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Structure of atom

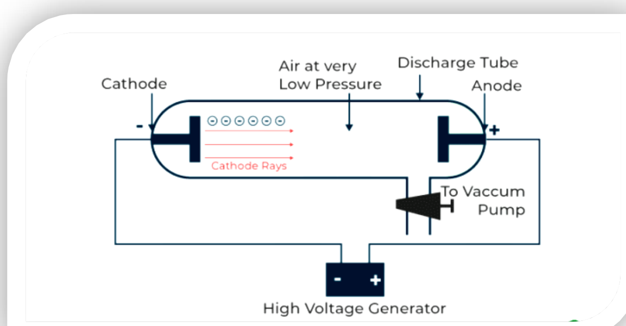
Discovery of fundamental particles

Dalton's atomic theory was able to explain the law of conservation of mass, law of constant composition and law of multiple proportions very successfully but it failed to explain the results of many experiments.

Discovery of electron

William Crookes in 1879 studied the electrical discharge in partially evacuated tubes known as cathode ray discharge tubes.

1. A discharge tube is made of glass, about 60cm long. This is known as crooke's tube.
2. The negative electrode is known as cathode and positive electrode is known anode.
3. When a gas enclosed at low pressure(10^{-4} atm) in discharge tube is subjected to a high voltage (10,000V).
4. Invisible rays originating from the cathode and producing a greenish glow behind the perforated anode on the glass wall coated with phosphorescent material ZnS is observed.
5. These rays were called cathode rays.



Properties

- They produce sharp shadow of the solid object in their path suggesting that they travel in straight line.
- They are deflected towards the positive plate in an electric field suggesting that they are negatively charged.
- They can make a light paddle wheel to rotate placed in their path. This means they possess kinetic energy.
- They have a charge to mass ratio = 1.75882×10^{11} C/kg.
- They produce X-rays when they strike a metallic target.
- The characterstics of cathode rays (electrons) do not depend on the material of electrodes and nature of the gas present in the cathode ray tube.

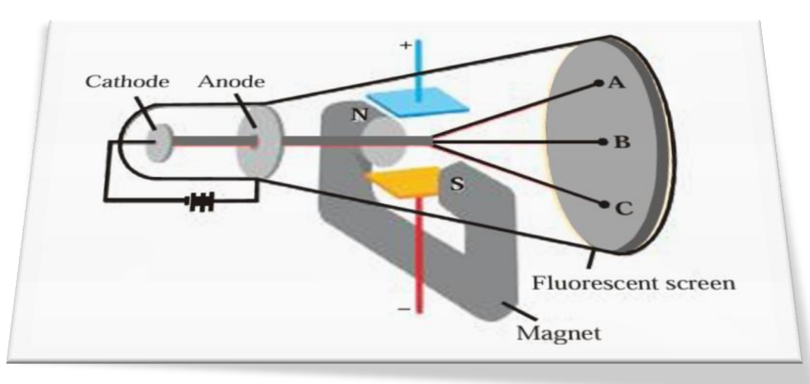
Charge to mass ratio of electron

In 1897 J.J. Thomson measured e/m ratio of electron by using cathode ray tube and applying electric and magnetic field perpendicular to each other as well as to the path of electrons. The extent of deviation of electrons from their path in the presence of electric and magnetic field depends on:

(a) Charge on the electron

(b) Mass of the particle

(c) The strength of electric or magnetic field

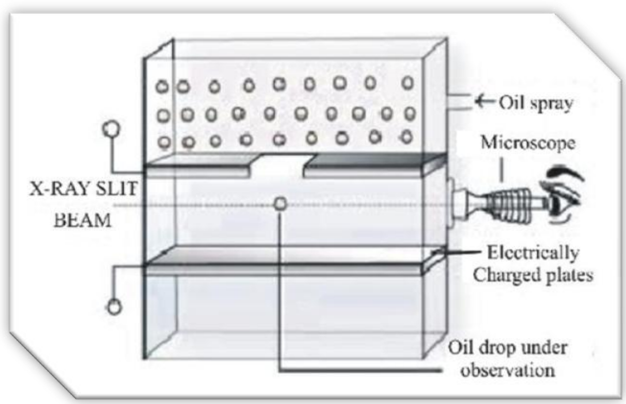


Oil droplets

R.A Millikan devised a method known as oil drop experiment to determine the charge on the electrons.

