6. Structure of atom

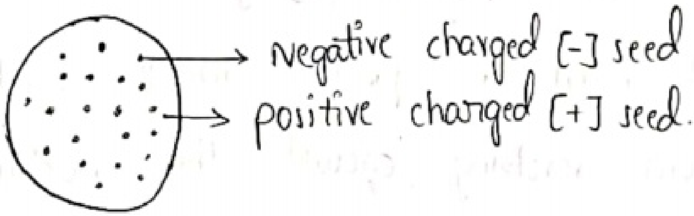
Dalton (1808).

* Dalton in 1808 he explained about atom.
* Atom is a smallest partical which we can't divide further.

J.J Thomson (1903) (1897)

He given first atomic model that is watermelon theory. this is also called as plum pudding model. Atom consist positive and negative charged particles like in a watermelon seeds are known as negative charged particles remaining part known as positive charged particles.

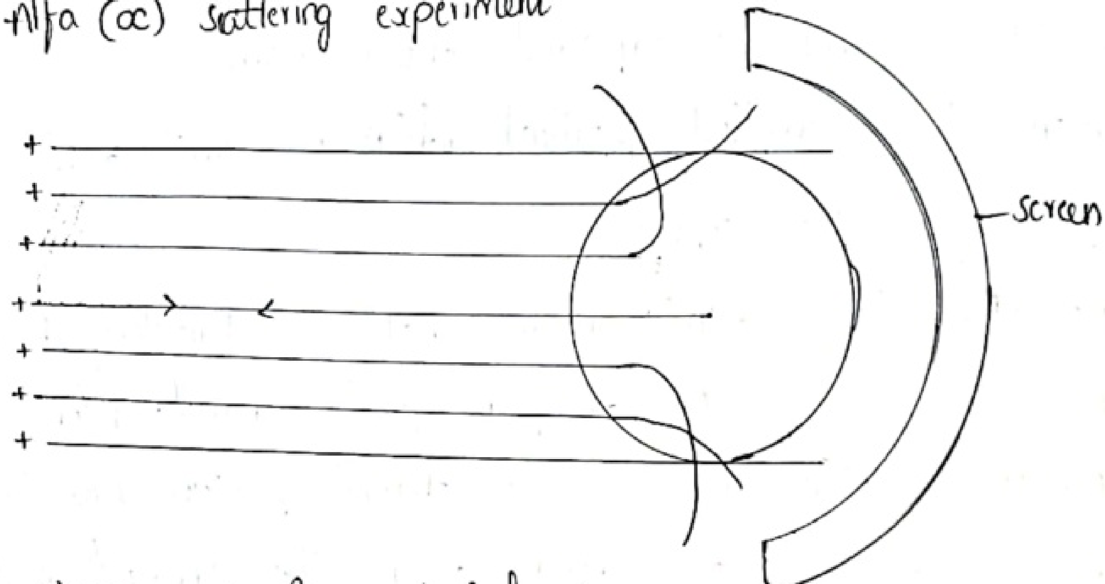
Diagram :



Limitations / Drawbacks : According to classical law of physics the positive and negative never combined each other.

* Ruther ford scattering experiment

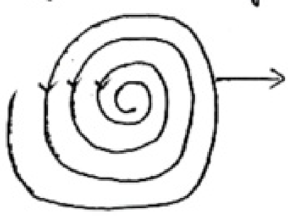
In 1911 Rutheford proposed planatery model based on Alfa (α) scattering experiment.



* the Atoms are in spherical shape mostly hollowspace.
* In its centre small nucleus present protons and neutrons present in the nucleus. [10^{-15}].
* The total mass is present in the nucleus.
* The electrons revolving around the nucleus with high velocities like planets revolving around the sun. so it is called planetary model.

Rutherford limitations

1) According to classical law of physics the rotating electron should loss energy continuosly and then fall into the nucleus.



If it loss energy continuously it should form continuous spectrum but it forms discreet line.

