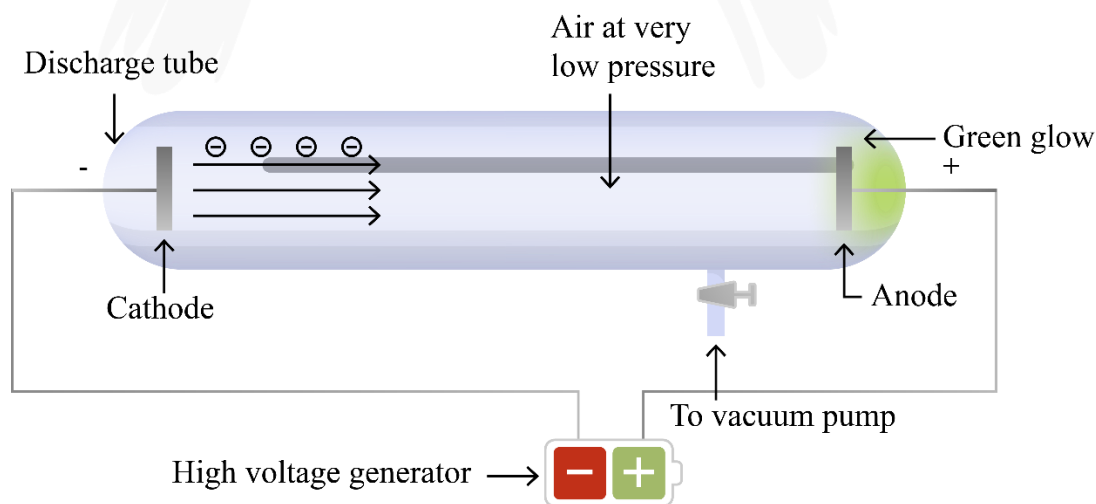


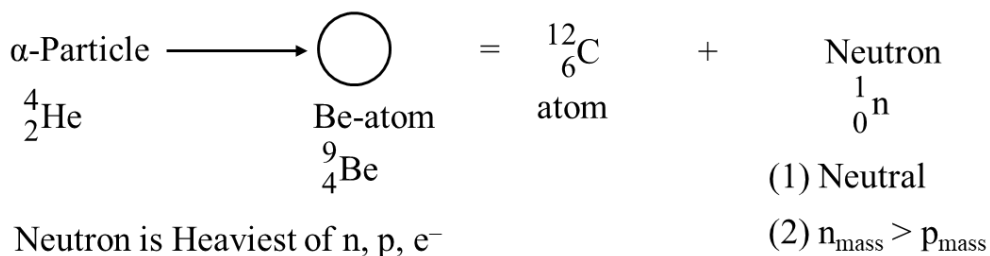
Electron	Proton
Electron: J.J Thomson Apparatus used: Cathode ray tube	Scientist: Goldstein Apparatus used: Cathode ray tube.
(1) Rays moving from <b>cathode to anode</b> are negatively charged $e^-$ . $\textcircled{C} \xrightarrow{e^-} \textcircled{A}$	(1) Rays moving from <b>anode to cathode</b> are positively charged anions. $\textcircled{A} \xrightarrow[ions]{+ve} \textcircled{C}$
(2) Effect of gases and electrode material: on changing the constituting gases and the material of electrode there is no change in the cathode rays.	(2) On changing the gases there is subsequent change observed for heavier and light gases accordingly. i.e. size of cation changes
<b>Characteristics:</b>	<b>Characteristics:</b>
(1) They move from cathode to anode.	(1) They move from anode to cathode.
(2) As they are not visible fluorescent screen is used (ZnS, Ag)	(2) As they are positively charged so they move from anode to cathode.
(3) The travel in a straight line in the absence of electric and magnetic field.	(3) It depends on the gas taken in the tube.
(4) Cathode rays has mass and velocity.	(4) The mass of the atomic positive particle is same as that of the atomic mass of gas taken.
(5) Ionisation occurs and thus produces electron $A \longrightarrow A^+ + e^-$	(5) $A \xrightarrow{1e^-} A^+$ ; $B \xrightarrow{2e^-} B^{+2}$ Depends on the electron lost. $H \longrightarrow H^+ + e^-$
(6) As they are negatively charged they get deflected due to electric and magnetic field.	
(7) They produce X-rays.	



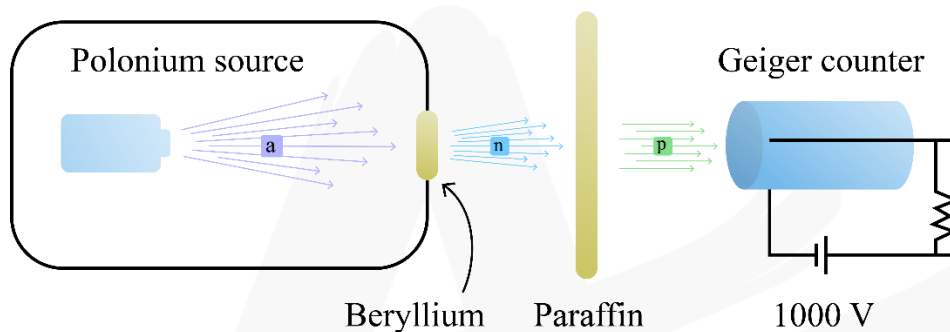
Production of cathode rays  
Cathode Rays Discharge Tube

## Discovery of Neutron.

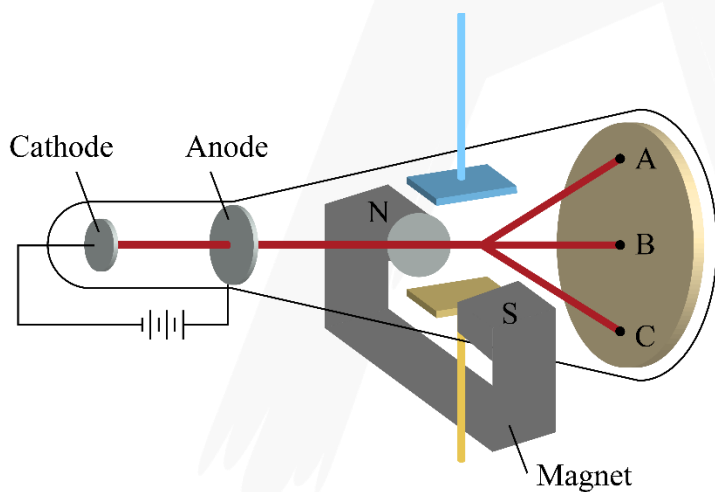
Scientist: Chadwick.



### Discovery of neutron



### Note



The apparatus to determine the charge to the mass ratio of electron

Charge to mass ratio:  $\frac{e}{m_e} = 1.75 \times 10^{11} \text{ C kg}^{-1}$ .

Charge on e<sup>-</sup>:  $-1.6 \times 10^{-19} \text{ C}$

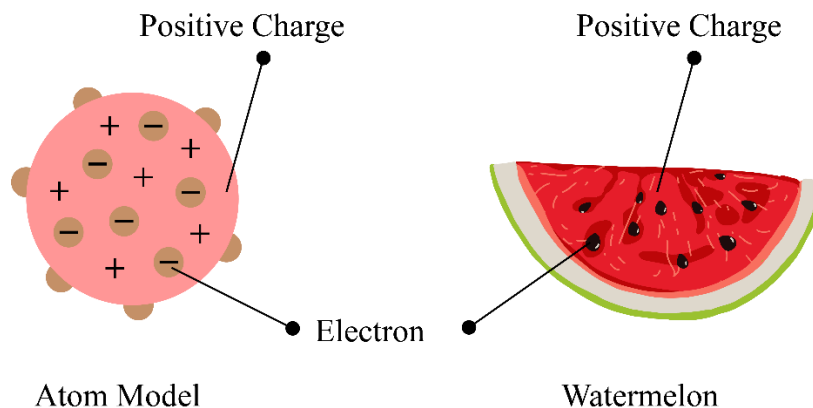
Mass of e<sup>-</sup>:  $9.1 \times 10^{-31} \text{ kg}$

Ratio of  $m_{e^-}/m_H$ : 1:1831

**Proton:** Positively charged particle having a positive unit charge and mass equal to that of 'H' atom.

### Atomic models:

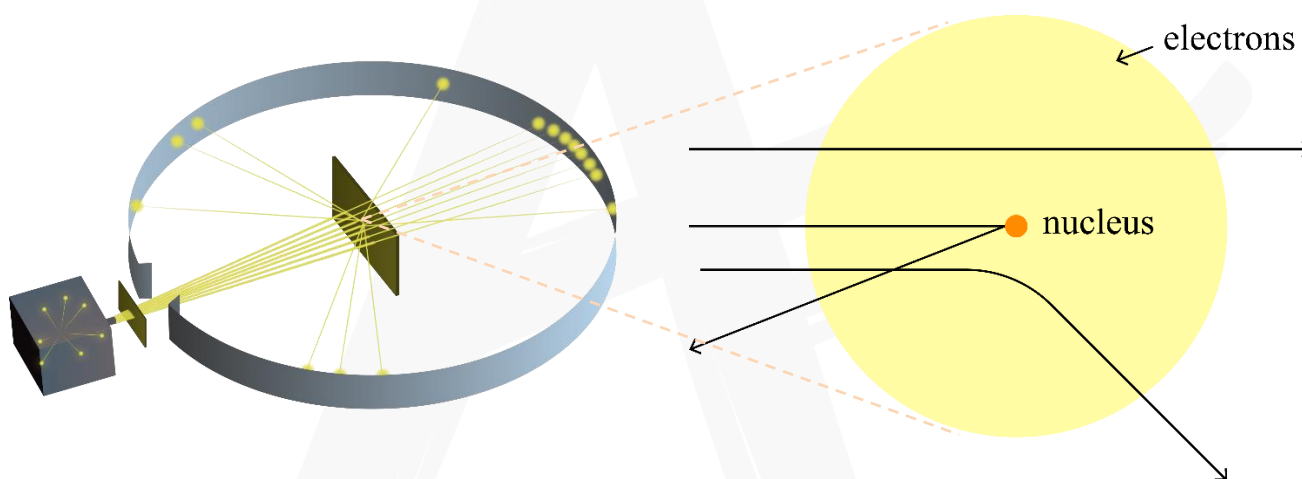
**Thomson's model:** (Plum pudding model)



Spherical model having diffused positive charge and negative charge which are unevenly distributed.

Drawback: It could not explain the Rutherford's experiment.

### Rutherford's model



#### Observation:

- (1) Most of the space in atom is empty.
- (2) Few are deflected at small angles.
- (3) Very few get reverted back at  $180^\circ$  angle.

**Experiment:**  $\alpha$ -particles were allowed to fall on gold foil, guarded by the circular screen.

**Conclusions:** Heavy (positively) charged body in a very small portion is concentrated at the centre called as nucleus.

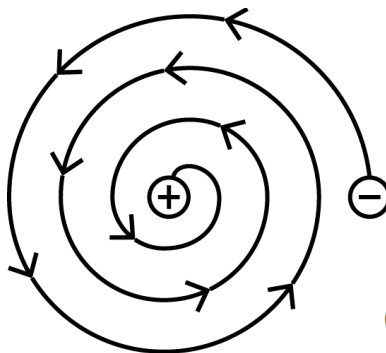
### Rutherford's atomic model

(i) **Nucleus:** Central positive charged body with density. It consists of protons and neutrons i.e. nucleons.

(ii) **Extra nuclear part:** Here  $e^-$  revolve around the nucleus like planets around the sun.

**Drawback:** It could not explain stability.

(1) **Stability of atoms:** According to Maxwell's electromagnetic theory any charged body under acceleration radiate energy in the form of electromagnetic radiations. As a result,  $e^-$  will loses its energy and will collapse spirally into the nucleus, but this do not happen i.e. failed to explain stability.