1. In this method, oil droplets in the form of mist, produced by the atomizer, were allowed to enter through a tiny hole in the upper plate of electrical condenser.
2. The downward motion of these droplets was viewed through the telescope, equipped with a micrometer eye piece.
3. By measuring the rate of fall of these droplets, Millikan was able to measure the mass of oil droplets.
4. Millikan concluded that the magnitude of electrical charge q, on the droplets is always an integral multiple of the electrical charge that is q=ne, where n = 1, 2…

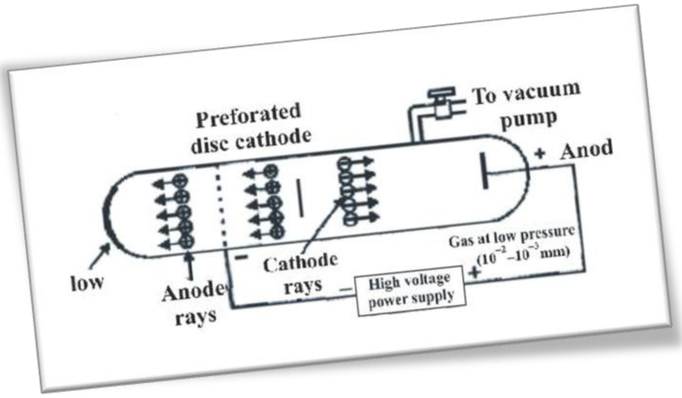
$$m = \frac{e}{e/m} = \frac{1.6022\times10^{-19}\,C}{1.758820\times10^{11}\,C/kg} = 9.1094\times10^{-31}\,kg.$$



Discovery of proton

During the experiments with cathode rays, the scientist Goldstein designed a special type of discharge tube. He discovered new rays called Canal rays.
The name canal rays are derived from the fact that the rays travelled in the opposite direction to cathode rays, they are also known as anode rays.

Properties



1. They travel in straight lines.
2. They carry a positive charge.
3. They are made up of material particles.
4. The charge to mass ratio (e/m) of the particles is also found to depend on the gas taken.
5. Their behavior in electric and magnetic field is opposite to that observed for electron.

Origin of anode rays

These anode rays are not emitted from the anode but are produced in the space between the anode and the cathode.

1. The lightest charged particles were obtained when the gas taken in the discharge tube was hydrogen.
2. They had minimum mass and unit positive charge.
3. The particle was called a proton.
4. Charge on a proton = $+ 1.6022 \times 10^{-19}C$

Discovery of neutron

The theoretical requirement for the existence of a neutron particle in the atomic nucleus was put forward by Rutherford in 1920.

1. It was proposed to be a particle with no charge and having mass almost equal to that of a proton.
2. He named it as neutron.
3. It had mass almost equal to that of a proton (1.674×10^{-27}kg) and carried no charge.

Name	Discoverer	Symbol	Charge	Relative Charge	Mass
Electron	J.J. Thomson	e	-1.6022×10^{-19} C	–1	9.1094×10^{-31} Kg
Proton	Goldstein	p	$+1.6022 \times 10^{-19}$C	+1	1.6726×10^{-27} kg
Neutron	Chadwick	n	0	0	1.6749×10^{-27} kg

THOMSON MODEL

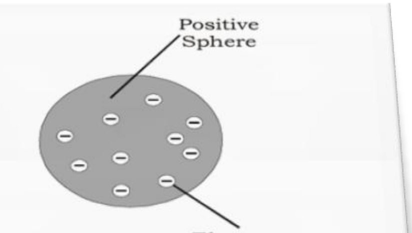
Sir J. J. Thomson, who discovered the electron, was the first to suggest a model of atomic structure.