<u>Max plank quantum theory</u>

the energy of a light is directly proportional to frequency of the light.

$$E \propto f$$

$$E \propto \nu$$

$$\boxed{E = h \cdot \nu}$$

E = energy

h = plank constant

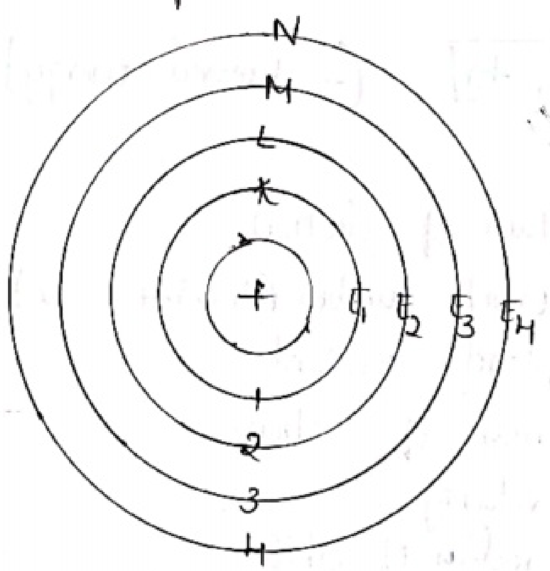
ν = frequency.

$h = 6.625 \times 10^{-27}$ erg·sec

$h = 6.625 \times 10^{-34}$ Joule x sec

<u>Bhor's Atomic model.</u>

In 1913 Bhor proposed atomic model based on max plank quantum theory.



1. stationary orbit

* electrons revolving around the nucleus with high velocities in a specified paths are called orbit.
* As long as the electron rotating in that orbit it should not gain and should not lose the energy. so it is a stationary orbit.

2. Naming the orbits.

* The stationary orbits are denoted by K, L, M, N (or) 1, 2, 3, 4.
* the orbit which is close to the nucleus has less energy compared to outer most orbit.

3) Jumping the electrons

* If electron jumped from higher energy level to lower energy level the energy released as radiation

$$\boxed{E_3 - E_1 = h\nu}$$

* If electron jumped from lower to higher energy absorbed as radiation.

$$\boxed{E_1 - E_4 = -h\nu}$$ (- observed energy)

4) Angular momentum

the angular momentum of electron

$$\boxed{MVR = \frac{nh}{2\pi}}$$

n - orbit number (1, 2, 3, 4 .... α)
h - planck constant
m - mass of electron.
v - velocity.
r - radius of orbit.

Note : The maximum no. of electrons in any orbit → $2n^2$.

Draw back :

* Bhor's atomic model explains only about single electron atomic spectral lines.

  Ex: $H, He^{+1}, Li^{+2}, Be^{+3}$

* It does not explain zeeman effect ; the atomic spectral lines spilts into several lines in the presence of magnetic field.
* It does not explains about starq effect; the splitting of atomic spectral lines into several lines in the presence of electric fields are called starq effect.
* It does not explains about chemical bonds.
* It does not explains about why angular momentum is quantised

$$mvr = \frac{nh}{2\pi}$$

$\therefore n = 1, 2, 3, 4, 5$ - - - - - - - why not 0 :

Bohr's sommerfields models

* elliptical orbit was introduced by sommerfield
* He introduced Azimuthal quantum number (l) orbital Quantum Number.
* Every stationary orbits contains some sub stationary are called orbitals. These are denoted by s, p, d, f.
* These sub-stationary orbits are called orbitals number depends on stationary orbit number.

Examples. NOTE: The maximum electron in any orbit is $2n^2$.

* 1^{st} orbit $= 2 \times 1^2 = 2$
2^{nd} orbit $2 \times 2^2 = 8$
3^{rd} orbit $2 \times 3^2 = 2 \times 9 = 18$
4^{th} orbit $2 \times 4^2 = 2 \times 16 = 32$.

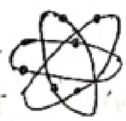
Note :-

* In s orbital '2' electrons can fill $\boxed{\uparrow\downarrow}$
* In p orbital '6' electrons can fill $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}$
* In d orbital '10' electrons can fill $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}$
* In f orbital 14 electrons can fill $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}$

we can find the no. of orbitals by using the formula: $[2l + 1]$.

Elliptical orbits

* This is called elliptical orbits. ←

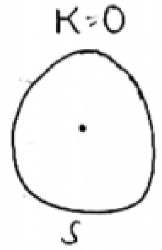


* Each orbit contains 2 electrons.

values of l:-

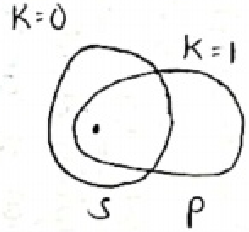
$s = 0$
$p = 1$
$d = 2$
$f = 3$

I 1 orbit [K] = 1 orbital [s], K = 0, n = 1, = K

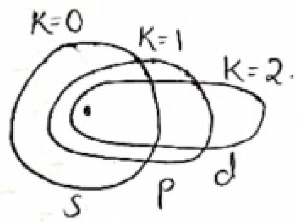


K = orbit
S = orbital.

