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ZOOM MEETING /VIDEO LECTURER

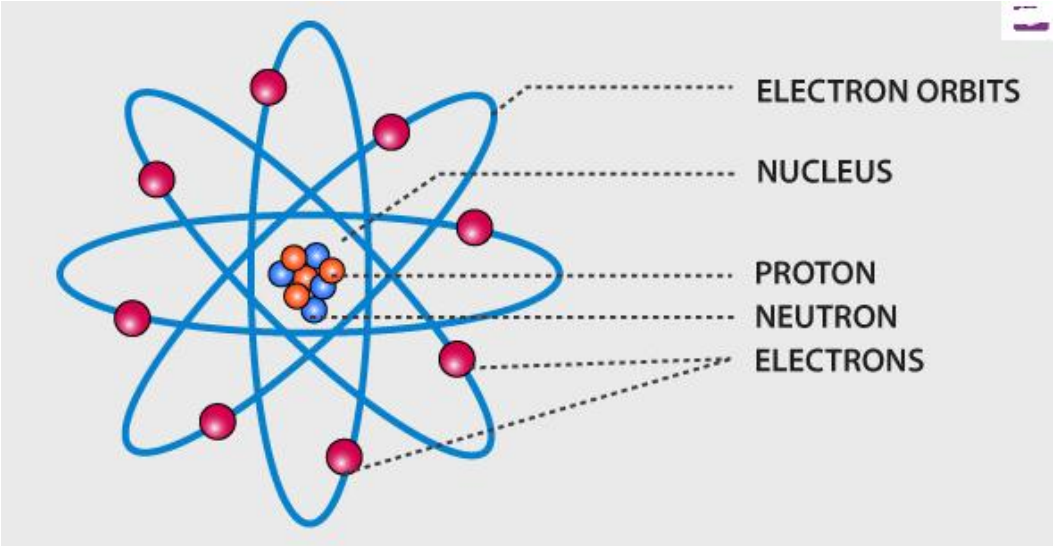
PHYSICS DEPARTMENT

Choice Based Credit System In F. Y. B. Sc. Sem I

PHY-112: Physics Principles and Applications-I Credits-02

ATOMS AND ATOMIC STRUCTURE

- **Atoms** are the basic building blocks of ordinary matter. Atoms can join together to form molecules, which in turn form most of the objects around us
- Atoms are composed of particles like **protons**, **electrons** and **neutrons**.
- **Protons** carry a positive electrical charge, **electrons** carry a negative electrical charge and **neutrons** carry no electrical charge at all.
- The protons and neutrons cluster together in the central part of the atom, called the **nucleus**, and the electrons **'orbit'** the nucleus.
- A particular atom will have the same number of protons and electrons and most atoms have at least as many neutrons as protons.



ATOMICS MODELS

- In the **18th and 19th** centuries, many scientists attempted to explain the structure of the **atom** with the help of **atomic models.**
- Each of these models had their own **merits and demerits** and which turn to the development of the **modern atomic model**.
- The most notable contributions to this field by the scientists **John Dalton, J.J. Thomson, Ernest Rutherford and Niels Bohr.** Their ideas on the structure of the atom are discussed .

Dalton's Atomic Theory

- He is English chemist **John Dalton** suggested that all matter is made up of atoms, which were indivisible and indestructible.
- He also stated that all the atoms of an element were exactly the same, but the atoms of different elements differ in size and mass.
- According to his postulates proposed the atomic structure comprised atoms, the smallest particle responsible for the chemical reactions to occur.
- **The following are the postulates of his theory:** Every matter is made up of atoms. Atoms are indivisible. Specific elements have only one type of atom. Each atom has its own constant mass that varies from element to element. Atoms undergo rearrangement during a chemical reaction. Atoms can neither be created nor be destroyed but can be transformed from one form to another.
- Dalton's atomic theory successfully explained the Laws of chemical reactions, namely, the Law of conservation of mass, Law of constant properties.
- Demerits of Dalton's Atomic Theory The theory was unable to explain the existence of isotopes. Nothing about the structure of atom was appropriately explained. Later, the scientists discovered particles inside the atom that proved, the atoms are divisible.

