Atomic Structure

Atomic Structure

Important Facts :

- Atom is derived from word 'atomos' which is Greek word means uncut or non-divisible.
- The first atomic model was proposed by John Dalton in 1808 depicting the atoms as ultimate and indivisible particle of matter.
- This theory did not hold long and it was proved through the experiments of research workers like J.J. Thomson (1897), Rutherford (1911), Neils Bohr (1913), Chadwick, Moseley etc. that atom itself has a complex structure.

- **Atom consists of particles like electrons, protons, neutrons, positron, neutrino, antiproton, meson etc. The first three are fundamental sub-atomic particles.**

Particle	Discoverer	Absolute Charge/C	Relative Charge	Mass (g)	Mass (amu)	Approx. Mass (u)	Location
Electron (e^-)	J.J. Thomson	-1.6022×10^{-19}	-1	9.11×10^{-28}	0.00054	0	Electron cloud
Proton (p^+)	Eugen Goldstein	$+1.6022\times10^{-19}$	+1	$1.672\ 6 \times 10^{-24}$	1.000727	1	Nucleus
Neutron (n^o)	James Chadwick	0	0	1.6749×10^{-24}	1.00867	1	Nucleus

- **Atom is electrically neutral because number of electrons present in any atom is equal to number of protons.**

Important Facts :

- **J. J. Thomson did not discover the cathode ray - he discovered that cathode rays are composed of negatively charged particles (which we now know as electrons-discovered in 1897).**
- **Sir William Crookes (1879) was the first to produce cathode rays, leading eventually to the discovery of the electron.**
- **Eugen Goldstein introduced the term "cathode ray" very first time. He was credited with the discovery of the anode rays and proton in 1886.**
- **Neutron was discovered by Chadwick in 1932 for which he was awarded Nobel Prize in 1935.**

- **e/m ratio for electron:**
- **The specific charge is the ratio of charge to the mass of an electron, denoted as e/m ratio:**
  e/m ratio = 1.758×10^8 C/g.
- **The e/m value of electron was determined by J.J. Thomson.**
- **The e/m ratio of electron was found to be same for all gases.**
- **Unlike cathode rays, the e/m ratio of anode rays depends upon the nature of the gas taken in the vacuum tube.**
- **The e/m value for anode rays is maximum when hydrogen gas is taken in the tube.**

Charge on the electron:

The charge on the electron, measured by R. A. Millikan with the help of his oil drop experiment, was found to be, $e = 1.60 \times 10^{-19}$ Coulombs or 4.8×10^{-10} esu.

- **Note:**

  This is the smallest measurable quantity of charge and is called one unit.

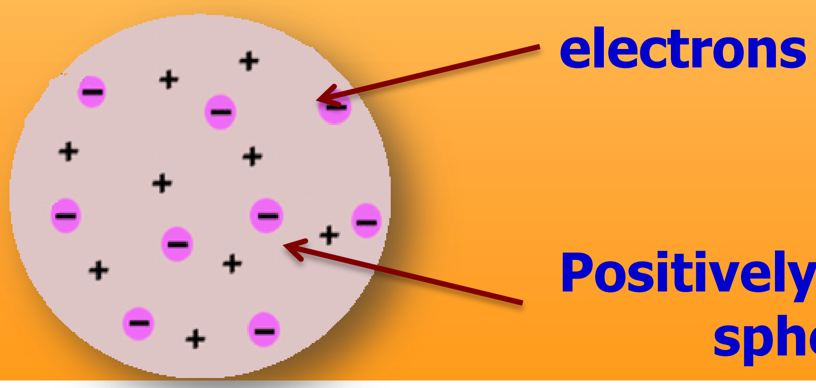
  The mass of the electron, 9.11×10^{-28} g is termed as the rest mass of the electron.

Mass of the moving electron: $m_{e^*} = \dfrac{m_e}{\sqrt{1-\dfrac{v^2}{c^2}}}$

Where, m_e = rest mass of an electron, v = velocity of the moving electron, c =speed of light= 3×10^8m/sec.

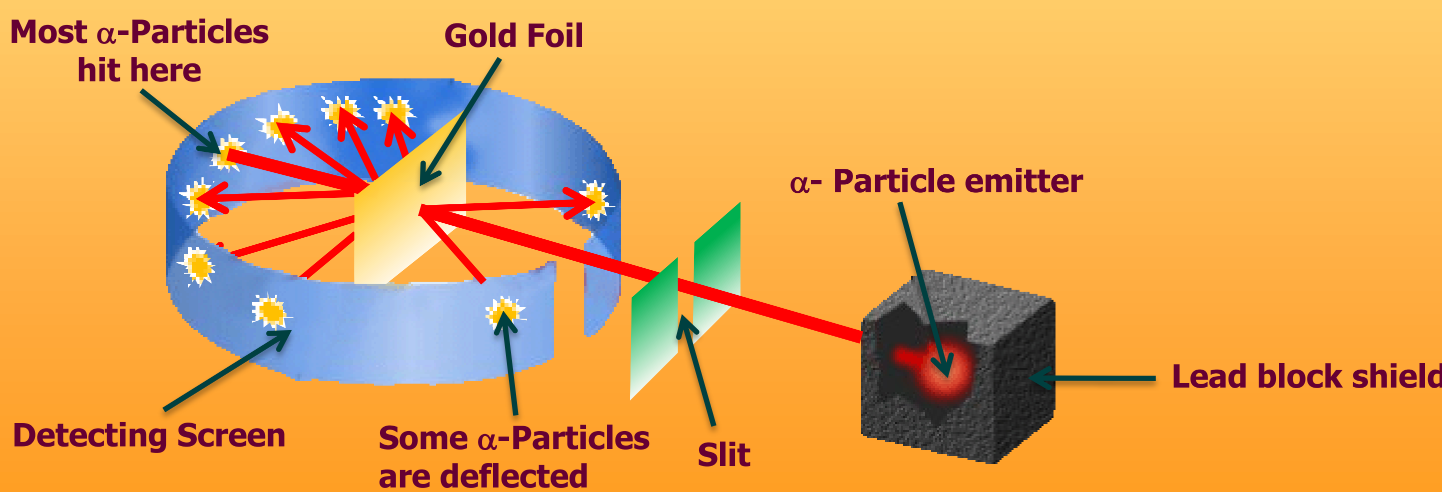
- **The mass of the proton is 1836 times the mass of electron and the mass of neutron is 1839 times the mass of electron.**
- **Atomic Models:**
- **Thomson's Model of Atom(1904):**

Thomson assumed that an atom is a sphere (radius approximately 10^{-10} m) of positively charges uniformly distributed, with the electrons scattered as points throughout the sphere. This was also called "plum pudding model".



- **Rutherford's Model of Atom(1911):**

Rutherford bombarded a thin foil of gold (thickness 100 nm) surrounded by fluorescent ZnS screen around it with high speed positively charged α-particles (charge +2, mass 4μ).

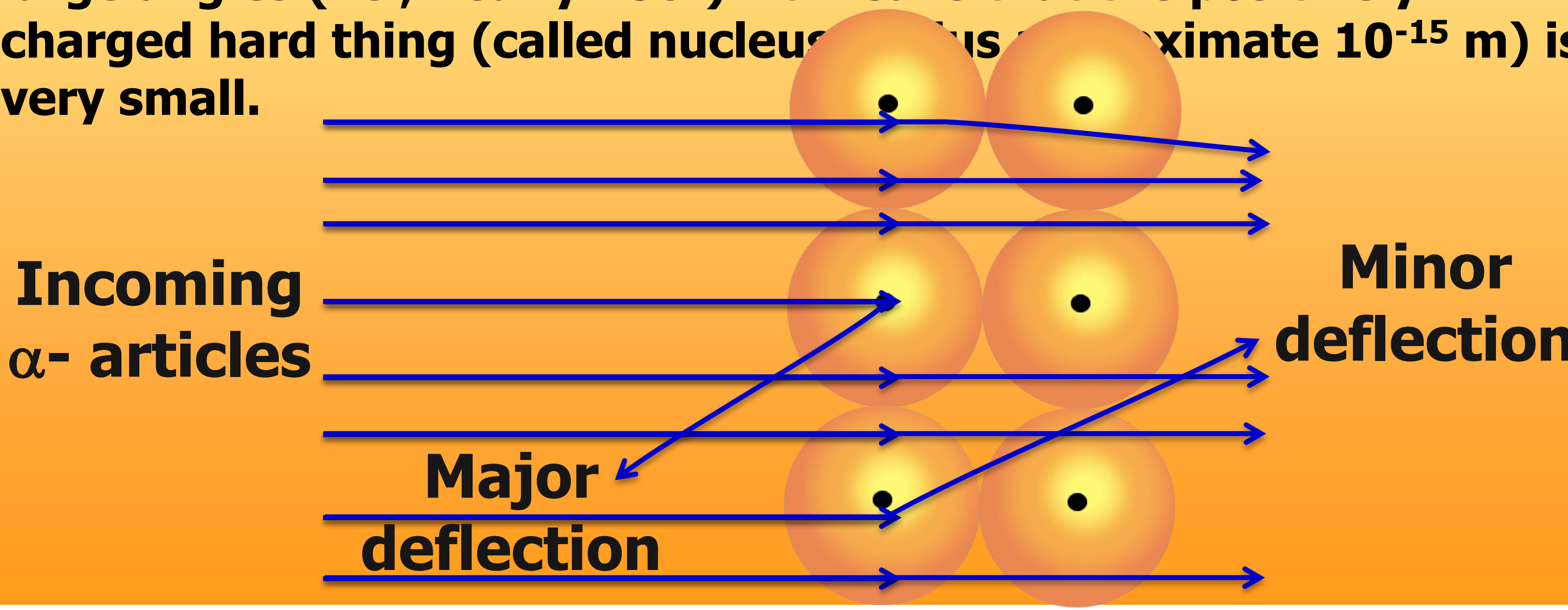


- The results of such an experiment were, however, quite unexpected as he observed that:

(i) Most of the α-particles passed through the foil undeflected, it means that most of the space in atom is empty.