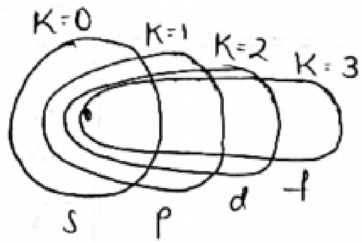
II 2 orbit [K, L] = 2 orbitals [s, p], K = 0, 1, n = 2 = L.



III 3 orbit [K, L, M] = 3 orbitals [s, p, d], K = 0, 1, 2, n = 3 = M.



IV 4 orbit [K, L, M, N] = 4 orbitals [s, p, d, f]; K = 0, 1, 2, 3, n = 4 = N.



Quantum Number

Slno	Principle Quantum Number	Azimuthal Quantum number	Magnetic Quantum number	spin Quantum number
i)	principle Quantum Number was introduced by Neils born	Azimuthal Quantum Number was introduced by sommer field	Magnetic Quantum number was introduced by landey	spin Quantum number introduced by ulhen beck and Goud smith
ii)	It is denoted by - n	It is denoted by - l	It is denoted by ml [or] m	It is denoted by sl [or] s [or] sm
iii)	The values of $n = 1, 2, 3 \ldots \infty$ limits: minimum = 1 maximum = ∞	The values of $l = 0, 1, 2 \cdots (n-1)$ limits: minimum = 0 maximum = $(n-1)$	The values of $m = (2l+1)$ values = $-l, 0, l$ limits: minimum = $-l$ maximum = l ex: If $l = 2$ $m = (2(2)+1)$ $= 5 =$ [5 values] $\therefore -2, -1, 0, 1, 2$	The values of $s = +\frac{1}{2}, -\frac{1}{2}$ limits minimum = $-\frac{1}{2}$ maximum = $+\frac{1}{2}$
(iv)	It explains about size and energy of stationary orbit	It explains the shape of the orbital $\therefore$ s = spherical shape p = dumble shape d = double dumble shape f = four folded dumble	it explains about orientation of orbital $\therefore$ s = s(1) p = p_x, p_y, p_z (3) d = $d_{xy}, d_{yz}, d_{xz}, d_{x^2y^2}, d_{z^2}$ [5] f = [7]	It explains about rotation of electrons $\therefore$ clock wise = $\frac{1}{2}\uparrow$ Anti-clock wise = $\frac{1}{2}\downarrow$

No. of

No of orbits	Name of orbit	No of orbital	Name of orbital	energy levels	No of electrons	Total ($2n^2$) electrons
1	K	1	S	1	2	2
2	L	2	s p	1 3(p_x, p_y, p_z)	2 6	8
3	M	3	s p d	1 3(p_x, p_y, p_z) [$d_{xy}, d_{yz}, d_{zx}, d_{x^2y^2}, d_{z^2}$]	2 6 10	18
4	N	4	s p d f	1 3(p_x, p_y, p_z) ($d_{xy}, d_{yz}, d_{zx}, d_{x^2y^2}, d_{z^2}$] = 5 [7]	2 6 10 14	32

Q. What are difference between orbit and orbitals?

	orbit	orbitals
1)	orbits have a definite boundary.	1) orbitals have no boundary
2)	They are circular in shape.	2) The shape of each orbital is different
3)	orbit is 2D	3) orbital is 3D
4)	It is represented by K, L, M, N.	4) It is represented by s, p, d, f.
5)	The path of electron around the nucleus	5) The maximum probability of finding the electron around the nucleus

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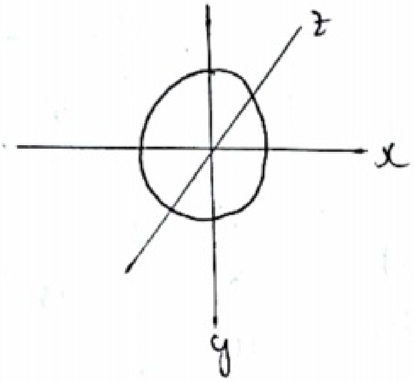
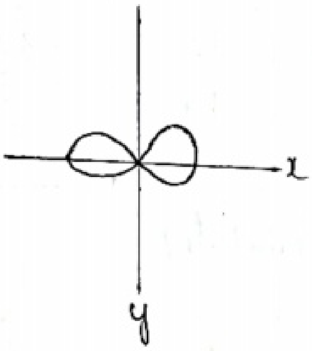
Shape of orbitals