(i) All atoms contain electrons.

(ii) The atom as a whole is neutral. The total positive charge and total negative charge must be equal.

(iii) Electrons were studded in the atom like plums in a pudding.

(iv) This model of the atom was often called the plum – pudding model, also the raisin pudding model or watermelon model.



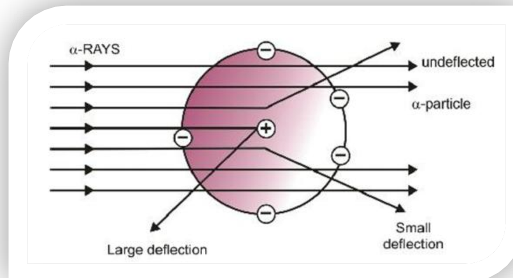
Thomson's proposed model of atom.

RUTHERFORD'S-SCATTERING EXPERIMENT

Alpha particles scattering experiments in 1909, in this experiment, a very thin foil of gold (0.004nm) is bombarded by a fine stream of alpha particles. A fluorescent screen (ZnS) is placed behind the gold foil, where points were recorded which were emerging from alpha-particles. Polonium was used as the source of Alpha-particles.

Observations

1. Most of the Alpha particles (99%) passes through it, without any deviation or deflection.
2. Some of the Alpha particles were deflected through small angles.
3. Very few Alpha particles were deflected by large angles and occasionally an Alpha particle got deflected by 180
4. Positively charged body must be at the centre of the atom which is called nucleus.



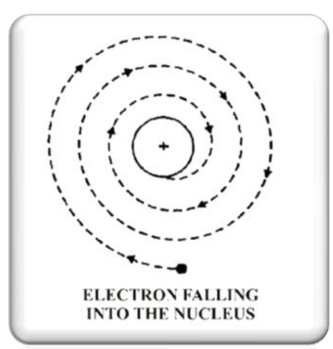
Rutherford's nuclear atomic model

The electrons are revolving around the nucleus in circular paths called orbits. Thus, an atom resembles the solar system in which the sun plays the role of nucleus and the planets that of revolving electrons and the model is known as planetary model.

Forces of attraction operating on the electron are exactly balanced by centrifugal forces.

Drawbacks

According to classical mechanics, any charged body in motion under the influence of attractive forces should radiate energy continuously. So, the electron will follow a spiral path and finally fall into the nucleus and the structure would collapse. This behavior is never observed.



ATOMIC NUMBER AND MASS NUMBER

Mass number (A) = Number of protons (Z) + Number of neutrons (n)

Isotopes are the atoms of the same element having identical atomic number but different mass number.

Isotope	Formula	Mass number	No. of protons	No. of neutrons
Protium	$^{1}_{1}H$ (H)	1	1	0
Deuterium	$^{2}_{1}H$ (D)	2	1	1
Tritium	$^{3}_{1}H$ (T)	3	1	2

Isotope number	Mass number	No. of protons	No. of neutrons
$^{16}_{8}O$	16	8	8
$^{17}_{8}O$	17	8	9
$^{18}_{8}O$	18	8	10

Isobars: Atoms of different elements having different atomic numbers but same mass numbers are called isobars.

Isobar	Atomic number	Mass number	No. of elctrons	No. of protons	No. of neutrons
$^{40}_{18}Ar$	18	40	18	18	22
$^{40}_{19}K$	19	40	19	19	21
$^{40}_{20}Ca$	20	40	20	20	20

Isotones: Atoms of different elements which contain the same number of neutrons are called isotones.

