nu$ and wavelength, $\lambda$ which are related as :




Characteristics of wave

- **Wavelength ( $\lambda$ ):** The distance between two nearest, crests or troughs is known as wavelength. It is expressed in meters, picometers (pm), nanometers (nm) or Angstrom (Å) units ($1\ \text{pm} = 10^{-3}\ \text{nm} = 10^{-2}\ \text{Å} = 10^{-10}\ \text{cm} = 10^{-12}\ \text{m}$).
- **Frequency ( $\nu$ ):** The number of times a wave passes a given point in one second is known as frequency of the wave. It is expressed in Hz (hertz) or cps (cycles per second) unit (1 Hz = 1 cps).
- **Velocity (c):** The distance travelled by the wave in one second is known as velocity of the wave, it is expressed in $ms^{-1}$ and related to v as: $c = \nu\lambda$ [value of $c$ is constant and equal to $3 \times 10^8$ m/s] or $\nu = \dfrac{c}{\lambda}$ ...(i)
- **Wave number ( $\bar{\nu}$ ) :** It is reciprocal of wavelength i.e. the number of wavelengths per centimetre or

$$\bar{\nu} = \frac{1}{\lambda}\ \text{m}^{-1} \quad \text{...(ii)} \qquad \bar{\nu} = \frac{\nu}{c}\ \text{m}^{-1} \quad \text{...(iii)}$$

- **Amplitude (A) :** It is the height of the crest or depth of the trough of a wave. It determines the intensity (brightness) of the beam of light.

When the different types of electromagnetic radiations are arranged in the increasing or decreasing order of wavelengths or frequencies, the pattern so obtained is called an electromagnetic spectrum.

### Electromagnetic Spectrum

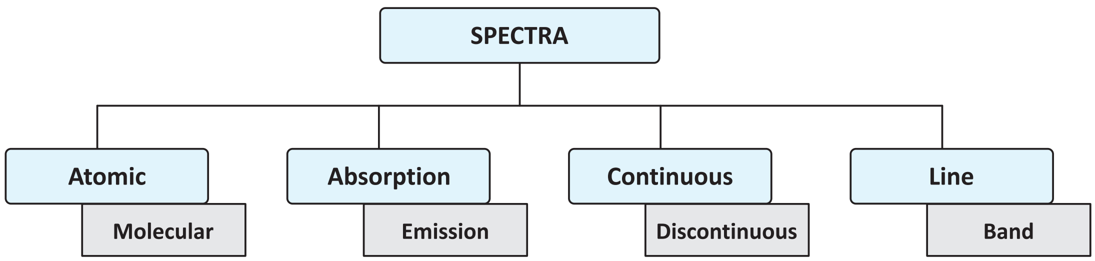
- Electromagnetic wave or radiation is not a single wavelength radiation, but a mixture of various wavelength or frequencies. All the frequencies have same speed (= $c$).
- If all the components of *Electromagnetic Radiation* (EMR) are arranged in order of decreasing or increasing wavelengths or frequencies, the pattern obtained is known as *Electromagnetic Spectrum*. The following table shows all the components of light.

Wavelength, $\lambda$ (in Å) 0.01 1.0 150 3800 7600 $6 \times 10^6$ $3 \times 10^9$

| Cosmic-rays | $\gamma$ rays | X-rays | UV rays | Visible light | Infra-red | Micro waves | Radio waves |
|---|---|---|---|---|---|---|---|

Frequency, $v$ (in Hz) $10^{17}$ $10^{15}$ $10^{14}$ $10^{11}$ $10^9$

**Types of Spectra :** Spectra can be classified into various type as expressed below :




**(i)** **Atomic and Molecular spectra :** Atomic spectra represents the radiation (energy emitted or absorb in space) by an atom whereas molecular spectra is associated with radiation (energy emitted or absorb in space) by a molecule. Molecular spectra is more complex than atomic spectra because molecule involve more electronic transitions than in atom.

**(ii)** **Absorption spectra and emission spectra :** Absorption spectra shows the wavelengths absorbed by an atom/molecule. The wavelengths absorbed appear as lines missing from the continuous spectra of radiation to which species is exposed. Each wavelength absorbed is characterstic of energy transition taking place in that species.

On the other hand, emission spectra indicates the wavelengths (or frequencies) emitted by an excited species. Each spectral line of emission spectra is characteristic of one electronic transition from higher energy level to a lower energy level (i.e., de-excitation). Absorption and emission spectra are complementary to each other. A line missing in absorption spectra will appear in emission spectra.