Thomson model

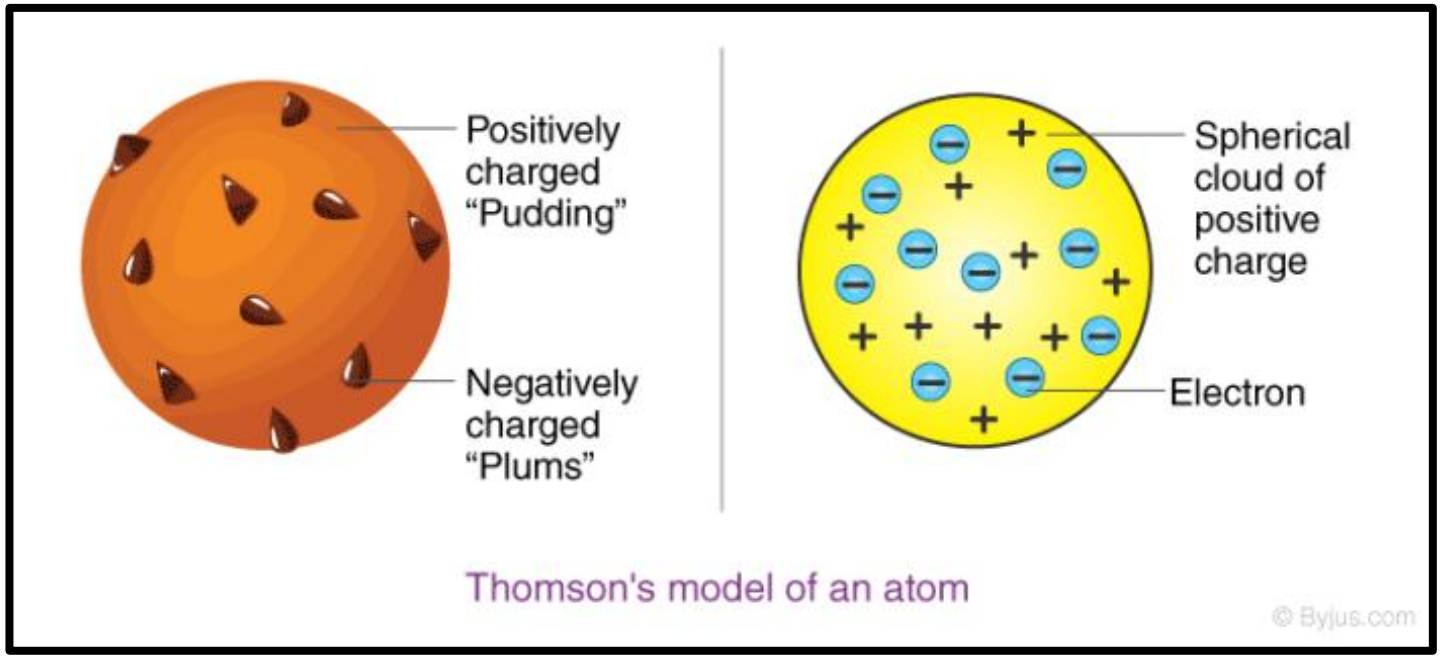
- was proposed by William Thomson in the year 1900. This model explained the description of an inner structure of the atom theoretically. It was strongly supported by Sir Joseph Thomson, who had. Thomson. This discovered the electron earlier.
- During cathode ray tube experiment, a negatively charged particle was discovered by J.J. That experiment took place in the year 1897. Cathode ray tube is a vacuum tube. The negative particle was called an electron.
- In this atomic structure model, he considered atoms surrounded by a cloud having positive as well as negative charges.
- Thomson's model of an atom is similar to a plum pudding.