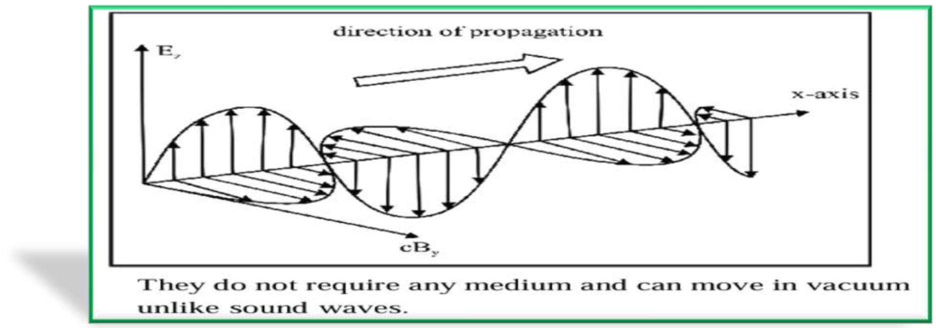
Isotones	Atomic	Mass number	No. of neutros
$^{36}_{16}S$	16	36	20
$^{37}_{17}Cl$	17	37	20
$^{38}_{18}Ar$	18	38	20
$^{39}_{19}K$	19	39	20
$^{40}_{20}Ca$	20	40	20

Isoelectronic:

The species (atoms or ions) containing the same numbers of electrons are called isoelectronic.

$O2^{-}$, F^{-}, Na^{+}, Mg^{+2}, Al^{+3}, Ne etc

Electromagnetic Radiations are waves which are formed as a result of oscillating magnetic and electric fields which are perpendicular to each other and both are perpendicular to direction of motion.



Electric fields propagate in x-y direction & Magnetic fields propagate in x-z direction.

Amplitude: It is height of the crest or trough (depth) of a wave.

Units: metre (m)

Frequency: The number of waves passing through a point in one second.

Units: Hertz (Hz) or s-1

Velocity: The distance travelled by a wave in one second is called velocity.

Units m/s I

Wavelength: The distance between two adjacent crests or troughs is called wavelength.

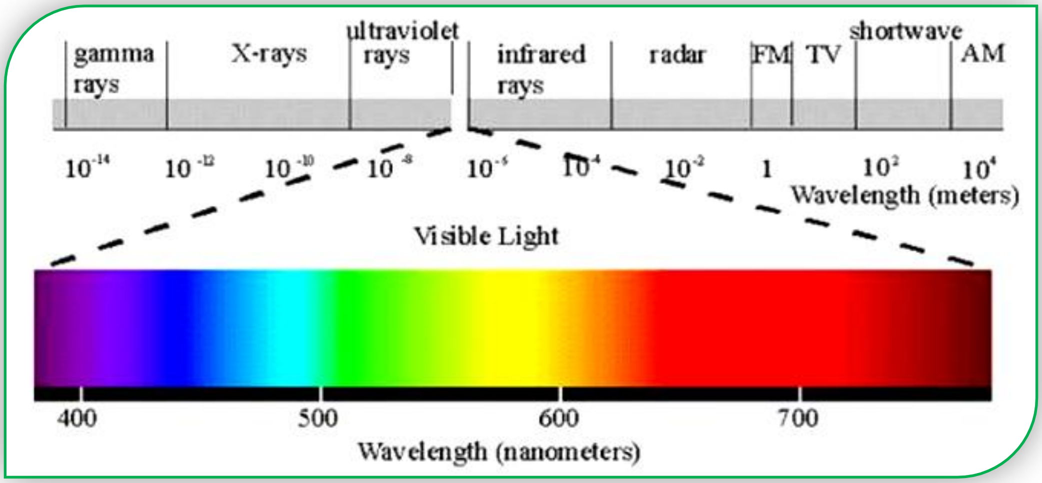
Units: Angstrom (Å) [1Å=10^{-10}m]

Wave Number: It is the number of wavelengths per centimetre of length.