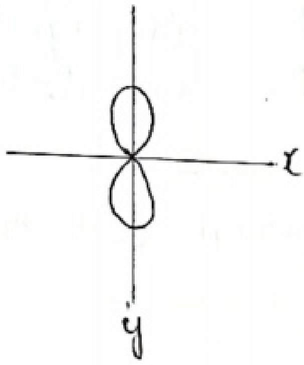
s- orbital , shape = spherical.



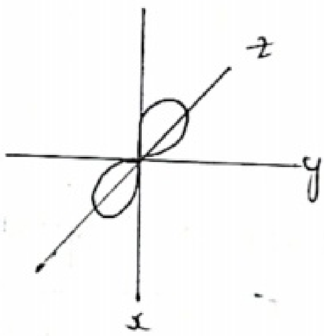
p-orbital, shape = Dumbell.



px

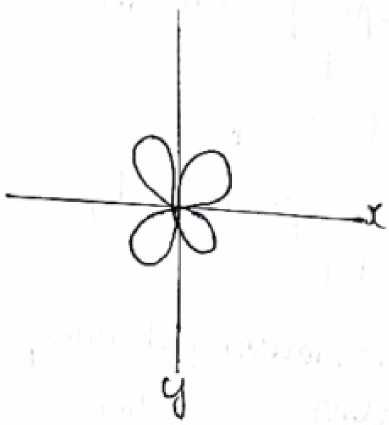


py

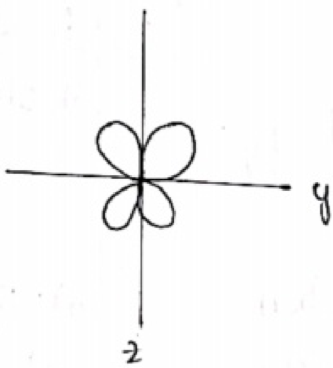


pz

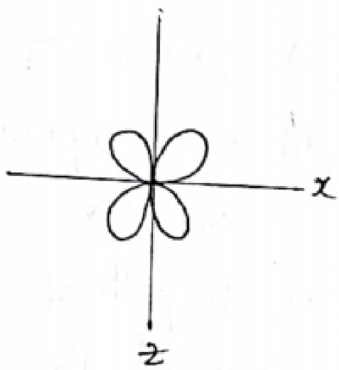
d-orbital , shape = double dumbell.



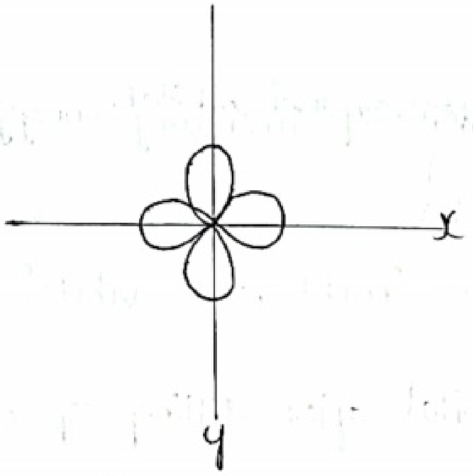
dxy



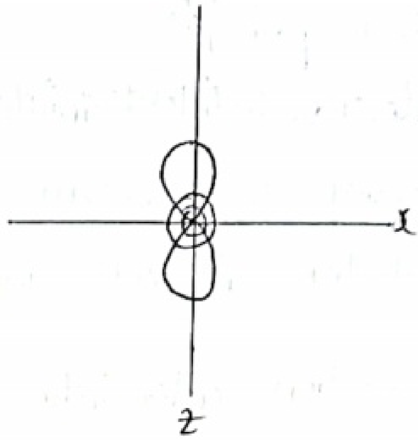
dyz



dzx



$d_{x^2y^2}$



d_{z^2}.

f - orbital

shape : four folded dumbell / complex shape.

Electronic configuration :-

The distribution of electrons among the atomic orbitals is called electronic configuration.

Ex :- Na [Z = 11]

$1s^2\ 2s^2 2p^6\ 3s^1$

nl^x method.

n = principle Quantum number

l = orbital Quantum number.

x = electrons number.