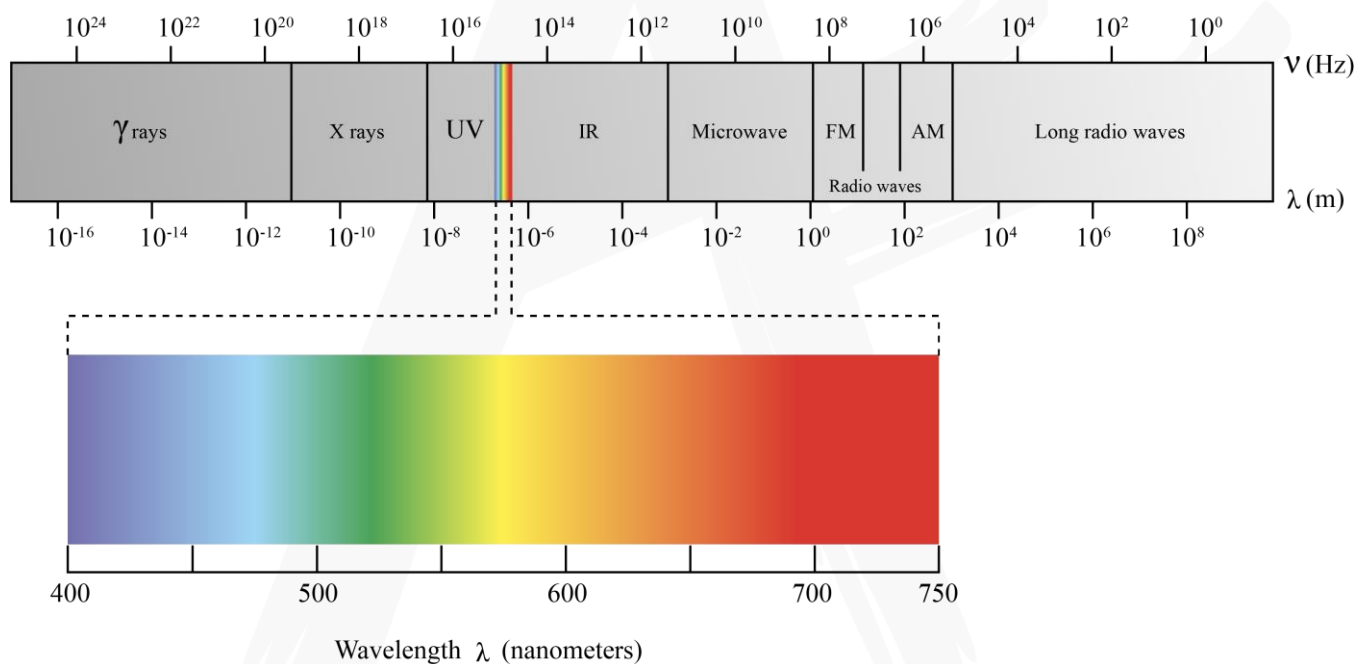


(Practice Question in the End, Q.8.9)

**(2) Electronic structure of atom:** He failed to explain how the electrons are distributed around the nucleus and what are the energies of this electrons.

### EM waves

- They are associated with electric and magnetic fields which moves with the velocity of light.
- They need no medium for propagation.



→ **Wavelength increases** →

→ **Frequency decreases** →

**Radioactivity:** Spontaneous emission of  $\alpha$ ,  $\beta$ ,  $\gamma$  radiations from some elements.

$\alpha \rightarrow$  He nuclei  $\text{He}_2^4$

$\beta \rightarrow$  Negative particles  $\beta_{-1}^0$

$\gamma \rightarrow$  form of energy (carries no mass)

Penetrating power:  $\gamma > \beta > \alpha$ .

### Some important terms:

(1) **Atomic no. (Z):** No. of protons.

(2) **Mass no. (A):** (No. of n + No. of p)

(3) **Isotopes of H:**  $H_1^1, H_1^2, H_1^3$

Atoms having same (Z) but different (A)

(4) **Isobars:** Atoms having different (Z) but same (A)

(5) **Isodiapheres:** Species having same isotopic excess i.e. have diff (Z) & (A) but same neutron excess i.e. Ex:  $Th_{234}$  & Uranium 238 (n - p)

(6) **Isosteres:** Species having same no of atoms as well as  $e^-$ .

Ex: CO,  $N_2$ .

(7) **Isotones:** Atoms having same no. of neutrons ( $A - Z$ )

(8) **Isoelectronic:** Species having same no of  $e^-$ . Ex:  $O^{2-}$ ,  $Na^+$ .

**Q.** Calculate the number of protons, neutrons and electrons in  ${}^{80}_{35}Br$

NCERT

**Sol.** In this case,  ${}^{80}_{35}Br$ ,  $Z = 35$ ,  $A = 80$ , species is neutral

Number of protons = number of electrons =  $Z = 35$

Number of neutrons =  $80 - 35 = 45$ , (equation 2.4)

**Q.** The number of electrons, protons and neutrons in a species are equal to 18, 16 and 16 respectively. Assign proper symbol to the species?

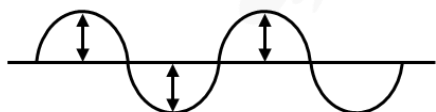
NCERT

**Sol.** The atomic number is equal to number of protons = 16. The element is sulphur (S).

Atomic mass number = number of protons + number of neutrons =  $16 + 16 = 32$

Species is not neutral as the number of protons is not equal to electrons. It is anion (negatively charged) with charge equal to excess electrons =  $18 - 16 = 2$ . Symbol is  ${}^{32}_{16}S^{2-}$ .

### Wave Characteristics:



(Practice Question in the End, Q.1)

(1) **Amplitude:** Height of crest or depth of trough. It decides brightness of the wave & intensity of water. Unit: m or cm.

(2) **Frequency ( $\nu$ ):** No. of waves per unit time  $1 \text{ Hz} = 1 \text{ cycle s}^{-1}$ .

(3) **Wave no ( $\bar{\nu}$ ):** No. of waves per unit length =  $1/\lambda$   
unit:  $m^{-1}$ .

(4) **Velocity ( $v$ ):** Distance travelled per unit time  $C = \nu\lambda$ .

(5) **Wavelength ( $\lambda$ ):** Distance between two adjacent points in same phase.

**Q.** 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? **NCERT**

**Sol.** The wavelength,  $\lambda$ , is equal to  $c/\nu$ , where  $c$  is the speed of electromagnetic radiation in vacuum and  $\nu$  is the frequency. Substituting the given values, we have

$$\begin{aligned}\lambda &= \frac{c}{\nu} \\ &= \frac{3.00 \times 10^8 \text{ ms}^{-1}}{1368 \text{ kHz}} \\ &= \frac{3.00 \times 10^8 \text{ ms}^{-1}}{1368 \times 10^5 \text{ s}^{-1}} = 219.3 \text{ m}\end{aligned}$$

This is a characteristic radiowave wavelength.

**Q.** The wavelength range of the visible spectrum extends from violet (400 nm) to red (750 nm). Express these wavelength in frequencies (Hz). ( $1 \text{ nm} = 10^{-9} \text{ m}$ ) **NCERT**

**Sol.** Using equation 2.5, frequency of violet light

$$\nu = \frac{c}{\lambda} = \frac{3.00 \times 10^8 \text{ ms}^{-1}}{400 \times 10^{-9} \text{ m}} = 7.50 \times 10^{14} \text{ Hz}$$

Frequency of red light

$$\nu = \frac{c}{\lambda} = \frac{3.00 \times 10^8 \text{ ms}^{-1}}{750 \times 10^{-9} \text{ m}} = 4.00 \times 10^{14} \text{ Hz}$$

The range of visible spectrum is from  $4.0 \times 10^{14}$  to  $7.5 \times 10^{14}$  Hz in terms of frequency units.

**Q.** Calculate (a) wavenumber and (b) frequency of yellow radiation having wavelength 5800 Å.

**Sol.** (a) Calculation of wavenumber ( $\bar{\nu}$ ) **NCERT**

$$\begin{aligned}\lambda &= 5800 \text{ Å} = 5800 \times 10^{-8} \text{ cm} \\ &= 5800 \times 10^{-10} \text{ m}\end{aligned}$$

$$\begin{aligned}\bar{\nu} &= \frac{1}{\lambda} = \frac{1}{5800 \times 10^{-10} \text{ m}} \\ &= 1.724 \times 10^6 \text{ m}^{-1} \\ &= 1.724 \times 10^4 \text{ cm}^{-1}\end{aligned}$$

(b) Calculation of the frequency ( $\nu$ )

*(Practice Question in the End, Q.2)*

$$\nu = \frac{c}{\lambda} = \frac{3 \times 10^8 \text{ ms}^{-1}}{5800 \times 10^{-10} \text{ m}} = 5.172 \times 10^{14} \text{ s}^{-1}$$

## Dual character of radiation Energy (Electromagnetic wave Theory)

Wave character (continuous)	Particle character (discontinuous)
Interference	Black Body Radiation
Diffraction	Photoelectric effect

**Interference:** It is the combination of two waves of same or different frequencies to give a new wave.

**Diffraction:** Bending of wave around obstacles.

### Limitations of Electromagnetic wave theory

EM wave theory was successful in explaining the properties of light such as interference and diffraction etc but it could not explain the following.

- (i) The phenomenon of the black body radiation.
- (ii) The photoelectric effect.
- (iii) The variation of heat capacity of solids.
- (iv) The line spectrum of atoms with special reference to hydrogen.

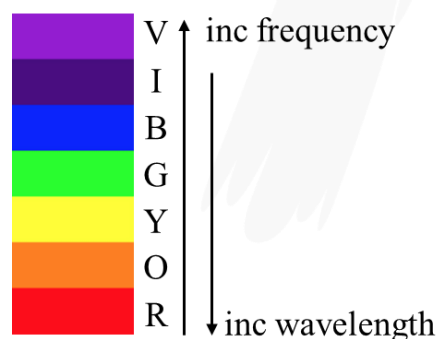
This can only be explained only if EM waves are supposed to have particle nature.

### (1) Black body radiation

If any substance with high melting point is heated it becomes hot and first becomes red, then yellow and finally white and blue in the end (an iron rod).

- If the substance being heated is a black body, the radiation emitted is called as black body radiation.

**Note:** Black body: perfect absorber and radiator.

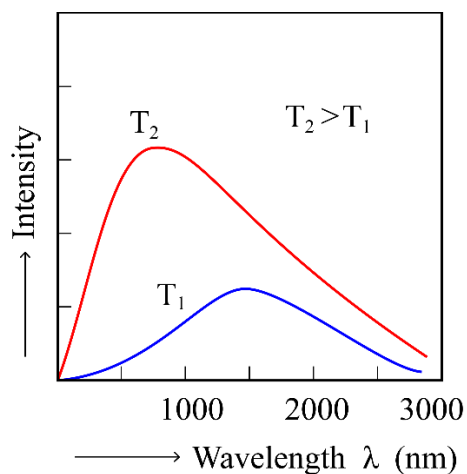


∴ Its clear that radiations emitted goes from higher wavelength to lower or lower frequency to higher frequency.

i.e. frequency  $\propto T \propto$  intensity

The change in colour shows that on heating the frequency of the radiation emitted increases with increase in temperature.

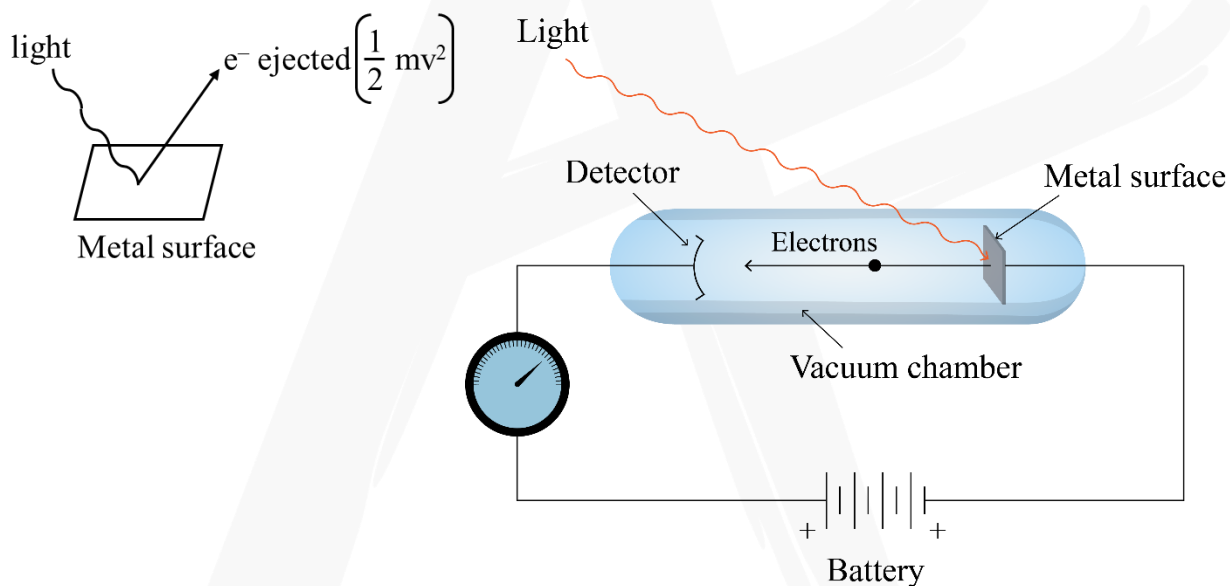




At a given temperature intensity of the radiation increases with decreases in wavelength reaches max at a particular ( $\lambda$ ) and starts decreasing with further decreases in ( $\lambda$ ).