**(iii)** **Continuous and discontinuous spectra :** A continuous spectra is that which contains all the wavelength lying in a particular region of spectra. e.g., VIBGYOR indicates a continuous spectra in visible region.
On the other hand, if some wavelengths are missing and spectra contain certain wavelengths, it is called discontinuous spectra. e.g., VBYR is discontinuous spectra in which I, G & O wavelengths are missing from visible radiation.

**(iv)** **Line and band spectra :** Spectra containing only one or a few lines is called line spectra. It is obtained in the spectra of those species where number of electronic transitions is less. On the other hand, spectra containing several wavelengths is wide and is called band spectra. It is complex and is shown by multiatomic species in which number of electronic transitions is more.

**Note :** Light emitted from atoms heated in a flame or excited electrically in gas discharge tube, does not contain a continuous spread of wavelengths (or frequencies). It contains only certain well-defined wavelengths (or frequencies). The spectrum pattern appears as a series of bright lines (separated by gaps of darkness) and hence called as *Line Spectrum*.

One notable feature observed is, that each element emits a characteristic spectrum, suggesting that there is direct relation between the spectrum characteristics and the internal atomic structure of an atom.

### The Quantum Theory of Radiation

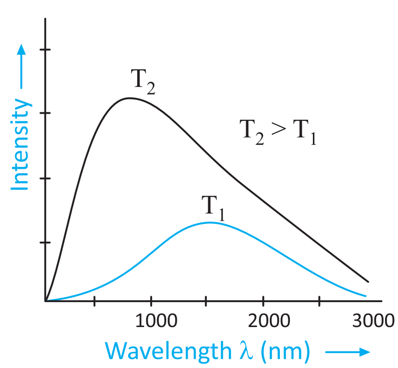
In 1901, Max Planck studied the distribution of the frequencies of radiations emitted from the hot bodies. He proposed a bold hypothesis that the radiant energy such as heat or light, is not emitted continuously but discontinuously in the form of small packets called as quanta. According to him, the energy of the electromagnetic radiation is directly proportional to the frequency of the radiation. The proportionality constant is called as *Planck's constant* ($h$). If energy of the radiation is $E$, and the frequency of the radiation is $\nu$, then we have :

$$E = h\,\nu \qquad \text{(The value of } h \text{ is } 6.626 \times 10^{-34}\ \text{J–s)}$$

If $n$ is the number of quanta of a particular frequency and TE be the total energy, then : $\text{TE} = n\,(h\nu)$

## BLACK BODY RADIATION (1900)

- In 1900, Max Planck was the first to give a concrete explanation for the phenomenon of black body radiation. According to the Plank's quantum theory, an ideal body is a perfect absorber and perfect emitter of radiation and called a black body.
- When such a body is heated it emits radiation over a wide range of wavelengths. For instance, when an iron rod is heated in a furnace, it firstly turns dull red, then progressively becomes more and more red as the temperature increases. On heating further, the radiation emitted becomes white and then blue as the temperature becomes very high.
- In terms of frequency, it means that radiation emitted goes from a lower frequency to higher frequency as the temperature increases. The red colour lies in the lower frequency area whereas blue light lies in the higher frequency area of the electromagnetic spectrum.
- The exact frequency distribution of emitted radiation from a black body depends only on its temperature. At a given temperature, the intensity of radiation increases with decrease in wavelength, reaches a maximum value at a given wavelength and then starts decreasing with further decrease in wavelength as shown in the figure below.




- Planck explained that atoms and molecules could emit (or absorb) energy only in discrete quantities (quantum) and not in anarbitrary manner as was believed at that time.

**Illustration - 1** *Find the ratio of frequencies of violet light ($\lambda_1$ = 4.10 × $10^{-5}$cm) to that of red light ($\lambda_2$ = 6.56 × $10^{-5}$ cm). Also determine the ratio of energies carried by them.*

**SOLUTION :**

Using $c = \nu\lambda$

where $c$: speed of light; $\nu$ : frequency; $\lambda$ : wavelength

$$\frac{\nu_1}{\nu_2} = \frac{\lambda_2}{\lambda_1} \qquad [1 : \text{violet and } 2 : \text{red}]$$

$$\Rightarrow \quad \frac{\nu_1}{\nu_2} = \frac{6.56 \times 10^{-5}}{4.10 \times 10^{-5}} = 1.6 : 1$$

Now the energy associated with electromagnetic radiation is given by $E = h\nu$

$$\Rightarrow \quad \frac{E_1}{E_2} = \frac{\nu_1}{\nu_2} = \frac{\lambda_2}{\lambda_1} = 1.6 : 1$$

Hence the ratio of energies is same as that of frequencies.