* It is used to write the electron configuration.

Ex :- He [Z = 2].

electrons = 2.

($1s^2$) → electron configuration.

$1s^2$
- 1 → principle Quantum number.
- s → orbital Quantum number
- 2 → electrons number.

* It is used to find the positions of electrons around the neucleus in an atom.

Aufbau principle

1) The electron first enter into least energy level orbital.

According to quantum numbers

2) The electron 1st enter into least [n+l] value orbital.

Ex: why electron enter into 4s orbital after filling 3p orbital but not 3d?

An According to Aufbau principle.

(n+l) value of 3d and 4s

3d	4s
[3+2]	[4+0]
= 5	= 4

The electron enters 4s orbital because (n+l) value of 4s is less than 3d.

3) If (n+l) values are same then the electron first enter into least 'n' value orbital.

Ex: After filling of 4s orbital electron enter into 3d but not 4p. why?

Ans: According to Aufbau principle.

(n+l) value of 3d	(n+l) value of 4p
(3+2)	= (4+1)
= 5	= 5

∴ The value of (n+l) are same and the 'n' value is least in 3d.

∴ electron first enters into 3d before 4p.

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II **Hund's Rule**

According to this rule election pairing in orbitals starts only when all available empty orbitals of the same energy [degenerate orbitals] are singly occupied.

Ex: N[Z=7]

$\frac{1s^2}{2}$ $\frac{2s^2 2p^3}{5}$

Black diagram: $\underset{1\ 2}{\boxed{\uparrow\downarrow}}$ $\underset{3\ 4}{\boxed{\uparrow\downarrow}}$ $\overset{5\quad 6\quad 7}{\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}}$

p_x p_y p_z

III **pauli's diagram:-**

According to pauli's exclusion principle no two electrons of the same atom can have all four quantum numbers are same.

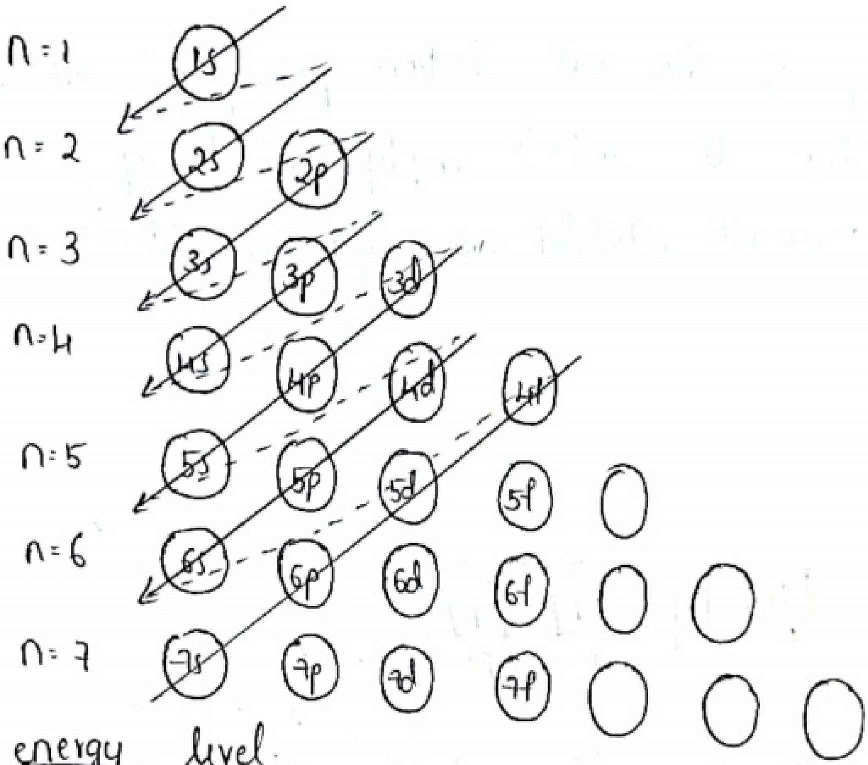
Ex: He [$1s^2$] Black diagram: $\boxed{\uparrow\downarrow}$

election	n	l	ml	sm
1	1	0	0	+1/2
2	1	0	0	-1/2.