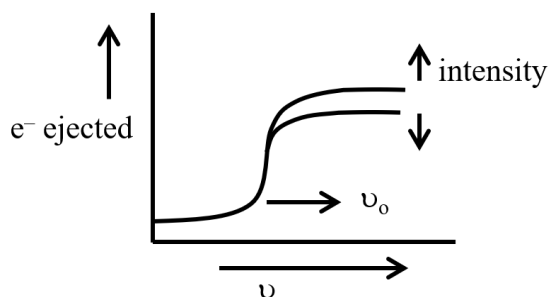
## (2) Photo electric effect:

When radiation with frequency greater than a certain minimum frequency ( $\nu_0$ ) strike the surface of a metal, the electrons are ejected photo electric effect. The emitted electrons are photoelectrons.

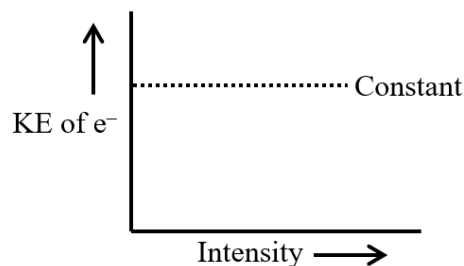
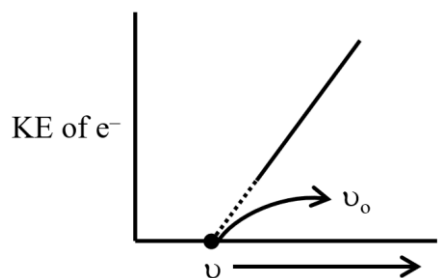


### Observation:

- (1) No time lag between the striking of light and the ejection of electrons.
- (2) **Threshold frequency:** Minimum frequency of light to cause photo electric effect ( $\nu_0$ ).
- (3) **No of electrons:** Photo electrons ejected  $\propto$  intensity of light striking the metal.



- (4) **Kinetic Energy of photo electrons:** It depends on the frequency of light used.



**Conclusion:** No of  $e^-$  ejected  $\propto$  intensity of light but KE of  $e^-$  ejected  $\propto$  frequency of light.

Ex: when red light of high intensity strikes the metal surface (k) it is not able to eject  $e^-$  but when yellow light having high frequency strikes the potassium surface  $e^-$  is ejected i.e.  $\nu_0$  of k atom is  $\uparrow$  greater than the frequency possessed by red light.

### Explanation for photo electric effect/ Einstein's photoelectric equation.

When a light source of particular  $\nu$  strikes the metal surface it transfer its energy to the  $e^-$  and thus the electron gets emitted. A certain  $\nu$  of light is required to eject this  $e^-$  it is termed as  $\nu_0$  (Threshold frequency). If the  $\nu > \nu_0$ , then the excess energy.  $(h\nu - h\nu_0)$  is imported to  $e^-$  as KE.

$$\therefore KE = h\nu - h\nu_0$$

$$\therefore \frac{1}{2}mv^2 = h\nu - h\nu_0$$

$$\therefore h\nu = \frac{1}{2}mv^2 + h\nu_0$$

$$\therefore h\nu = \frac{1}{2}mv^2 + w_0 \quad \{\because w_0 \text{ is the work function}\}$$

This is the Einstein photoelectric equation.

### Note:

1. Energy acquired by an  $e^-$  when it is accelerated through a potential difference of 1V is called as "1eV"

$$\begin{aligned} 1\text{eV} &= \text{charge on } e^- \times v = 1.602 \times 10^{-19} \text{ CV} \\ &= 1.602 \times 10^{-19} \text{ J.} \end{aligned}$$

2. The energy possessed by one mole of quanta (photons) ie  $N_A$  quanta's is called as "one Einstein"

$$1 \text{ Einstein of energy} = N_A h\nu = N_A h \frac{c}{\lambda}$$

$h$  = plank's constant

$$= 6.626 \times 10^{-27} \text{ ergs}$$

$$= 6.626 \times 10^{-34} \text{ Joule sec}$$

"Quantum": The radiant energy is emitted or observed not continuously but discontinuously in the form of small discrete packets of energy. This each packet is called as "quantum".

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

**Sol.** Power of the bulb = 100 watt

$$= 100 \text{ J s}^{-1}$$

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

$$= \frac{6.626 \times 10^{-34} \text{ J s} \times 3 \times 10^8 \text{ ms}^{-1}}{400 \times 10^{-9} \text{ m}} = 4.969 \times 10^{-19} \text{ J}$$

Number of photons emitted

$$\frac{100 \text{ J s}^{-1}}{4.969 \times 10^{-19} \text{ J}} = 2.012 \times 10^{20} \text{ s}^{-1}$$

**Q.** When electromagnetic radiation of wavelength 300 nm falls on the surface of sodium, electrons are emitted with a kinetic energy of  $1.68 \times 10^5 \text{ J mol}^{-1}$ . 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? NCERT

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

$$h\nu = hc / \lambda$$

$$= \frac{6.626 \times 10^{-34} \text{ J s} \times 3.0 \times 10^8 \text{ m s}^{-1}}{300 \times 10^{-9} \text{ m}} = 6.626 \times 10^{-19} \text{ J}$$

The energy of one mole of photons

$$= 6.626 \times 10^{-19} \text{ J} \times 6.022 \times 10^{23} \text{ mol}^{-1}$$

$$= 3.99 \times 10^5 \text{ J mol}^{-1}$$

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

$$= (3.99 - 1.68) \times 10^5 \text{ J mol}^{-1}$$

$$= 2.31 \times 10^5 \text{ J mol}^{-1}$$

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

$$\therefore \lambda = \frac{hc}{E}$$

$$= \frac{6.626 \times 10^{-34} \text{ J s} \times 3.0 \times 10^8 \text{ ms}^{-1}}{3.84 \times 10^{-19} \text{ J}} = 517 \text{ nm}$$

(This corresponds to green light)

*(Practice Question in the End, Q.3, 10)*

## Dual Nature of Electromagnetic Radiations

From the above theory it can be concluded that some properties of light i.e. interference, diffraction can be explained if the light has wave nature while, some other properties of light like Black Body Radiation & Photoelectric Effect can be explained if light have particle nature thus light have a dual nature.

### Emission and adsorption spectra

An instrument used to separate the radiations of different wavelength or frequencies called spectroscopy which contains prism which produce spectra and the study of this spectra is called as spectroscopy

This spectra is broadly classified into

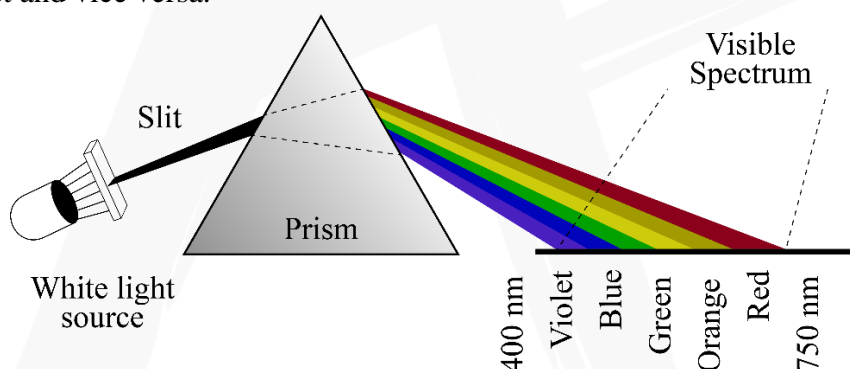
#### 1. Emission spectra

#### 2. Absorption spectra.

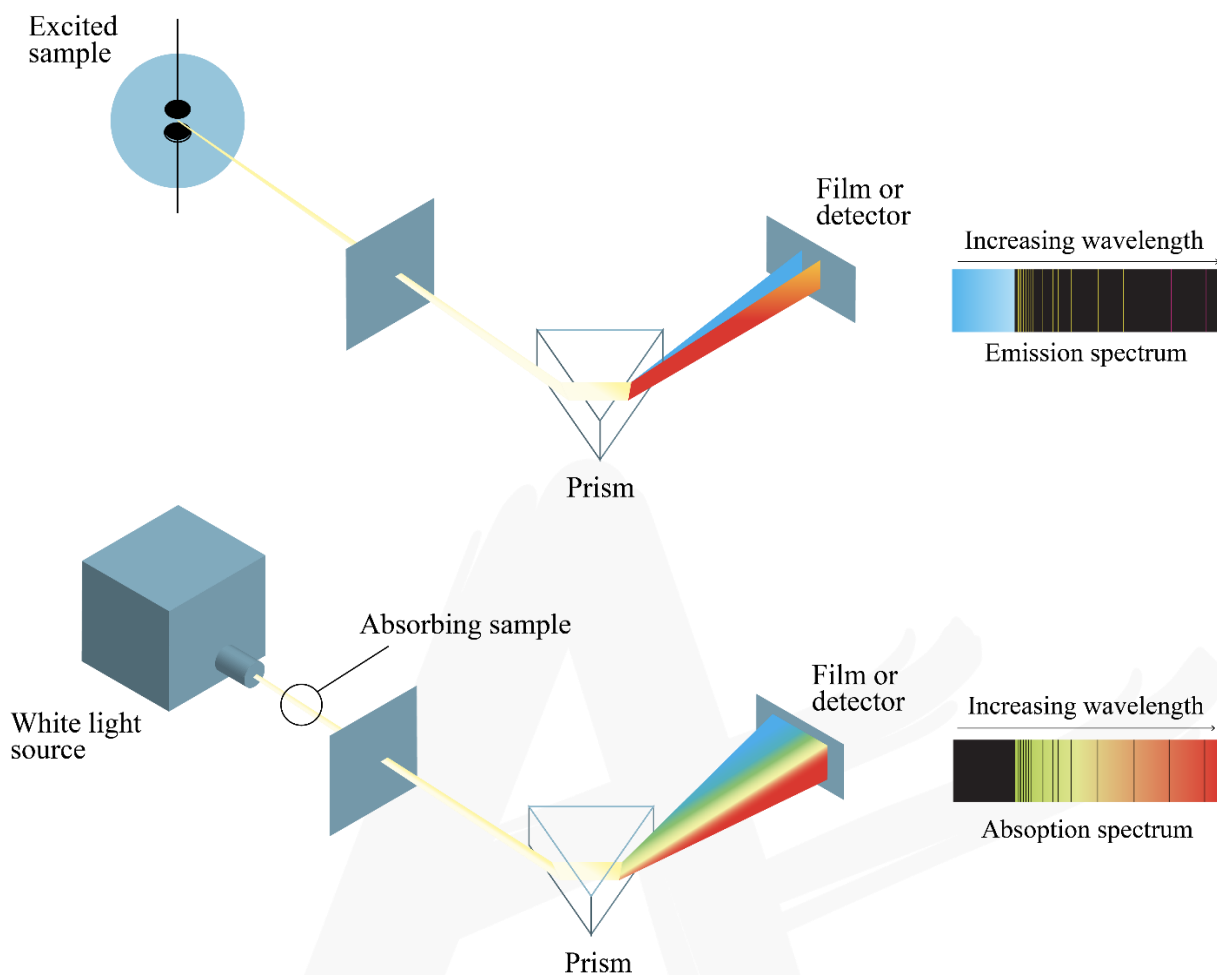
1. **Emission spectra:** When radiation passed through prism and then received on the photo graphic plate, the spectrum obtained is “emission spectrum”.