Units: m-1

ELECTROMAGNETIC SPECTRUM

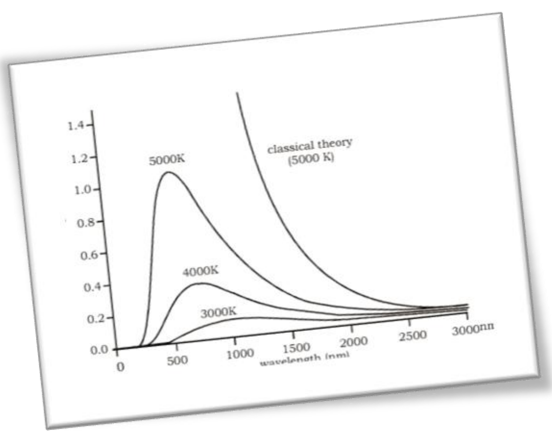
When all the electromagnetic radiations are arranged in increasing order of wavelength or decreasing frequency the band of radiations obtained is termed as electromagnetic spectrum.



Gamma Rays < X-rays < Ultra-violet rays < Visible< Infrared < Micro-waves <Radio waves

Black Body Radiation:

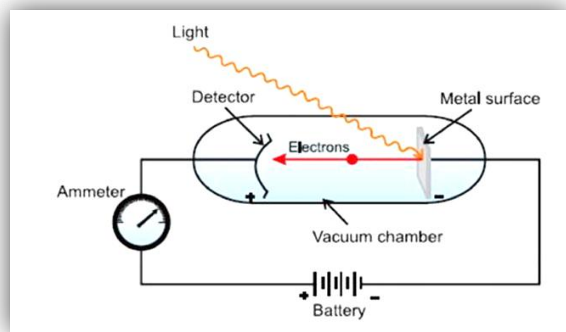
An ideal body, which emits and absorbs radiations of all frequencies, is called black body and radiation emitted by a black body is called black body radiation.



That is at higher temperatures, though the intensity rises as predicted by Maxwell's theory but the wavelength decreases.

If T1 >T2 >T3 then 1 < 2 < 3.

Photoelectric Effect when radiations with certain minimum frequency (V_0) strike the surface of a metal, the electrons are ejected from the surface of the metal. This phenomenon is called photoelectric effect. The electrons emitted are called photoelectrons.



For each metal, there is a characteristic minimum frequency, V_0 (also known as threshold frequency) below which photoelectric effect is not observed.

PLANCK'S QUANTUM THEORY

The energy is emitted or absorbed not continuously but discontinuously in the form of small discrete packets of energy. Each such packet of energy is called a 'quantum'. In case of light this quantum of energy is called a photon.

$$E \propto \nu \text{ or } E = h\nu = \frac{hc}{\lambda}$$

$$h = \text{Planck's constant} = 6.626 \times 10^{-34}\,\text{Js}$$

SPECTRUM

A spectrum is a group or band of wavelengths/colours and the study of emission or absorption spectra is known as spectroscopy.

Types of spectrum

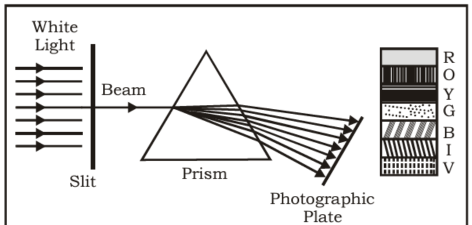
1) Emission Spectrum

2) Absorption Spectrum

Emission Spectrum

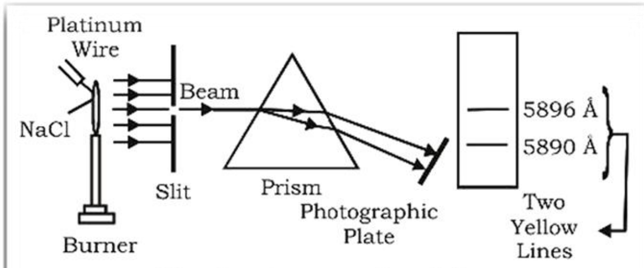
When radiations emitted from a source are incident on a prism and are separated into different wavelengths and obtained on a photographic plate.