2) Na [$1s^2\ 2s^2\ 2p^6\ 3s^1$]. $\overset{1s^2}{\boxed{\uparrow\downarrow}}$ $\overset{2s^2}{\boxed{\uparrow\downarrow}}$ $\overset{2p^6}{\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}}$ $\overset{3s^1}{\boxed{\uparrow\downarrow}}$

election	n	l	ml	sm
3	2	0	0	+1/2
5	2	1	-1	+1/2
7	2	1	+1	+1/2
9	2	1	0	-1/2
10	2	1	+1	-1/2
11	3	0	0	+1/2

Moellar chart



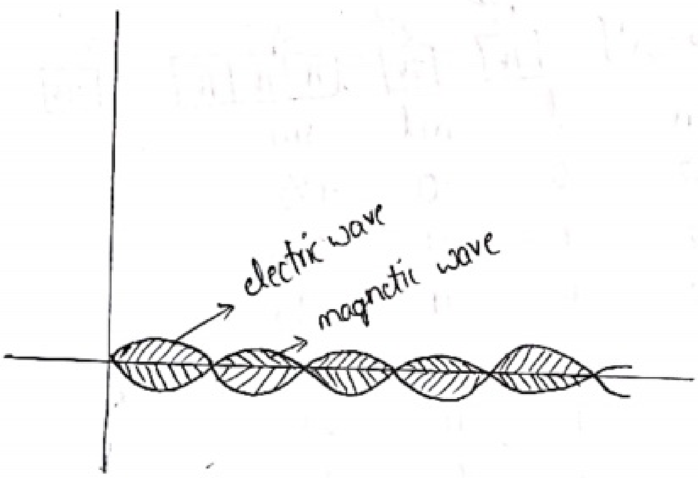
energy level

1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s

Spectrum

The group of wave lengths or frequencies is called spectrum.

⟹ electro magnetic waves



properties

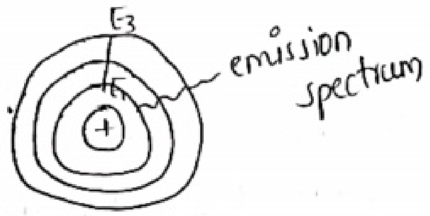
1) Electromagnetic waves can be described by means of vibrating electrical and magnetic fields which are mutually perpendicular to each other.

2) The nature of these waves is transverse.

3) These waves travel with a speed equal to the speed of light in vaccum.

Emission spectrum

* The spectrum formed by emission of radiation.
* when electron jump from higher energy level to lower energy level then the energy is released as radiation. Due to this radiation the emission spectrum is formed.



Absorption spectrum

* The spectrum formed by the absorption of energy when electron jumps from lower energy level to higher energy level is called absorption spectrum
* It contains dark lines on bright background.

wave length

$v = \gamma \lambda$

C or [v]: velocity / high velocity

γ = frequency

λ = wave length

Q. The wave length of radiowave is 1 meter ? Find its frequency ?

$C = \gamma \lambda$

$\gamma = C/\lambda$

$\gamma = 3 \times 10^8$.

$\gamma = 3 \times 10^8 \, Hz$

Flame test

* when we heat an iron rod some of the heat is emitted as light.
* First it turns into red colour because red colour has more wave length, least frequency and less energy
* As the temperature increases the iron rod turn into orange, yellow, blue ... more energy and lower frequency.
* If the temperature increases more findly it turns into white colour.

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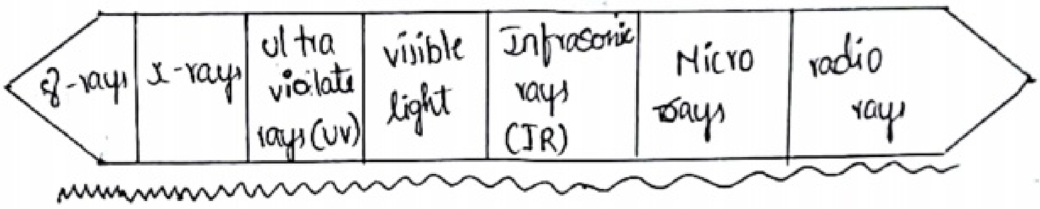
Electromagnetic spectrum

Electro magnetic waves can have a wide varidy of wavelengths. The entire range of wave length is known as electromagnetic spectrum.



The electro magnetic spectrum consists of a continuous range of wavelengths of gamma rays at the shorts wavelength of ratio waves at the longer wavelength. But our eyes are sensitive only to visible light.