Depending upon the source of radiation emission spectra are of two types.

- i. **Continuous spectra:** When white light from any process is passed through a prism, it splits into seven different wide bands of colours from violet to red. (VIBGYOR) this spectra is so continuous that each band merge into other. It is noted that red colour with the longest wavelength deviates least and vice versa.



- ii. **Line Spectrum or line spectrum:** When some volatile salt ( $\text{NaCl}$ ) is placed in Bunsen flame light is emitted. The colour of the light emitted depends on the nature of substance ex: Na emits yellow while K emits violet light, when this light is passed the spectroscopy it is observed that no continuous spectrum is obtained but some isolated coloured lines are observed on the photographic plate separated from each other by dark spaces this is line spectrum. The spectra of these lines of the elements is termed as finger prints.
2. **Absorption Spectra:** When white light from any source is passed through a solution or vapours of the substance and then analysed by the spectroscopy, it is observed that some dark lines are seen on the plate, this is due to the absorption of certain light of particular frequency or wavelength is observed in the solution. These dark spaces are on the same place where colour was observed in continuous spectrum, this can be termed as the photographic negative of the continuous spectrum. The spectrum thus obtained is called as absorption spectrum.



**Note:** Difference between the adsorption and emission spectra.

Emission		Adsorption	
1.	Emission Spectra is obtained when radiation is directly analysed by source to spectrocope	1.	Adsorption Spectra is obtained when white light is passed through the substance and then analysed by the spectrocope
2.	It consist of bright coloured lines separated by dark spaces	2.	It consist of dark lines separated by the coloured lines.
3.	It can be continuous or discontinuous	3.	It is always discontinuous.

### Emission spectrum of Hydrogen:

When hydrogen gas at low pressure is taken in the discharge tube and the light emitted on passing electric discharge is examined with spectrocope the spectrum obtained is called the emission spectrum of hydrogen. It is found that they have large number of lines which are grouped into different series name after the discoveries.

### Rydberg's formula:

Although a large no. of lines are present in hydrogen spectrum, Rydberg in 1980's gave a simple equation for the calculation of wavelengths of these lines

$$\bar{\nu} = R \left[ \frac{1}{n_1^2} - \frac{1}{n_2^2} \right]$$

R – Rydberg's constant

This formula is known as Rydberg's formula

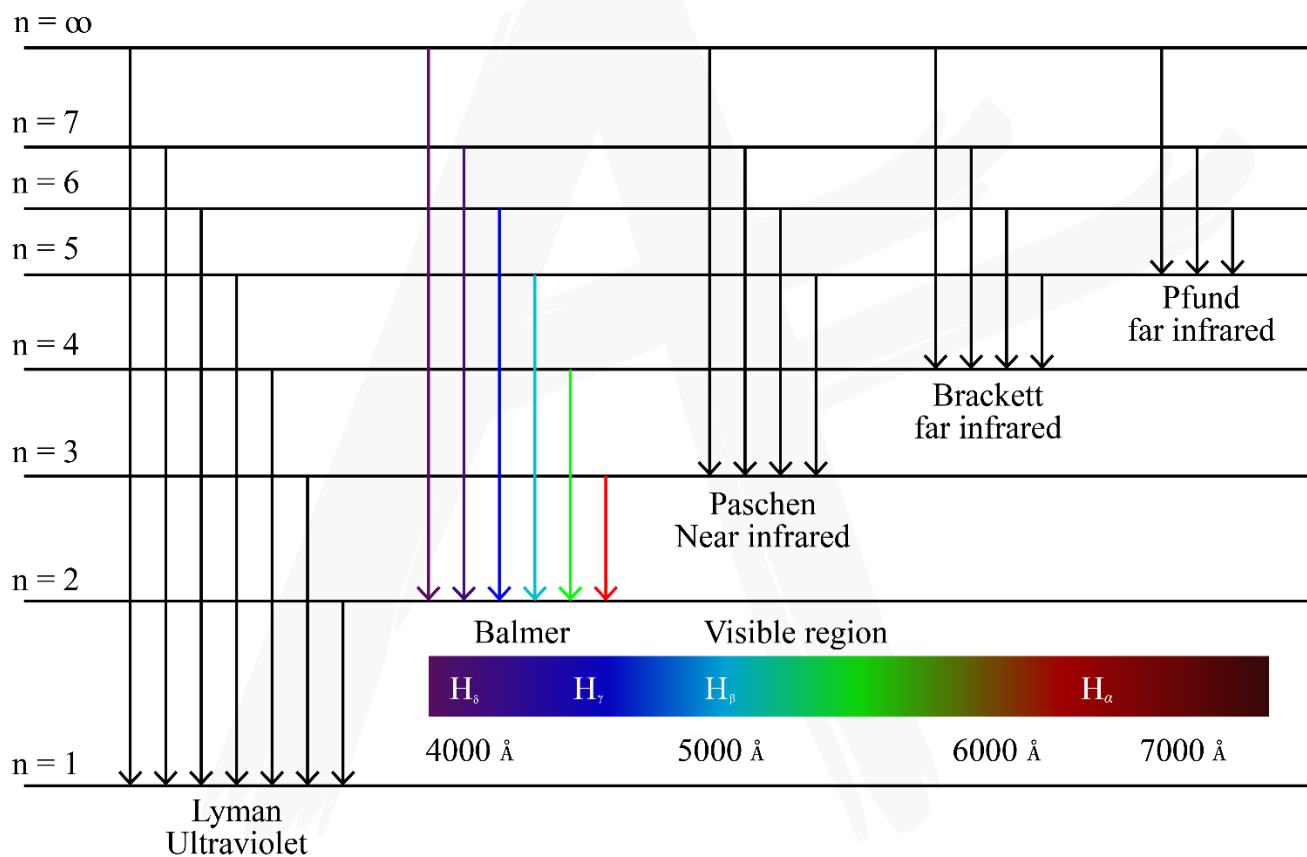
When  $n_1 = 1$

$n_2 = 2$

Balmer's formula and so on

**Note:**

**Limiting line:** The limiting line for any spectral series in the hydrogen spectrum is the line when  $n_2$  in the Rydberg's formula is infinity i.e.  $n_2 = \infty$



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
Humphrey	6	7, 8, ...	Far infrared

Neil Bohr was first to explain quantitatively the general features of hydrogen atom i.e., structure and its spectrum.

He extended the work of Rutherford to give the structure of atom. Once, the existence of the nucleus was proved through Rutherford's model of atom, now the work left was to concentrate on the arrangement of electrons around the nucleus. Neil Bohr was successful in answering the questions like:

- **Why line spectrum of atoms is obtained?**
- **The explanation of spectral lines of hydrogen atom etc.**

Let us study the postulates on which Bohr's model for hydrogen atom is based:

- (i) Nucleus:** Hydrogen atom has small heavy positively charged nucleus in the centre (this postulate is the retention of Rutherford's model)
- (ii) Orbits:** The electron around the nucleus revolves only in certain fixed circular path of fixed radius and energy. These circular paths are also known as orbits, stationary states or allowed energy states or energy levels.
- (iii) Energy of an electron in the orbit does not change with Time:** As the energy of an electron revolves in a fixed circular path called orbit is fixed. Therefore, orbits are called energy states or energy shells.

Therefore, electrons when absorb energy move from lower energy level to higher energy level, to its opposite electron moves from higher energy state to a lower energy state by emitting energy.

Thus, we can say that energy of the electron only changes when it undergoes transitions from one energy level to the other. But while revolving in a fixed orbit, they do not radiate or lose energy. In this manner, Bohr overcame Rutherford's difficulty to account for the stability of atom that why electrons do not lose energy while revolving and fall into the nucleus.

In the Bohr model, the lowest energy state available for an electron is the ground state and all higher energy states are called excited states.

- (iv) Frequency of the radiation:** The frequency of the radiation absorbed or emitted during the transition of electron from one energy state to another is given by

$$\nu = \frac{\Delta E}{h}$$

Where,  $\Delta E$  is the difference in the energies of the two energy states i.e.,  $\Delta E = E_2 - E_1$ .  
where,

$E_1$  = Energy of the lower allowed energy state

$E_2$  = Energy of higher allowed energy state

The above expression is commonly known as Bohr's frequency rule.

(v) **Angular Momentum of an electron in an orbit:** The angular momentum of an electron is fixed orbit in expressed as

$$\underset{\substack{\text{Angular} \\ \text{momentum}}}{m_e v r} = n \cdot \frac{h}{2\pi} \text{ where } n = 1, 2, 3, \dots$$

$m_e$  = mass of electron

$v$  = speed of electron

$r$  = radius of orbit

So, from the above expression, it was concluded that the permitted orbits are those for which the angular momentum of an electron is an integral multiple of  $\frac{h}{2\pi}$ .

From the above expression, the concept of quantised energy levels is also clear. Therefore, this equation is also known as the **Bohr quantization postulate**.

### Bohr's Model for hydrogen atom

1. The stationary states or the energy states in which the electrons move are numbered  $n = 1, 2, 3, 4$

These are also called as principle quantum no's or as

K, L, M, N ....

1 2 3 4....

2. Radii of stationary states : It depends on principle quantum no.(n)

$$r_n = n^2 a_0$$

Ex. for H atom

Where  $a_0 = 52.9 \text{ pm}$

$$r_H = (1)^2 \times 90 = 52.9 \text{ pm.}$$

3. Energy

$$E_n = -R_H \left( \frac{1}{n^2} \right)$$

$n = 1, 2, 3 \dots$

Note  $n \uparrow r \uparrow$

Ex. for  $n = 1$

$$E_1 = -R_H \left( \frac{1}{1^2} \right)$$

$$E_1 = -2.18 \times 10^{-18} \left( \frac{1}{1} \right)$$

$$E_1 = -2.18 \times 10^{-18} \text{ J/atom}$$

4. Bohr's theory for ions containing one  $e^-$  or hydrogen like species

Ex :  $\text{He}^+$ ,  $\text{Li}^{+2}$ ,  $\text{Be}^{+3}$  ie containing only one  $e^-$

(a) Energy ( $E_n$ ) :

$$E_n = -2.18 \times 10^{-18} \left( \frac{Z^2}{n^2} \right) \text{ J/atom}$$

(b) Radius (R) :

$$r_n = 52.9 \times \left( \frac{n^2}{Z} \right) \text{ pm}$$

$Z$  = atomic



**Note :**

- $Z \uparrow$  value for energy  $\downarrow$  (ie more -ve value radius  $\downarrow$ )
- Bohr's orbit: The first orbit with radius 52.9 pm in which an  $e^-$  of H atom is present is called Bohr's orbit.

5. Velocity of a revolving  $e^-$  depends upon,

**(a) Positive charge on nucleus :** As positive charge increase in nucleus velocity increases to counter balance the charge.

$$v \propto \text{positive charge on nucleus.}$$

**(b) Principle quantum number : (n)**

$$\text{Velocity} \propto \frac{1}{\text{Principle Quantum Number}}$$

**Note:**

$$1. \text{ KE} = 13.6 \times \frac{Z^2}{n^2} \text{ ev.}$$

$$2. \text{ PE} = -2 \times 13.6 \times \frac{Z^2}{n^2} \text{ ev.}$$

$$3. \text{ TE} = -13.6 \times \frac{Z^2}{n^2} \text{ ev.}$$

$$4. R = 109677 \text{ cm}^{-1} \\ = -2.18 \times 10^{-18} \text{ J.}$$

**Explanation for Line Spectrum of Hydrogen**

Bohr was successful in explaining the line spectrum of hydrogen. The spectrum is not continuous but discontinuous (line spectrum) because a definite amount of energy is emitted when an electron moves from a higher energy state to a lower energy state. Whereas a definite amount of energy is absorbed when electron moves from lower energy state to higher energy state.