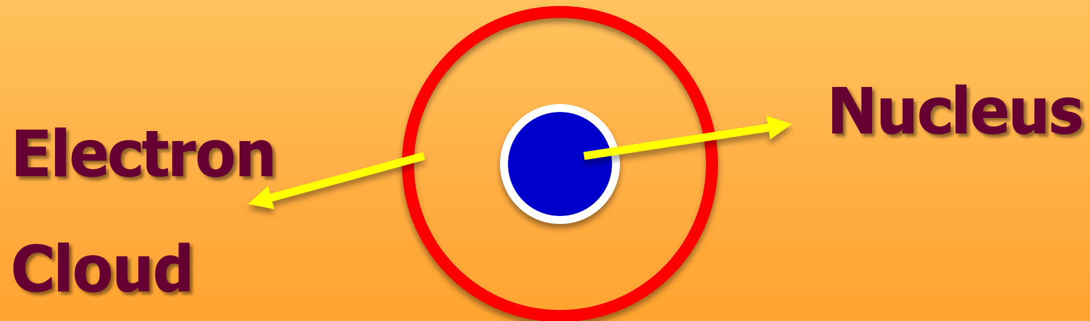
(ii) Some of them were deflected at small angles. This shows that there is something positively charged at the centre.

(iii) A very few particles (1 in 20,000 approximate) were deflected at large angles (i.e., nearly 180°). It means that the positively charged hard thing (called nucleus radius approximate 10^{-15} m) is very small.



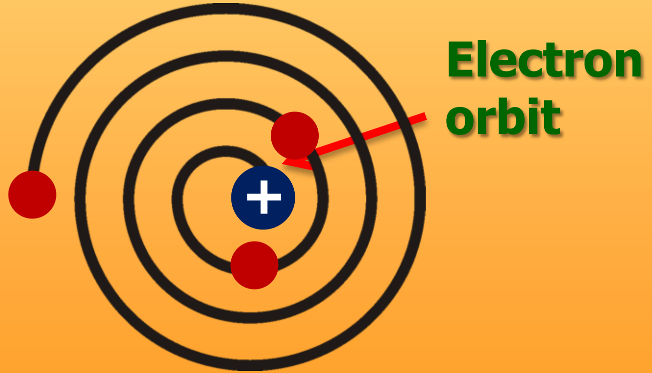
- **Thus, Rutherford was credited with the discovery of nucleus.**

The electrons revolve around the nucleus in closed orbits with a very high speeds (like planets around the Sun) and almost all the space around the nucleus is occupied by revolving electrons (in the diameter of about 10^{-10} m).



- **But the last assumption was the main weakness of Rutherford's model, because according to classical theory of electromagnetism, if a charged particle revolves around oppositely charged particle, the former one will radiate energy resulting to decrease in its speed. The circular orbit becomes spiral gradually due to such release of energy and electron finally fall on nucleus.**

Rutherford's model of atom: Electrons revolving around nucleus

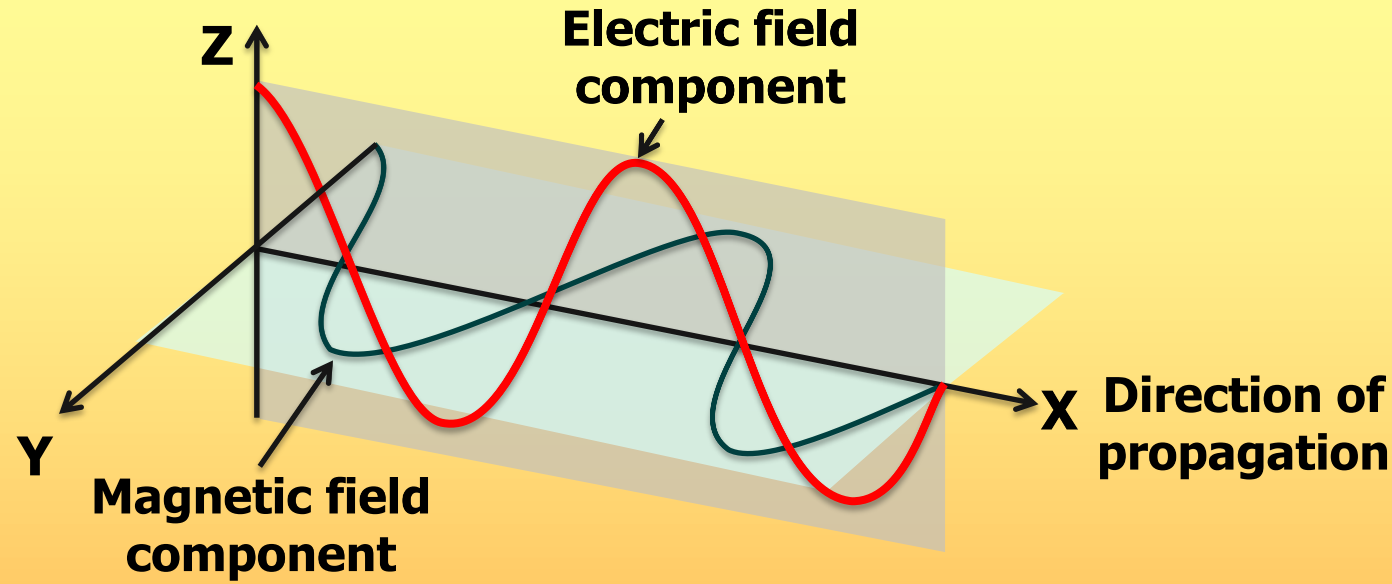


Revolving electrons would radiate energy and spiral into the nucleus

- **Rutherford’s nuclear Model of atom could not explain the characteristics line spectra of elements.**
- **In order to understand the line spectra, it is essential to understand the nature of light which in turn was explained on the basis on the “Electromagnetic wave theory” and “Planck Quantum Theory”.**

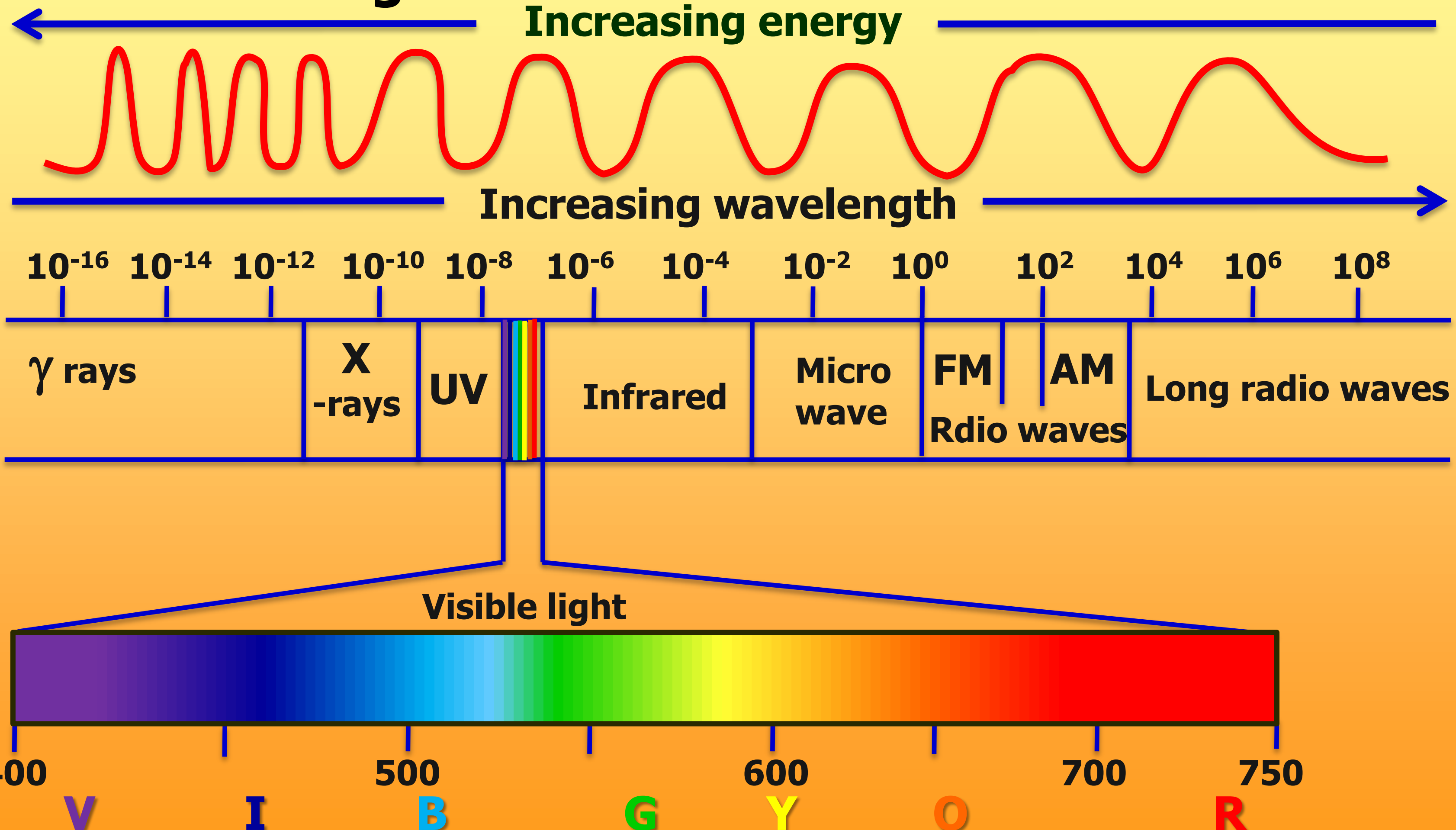
Electromagnetic wave theory:

- **It was Maxwell in 1870, who showed that visible light consists of electromagnetic waves, i.e., waves consisting of oscillating electric and magnetic fields. These two fields components have the same wavelength and frequency and travel in planes perpendicular to each other.**

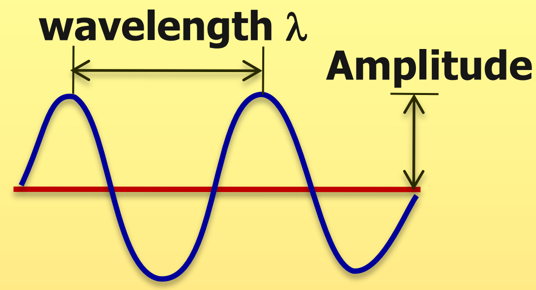


- **It is now well known that there are many types of electromagnetic radiations and visible light is one of them.**
- **The Electromagnetic Spectrum:**
  Cosmic rays < γ - rays < X - rays < ultra-violet rays < visible rays < infrared waves < micro-waves < radio waves (increasing wavelengths).