Postulates of Thomson's atomic model

- **Postulate 1**: An atom consists of a positively charged sphere with electrons embedded(fixed firmly and deeply in a surrounding mass in it.)
- **Postulate 2**: An atom as a whole is electrically neutral because the negative and positive charges are equal in magnitude.
- Thomson atomic model is compared to watermelon.
- Watermelon seeds as negatively charged particles.
- The red part of the watermelon as positively charged.

Limitations of Thomson's atomic model

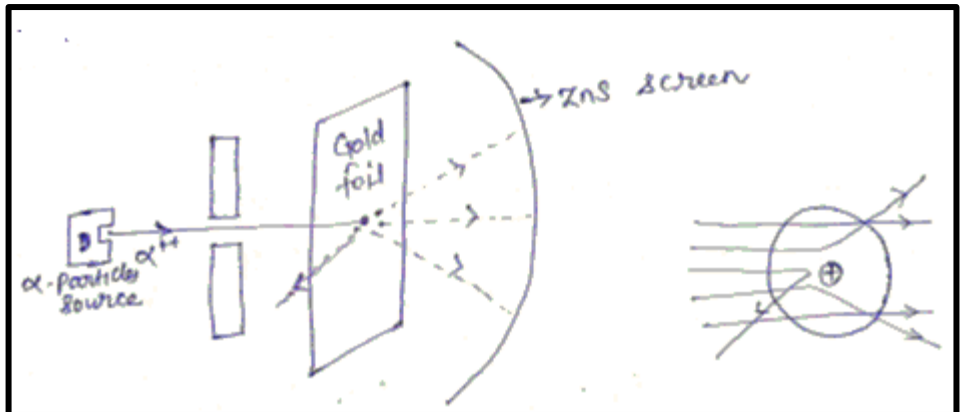
- It failed to explain the stability of an atom because he failed to explain how a positive charge holds the negatively charged electrons in an atom.
- His theory also failed to account for the position of the nucleus in an atom.
- Thomson's model failed to explain the scattering of alpha particles by thin metal foils No experimental evidence in its support
- Although Thomson's model was not an accurate model to account for the atomic structure, it proved to be the base for the development of other atomic models.