The difference in the energy between two energy states or orbits is

$$\Delta E = E_f - E_i \quad \dots(i)$$

where  $E_f$  = energy of final state and  $E_i$  is the energy of initial state.

$$\text{We know } E_n = -R_H \left( \frac{1}{n^2} \right)$$

$$\text{Therefore } E_f = -R_H \left( \frac{1}{n_f^2} \right) \quad \dots(ii)$$

$$E_i = -R_H \left( \frac{1}{n_i^2} \right) \quad \dots(iii)$$

Combining (ii) and (iii) and putting in (i), we get

$$\Delta E = \left(-\frac{R_H}{n_f^2}\right) - \left(-\frac{R_H}{n_i^2}\right)$$

$$\Delta E = R_H \left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$R_H (\text{Rydberg constant}) = 2.18 \times 10^{-18} \text{ J}$$

$$\Delta E = 2.18 \times 10^{-18} \text{ J} \left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

During the transition, energy (radiation) is emitted with a particular frequency

$$E = h\nu \text{ (Planck's theory)}$$

Here  $E = \Delta E$  (Energy gap)

$$\Delta E = h\nu$$

$$\nu (\text{Frequency}) = \frac{\Delta E}{h}$$

$$= \frac{R_H}{h} \left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$\nu = \frac{2.18 \times 10^{-18}}{6.626 \times 10^{-34}} \left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$= 3.29 \times 10^{15} \left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) \text{ Hz}$$

$$\text{Wave number } (\bar{\nu}) = \frac{\nu}{c}$$

$$\bar{\nu} = \frac{R_H}{hc} \left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$= \frac{3.29 \times 10^{15} \text{ s}^{-1}}{3 \times 10^8 \text{ ms}^{-1}} \left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

$$\bar{\nu} = 1.09677 \times 10^7 \left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) \text{ m}^{-1}$$

In case of absorption spectrum:  $n_f > n_i$

$$n_f = 2$$

$$n_i = 1$$

Therefore,  $n_f > n_i$

In case of emission spectrum:  $n_i > n_f$

$$n_i = 2$$

$$n_f = 1$$

Therefore,  $n_i > n_f$

Bohr was able to quantitatively explain the line spectrum of hydrogen because the expression for wave number ( $\bar{\nu}$ ) given by Bohr

$$\bar{\nu} = 1.09677 \times 10^7 \left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) \text{ m}^{-1}$$

It is similar to the expression given by Johannes Rydberg for wave number from the experimental data available.

$$\bar{\nu} = 109677 \left( \frac{1}{n_i^2} - \frac{1}{n_f^2} \right) \text{cm}^{-1}$$

He gave reasons for the line spectrum because in the emission spectra, each spectral line is due to a particular transition associated with a particular frequency or wavelength.

In large number of hydrogen atoms, different possible transitions can be observed, which leads to large number of spectral lines.

### Brightness or Intensity of Spectral Lines

It depends upon the number of photons of same frequency or wavelength absorbed or emitted.

**Q.** What are the frequency and wavelength of a photon emitted during a transition from  $n = 5$  state to the  $n = 2$  state in the hydrogen atom?

**Sol.** Since  $n_i = 5$  and  $n_f = 2$ , this transition gives rise to a spectral line in the visible region of the Balmer series. From equation (2.17)

$$\Delta E = 2.18 \times 10^{-18} \text{J} \left[ \frac{1}{5^2} - \frac{1}{2^2} \right]$$

$$= -4.58 \times 10^{-19} \text{J}$$

It is an emission energy

The frequency of the photon (taking energy in terms of magnitude) is given by

$$\nu = \frac{\Delta E}{h}$$

$$= \frac{4.58 \times 10^{-19} \text{J}}{6.626 \times 10^{-34} \text{Js}} = 6.91 \times 10^{14} \text{Hz}$$

$$\lambda = \frac{c}{\nu} = \frac{3.0 \times 10^8 \text{ m s}^{-1}}{6.91 \times 10^{14} \text{Hz}} = 434 \text{ nm}$$

**Q.** Calculate the energy associated with the first orbit of  $\text{He}^+$ . What is the radius of this orbit?

$$\text{Sol. } E_n = - \frac{(2.18 \times 10^{-18} \text{J}) Z^2}{n^2} \text{ atom}^{-1}$$

For  $\text{He}^+$ ,  $n = 1$ ,  $Z = 2$

$$E_1 = - \frac{(2.18 \times 10^{-18} \text{J})(2^2)}{1^2} = -8.72 \times 10^{-18} \text{J}$$

The radius of the orbit is given by equation (2.15)

$$r_n = \frac{(0.0529 \text{ nm}) n^2}{Z}$$

*(Practice Question in the End, Q.4, 11)*

$$\text{Since } n = 1, \text{ and } Z = 2; \quad r_n = \frac{(0.0529 \text{ nm})1^2}{2} = 0.02645 \text{ nm}$$

### Limitation of Bohr's Model

Though Neil Bohr was able to improve Rutherford's model of atom as Bohr's model could successfully

- (i) Account for the stability of atom.
- (ii) Explain the origin of discontinuous line spectrum.

**But Bohr's model too faced many objections. Major limitations of Bohr's model are as follows:**

1. It could only explain the line spectra of atoms containing single electron. It failed to explain the line spectra of multielectron atoms.
2. With the help of high resolving spectroscopes, it was found that a single line in the spectra consists of number of closely spaced lines (like doublets, triplets etc.)
3. It could not explain the splitting of spectral lines in the magnetic field. This splitting of spectral lines in the presence of magnetic field is called **Zeeman effect**.
4. Bohr was even unable to explain the splitting of spectral lines in the strong electric field (Stark effect).

In the later section we will see that even Bohr was wrong in thinking that an electron's position and motion can be specified at a given time, as well as thinking that electrons revolve in orbits of fixed radii.

### Note:

How to calculate no of spectral lines?

$$\text{No of spectral lines} = \frac{(n_2 - n_1)(n_2 - n_1 + 1)}{2}$$

For example  $n_2 = 4, n_1 = 1$

$$\therefore \text{No of spectral lines} = \frac{(4-1)(4-1+1)}{2} = 6$$

### Development leading to Quantum mechanical model of atom.

To overcome the limitations of Bohr's model new model was developed on the basis of two concepts.

1. **De-Broglie's hypothesis**
2. **Heisenberg's uncertainty principle**

This model was called as quantum or wave mechanical model of atom.

#### 1. De-Broglie Hypothesis of dual nature of matter.

- i. According to plank's theory, light has dual nature i.e. wave and particle nature.
- ii. de Broglie's says that all microscopic particle which has some mass i.e.  $e^-$ ,  $p^+$  etc also possess dual character.

iii. Wavelength of wave associated with any microscopic particle calculated as

(a)  $E = h\nu$  (plank's wave)

(b)  $E = mc^2$  (Einstein's eq) (particle)

From (a) & (b)

$$h\nu = mc^2$$

$$h \frac{c}{\lambda} = mc^2$$

$$\therefore \lambda = \frac{h}{mc} \quad \text{or} \quad \lambda = \frac{h}{mv} \quad \text{or} \quad \lambda = \frac{h}{p}$$

c = velocity of photon.

v = velocity of material particle

p = momentum of particles

$\lambda$  = de-Broglie wavelength

m = mass of particle

h = plank's constant

These equations are called as de-Broglie's equation.

iv. we know,

$$KE = \frac{1}{2}mv^2$$

Multiply m on both sides

$$m \cdot KE = \frac{1}{2}m^2v^2$$

$$m \cdot KE = \frac{1}{2}p^2 \quad \because (p = mv)$$

$$2m \cdot KE = p^2$$

$$p = \sqrt{2mKE}$$

$$\therefore \lambda = \frac{h}{\sqrt{2m \cdot KE}} \quad \left\{ \because \lambda = \frac{h}{p} \right.$$

$$\text{or } \lambda = \frac{h}{\sqrt{2mev}}$$

**Note:**

KE = ev

(1ev =  $1.602 \times 10^{-19}$  J)

**De-Broglie hypothesis is applicable for**

**1. Verification of wave character**

By Davison's experiment in 1927

He observed that beam of  $e^-$  fall on Ni crystal and reflected rays received on photographic plate, diffraction pattern observed is similar to that of X-ray.

X-ray has wave character hence it was concluded that  $e^-$  also have wave character.

## 2. Heisenberg's Uncertainty principle

### Principle:

1. It is possible to measure simultaneously the position and momentum of small, microscopic moving particle.
2. If we measure one quantity with higher accuracy other becomes less accurate

$$\Delta x \cdot \Delta p \geq \frac{h}{4\pi m}$$

$\Delta x \rightarrow$  uncertainty in position

$\Delta p \rightarrow$  uncertainty in momentum



### Explanation Beyond the Heisenberg's Uncertainty Principle

If we suppose that we try to measure the position of the electron. To see its position, we need small wavelength light. We know that light consists of photons. So, photons strike the electron, the reflected photon in the form of light reaches our eye and we are able to see the electron. No doubt, in this way we are able to know the position of the electron.

We know that the electron is a point charge and almost dimensionless, therefore to know its position we must use wavelength smaller than the dimensions of the electron. Therefore, shorter the wavelength, greater the accuracy.

But shorter wavelength is associated with high frequency and high energy.

$$\text{Energy} \propto \frac{1}{\text{Wavelength}}$$

Therefore, when this high energy photon strikes the electron, it changes the speed as well as direction of the electron i.e., momentum changes. Thus, it is clear that momentum cannot be found with certainty in this case, but position is known.

### Significance of Uncertainty Principle

1. **Well Defined Trajectories are not Possible:** To find the trajectories, we need to know both the position and velocity of an electron at various moments. If we know the position of an electron at a particular instant and along with it if we even know the velocity at that instant then it becomes easy to tell where the electron will be after some time, but such thing is not possible in case of electron.

Therefore, talking about the precise position of the electron in the orbit is not correct.

2. **Uncertainty Principle is Only Significant to Microscopic Objects:** Uncertainty principle is only significant to microscopic objects but insignificant for macroscopic objects. Let us see by taking examples:

#### Case of Macroscopic Moving Object

Let mass of a macroscopic moving object =  $10^{-3}$  g (1 milligram)  
 $= 10^{-6}$  kg

$$\Delta v \times \Delta x = \frac{h}{4\pi m}$$

$$= \frac{6.626 \times 10^{-34} \text{ Js}}{4 \times 3.1416 \times 10^{-6} \text{ kg}}$$

$$\Delta v \times \Delta x = 0.5 \times 10^{-28} \text{ m}^2 \text{ s}^{-1}$$

This value is extremely small and insignificant. Therefore, when we deal with heavier objects (macroscopic), then the uncertainties associated with these moving objects are hardly valuable and they make no real effect on the calculations.

Case of Microscopic Moving objects

Mass of electron =  $9.11 \times 10^{-31}$  kg

$$\Delta v \times \Delta x = \frac{h}{4\pi m}$$

$$= \frac{6.626 \times 10^{-34} \text{ Js}}{4 \times 3.1416 \times 9.11 \times 10^{-31} \text{ kg}}$$

$$\Delta v \times \Delta x = 5 \times 10^{-5} \text{ m}^2 \text{ s}^{-1}$$

Here in the case of electron  $\Delta x$  .  $\Delta v$  is large and thus significant.

This uncertainty gives an evidence that the electrons moving in Bohr's orbits cannot hold good. Therefore, quantum mechanical model of atom is established which talk about the probability of finding the electron at a given position and with certain velocity and disobeys the fact that electrons are present in fixed orbits.

**Q.** A microscope using suitable photons is employed to locate an electron in an atom within a distance of  $0.1\text{\AA}$ . What is the uncertainty involved in the measurement of its velocity? **NCERT**

**Sol.**  $\Delta x \Delta p = \frac{h}{4\pi}$  or  $\Delta x m \Delta v = \frac{h}{4\pi}$

$$\Delta v = \frac{h}{4\pi \Delta x m}$$

$$\Delta v = \frac{6.626 \times 10^{-34} \text{ Js}}{4 \times 3.14 \times 0.1 \times 10^{-10} \text{ m} \times 9.11 \times 10^{-31} \text{ kg}}$$

$$= 0.597 \times 10^7 \text{ m s}^{-1} \text{ (1J = 1kg m}^2 \text{ s}^{-2}\text{)}$$

$$= 5.79 \times 10^6 \text{ m s}^{-1}$$

**Q.** A golf ball has a mass of 40g, and a speed of 45m/s. If the speed can be measured within accuracy of 2%, calculate the uncertainty in the position. **NCERT**

**Sol.** The uncertainty in the speed is 2%, i.e.,

$$45 \times \frac{2}{100} = 0.9 \text{ m s}^{-1}$$

Using the equation (2.22)

$$\begin{aligned}\Delta x &= \frac{h}{4\pi m \Delta v} \\&= \frac{6.626 \times 10^{-34} \text{ Js}}{4 \times 3.14 \times 40 \times 10^{-3} \text{ kg g}^{-1} (0.9 \text{ ms}^{-1})} \\&= 1.46 \times 10^{-33} \text{ m}\end{aligned}$$

This is nearly  $\sim 10^{18}$  times smaller than the diameter of a typical atomic nucleus. As mentioned earlier for large particles, the uncertainty principle sets no meaningful limit to the precision of measurements.