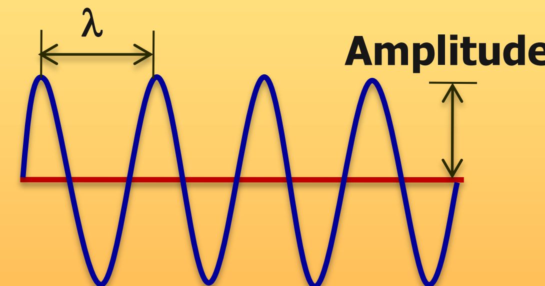
All electromagnetic radiations are characterized by their wavelengths.



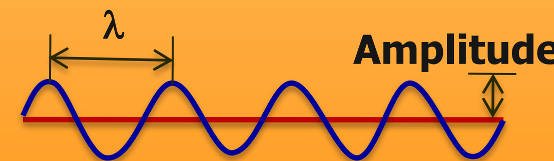
- **Here ν = frequency of wave (unit hertz) = number of waves that pass a given point in one second, i.e., number of oscillations per second.**
- **λ = wavelength normally given in Å = the distance between two successive crests and troughs.**
- **A = amplitude, it is the height of crest or the depth of trough.**
- **$\bar{\nu}$ = wave number, defined as "number of wave-length per unit length".** $\bar{\nu} = \dfrac{1}{\lambda}$



(a) Two complete cycles of wavelength λ



(b) Wavelength half of that in (a); frequency twice as great as in (a)



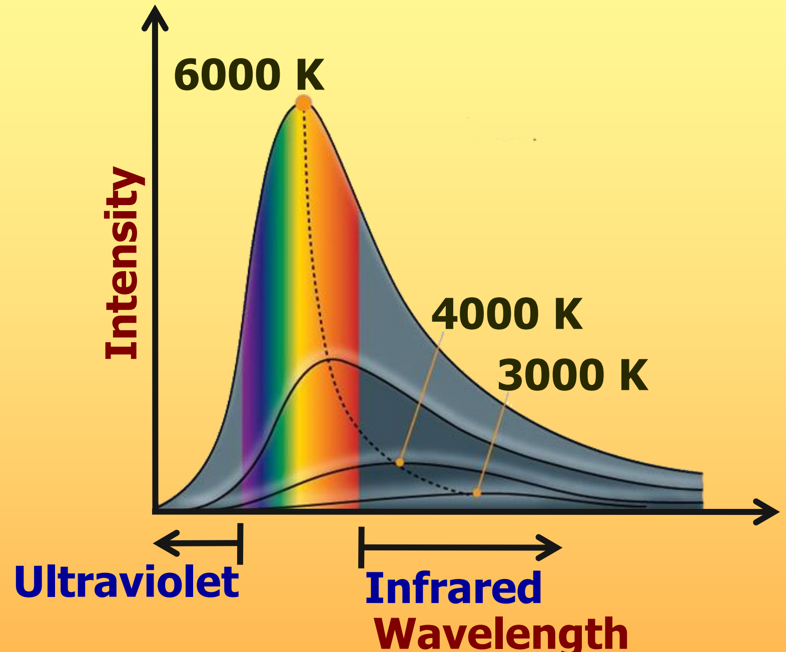
(a) Same frequency as (b), smaller amplitude

- The wave nature of EM radiations can explain the phenomenon of diffraction and interference but unable to explain the black body radiations and photoelectric effect.

Black Body Radiations:

- The ideal body, which emits and absorbs all possible frequencies is called a black-body and the radiation emitted by this body are called black-body radiation.
- When black body is heated, it first becomes red, then yellow, then white with glow and finally blue at very high temperature.
- It means that the radiation emitted goes from a lower frequency (red colour) to a higher frequency (blue colour) as the temperature increases.

- The variation of intensity with wavelength is shown in the figure:

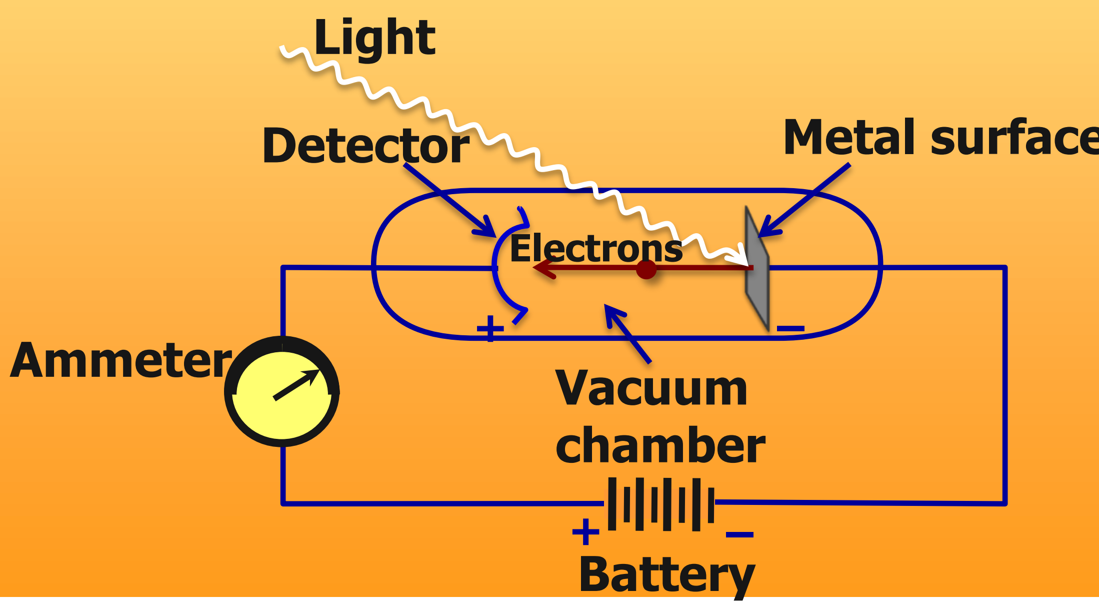


- According to EM wave theory, the energy is emitted or absorbed continuously, thus, the radiations emitted by the body being heated should have the same colour, although its intensity may vary as the heating is continued.

Photoelectric Effect

- **In 1887, H. Hertz performed an 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 of a particular frequency as shown in Figure. The phenomenon is called Photoelectric effect.**

Equipment for studying the photoelectric effect.



- **According to EM wave theory of light, both the numbers of electrons ejected and their energies should depend upon the intensity of incident light, however, in practice, it is found that only the number of ejected electrons depend upon the incident light intensity while their energies do not.**

Particle nature of EM Radiations: Planck's Quantum Theory

- **In 1900, Max Planck provided "Planck's Quantum Theory":**
- **The radiant energy is emitted and absorbed not continuously but discontinuously in the form of small discrete packets of energy, called 'quantum'. In case of light, the quantum of energy is called a 'photon'.**

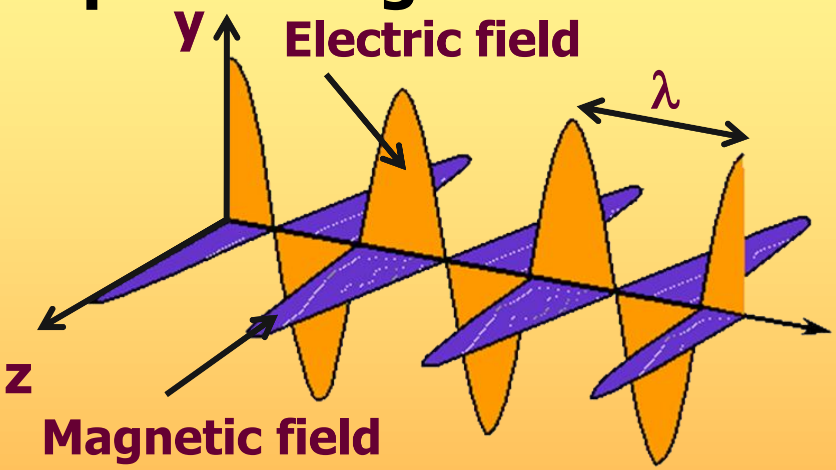
- The energy of a **quantum** is proportional to its frequency as:

$$E = h \cdot \nu$$

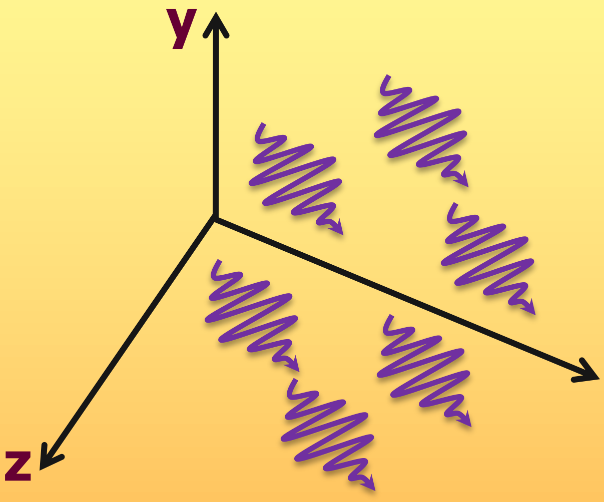
where *h* is Planck's constant (6.626×10^{-34} J s or 6.626×10^{-27} erg s) .