**Illustration - 2** *A 100 W power source emits green light at a wavelength $\lambda$ = 5000 Å. How many photons per minute are emitted by the source ?*

**SOLUTION :**

Energy given out by the source per sec = Power (P)

$\Rightarrow$ Energy given by source in $t$ sec = P × t

As $\lambda$ = 5000 Å, the energy per photon of green light is given by : $h\nu = \dfrac{hc}{\lambda}$

$\Rightarrow$ Number of photons ($n$) emitted in time $t$ sec is given by : $n = \dfrac{Pt}{(hc/\lambda)} = \dfrac{Pt\lambda}{hc}$

Using P = 100 J/s, $\lambda = 5000 \times 10^{-10}$ m and t = 60 s

$\Rightarrow$ Number of photons ($n$) :

$$= \frac{100(60)\left(5000 \times 10^{-10}\right)}{\left(6.626 \times 10^{-34}\right)\left(3 \times 10^{8}\right)} = 1.5 \times 10^{22}$$

## Photoelectric Effect :

It was observed by Hertz and Lenard around 1880 that when a clean metallic surface is irradiated by monochromatic light of proper frequency, electrons are emitted from it. This phenomenon of ejection of the electrons from metal surface was called as *Photoelectric Effect*.

- It was observed that if the frequency of incident radiation is below a certain minimum value (*threshold frequency*), no emission takes place however high the intensity of light may be.
- Another important feature observed was that the kinetic energy of the electrons emitted was independent of the intensity of the light. The kinetic energy of the electrons increase linearly with the frequency of incident light radiation. This was highly contrary to the laws of Physics at that time i.e. *the energy of the electrons should have been proportional to the intensity of the light, not to the frequency*.

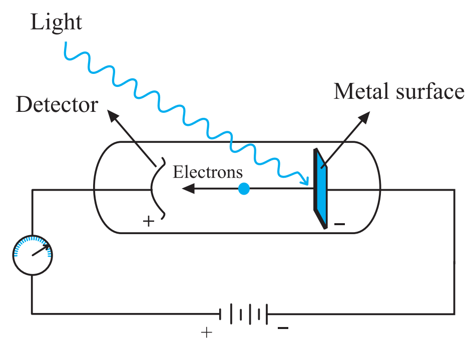
*These features could not be properly explained on the basis of Maxwell's concept of light i.e. light as electromagnetic wave.*

In 1905, Einstein applied Planck's quantum theory of light to account for the extraordinary features of the photoelectric effect. *He introduced a new concept that light shows dual nature. In phenomenon like reflection, refraction and diffraction, it shows wave nature and in phenomenon like photoelectric effects, it shows particle nature*. According to the particle nature, the energy of the light is carried in discrete units whose magnitude is proportional to the frequency of the light wave. These units were called as *photons* (or *quanta*).

According to Einstein, when a quantum of light (photon) strikes a metal surface, it imparts its energy to the electrons in the metal. In order for an electron to escape from the surface of the metal, it must overcome the attractive force of the positive ions in the metal. So a part of the photon's energy is absorbed by the metal surface to release the electron, this is known as *work function* of the surface and is denoted by $W_0$. The remaining part of the energy of the photon goes into the kinetic energy of the electron emitted. If $E_i$ is the energy of the photon, $KE$ is the kinetic energy of the electron and $W_0$ be the work function of the metal then we have ;

$$E_i = KE + W_0$$ (This is known as *Einstein's photoelectric equation*)

For each metal, there is a characteristic minimum frequency known as the *threshold frequency* ($\nu_0$) below which the photoelectric effect does not occur. Electrons are emitted only after the frequency of light is equal to or above the threshold frequency. The threshold frequency is proportional to the work function of the metal. If $\nu_0$ be the threshold frequency and $\nu$ the frequency of incident light, $E$ is energy of incident light, then we have :