Rutherford Atomic Model

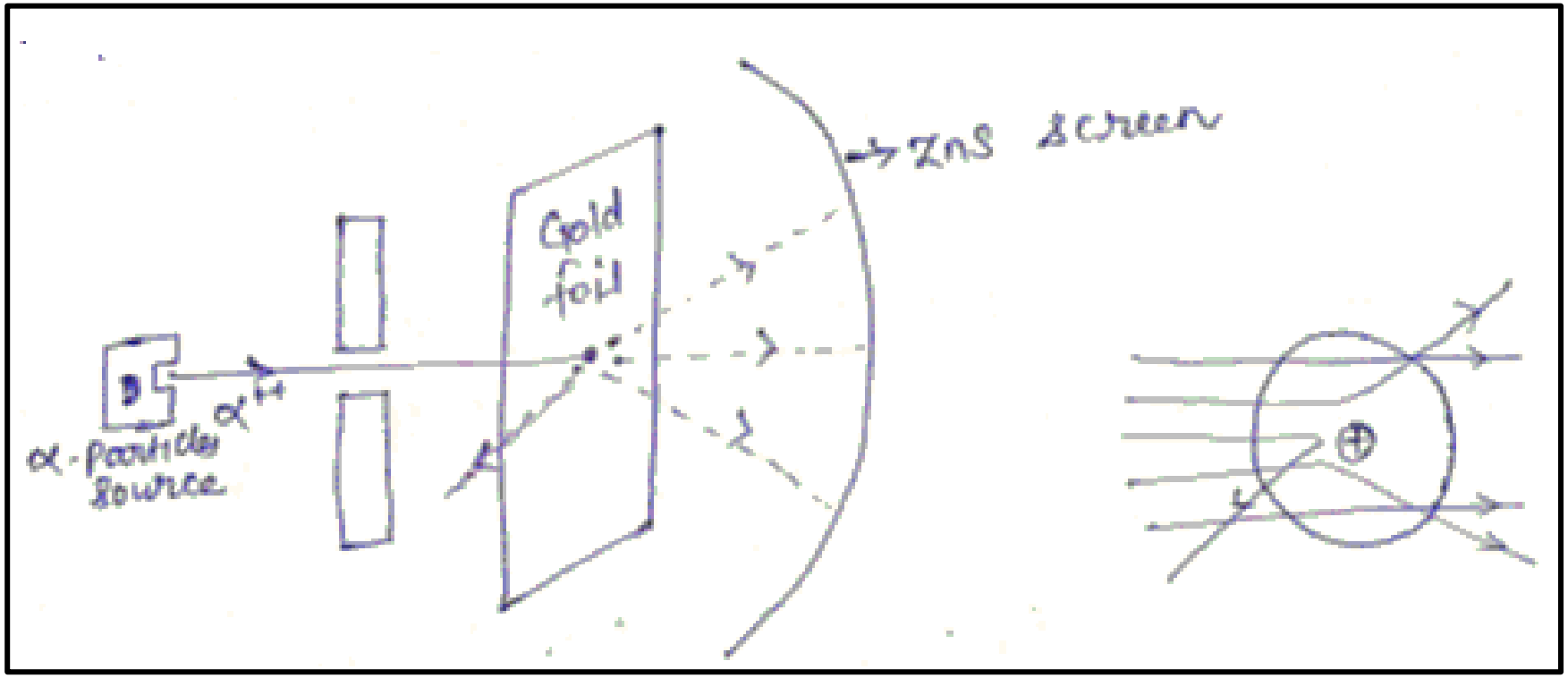
- J. J. Thomson failed to explain the experimental results which is associated with the atomic structure of elements.
- Ernest Rutherford, a British scientist conducted an experiment and based on the observations of this experiment.
- Rutherford's conducted an experiment by bombarding α-particles on a thin sheet of gold with and studied the trajectory of these particles after their interaction with the gold foil.
- Rutherford, in his experiment, directed high energy streams of α-particles from a radioactive source at a thin sheet (100 nm thickness) of gold.
- He study the deflection caused to the α-particles, he placed a fluorescent zinc sulphide screen around the thin gold foil.
- Rutherford made certain observations that contradicted Thomson's atomic model.

Observations of Rutherford's Alpha Scattering Experiment



Observations of Rutherford's Alpha Scattering Experiment.

- Most of the rays passed through undeviated which proves that the **most of the space in an atom is empty**
- Some particle deflated by small angle which proves that there were obstruction in their path.
- Scattered with large deviation shows that. Centre atoms is +charged .
- Some rays deflected back without any deviation which proves that whole mass of atom is concentered in the center.



Rutherford Atomic Model

- Rutherford Atomic Model Based on the above observations and conclusions,
- **According to the Rutherford atomic model:**
- The positively charged particles and most of the mass of an atom was concentrated in an extremely small volume. He called this region of the atom as a **nucleus**.
- Rutherford model proposed that the negatively charged electrons surround the nucleus of an atom.
- He also claimed that the electrons surrounding the nucleus revolve around it with very high speed in circular paths. He named these circular paths as orbits.
- Electrons being negatively charged and nucleus being a densely concentrated mass of positively charged particles are held together by a strong electrostatic force of attraction.

Limitations of Rutherford Atomic Model

- Although the Rutherford atomic model was based on experimental observations it failed to explain certain things.
- Rutherford proposed that the electrons revolve around the nucleus in fixed paths called orbits.
- According to Maxwell, accelerated charged particles (e^-)emit electromagnetic radiations and hence an electron revolving around the nucleus should emit electromagnetic radiation.
- This radiation would carry energy from the motion of the electron which would come at the cost of shrinking of orbits.
- Ultimately the electrons would collapse in the nucleus. Calculations have shown that as per the Rutherford model, an electron would collapse in the nucleus in less than 10-8 seconds.
- So Rutherford model was **not in accordance with Maxwell's theory and could not explain the stability of an atom**.
- One of the drawbacks of the Rutherford model was also that he **did not say anything about the arrangement of electrons in an atom** which made his theory incomplete.
- Although the early atomic models were inaccurate and failed to explain certain experimental results, **they were the base for future developments in the world of quantum mechanics**.