- Thus, according to Planck's theory, energy is always emitted in integral multiples of $h\nu$, $2h\nu$, $3h\nu$, etc.
- Planck was unable to explain that why energies are quantized in this manner. However, Einstein (1905) explained the quantized nature of light while explaining the phenomenon of photoelectric effect.

- Einstein suggested that the light consists of streams of particles called **photon** which move with the speed of light.



A classical view of a travelling light wave



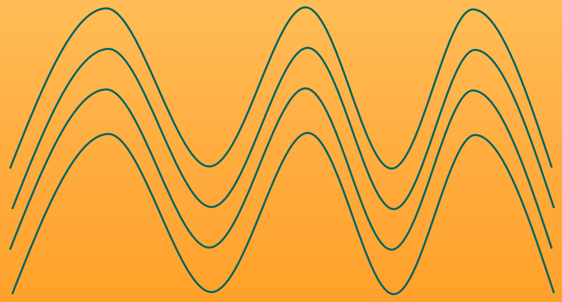
Einstein's photon picture of a travelling light wave

- Using Planck's quantum theory of radiation as a starting point, he deduced that each photon possess energy E which is equal to $E = h\nu = \frac{hc}{\lambda} \quad \left(\because \ \nu = \frac{c}{\lambda}\right)$

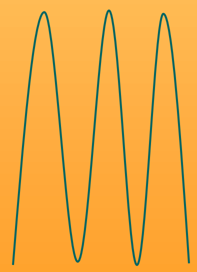
- Thus, shooting a beam of light on a metal surface means shooting a beam of photons. An electron is ejected when a photon strikes that electron and transfers all of its energy to electron.

Remember:

- The intensity of light depends upon the number of photons only, the more intense light, more photons it contains. However, the energy of ejected electron depends upon the energy of photon, i.e., the more energetic the photon, the more energy it transfers to electron resulting to greater energy of electron ejected.



More intense light



More energetic light

- The energy possessed by one mole of quanta (or photons), i.e., Avogadro's number (N_0) of quanta is called one einstein of energy.
- 1 Einstein of energy (E) = $N_0 h\nu = N_0 hc/\lambda$
- The energy acquired by an electron when accelerated through a potential difference of one volt is called one-electron volt ($1eV=1.602 \times 10^{-19}$ J).
- There is a characteristics minimum frequency for each metal called threshold frequency (ν_0) below which the photoelectric effect does not occur. We can calculate, minimum amount of energy required to remove an electron from the metal surface, with the help of frequency as :$E = h \cdot v_0$
- This is called work function (w) of metal.

- If the energy of the incident light (ν) is more than the threshold frequency (ν_0), the excess of energy ($h\nu$ - $h\nu_0$) is imparted o the electron as kinetic energy, i.e., K.E. of the ejected electron,

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

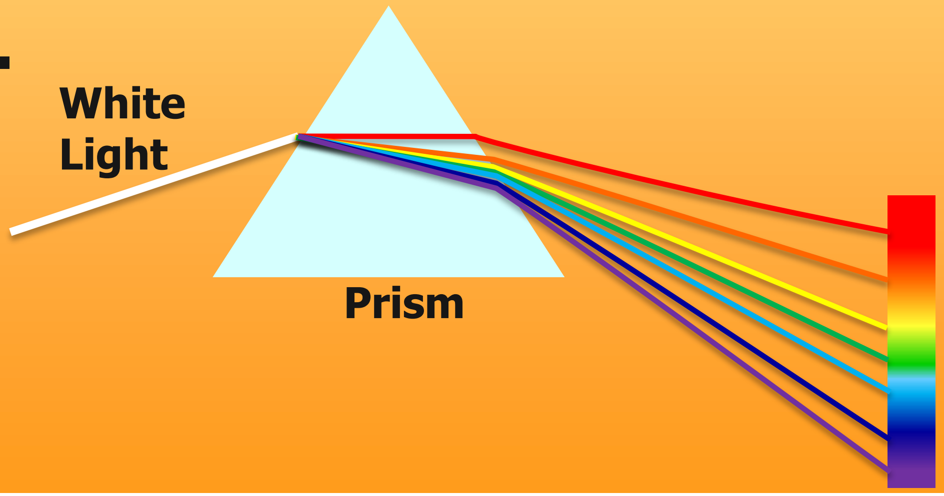
$$h\nu = h\nu_0 + \frac{1}{2}mv^2 = W_0 + \frac{1}{2}mv^2$$

- Hence, greater the frequency of the incident light greater is the kinetic energy of the emitted electron.
- **Note:** Light is said to have a dual nature, i.e., it behaves as a wave as well as a stream of particles. These ideas was put forward by Einstein in 1905.

Emission and Absorption spectra

- A spectra can be defined as “Pictorial representation of arrangement of radiations in increasing order of wavelengths or decreasing order of frequencies”.
- The evidence for quantized electronic energy levels comes from atomic spectra.
- The spectrum may be continuous, when one radiation merges with other at its boundary, i.e., when no sharp boundaries are between adjacent radiations. On the other hand, a spectra which contains sharp lines with sharp boundaries is called line spectra or discontinuous spectra.
- The branch of science dealing with the study of spectra is called spectroscopy.

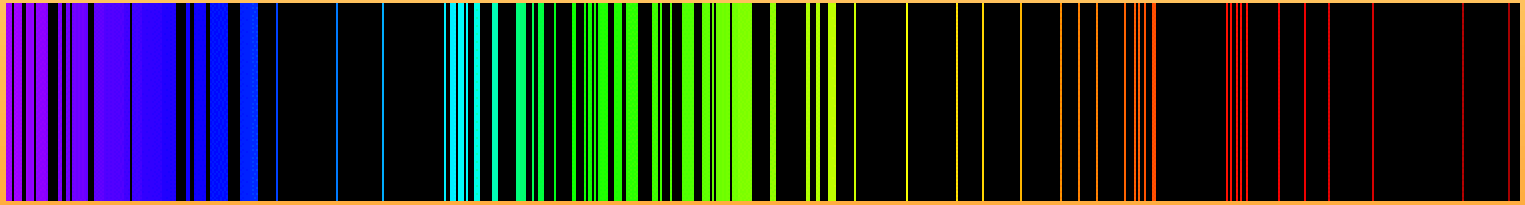
- **Emission spectra: the spectra of radiations emitted by a substance (atoms, molecules or ions) after absorbing energy is called emission spectrum. It is of two types:**
- **The emission spectra of white light from sun, a bulb or any hot glowing body are continuous, i.e., one radiation merges with other at its boundary and no sharp boundaries are found between adjacent radiations.**



- **The emission spectra of atoms in gaseous phases (after being excited, i.e., after absorbing energy) do not show a continuous spread of wavelength. Here they emit radiation only at specific wavelengths. Such an spectra which contains bright lines with sharp boundaries is called line spectra or discontinuous spectra.**
- **Note: The line spectra are also called atomic spectra as they are obtained as a result of absorption and subsequent emission of energy by electrons in the individual atoms of the elements.**
- **Each element has a unique line emission spectra and the characteristic line can be used in chemical analysis to identify unknown atoms.**

- Hence, the spectra of the elements are their **finger prints** differing them from each-other. Elements like Rb, Cs, Tl, In, Ga and Se were discovered when their minerals were analysed by spectroscopic methods.

Emission spectrum of Hydrogen

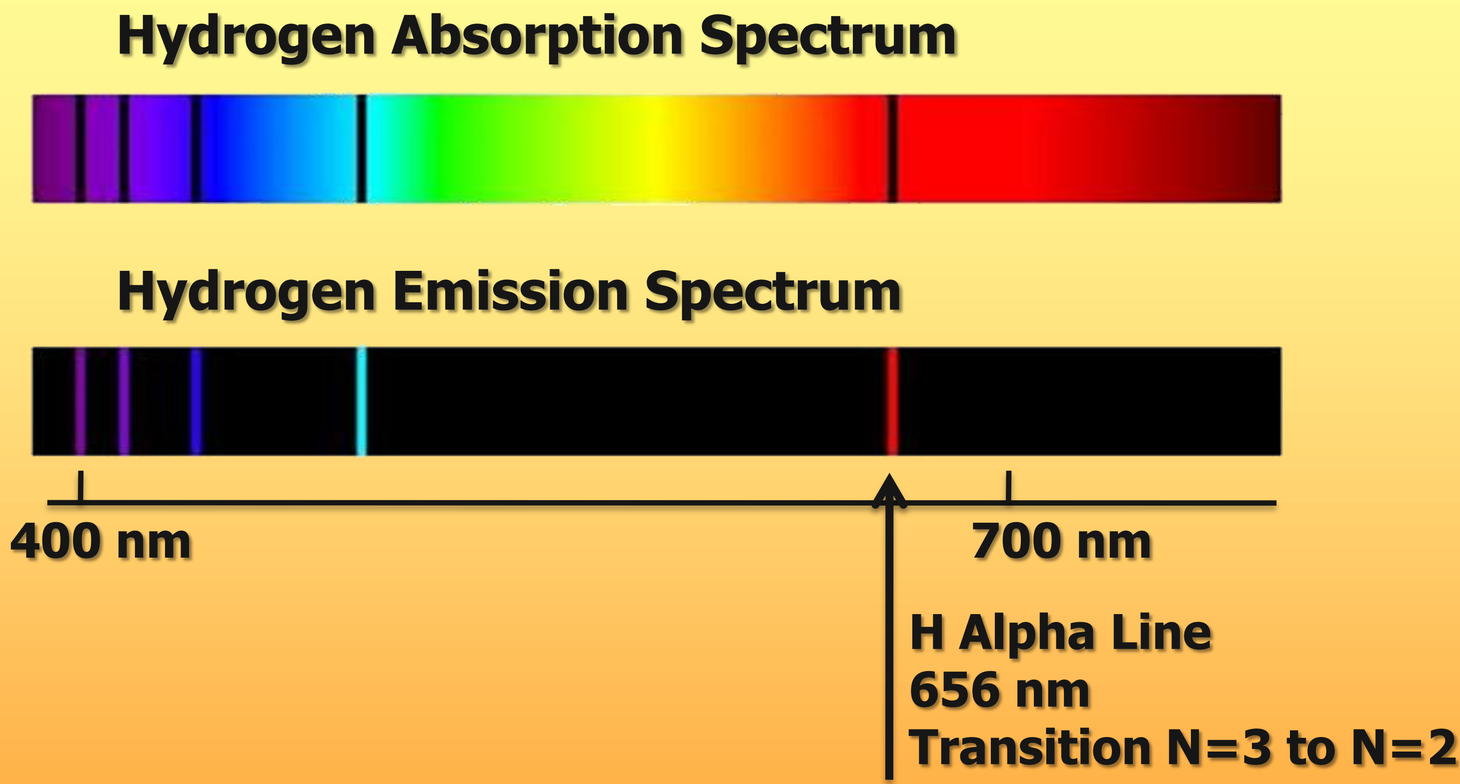


Emission spectrum of Iron

- Helium was discovered after spectroscopic study of sun.