(a) **Continuous Emission Spectra:** There are no gaps between various wavelengths, one wavelength merges into another.



Continuous spectrum of white light

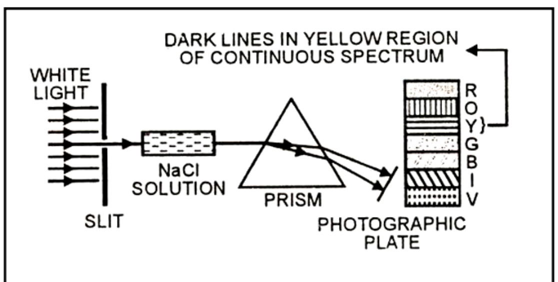
(b) **Discontinuous Emission Spectra:** It is also known as Line Spectra or atomic spectra. In this, certain wavelengths go missing from a group and that leaves dark spaces in between giving discontinuity to the spectrum. It is also known as fingerprint of an element.



Line Spectrum produced from a volatile salt placed in a flame.

Absorption Spectra

When light from any source is first passed through the solution of a chemical substance and then analysed, it is observed that there are <u>some dark lines</u> in the otherwise continuous spectra.



Production of absorption spectrum

BOHR'S MODEL

1. An atom consists of a small, heavy, positively charged nucleus in the centre and the electrons revolve around it in circular orbits.
2. Electrons revolve only in those orbits which have a fixed value of energy. Hence, these orbits are called energy levels or stationary states.

They are numbered as 1, 2, 3...... These numbers are known as Principal quantum Numbers.

- ✓ Energy of an electron is given by: $E_n = -R_H (Z^2/n^2)$ n = 1, 2, 3...
- ✓ where R_H is Rydberg's constant and its value is 2.18×10^{-18} J
- ✓ K < L < M < N...... and so on.
- ✓ Energy of the lowest state (n=1) is called ground state.

Radii of the stationary states:

$$r_n = \frac{52.9n^2}{Z} \text{ pm}$$

For H-atom (Z = 1), the radius of first stationary state is called Bohr orbit.

Velocities of the electron in different orbits:

$$v_n = \frac{2.188 \times 10^6 Z}{n} \text{ m/s}$$

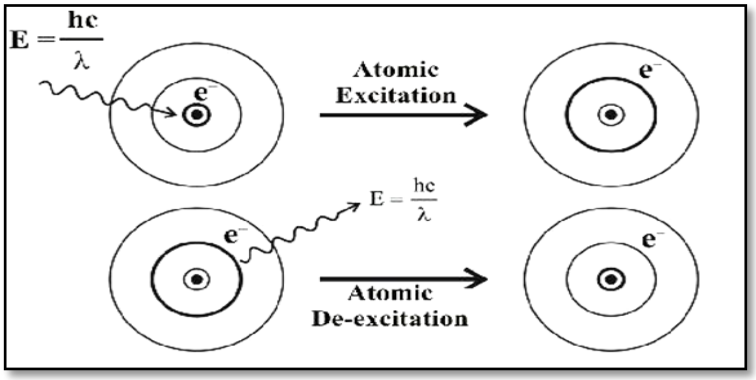
3. Since the electrons revolve only in those orbits which have fixed values of energy.
4. The angular momentum of an electron in an atom can have certain definite values and not any value of their own.

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

Where n=1,2,3...... and so on.

5. An electron does not lose or gain energy when it is present in the same shell.

6. When an electron gains energy, it gets excited to higher energy levels and when it de-excites, it loses energy in the form of electromagnetic radiations and comes to lower energy values.