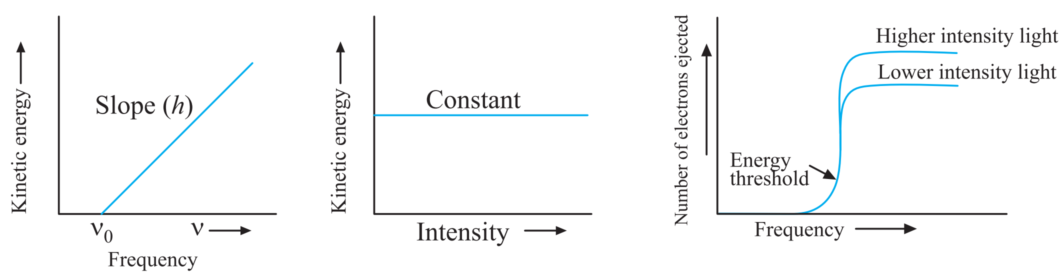

$$W_0 = h\nu_0 \quad \text{and} \quad E_i = h\nu$$

$$\Rightarrow \quad KE = E_i - W_0 \quad \text{or} \quad KE = h\nu - h\nu_0 = h(\nu - \nu_0)$$

Also, if $m$ be the mass and $v$ be the velocity of the electron ejected then $KE = \frac{1}{2}mv^2 = h(\nu - \nu_0)$

***Note :***

- The Electromagnetic radiation (or wave) now emerges as an entity which shows dual nature i.e. sometimes as *Wave* and sometimes as *Particle* (quantum aspect).
- The energy of an individual photon depends only on its frequency and not on the intensity of the light beam. The intensity of a light beam is a measure of the number of photons in the beam and not of the energies of those photons. A low-intensity beam of high-energy photons might easily knock out electrons from a metal but a high intensity beam of low energy photons might not be able to knock out a single electron.




- Sometimes, it is convenient to calculate energy (in eV) of a photon in short form using :

$$E_p = \frac{hc}{\lambda} = \frac{12400}{\lambda(\text{in Å})}\,eV \equiv \frac{1240}{\lambda(\text{in nm})}\,eV$$

**Illustration - 3** *Calculate the velocity of electron ejected from platinum surface when radiation of* 200 *nm falls on it. Work function of platinum is* 5 *eV*. (1 *eV* = $1.6 \times 10^{-19}$ *J*)

**SOLUTION :**

Using Einstein's photoelectric equation :

$E_i = KE + W_0$

where $E_i$ : energy of incident radiation ;

KE : kinetic energy of ejected electron

$W_0$ : work function of metal

$E_i = \frac{1240}{200} eV = 6.2 eV$ ; and $W_0 = 5$ eV

$\Rightarrow \quad KE = E_i - W_0 = (6.2 - 5)$ eV $= 1.2$ eV

$= 1.2 \times 1.6 \times 10^{-19} J = 1.94 \times 10^{-19}$

Now, $KE = \frac{1}{2}mv^2 \quad \Rightarrow \quad v = \sqrt{\frac{2KE}{m}}$

$\Rightarrow \quad v = \sqrt{\frac{2(1.94 \times 10^{-19})}{(9.1 \times 10^{-31})}} = 6.52 \times 10^5$ m/s

**Illustration - 4** *A photon of light with $\lambda = 400$ nm falls on a metal surface. As a result, photoelectrons are ejected with a velocity of $6.4 \times 10^5$ m/s. Find :*

**(a)** *the kinetic energy of emitted photoelectrons,* **(b)** *the work function (in eV) of the metal surface.*

**SOLUTION :**

**(a)** Kinetic energy of electron $= \frac{1}{2}mv^2$

$\Rightarrow \quad KE = \frac{1}{2}(9.1 \times 10^{-31})(6.4 \times 10^5)^2 = 1.86 \times 10^{-19}$ J

$= 1.16$ eV

**(b)** From Einstein's photoelectric equation :

$E_i = KE + W_0 \quad \Rightarrow \quad W_0 = E_i - KE$

$\Rightarrow \quad W_0 = \frac{1240}{400} - KE = 3.1 - 1.16 = 1.94$eV

$\Rightarrow \quad W_0 = 1.94$eV

## IN-CHAPTER EXERCISE - A

**1.** *Suppose $10^{-17}$ J of light energy is needed by the interior of the human eye to see on object. How many photons of green light ($\lambda = 550$ nm) are needed to generate this minimum amount of energy ?*

**2.** *A photon of 300 nm is absorbed by a gas and then re-emits two photons. One re-emitted photon has wavelength 400 nm. Calculate energy of other photon re-emitted out.*

**3.** *The energy required to stop the ejection of electrons from Cu plate is 0.24 eV, when the radiation of wavelength $\lambda = 300$ nm strikes the plate. Calculate the work function of Cu.*