Absorption Spectra:

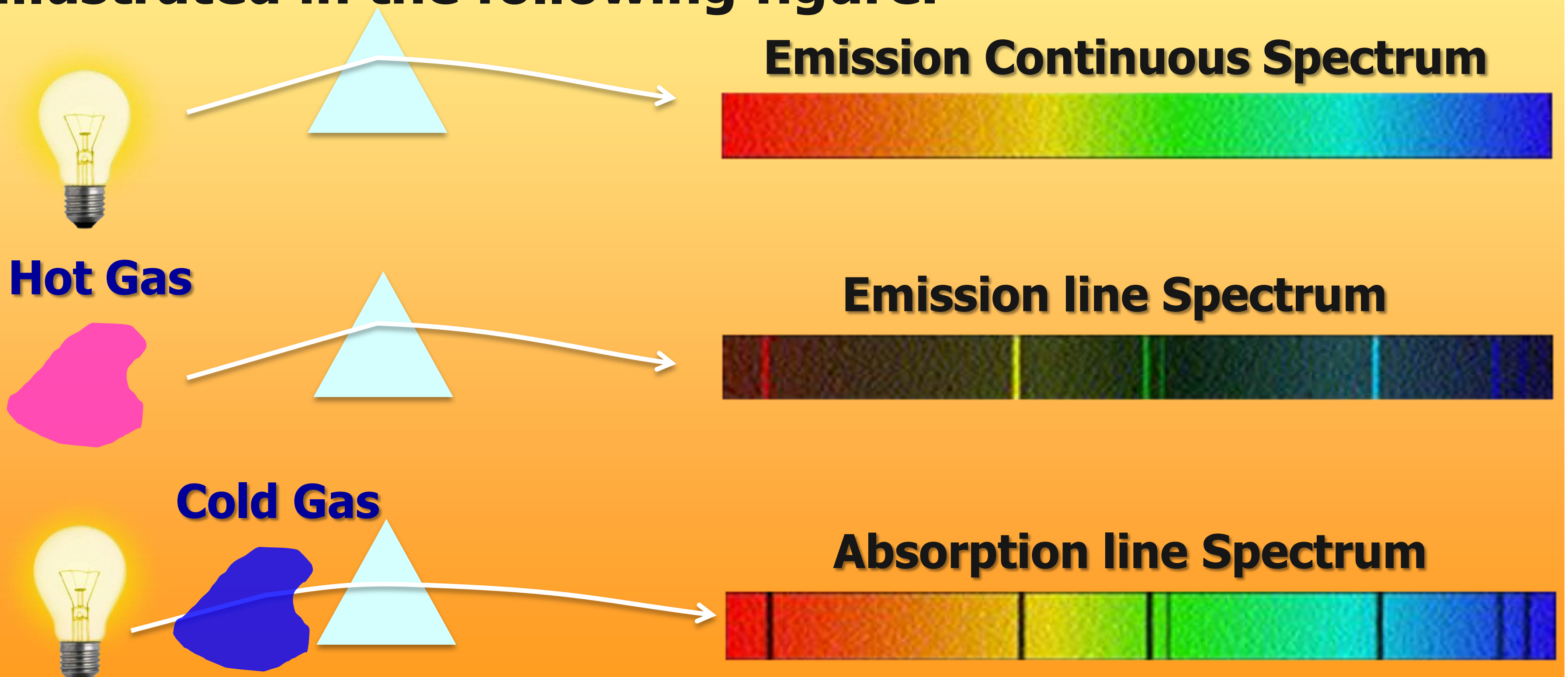
- An absorption spectra is like the photographic negative of an emission spectra. Here a continuum of radiations is passed through a sample which absorbs the radiations of certain wavelengths. The missing wavelengths (which corresponds to the radiations absorbed), leave dark lines in the bright continuum spectrum.
- It is observed that the dark lines are at the same place where coloured lines are obtained in the emission spectra for the same substance. This shows that the wavelengths absorbed were the same as were emitted in the emission spectra.



- **Note:** **Absorption spectrum is always discontinuous spectrum consisting of dark lines.**

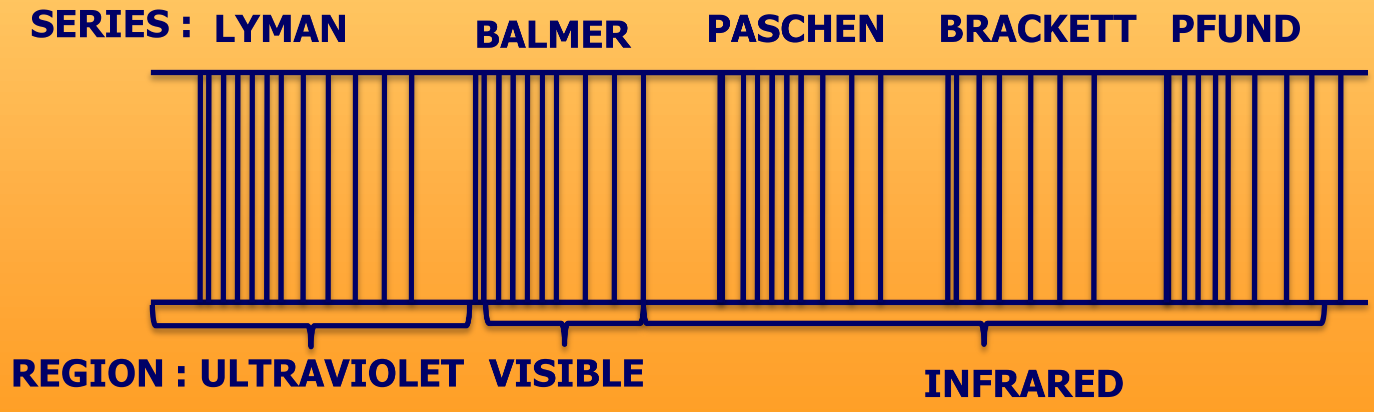
Origin of Continuum, Emission, and Absorption Spectra

The origins of these three types of spectra are illustrated in the following figure.



Line Emission (Atomic) Spectrum of hydrogen

- It can be obtained by taking hydrogen gas in the discharge tube at low pressure which dissociates into energetically excited hydrogen atoms on electric discharge and by passing the light emitted by it through spectrograph.
- The spectrum of hydrogen consists of large number of sharp lines, which are grouped into different series, named after the discoverers. Each of which corresponds to a particular frequency.



- Balmer, in 1885 showed that if spectral lines are expressed in terms of wave number $\bar{\nu}$, then the visible lines of hydrogen spectrum obey the following formula:

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

- Where n= an integer equal to or greater than 3.
- Swedish spectroscope Rydberg gave above formula in more generalized from which is applicable to all the series of hydrogen spectrum and can be written as:

$$\bar{\nu} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)\ \text{cm}^{-1}$$

Where, R = Rydberg's constant = 109,677 cm^{-1}, n_1 and n_2 are two intergers called Rydberg's integers and $n_2 > n_1$. For a particular series n_1 is constant but n_2 varies.

- For example:

Series			Value of n_1	Value of n_2
Lyman	-	UV	1	2,3,4,5,….
Balmer	-	Visible region	2	3,4,5,6,….
Paschen	-	IR	3	4,5,6,7,….
Brackett	-	IR	4	5,6,7,8,….
Pfund	-	IR	5	6,7,8,9,….

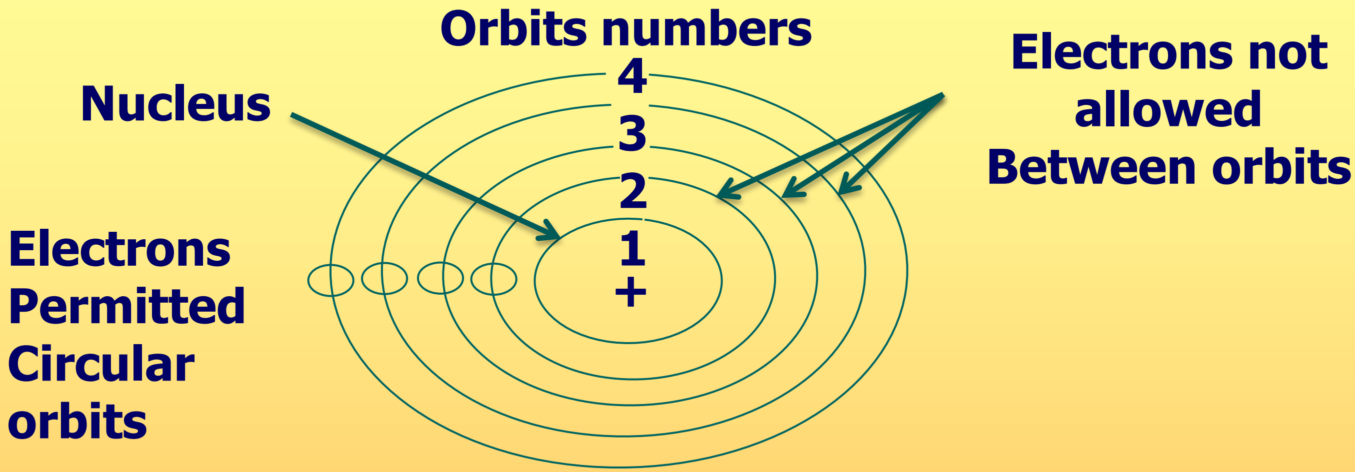
- This formula is only valid for hydrogen (with z = 1). For other elements the formula is written as:

$$\bar{\nu} = \frac{1}{\lambda} = R.Z^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

- Where Z = atomic number of element.
- Limiting line: The line in the hydrogen spectrum when n_2 in the Rydberg formula is infinity ($n_2 = \infty$)

Bohr's Model of atom

- To overcome the drawbacks of Rutherford's atomic model and to explain why elements produce line spectra, Neils Bohr (1913) proposed a model of atom based upon Planck's quantum theory. According to which:
- Electrons revolve around the nucleus in specific circular orbits (situated at a definite distance from the nucleus) with a definite velocity.
- Each orbit has a definite energy associated with it. These are called stationary orbits or energy levels.
- Different energy levels are denoted as n= 1,2,3…….. (or K,L,M,N etc) and called principal quantum number by Bohr.