### Reasons for the Failure of the Bohr Model

The acceptance of successful work of de-Broglie and Heisenberg led to the failure of Bohr's model of atom.

1. **Dual nature of matter (electron)** given by de-Broglie put an obligation on the Bohr's model which did not consider wave nature of matter (electron). Bohr treated the electron only as particle.

$$\lambda = \frac{h}{mv} = \frac{h}{p}$$

Where  $m$  is the mass of the particle,  $v$  is velocity and  $p$  is momentum.

**Q.** What will be the wavelength of a ball of mass  $0.1 \text{ kg}$  moving with a velocity of  $10 \text{ m s}^{-1}$ ?

**Sol.** According to de Broglie equation (2.22)

NCERT

$$\begin{aligned}\lambda &= \frac{h}{mv} = \frac{(6.626 \times 10^{-34} \text{ Js})}{(0.1 \text{ kg})(10 \text{ ms}^{-1})} \\&= 6.626 \times 10^{-34} \text{ m (J = kg m}^2 \text{ s}^{-2})\end{aligned}$$

**Q.** The mass of an electron is  $9.1 \times 10^{-31} \text{ kg}$ . If its K.E. is  $3.0 \times 10^{-25} \text{ J}$ , calculate its wavelength.

**Sol.** Since K.E. =  $\frac{1}{2} mv^2$

NCERT

$$\begin{aligned}v &= \left( \frac{2 \text{ K.E.}}{m} \right) = \left( \frac{2 \times 3.0 \times 10^{-25} \text{ kg m}^2 \text{ s}^{-2}}{9.1 \times 10^{-31} \text{ kg}} \right)^{1/2} \\&= 812 \text{ m s}^{-1}\end{aligned}$$

$$\lambda = \frac{h}{mv} = \frac{6.626 \times 10^{-34} \text{ Js}}{(9.1 \times 10^{-31} \text{ kg})(812 \text{ ms}^{-1})}$$

(Practice Question in the End, Q.5, 12)

$$= 8967 \times 10^{-10} \text{ m} = 896.7 \text{ nm}$$

2. **Heisenberg's uncertainty principle** contradicted the concept of Bohr's orbit (orbit with  $n=1$  in which the single electron of the hydrogen atom is considered to be present). According to Heisenberg, the electron cannot exhibit a defined path or fixed orbit as both position and velocity



cannot be simultaneously known with certainty. The postulate of Bohr that “electrons revolve in well-defined orbits around the nucleus with well-defined velocities” is thus not tenable.

### Quantum Mechanical model

1. Classical mechanics based on newtons law of motion was successful in explaining the motion of macroscopic bodies not microscopic.
2. Quantum mechanics came into picture as it is a branch of science that accounts for dual behaviour of matter like (electrons).
3. Schrödinger in 1926 give an equation based on wave motion associated with particle known as Schrödinger wave equation.

$$\frac{\delta^2 \psi}{dx^2} + \frac{\delta^2 \psi}{dy^2} + \frac{\delta^2 \psi}{dz^2} + \frac{8\pi^2 m}{n^2} (\epsilon - v) \psi = 0.$$

In short  $\hat{H} \psi = \epsilon \psi$

$\hat{H}$  : Hameltonian operator.

E = Energy of  $e^-$

$\Psi$  = wave function

x, y, z  $\rightarrow$  axis (as  $e^-$  revolves in 3D)

**Wave function ( $\psi$ ):** It is a mathematical function whose value depends on the co-ordinates of the electron in atom.

The wavefunction is case of hydrogen or hydrogen like atom, wave functions pertaining to one electron is called as one electron system

Schrödinger equation does not talk about the presence of an electron in specified orbit. It talks about the probability of finding the electron i.e.  $\psi^2$

**For Multielectron system or atoms:** Schrodinger equation applied to multi-electron atoms, the results obtained are not correct. This is due to the increased nuclear charge. This results in contraction of the orbitals.

**Quantum Numbers:** It is defined as set of four numbers which give complete information about electrons in orbital i.e. location, shape, energy, orientation of orbital.

1. **Principal Quantum Number (n)**
2. **Azimuthal Quantum Number (l)**
3. **Magnetic Quantum Number (m)**
4. **Spin Quantum Number (s)**

### 1. Principal Quantum Number: (n)

(a) Gives information about shell in which electron is present.

n = 1	K	
n = 2	L	
n = 3	M	size and energy inc ↑
n = 4	N	
n = 5	O	

### 2. Azimuthal / Angular momentum.

(a) It gives information about subshell in which electron present

(b) Tells about shape of subshell.

<i>l</i>	subshell	shape
0	s	spherical
1	p	dumbbell shape
2	d	double dumbbell.
3	f	complex shape
4	g	8 fold structure

(c) value of *l* is  $0 \leq l \leq n - 1$

Shell (n)	subshell ( <i>l</i> )	
1	0	1s
2	0, 1	2s 2p
3	0, 1, 2	3s 3p 3d
4	0, 1, 2, 3	4s 4p 4d 4f

Note:

(d) Orbital angular momentum =  $\frac{h}{2\pi} \sqrt{l(l+1)}$  or  $h \sqrt{l(l+1)}$

(e) It gives into about energies of different sub-shell  $s < p < d < f$

### 3. Magnetic Quantum Numbers: (m)

It gives information about orientation of orbitals and number of orbitals in a given subshell

$$(-l \leq m \leq +l)$$

	<i>l</i>	<b>m</b>	<b>No of orbitals</b>
s	0	0	1
p	1	-1 0 1	3
d	2	-2 -1, 0, 1, 2	5
f	3	-3, -2, -1, 0, 1, 2, 3	7

- (a) No of orbitals in a shell =  $n^2$   
 (b) No of orbitals in a subshell =  $2l + 1$

#### 4. Spin Quantum Number ( $s$ or $m_s$ )

(a) It gives information about direction of spin of electron in any orbital ( $e^-$  rotates about its own axis may be clockwise or anticlockwise)

$\therefore$  It has a value of  $+\frac{1}{2}$  or  $-\frac{1}{2}$  of  $\uparrow\downarrow$

Pauli's Exclusion Principle.

No two electrons in the same atom can have same set of 4 Quantum no's.

n	l	m	s
1	0	0	$+\frac{1}{2}, -\frac{1}{2}$

Only two  $e^-$  with opposite spin can be present in an orbital

**Note:** Spin Magnetic Moment: ( $\mu$ )

$$\mu = \sqrt{n(n+2)} \text{ BM}$$

n = no of unpaired electrons

#### Formula's

- No of subshell in a shell = No of shells
- No of orbitals in a shell =  $n^2$
- No of orbitals in a subshell =  $2l + 1$
- No of electron in orbital = 2
- No of electrons in subshell =  $2(2l + 1)$
- No of electrons in shell =  $2n^2$

#### Note:

Number of spherical / Radial modes =  $n - l - 1$

Number of planar / angular node =  $l$

Total no of nodes =  $n - l - 1 + l = n - 1$

Spherical node is only for s orbital p & d orbital has both spherical / angular.

**Q.** What is the total number of orbitals associated with the principal quantum number  $n = 3$ ? **NCERT**

**Sol.** For  $n = 3$ , the possible values of  $l$  are 0, 1 and 2. Thus there is one 3s orbital ( $n = 3, l = 0$  and  $m_l = 0$ ); there are three 3p orbitals ( $n = 3, l = 1$  and  $m_l = -1, 0, +1$ ); there are five 3d orbitals ( $n = 3, l = 2$  and  $m_l = -2, -1, 0, +1, +2$ ).

Therefore, the total number of orbitals by using the relation; number of orbitals =  $n^2$ , i.e.  $3^2 = 9$ .

**Q.** Using s, p, d, f notations, describe the orbital with the following quantum numbers

(a)  $n = 2, l = 1$ ,

(b)  $n = 4, l = 0$ ,

(c)  $n = 5, l = 3$ ,

(d)  $n = 3, l = 2$

NCERT

**Sol.**     $n$          $l$         orbitals

(a)    2        1        2p

(b)    4        0        4s

(c)    5        3        5f

(d)    3        2        3d

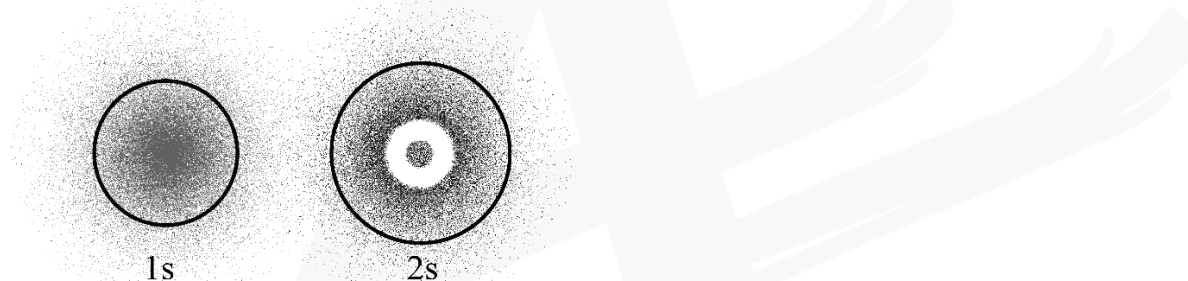
*(Practice Question in the End, Q.6, 7)*

### Shapes of orbitals:

#### I. Shape of s-orbital:

(a) Spherical shape

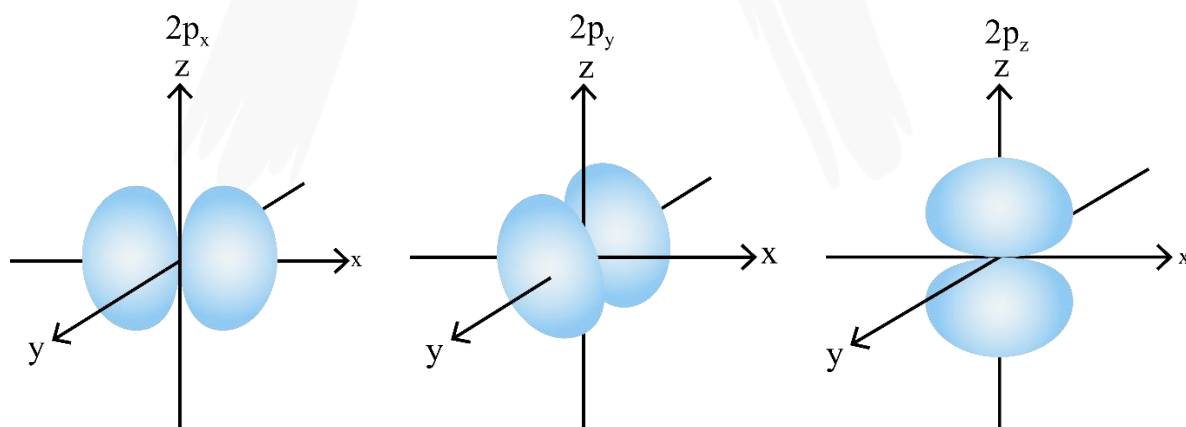
(b) Non-directional



#### II. Shape of p-orbital: (dumbbell shaped)

(a) p-orbitals have 3 different orientation i.e.  $p_x$ ,  $p_y$ ,  $p_z$  depending upon max electron density along x, y, z axis