Bohr's Atomic Model

- Bohr model of the atom was proposed by Neil Bohr in 1915. It came into existence with the modification of Rutherford's model of an atom.
- Rutherford's model introduced the nuclear model of an atom, in which he explained that a nucleus (positively charged) is surrounded by negatively charged electrons.
- Bohr modified this atomic structure model by explaining that electrons move in fixed orbitals (shells) and not anywhere in between and he also explained that each orbit (shell) has a fixed energy level.
- Rutherford basically explained the nucleus of an atom and Bohr modified that model into electrons and their energy levels.

Bohr's Postulates of an Atom

- In an atom, electrons (negatively charged) revolve around the positively charged nucleus in a definite circular path called orbits or shells.
- The energy levels are represented by an integer (n=1, 2, 3...) known as the quantum number. This range of quantum number starts from nucleus side with n=1 having the lowest energy level. The orbits n=1, 2, 3, 4... are assigned as K, L, M, N.... shells and when an electron attains the lowest energy level, it is said to be in the ground state.
- The electrons in an atom move from a lower energy level to a higher energy level by gaining the required energy and an electron moves from a higher energy level to lower energy level by losing energy.

Limitations of Bohr's Model of an Atom

Bohr's model of an atom failed to explain the Zeeman Effect (effect of magnetic field on the spectra of atoms).

It also failed to explain the Stark effect (effect of electric field on the spectra of atoms).

It violates the Heisenberg Uncertainty Principle.

It could not explain the spectra obtained from larger atoms.

It is valid only for one electron atoms, eg: H, He^{+},Li^{+2},Na^{+1} etc.

Orbits were taken as circular but according to Somerfield these are elliptical.

Intensity of spectral lies could not be explained. Nucleus was taken as stationary but it also rotates on its own axis.

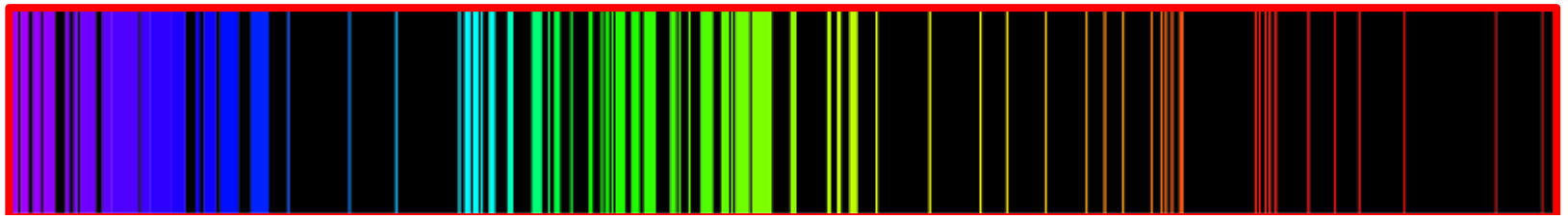
It could not be explained the minute structure in spectrum line.

This does not explain the Zeeman effect (splitting up of spectral lines in magnetic filed) and Stark effect (splitting up in electric filed)

This does not explain the doublets in the spectrum of some of the atoms.

Atomic spectra

- *Atomic spectra is the study of atoms (and atomic ions) through their interaction with electromagnetic radiation.* (EMR)
- We all know about the refraction of light. When light travels from one medium to another, it either bends towards the normal or away from the normal.
- The phenomenon of refraction is mainly attributed to the difference in the speed of light in various mediums.
- The speed of light depends upon the nature of the medium through which it passes
- phenomenon of dispersion of white light through a prism. Also, learn about the *emission spectrum and absorption spectrum.*
- Spectrum is obtained by interaction of matter with electro magnetic radiation(EMR).When an atom or molecule subjected to EMR it absorbs certain frequency of radiation which corresponds to it's energy, this absorbed energy is plotted as absorbtion spectrum. when this absorbed energy is emitted by that corresponding atom or molecule to attain stable energy



Emission spectra of Iron



spectra



spectra

- The **emission spectrum** of a gas is represented by a collection of separate colored lines, with dark spaces between them.
- These lines are parts of the spectrum where emission occurs and photons are emitted, while the dark spaces are the parts where there is no emission, hence the darkness.
- The difference in colors is due to the variation of the energy levels of the electrons.