- **The orbital angular momentum of the electron moving in a circular orbit is quantized (can only take discrete values) and an integral multiple of $\frac{h}{2\pi}$.**

$$\text{Angular momentum (mvr)} = \frac{nh}{2\pi}$$

- **m = mass of the electron, v = velocity of electron, r = radius of the orbit, n = number of the orbit.**

- **Energy of the electron in the n^{th} Bohr's orbit of H atom**

$$E_n = -\frac{2\pi^2 me^4}{n^2h^2}$$

- **On substituting the values of m, e and h, we get**

$$E_n = -\frac{1312}{n^2}\,\text{kJ/mol}$$

$$\text{or } E_n = -\frac{13.6}{n^2}\,\text{eV/atom}$$

- **Thus, the Ist energy level (n=1) has lowest energy.**
- **For H-like particles, e.g., He^+, Li^{2+} etc., the expression for energy is,**

$$E_n = -\frac{1312\,Z^2}{n^2}\,\text{kJ/mol}$$

- **Z = atomic number of the element.**

Radius of the n^{th} Bohr's orbit of H atom

$$r_n = \frac{n^2h^2}{4\pi^2kme^2}$$

- By putting the values of all the constants we get

$$r_n = n^2a_0$$

- where a_0 = 0.53 X 10^{-8} cm = 0.53Å.
- a_0 is the radius of the first stationary state and is called Bohr radius.
- For H- like particles, the radii of the stationary states are given by the expression:

$$r_n = \frac{n^2a_0}{Z}$$

- Z = Atomic number of the element.

Velocity of the n^{th} Bohr's orbit of H atom

$$V_n = \frac{2\pi k e^2}{nh}$$

- By putting the values of all the constants we get

$$V_n = \frac{2.184 \times 10^8}{n}\ \text{cm s}^{-1} \text{ or } V_n = \frac{V_0}{n}$$

- Where $V_0 = 2.184 \times 10^8$ cm s^{-1} is the velocity of the electron in the first orbit of hydrogen atom.

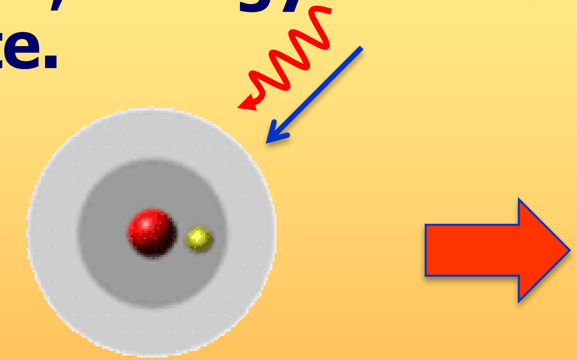
For H – like atom :

$$V_n = \frac{2.184 \times 10^8}{n} \times Z\ \text{cm s}^{-1}$$

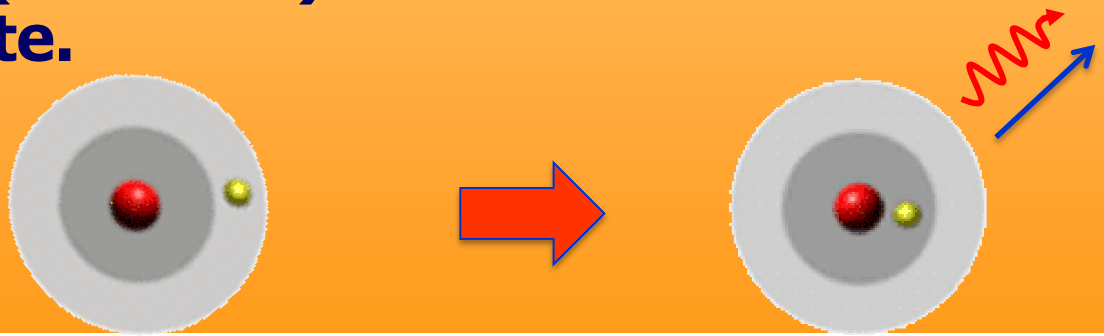
- Z = Atomic number of the element.

Electronic Transition between Energy levels

- While in orbit, e- can neither gain or lose energy.
- When an e^- jumps from lower energy level to higher energy level, energy is absorbed. Now e^- is called in exited state.



- When e- falls from higher state to lower state, energy is emitted (Emission). Now electron is said to be in ground state.



Explanation of line spectrum of Hydrogen with the help of Bohr's Model

- When an electron moves from one energy level to another, the energy gap between the two orbits is

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

- The frequency, $\nu = \frac{\Delta E}{h} = \frac{R_H}{h}\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$

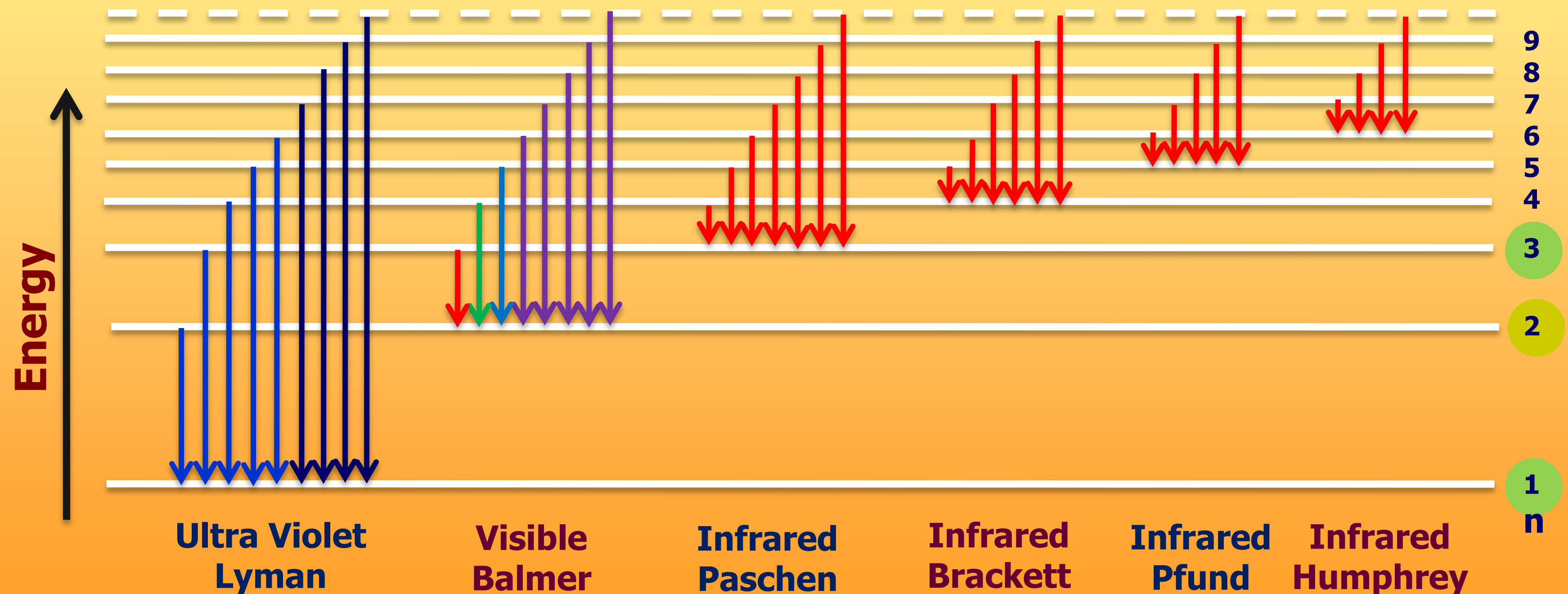
- Wave number, $\bar{\nu} = \frac{\nu}{c} = \frac{R_H}{hc}\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$

- By putting the values of constants,

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

- Thus, the expression is similar to that of Rydberg formula.