(b) Directional nature

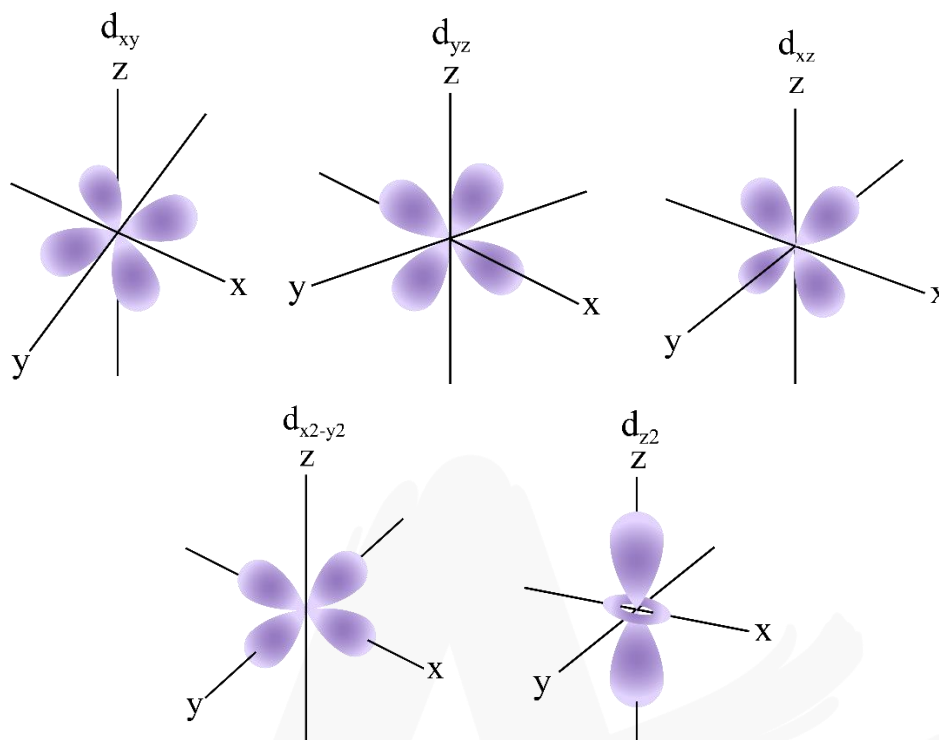


#### III. Shape of d-orbital:

(a) double dumbbell shape

(b) directional in nature

(c) 5d orbitals



$dx^2 - y^2$  &  $dz^2$  have  $e^-$  density along the axis.

## Energies of Atomic Orbitals

### Energy of Electron in Hydrogen Atom

1. For a hydrogen atom which contains only one electron, the energy of that electron is just dependent upon the value of principal quantum number's.
2. As the principal quantum number increases, the energy of the electron also increases. Therefore, we can show the order of energies given below:

$$1s < 2s = 2p < 3s = 3p = 3d < 4s = 4p = 4d = 4f < \dots$$

Here as you can see the energy of electron present in third shell is same, irrespective of the fact that whether it is present in 3s, 3p or 3d orbital inspite of the difference in the shapes of s, p and d orbitals.

3. The reason for the above sequence of energy is because of the fact that in hydrogen atom only one electron is present, therefore the only electrical interaction present is the attraction between the negatively charged electron and the positively charged nucleus.
4. **Degenerate Orbitals:** Degenerate orbitals are those orbitals which have same energy. As in the above case 3s, 3p and 3d orbitals are degenerate.
5. **Ground State:** It is the state where the electron is most stable. Ground state is the state where the energy of the electron is minimum and is tightly held by the nucleus. Therefore, 1 s-orbital in a hydrogen atom is the ground state.
6. **Excited State:** Electron present in high energy orbitals is said to be in the excited state ex. in 2s, 2p or higher orbitals.

## Energy of an Electron in a Multi-electron Atom

1. In case of multi-electron atoms, the energy of an electron depends both on the principal quantum number 'n' and azimuthal quantum number 'l'. This means that the energy of electron depends upon the sub-shell of the principal shell to which the electron belongs i.e... whether the electron is present in s, p, d or f subshell in a given shell.
2. The dependence of the energy of the electron on both 'n' and 'l' value is because unlike hydrogen atom, in multi-electron atoms two types of electrical interactions are present:
  - (i) Attraction between nucleus and electron of atom.
  - (ii) Repulsion between concerned electron and rest of the electron present in the atom.

Thus, the stability of an electron in a multi-electron atom is because the total attractive interactions are more than the total repulsive interactions.

3. **Shielding Effect:** We know that more the number of protons more is the positive charge ( $Z_e$ ) on the nucleus. With the increase in the positive charge on the nucleus, the attractive interactions between the nucleus and electron also increases. But there are some other factors which also play a major role in establishing the energy of the electron in an orbital. The major factor is shielding or screening effect.

**Shielding Effect ( $\sigma$ ):** It is the shielding of the outer shell electrons from the nucleus by the inner shell electrons. Because of shielding effect, the nuclear force being felt by the outershell electrons decreases. Now, in such case, the nuclear charge being felt by the outershell electron is called effective nuclear charge ( $Z_{eff}$ .)

$$Z_{eff} = Z - \sigma$$

But despite the shielding of the outer electrons from the nucleus by the inner shell electrons, the attractive force being felt by the outershell electrons increases with the increase in the nuclear charge. Moreover we know that the more attractive force results in the decrease in energy (i.e., more negative).

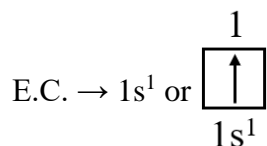
Shielding effect depends upon the type of orbital in which the inner shell electrons are present.

The sequence of shielding effect of orbitals is  $s > p > d > f$ .

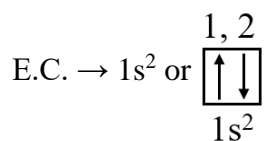
The electrons present in the s-orbital, shields the outershell electrons from the nucleus to maximum extent. The shielding effect of electrons in p-orbital is not that effective but more than that of the electron present in the d-orbitals. Further, the shielding effect of innershell electrons present in the f-orbital on outershell electrons from the nucleus is less than that of the d-orbitals. This sequence is true, when all the orbitals are present in the same energy shell.

## Electronic Configuration of Various Elements

1. Hydrogen (H) ( $Z = 1$ ): It has only one electron.

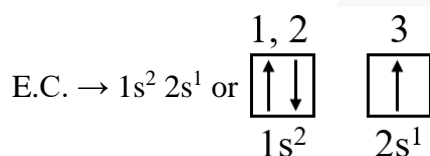


2. Helium (He) ( $Z = 2$ ): It has two electrons.

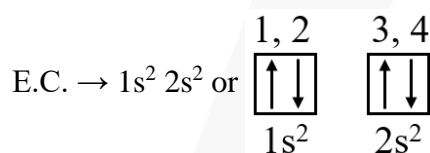


3. Lithium (Li) ( $Z = 3$ ): It has three electrons. One orbital can accommodate a maximum of two electrons. So, the third electron of lithium is not allowed in the 1s orbital, following Pauli's exclusion principle.

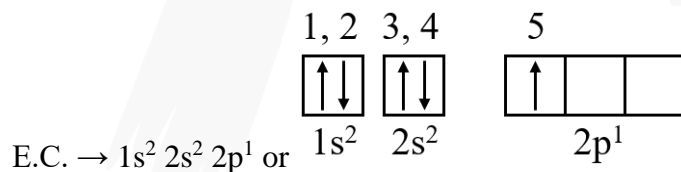
Therefore, it goes into the higher energy orbital which is 2s orbital.



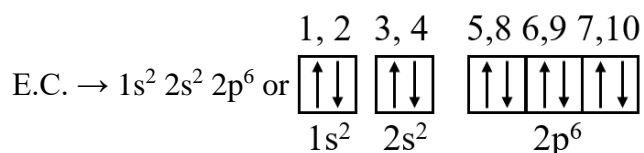
4. Beryllium (Be)  $Z = 4$ :



5. Boron (B)  $Z = 5$ :



6. Neon (Ne)  $Z = 10$ :



Electronic configuration of the elements from Sodium to Argon

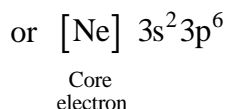
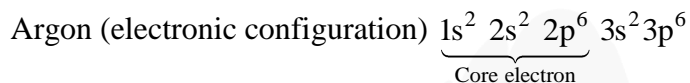
Sodium (Na)  $Z = 11$ : Electronic configuration is  $\underbrace{1s^2 2s^2 2p^6}_{\text{E.C. of Neon (Ne)}} 3s^1$

Argon (Ar)  $Z = 18$ : Electronic configuration is  $\underbrace{1s^2 2s^2 2p^6}_{\text{E.C. of Neon (Ne)}} 3s^2 3p^6$

The electronic configuration of the elements sodium to argon follow exactly the same pattern as the elements from lithium to neon. The only difference is that now 3s and 3p orbitals are getting filled instead of 2s and 2p orbitals. Then the simplified way of writing electronic configuration is to represent the total number of electrons in the first two shells by the name of the element neon (Ne). So, the electronic configuration of sodium is simply written as [Ne]3s<sup>1</sup>. (Condensed electronic configuration) Electronic configuration of Argon as [Ne]3s<sup>2</sup>3p<sup>6</sup>.

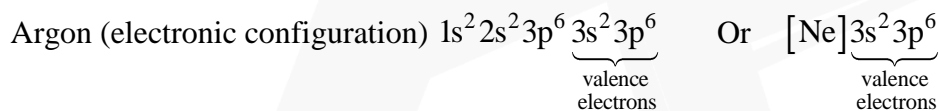
Core electrons: These are the innershell electrons.

Example:

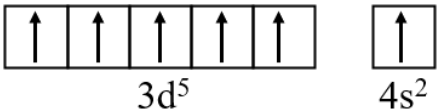
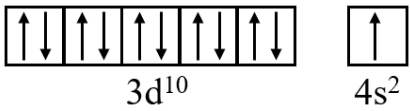


Valence electrons: These are the outer electrons i.e., the electrons which are present in the shell with highest principal quantum number.

Example:



Electronic configuration of Scandium (Z = 21) to Zinc (Z = 30): In these elements 3d-orbital gets filled as it is lower in energy than 4p-orbital.

Element	Symbol	Atomic number	Electronic configuration	Orbital notation
Chromium	Cr	24	[Ar] 3d <sup>5</sup> 4s <sup>1</sup>	 3d <sup>5</sup> 4s <sup>2</sup>
Copper	Cu	29	[Ar] 3d <sup>10</sup> 4s <sup>1</sup>	 3d <sup>10</sup> 4s <sup>2</sup>

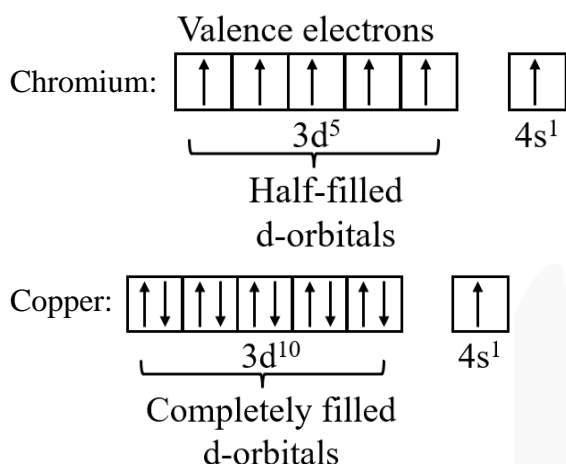
#### Exceptional Configurations of Chromium and Copper

The electronic configuration of chromium (Cr) and Copper (Cu) are different from expected electronic configuration.

	Expected	Actual
<b>Chromium</b>	[Ar]3d <sup>4</sup> 4s <sup>2</sup>	[Ar]3d <sup>5</sup> 4s <sup>1</sup>
<b>Copper</b>	[Ar]3d <sup>9</sup> 4s <sup>2</sup>	[Ar]3d <sup>10</sup> 4s <sup>1</sup>



These actual configurations are due to the fact that the half-filled and full-filled orbitals are very stable and nature loves stability and so do the elements. So, the unexpected electronic configuration are acquired to gain stability. As in case of chromium, in its actual electronic configuration, d-orbitals are half-filled.



#### Causes of Stability of Completely Filled and Half-filled Subshells

The reasons for the stability of completely filled and half-filled subshells are as follows:

- 1. Symmetrical distribution of electrons:** It is a known fact that the symmetry leads to stability. So, the half-filled and fully-filled orbitals are symmetrical, hence stable. Moreover, this type of electronic configuration is even favoured by energy factor. The electrons in the 3d subshell have equal energy, but they just differ in their special distribution. This results in less shielding of one another. Due to small shielding, the electrons are pulled closer to the nucleus. And as the electrons move closer to the nucleus, their energy decreases (more negative). Further this decrease in energy leads to stability.
- 2. Exchange energy:** Electrons with the same spin have a tendency to exchange their positions when they are present in the degenerate orbitals of a subshell. The energy released during this exchange is called exchange energy.