**4.** *The minimum energy required to overcome the attractive forces between electron and the surface of Ag metal is $7.52 \times 10^{-19}$ J. What will be maximum KE of an electron ejected out from Ag which is being exposed to UV radiations of $\lambda = 360 \times 10^{-10}$ m ?*

**5.** *An oil drop has $6.39 \times 10^{-19}$ C charge. How many electrons does this oil drop has ?*

**Choose the correct option for each of the following.**

**6.** *The ratio of e/m, i.e., specific charge for a cathode ray :*

**(A)** *has the smallest value when the discharge tube is filled with $H_2$*
**(B)** *is constant*
**(C)** *varies with the atomic number of gas in the discharge tube*
**(D)** *varies with the atomic number of an element forming the cathode*

**7.** *Consider a 20 W light source that emits monochromatic light of wavelength 600 nm. The number of photons ejected per second in form of Avogadro's constant $N_{AV}$ is approximately :*

**(A)** $N_{AV}$ **(B)** $10^{-2} N_{AV}$ **(C)** $10^{-4} N_{AV}$ **(D)** $10^{-6} N_{AV}$

**8.** *Rutherford's experiment, which established the nuclear model of the atom, used a beam of :*

**(A)** *$\beta$-particles, which impinged on a metal foil and got absorbed*
**(B)** *$\gamma$-rays, which impinged on a metal foil and ejected electrons*
**(C)** *Helium atoms, which impinged on a metal foil and got scattered*
**(D)** *Helium nuclei, which impinged on a metal foil and got scattered.*

**9.** *Of the following, radiation with maximum wavelength is :*

**(A)** *UV* **(B)** *Radio wave* **(C)** *X-rays* **(D)** *IR*

**10.** *When a certain metal was irradiated with light of frequency $3.2 \times 10^{16}$ Hz, photoelectrons emitted had twice the kinetic energy as did photoelectrons emitted when the same metal was irradiated with light of frequency $2.0 \times 10^{16}$ Hz. Hence threshold frequency is :*

**(A)** $0.8 \times 10^{15}\,Hz$ **(B)** $8.0 \times 10^{15}\,Hz$ **(C)** $0.8 \times 10^{14}\,Hz$ **(D)** $6.4 \times 10^{16}\,Hz$

**11.** *How may photons are emitted per second by a 10 mW laser source operating at 626 nm ?*

**(A)** $1.6 \times 10^{16}$ **(B)** $1.6 \times 10^{18}$ **(C)** $3.2 \times 10^{16}$ **(D)** *None of these*

**12.** *If the frequency of light in a photoelectric experiment is doubled, the stopping potential will :*

**(A)** *be doubled* **(B)** *be halved*
**(C)** *become more than double* **(D)** *become less than double*

## ATOMIC SPECTRA OF HYDROGEN AND BOHR'S MODEL SECTION - 3

It is observed that the atoms of hydrogen in gas discharge tube emit radiations whose spectrum shows line characteristics (line spectra). The line spectra of hydrogen lies in three regions of Electromagnetic Spectrum: *Infra-red*, *Visible and UV* region. In all there are five sets of discrete lines.

The set of lines in the *Visible* region are known as *Balmer Series*, those in *Ultra-Violet* as *Lyman series* and there are three sets of lines in *Infra-red* region : *Paschen*, *Brackett* and *Pfund series*. Balmer and Rydberg gave an empirical relation to define the wavelength of the lines in each series in terms of a parameter called as *Wave Number* denoted by $\bar{\nu}$. The wave number is defined as reciprocal of the wavelength i.e., $\bar{\nu} = \frac{1}{\lambda}$

$$\bar{\nu} = RZ^2 \left( \frac{1}{n^2} - \frac{1}{m^2} \right)$$

where *n* and *m* are whole numbers; $\lambda$ : wavelength of spectral line ; $\bar{\nu}$ : wave number of spectral line
*R* : Rydberg constant. The values of *n* and *m* for different spectral lines for each series are listed below.

| Region | Spectral line | *n* | *m* |
|---|---|---|---|
| *UV* | *Lyman Series* | 1 | 2, 3, 4, . . . |
| *Visible* | *Balmer Series* | 2 | 3, 4, 5, . . . |
| *Infra-red* | *Paschen Series* | 3 | 4, 5, 6, . . . |
| *Infra-red* | *Brackett Series* | 4 | 5, 6, . . . . |
| *Infra-red* | *Pfund Series* | 5 | 6, 7, . . . . |
| | *Humphry Series* | 6 | 8, 7, 8 . . . |