- **Each spectral line in a spectrum can be associated to the particular transition in hydrogen atom. Different possible transitions can lead to the large number of spectral lines.**

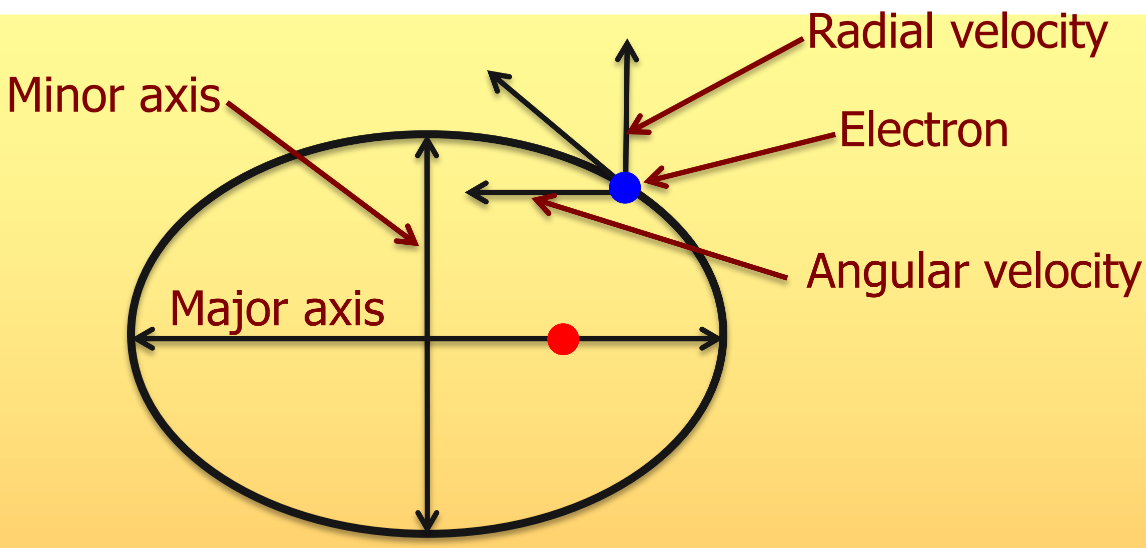


Limitations of Bohr's Model

- **Bohr left the following facts unexplained:**
- **Fine spectrum of atom.**
- **Spectrum of multielectron system.**
- **Zeeman effect and Stark effect.**
- **Three dimensional existence of atom.**
- **Dual nature of electron.**

- **The Sommerfeld model or Bohr-Sommerfeld model:**

- **Sommerfeld (1915) suggested the existence of elliptical orbits along with circular orbits (with situation of nuclei at one of the foci of ellipse) to explain the fine structure in hydrogen spectrum.**



$$\text{Radial momentum} = n_r \times \frac{h}{2\pi}$$

$$\text{Azimuthal momentum} = n_\phi \times \frac{h}{2\pi}$$

- **n_r = radial quantum number**
- **n_ϕ = azimuthal quantum number**

These n_r and n_Φ are related to principal quantum number of Bohr (n) as :

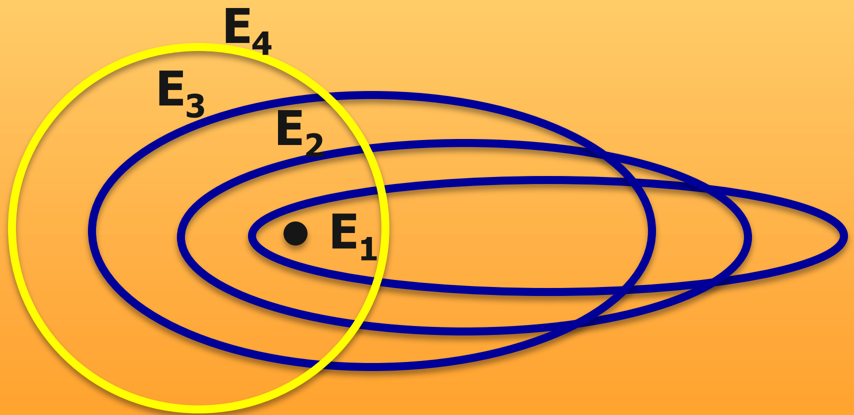
$$\mathbf{n = n_r + n_\varphi}$$

$$\frac{\mathbf{n}}{\mathbf{n_\varphi}} = \frac{\mathbf{n_r + n_\varphi}}{\mathbf{n_\varphi}} = \frac{\textbf{length of major axis}}{\textbf{lengthof minor axis}}$$

Note

1. The n_Φ can never be zero.
2. n_Φ cannot be greater than n.
3. n_Φ equal to n is possible.
4. n_Φ smaller than n is possible.

- Thus, for n = 4, 4 values of n_ϕ are possible as 1,2,3,4. (shown as E_1, E_2, E_3 and E_4 in the following figure.)



n = 4

k = E_1, E_2, E_3, E_4

De Broglie wave equation

- De-Broglie (1924), a French physicist extended the idea of dual nature of light to matter particles and suggested that all matter particles and in motion have a dual nature. The wave associated with matter particles is called de-Broglie's matter wave.
- De Broglie wave equation:

De Broglie wave equation is $\lambda = h/mv$

In case of kinetic energy E, $\lambda = \frac{h}{\sqrt{2mE}}$

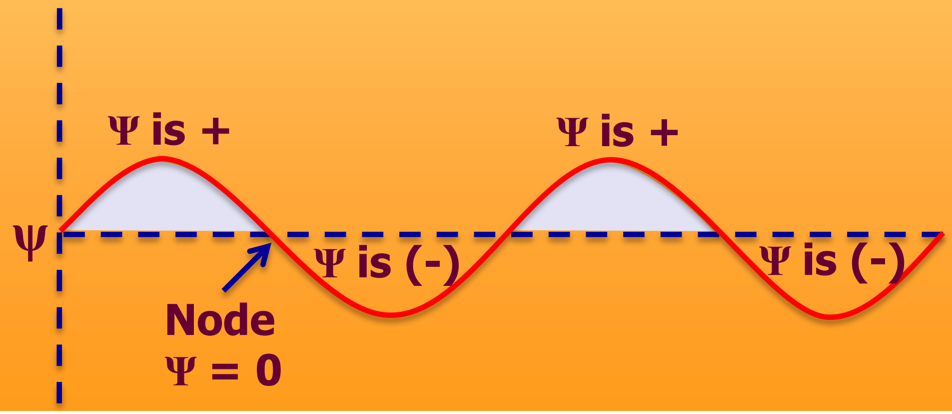
Heisenberg's Uncertainty Principle

It is impossible to determine exactly both the position and the momentum of an electron or of any other small moving particle at the same time. This is **Heisenberg's Uncertainty Principle.**

$$\Delta x \times \Delta p \geq \frac{h}{2\pi};$$

- Δx = Uncertainty in position, Δp = Uncertainty in momentum
- Here it must be keep in mind that this principle applies to location and momentum along the same axis.The relation holds true for microscopic particles.

- The findings of de-Broglie and Heisenberg led the foundation of probability concept about electrons.
- According to the concept of probability, we can only predict the probability or relative chance of finding or locating an electron with a probable velocity in a particular region or space around the nucleus, thus ruling out the possibility of strictly defined orbit as given by Bohr.
- Erwin Schrodinger, an Austrian Physicist, in 1926 developed a new branch of science called quantum mechanics or wave mechanics and gave a model for atom called 'wave mechanical model' of atom.



- **In his model he visualised electron as a three dimensional 'wave in the electronic field' of a positively charged nucleus inside the atom.**
- **Schrodinger wave equation is fundamental equation:**

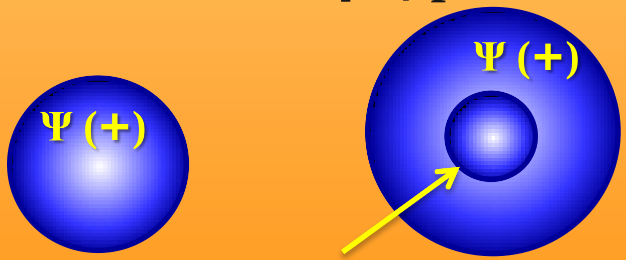
$$\frac{\partial^2\psi}{dx^2}+\frac{\partial^2\psi}{dy^2}+\frac{\partial^2\psi}{dz^2}+\frac{8\pi^2 m}{h^2}(E-PE)\psi=0;$$

Where PE is Potential Energy

Ψ **= wave function**

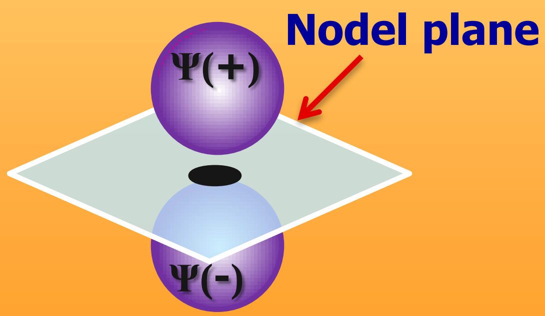
Ψ^2 **= Probability of finding electrons.**

- **At nodes the value of** Ψ^2 **becomes zero.**
- s **orbital for which** $l = 0$**, is spherical in shape,** p **orbital,** $l = 1$**, is dumb – bell in shape.**

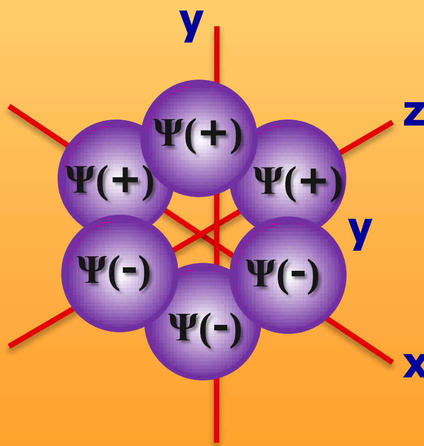


Orbital **Nodel surface** **2s orbital**

- There are three *p* orbitals, perpendicular to each other: p_x, p_y, p_z.
- The two lobes of a p orbital may be separated by a plane that contains the nucleus and is perpendicular to the corresponding axis. Such plane is called a nodal plane.