Transition of Electron

When energy is absorbed or emitted when electron excites or de-excites respectively. The energy gap between the two orbits is

$$\Delta E = E_f - E_i$$

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

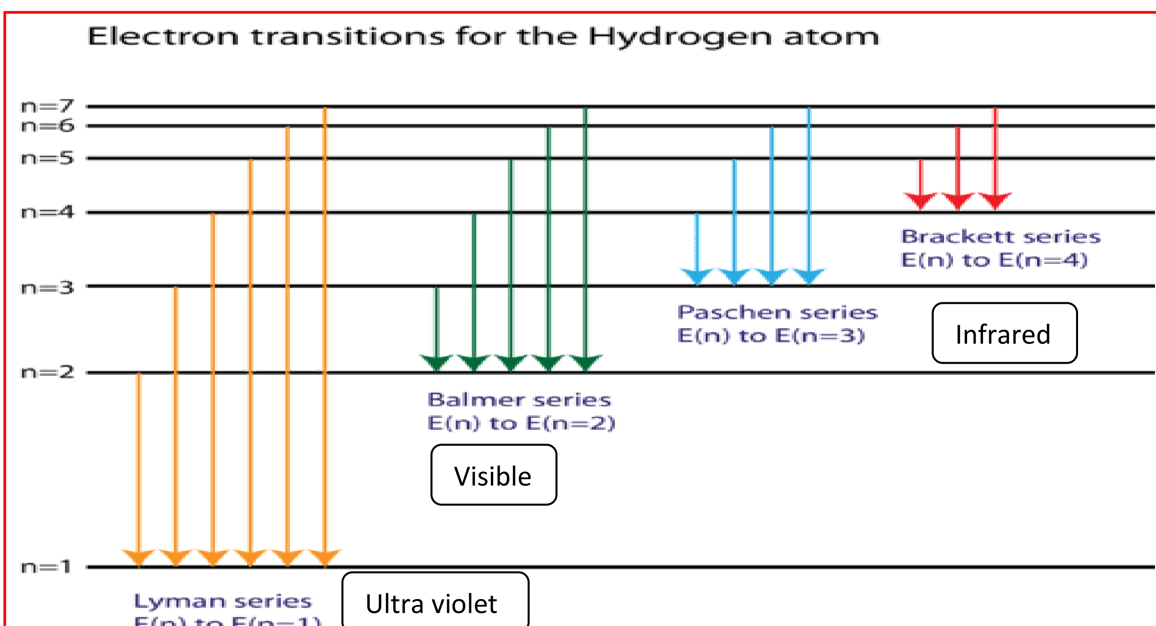
$$\Delta E = R_H\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) = 2.18\times10^{-18}\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) \text{J/atom}$$

The wavelength associated with the absorption or emission of the photon is:

$$\frac{1}{\lambda} = \frac{\Delta E}{hc} = \frac{R_H}{hc}\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) = 1.09677\times10^7\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) m^{-1}$$

This is known as Rydberg's formula.

Line Spectrum of Hydrogen



Limitations of Bohr's Model

1. It fails to account for the finer details (doublet-two closely spaced lines) of the hydrogen spectra.
2. Inability to explain splitting of lines in the magnetic field (Zeeman Effect) and in the electric field (Stark Effect)
3. It could not explain the ability of atoms to form molecules by covalent bonds.

Splitting of spectral lines in magnetic field is known as Zeeman Effect

Splitting of spectral lines in electric field is known as Stark Effect.

Dual behavior of matter

De Broglie proposed that matter should also exhibit dual behavior, Example both particle and wave like properties.

$$\boxed{\lambda = \frac{h}{mv} = \frac{h}{p} = \frac{h}{\sqrt{2m(KE)}} = \frac{h}{\sqrt{2mqV}}}$$

Heisenberg's Uncertainty Principle

It is impossible to measure simultaneously the position and momentum of a small particle with absolute accuracy.

The product of the uncertainty in the position (Δx) and the uncertainty in momentum (Δp) is always a constant and is equal to or greater than $h/4\pi$.

$$\boxed{\begin{aligned} &(\Delta x).\,(\Delta p) \geq h/4\pi \\ &\text{Or } (\Delta x).\,(m\Delta v) \geq h/4\pi \\ &\text{Or } (\Delta x).\,(\Delta x) \geq h/4\pi m \end{aligned}}$$