More are the number of electrons with identical spin, more are the number of ways of exchanging with other electrons, more will be the energy released called exchange energy. Moreover, we know that decrease in energy leads to stability. *(Practice Question in the End, Q.13, 14)*

**Notes End**

### Important NCERT Questions

**Q1. (i) Calculate the number of electrons which will together weigh one gram.**

**(ii) Calculate the mass and charge of one mole of electrons.**

**[NCERT Exercise]**

**Sol.** (i) Mass of 1 electron =  $9.108 \times 10^{-28}$  g

Hence,

$$1 \text{ g} = 1/(9.108 \times 10^{-28}) = 1.098 \times 10^{27} \text{ electrons}$$

(ii) Mass of one mole of electron =  $9.108 \times 10^{-28} \times 6.022 \times 10^{23}$

We get,

$$= 5.48 \times 10^{-4} \text{ g}$$

Charge on one mole of electron =  $1.6 \times 10^{-19} \text{ C} \times 6.022 \times 10^{23}$

We get,

$$= 9.63 \times 10^4 \text{ C}$$

**Q2. Yellow light emitted from a sodium lamp has a wavelength ( $\lambda$ ) of 580 nm. Calculate the frequency ( $\nu$ ) and wavenumber ( $\bar{\nu}$ ) of the yellow light.**

**[NCERT Exercise]**

**Sol.** Rearranging the expression,

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

the following expression can be obtained,

$$\nu = \frac{c}{\lambda} \dots\dots\dots (1)$$

Here,  $\nu$  denotes the frequency of the yellow light

$c$  denotes the speed of light ( $3 \times 10^8$  m/s)

$\lambda$  denotes the wavelength of the yellow light (580 nm,  $580 \times 10^{-9}$  m/s)

Substituting these values in eq. (1):

$$\nu = \frac{3 \times 10^8}{580 \times 10^{-9}} = 5.17 \times 10^{14} \text{ s}^{-1}$$

Therefore, the frequency of the yellow light which is emitted by the sodium lamp is:

$$5.17 \times 10^{14} \text{ s}^{-1} \text{ The wave number of the yellow light is } \bar{\nu} = \frac{1}{\lambda} = \frac{1}{580 \times 10^{-9}} = 1.72 \times 10^6 \text{ m}^{-1}$$

**Q3. Electromagnetic radiation of wavelength 242 nm is just sufficient to ionise the sodium atom. Calculate the ionisation energy of sodium in  $\text{kJ mol}^{-1}$ .**

**[NCERT Exercise]**

**Sol.** Ionization energy (E) of sodium =  $\frac{N_A hc}{\lambda}$ .

$$= \frac{(6.023 \times 10^{23} \text{ mol}^{-1})(6.626 \times 10^{-34} \text{ Js})(3 \times 10^8 \text{ m s}^{-1})}{242 \times 10^{-9} \text{ m}}$$

$$= 4.947 \times 10^5 \text{ J mol}^{-1}$$

$$= 494.7 \times 10^3 \text{ J mol}^{-1}$$

$$= 494 \text{ kJ mol}^{-1}$$

**Q4. What is the maximum number of emission lines when the excited electron of a H atom in  $n = 6$  drops to the ground state?** [NCERT Exercise]

**Sol.** A total number of 15 lines ( $5 + 4 + 3 + 2 + 1$ ) will be obtained in this hydrogen emission spectrum.

The number of spectral lines emitted when an electron in the  $n^{\text{th}}$  level drops down to the ground state is given by

$$\frac{n(n-1)}{2}$$

$$\text{Since } n = 6, \text{ total no. spectral lines} = \frac{6(6-1)}{2} = 15$$

Hence, the maximum number of emission lines when the excited electron of a H atom in  $n = 6$  drops to the ground state is 15.

**Q5. Calculate the wavelength of an electron moving with a velocity of  $2.05 \times 10^7 \text{ m s}^{-1}$ .**

**Sol.** As per de Broglie's equation, [NCERT Exercise]

$$\lambda = \frac{h}{mv}$$

Where,  $\lambda$  denotes the wavelength of the moving particle

$m$  is the mass of the particle

$v$  denotes the velocity of the particle

$h$  is Planck's constant

$$\text{Substituting these values we get, } \lambda = \frac{(6.626 \times 10^{-34}) \text{ Js}}{(9.10939 \times 10^{-31} \text{ kg})(2.05 \times 10^7 \text{ ms}^{-1})} = 3.548 \times 10^{-11} \text{ m}$$

Therefore, the wavelength of an electron moving with a velocity of  $2.05 \times 10^7 \text{ ms}^{-1}$  is  $3.548 \times 10^{-11} \text{ m}$ .

**Q6. An electron is in one of the 3d orbitals. Give the possible values of  $n$ ,  $l$  and  $m_l$  for this electron.** [NCERT Exercise]

**Sol.** For the 3d orbital:

Possible values of the Principal quantum number ( $n$ ) = 3

Possible values of the Azimuthal quantum number ( $l$ ) = 2

Possible values of the Magnetic quantum number ( $m_l$ ) =  $-2, -1, 0, 1, 2$ .

**Q7. Explain, giving reasons, which of the following sets of quantum numbers are not possible.** [NCERT Exercise]

(a)  $n = 0, l = 0, m_l = 0, m_s = +\frac{1}{2}$

(b)  $n = 1, l = 0, m_l = 0, m_s = -\frac{1}{2}$

$$(c) n = 1, l = 1, m_l = 0, m_s = +\frac{1}{2}$$

$$(d) n = 2, l = 1, m_l = 0, m_s = -\frac{1}{2}$$

$$(e) n = 3, l = 3, m_l = 3, m_s = +\frac{1}{2}$$

$$(f) n = 3, l = 1, m_l = 0, m_s = +\frac{1}{2}$$

**Sol.** (a) Not possible. The value of  $n$  cannot be 0.

(b) Possible.

(c) Not possible. The value of  $l$  cannot be equal to that of  $n$ .

(d) Possible.

(e) Not possible. For  $n = 3$ ,  $l$  cannot be 3

(f) Possible.

**Q8. In Rutherford's experiment, generally the thin foil of heavy atoms, like gold, platinum etc. have been used to be bombarded by the  $\alpha$ -particles. If the thin foil of light atoms like Aluminium etc. is used, what difference would be observed from the above results?**

[NCERT Exercise]

**Sol.** Rutherford used alpha rays scattering to reveal the structure of the atom in 1911. The nucleus of heavy atoms is big and carries a lot of positive charge. As a result, when certain alpha particles hit the nucleus, they are easily deflected back. A number of alpha particles are also deflected at tiny angles due to the nucleus's substantial positive charge. If light atoms are used, their nuclei will be light and their nuclei will have a modest positive charge. As a result, the number of particles deflected back and those deflected at an angle will be insignificant.

**Q9. Arrange the following type of radiations in increasing order of frequency: (a) radiation from microwave oven (b) amber light from traffic signal (c) radiation from FM radio (d) cosmic rays from outer space and (e) X-rays.**

[NCERT Exercise]

**Sol.** The increasing order of frequency is as follows:

Radiation from FM radio < amber light < radiation from microwave oven < X-rays < cosmic rays.

The increasing order of a wavelength is as follows:

Cosmic rays < X-rays < radiation from microwave oven < amber light < radiation from FM radio.

**Q10. The longest wavelength doublet absorption transition is observed at 589 and 589.6 nm. Calculate the frequency of each transition and energy difference between two excited states.**

[NCERT Exercise]

**Solu.**  $\lambda_1 = 589\text{nm} = 589 \times 10^{-9} \text{ m}$

$$\therefore \text{Frequency, } (\nu_1) = \frac{c}{\lambda_1} = \frac{3.0 \times 10^8 \text{ ms}^{-1}}{589 \times 10^{-9} \text{ m}} = 5.093 \times 10^{14} \text{ s}^{-1}$$

$$\lambda_2 = 589.6 \text{ nm} = 589.6 \times 10^{-9} \text{ m}$$

$$\therefore \text{Frequency, } (\nu_2) = \frac{c}{\lambda_2} = \frac{3.0 \times 10^8}{589.6 \times 10^{-9}} = 5.088 \times 10^{14} \text{ s}^{-1}$$

The different in energy ( $\Delta E$ ) =  $E_1 - E_2 = h(\nu_1 - \nu_2)$

$$= 6.626 \times 10^{-34} \times (5.093 - 5.088) \times 10^{14} = 3.313 \times 10^{-22} \text{ J}$$

**Q11. Emission transitions in the Paschen series end at orbit  $n = 3$  and start from orbit  $n$  and can be represented as  $\nu = 3.29 \times 10^{15} \text{ (Hz)} [1/3^2 - 1/n^2]$  Calculate the value of  $n$  if the transition is observed at 1285 nm. Find the region of the spectrum. [NCERT Exercise]**

**Sol.** Here,  $\nu = 3.29 \times 10^{15} \text{ (Hz)} \left[ \frac{1}{3^2} - \frac{1}{n^2} \right] \dots\dots\dots (1)$

$$\text{Also, } \nu = \frac{c}{\lambda} = \frac{3.0 \times 10^8 \text{ ms}^{-1}}{1285 \times 10^{-9} \text{ m}} \dots\dots\dots (2)$$

From equation (1) and (2) we have

$$3.29 \times 10^{15} \left[ \frac{1}{3^2} - \frac{1}{n^2} \right] = \frac{3.0 \times 10^8}{1285 \times 10^{-9}}$$

$$\text{or } \frac{1}{9} - \frac{1}{n^2} = \frac{3.0 \times 10^8}{1285 \times 10^{-9} \times 3.29 \times 10^{15}} = 0.071$$

$$\text{or } \frac{1}{n^2} - \frac{1}{9} = 0.111 - 0.071$$

$$= 0.04 = \frac{1}{25}$$

$$\text{or } n^2 = 25 \text{ or } n = 5$$

The radiation corresponding to 1285 nm lies in the infrared region.

**Q12. The velocity associated with a proton moving in a potential difference of 1000 V is  $4.37 \times 10^5 \text{ ms}^{-1}$ . If the hockey ball of mass 0.1 kg is moving with this velocity, calculate the wavelength associated with this velocity. [NCERT Exercise]**

**Sol.** As per de Broglie's expression,

$$\begin{aligned} \lambda &= \frac{h}{mv} \\ &= \frac{(6.626 \times 10^{-34})}{0.1 \text{ kg} (4.37 \times 10^5 \text{ ms}^{-1})} \\ &= 1.516 \times 10^{-38} \text{ m} \end{aligned}$$

**Q13. The unpaired electrons in Al and Si are present in 3p orbital. Which electrons will experience more effective nuclear charge from the nucleus? [NCERT Exercise]**

**Sol.** The nuclear charge can be defined as the net positive charge that acts on an electron in the orbital of an atom that has more than 1 electrons. The greater the atomic number, the greater the nuclear charge. Silicon holds 14 protons while aluminium holds only 13. Therefore, the nuclear charge of silicon is greater than that of aluminium, implying that the electrons in the 3p orbital of silicon will experience a more effective nuclear charge than aluminium.

**Q14. (a) How many sub-shells are associated with  $n = 4$ ?**

**(b) How many electrons will be present in the subshells having  $m_s$  value of  $-1/2$  for  $n = 4$ ?**

**Sol.** (a)  $n = 4$  (Given)

[NCERT Exercise]

For a given value of 'n', the values of 'l' range from 0 to  $(n - 1)$ .

Here, the possible values of l are 0, 1, 2, and 3

Therefore, 4 subshells are associated with  $n = 4$ , which are s, p, d and f.

(b) Number of orbitals in the  $n^{\text{th}}$  shell  $= n^2$

For  $n = 4$

Therefore, number of orbitals when  $n = 4$  is 16

Each orbital has 1 electron with  $m_s$  value of  $-1/2$ .

Therefore, total number of electrons present in the

subshell having  $m_s$  value of  $\left(-\frac{1}{2}\right)$  is 16.

**Ab Phod Do!**