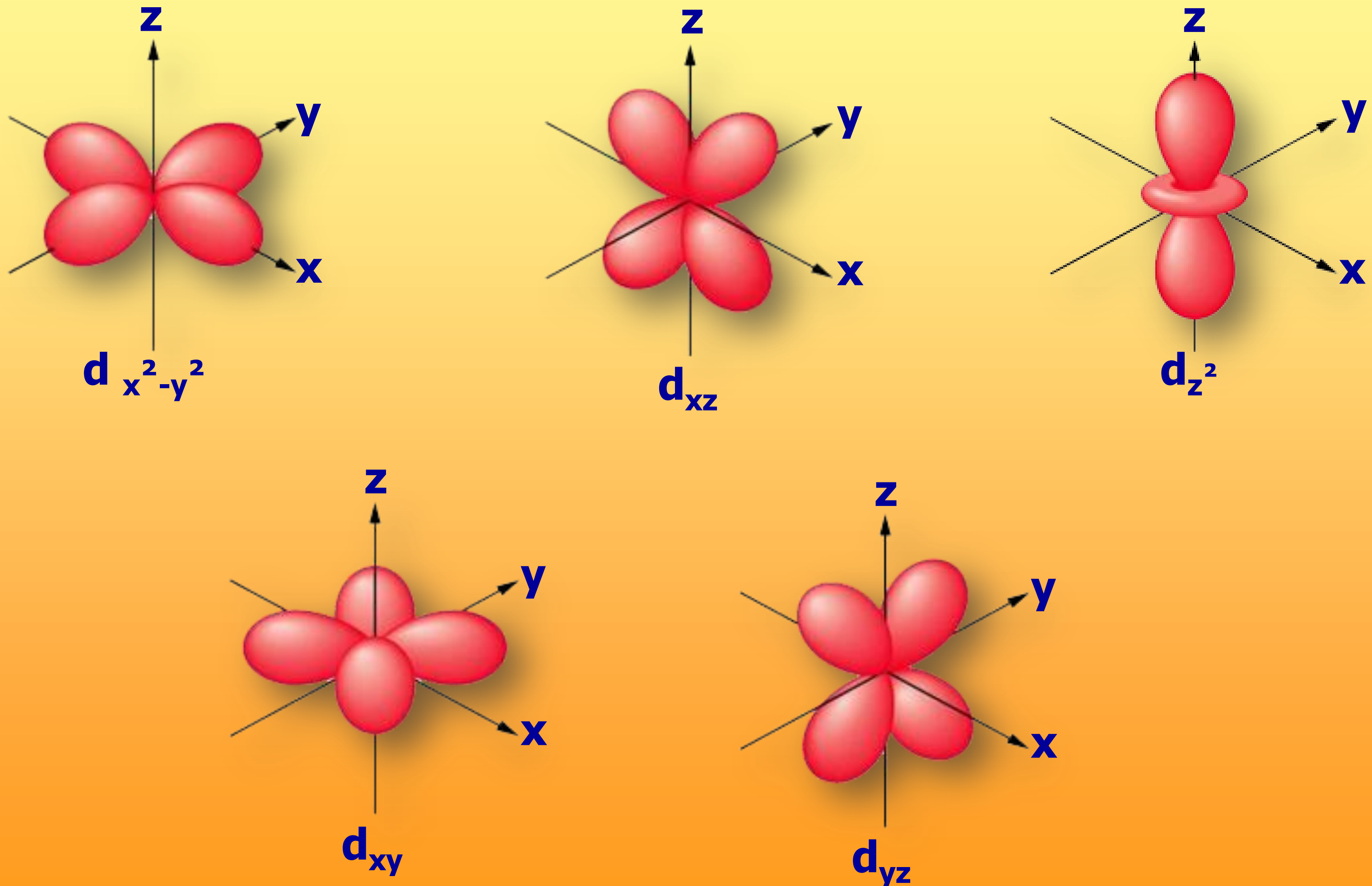


2p orbital



Three mutually perpendicular (orthogonal) 2p orbital

- **There are five d orbitals** d_{xy} , d_{yz} , d_{zx} , d_{z^2} , $d_{x^2-y^2}$.



Quantum Numbers

- Quantum numbers are also known as identification number. Every electron in an atom has a unique set of quantum numbers, which give idea about position of electron in orbital, energy, orientation of orbital in space, spinning of electron. These are of four types.
- Principal quantum numbers (n̂) describes size of an orbital, approximate energy of electron, position of electron in an orbital. Its value can be 1,2,3...... But cannot be zero.
- Azimuthal quantum number (ι):- it gives idea about the shape of an orbital 1 = 0, to (n-1)

n = 1, 1 = 0(s);
n = 2, 1 = 0,1(s,p);
n = 3, 1 = 0, 1, 2 (s, p, d)

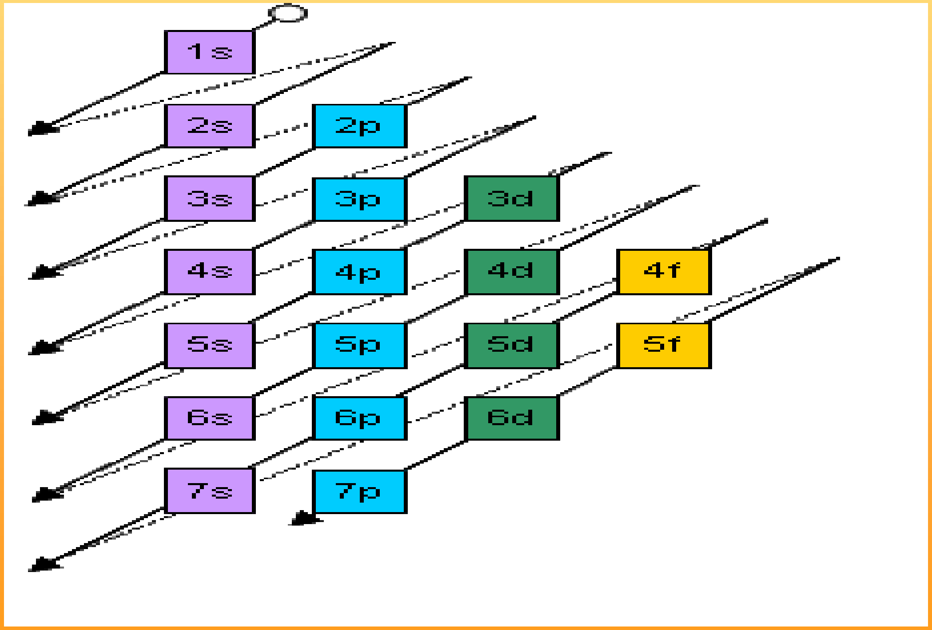
- **Magnetic quantum number (m):- denoted by m, describes orbital's orientation in space.**

n = 1,	**1 = 0,**	**m = 0**
n = 2,	**1 = 0, 1**	**m = 0, ± 1**
n = 3,	**1 = 0, 1, 2**	**m = 0, ± 1, ± 2**

- **Spin quantum numbers(s):- describes the spin or direction in which an electron spins, it can have two values,** $+\frac{1}{2}, -\frac{1}{2}.$

Electronic configuration of Elements

- **The Aufbau Principal**
- **This principle states that in addition to adding protons and neutrons to the nucleus, one simply adds electrons to the hydrogen-like atomic orbitals in order of their increasing energy.**



Electronic configuration of Elements

- Pauli’s Exclusion Principle:
- It is impossible for two electrons in a given atom to have all the four quantum numbers identical. This is Pauli’s Exclusion Principle.

- **Hund's Rule -**

- Electrons occupy all the orbitals of a given sublevel singly before pairing begins. Spins of electrons in different incomplete orbitals are parallel in the ground state.The most stable arrangement of electrons in the subshells is the one with the greatest number of parallel spins.

Some time-saving Shortcut for numericals

- **The number of the total spectral lines when an electron drops from nth shell to the ground state=** $\frac{n(n-1)}{2}$

- **If the maximum wave number(minimum wave length) for any spectral series is asked, the following formula is used:**

$$\therefore \quad \bar{\nu}_{max} = \frac{R_H Z^2}{n_1^2}$$

- **Minimum wave number (maximum wave length) is obtained for the case when:**

$$\therefore \quad \bar{\nu}_{min} = \frac{1}{\lambda_{max.}} = R_H Z^2 \left[\frac{1}{n_1^2} - \frac{1}{(n_1+1)^2} \right]$$

- **The relationship between the energy, wavelength, frequency and momentum of two different radiations or matter waves:**

$$\frac{E_1}{E_2} = \frac{\nu_1}{\nu_2} = \frac{\lambda_2}{\lambda_1} = \frac{p_1}{p_2}$$

- **The radius of Ist orbit for H-atom (r_1) = 0.529** Å
- **The radius of nth orbit for H-atom (r_n) = $n^2 \times$0.529** Å
- **For H-like particles,**
  radius of the first orbit $r_1 = \frac{0.529}{Z}$ Å

  radius of the nth orbit $r_n = \frac{0.529\ n^2}{Z}$ Å

The velocity of electron in different orbits:

- For H-atom velocity of electron in I^{st} orbit (v_1)= 2.188×10^8 cm s^{-1}
- For H-atom, velocity in n^{th} orbit: $v_n = \frac{v_1}{n}$
- For H-like particles, velocity in n^{th} orbit: $v_n = v_1 \times \frac{Z}{n}$
- If the velocity of electron in any one orbit is given and velocity in other orbit is asked, the following formula is used:

$$\frac{V_2}{V_1} = \frac{1}{2}, \frac{V_3}{V_1} = \frac{1}{3} \text{ and so on}$$

The energy of electron in different orbits

- For H-atom energy of electron in Ist orbit (E_1) = -13.6 eV/atom
- Ionisation energy of any electron is the same as its Bohr's energy in any orbit with the opposite sign.
- For H-atom the ionisation energy of electron in Ist orbit (E_1) = 13.6 eV/atom
- I.E. for H-like particles= $Z^2 \times$ I.E.$_H$
- For He^+ = $4 \times$ I.E.$_H$
- For Li^+ = $9 \times$ I.E.$_H$
- If the velocity of an electron in any orbit of the H-like particle is given and the radius of the orbit is asked the following formula is used: $v = \frac{1}{\sqrt{r}}$

- No. of subshells in n^{th} shell = n
- A subshell is only possible if $l < n$.
- No. of orbitals in a subshell (or the total values of magnetic quantum numbers) = (2l +1)
- The total number of lines from the splitting of a single spectral line in the Zeeman or Stark effect= (2l +1)
- Maximum number of electrons in a subshell=2 (2l +1)
- No. of orbitals in n^{th} shell = n^2
- Maximum no. of electrons in n^{th} shell = $2n^2$
- No. of spherical/radial nodes in a orbital = (n-l-1)
- No. of angular nodes in a orbital = l
- No. of total nodes = (n-1)
- Due to the Zeff, the similar orbitals of different atoms have different energy:
- $E_{2s}(H) < E_{2s}(Li) < E_{2s}(Na) < E_{2s}(K)$

Thanks…