Emission spectra

- Produced when atoms release energy
- Emission Spectrum means that only **certain frequency of light** are emitted each frequencies are related to energy of formula E=hv
- This concludes that only photons with specific energies are emitted by atoms.
- Whenever electromagnetic radiation interacts with atoms and molecules of matter, the electrons in these atoms may absorb energy and jump to a higher energy state, losing their stability.
- Each element produces a unique set of spectral lines
- electromagnetic spectrum emit radiation in various regions of the. This spectrum of radiation emitted by electrons in the excited atoms or molecules is known as an *emission spectrum.*

- What is Emission Spectrum. Produced when atoms release energy
- When energy is absorbed by electrons in an atom, electrons move from lower energy levels to higher energy levels.
- These excited electrons have to radiate energy when they return to ground states from the excited state.
- There are 2 types of emission spectra: 1) Line spectrum– When the spectrum appears as a series of lines, which are separated by black spaces, it is called a line spectrum. which contains well separated coloured lines in the spectrum and also called as fingerprint of atom.
- continuous spectrum. When the spectrum consists of a wide range of colors in a particular wavelength range or interval, it is called continuous spectrum.
- Emission spectra can emit all the colours in an electromagnetic spectrum.

Absorption spectrum

- On the other hand, absorption spectrum is noticed on absorption by transitioning electrons in an element. As a result, a set of black lines will appear on a colorful continuum background.
- When it comes to transitions, they are exactly between the same electronic levels for an element regardless of the spectrum being absorption or emission.
- The only difference is that the transitions are from higher to lower levels in an emission spectrum and the transitions are from lower to higher levels in an absorption spectrum.
- **Absorption spectrum** of an element is represented by a continuous band of colors with separate dark lines between them. The entire band represents the total light that is focused on the element.
- The dark lines are the parts of the spectrum where the electrons absorb light photons, hence, there is absence of light at these parts.
- The remaining colored parts of the spectrum represent the parts of the incident light that has not been absorbed, and hence, appear as wavelength-specific colors.
- The reason for this pattern is that, all electrons of an atom are at different energy levels at any given time. The energy difference between two energy levels of each electron is different.
- Absorption spectra can be measured in terms of their frequency, wavelength, or wave number.
- There are two types of absorption spectra: atomic absorption spectrum and molecular absorption spectrum.

Common Applications of Emission and Absorption Spectroscopy

Emission Spectrum

- **To identify a substance:** Every substance emits lights of different wavelengths. To identify the given substance, light is focused on it or the substance when it is heated.
- This causes the electrons to get excited and jump to a higher orbit. The energy emitted by these electrons while it returning to their ground states.
- It will compared to the characteristic colors of the elements, and the chemical composition of the determine.
- **To study the composition of stars**: The emission spectra of stars can be recorded and then compared with standard emission spectra of known elements to determine their chemical composition.

Absorption Spectrum

- **To identify a substance and determine its concentration:** An unknown substance can be identified by focusing light of a particular wavelength on it, and then studying the absorption spectrum of the substance.
- Since substances absorb light only from a particular wavelength or wavelength range, the wavelength of light focused on them is important.
- This spectrum can be compared with a set of reference values for identification. These reference values are known absorbance values of common elements and compounds.
- The concentration of the substance in the sample can also be determined.

To study the composition of stars:

- The light emitted by stars and planets passes through their atmosphere, where some of it is absorbed by the gases. When the absorption spectra of these gases is recorded and compared to the reference spectra values of gases, the composition of these planets or stars can be determined.

- **Remote sensing:** Absorption spectroscopy can be used to collect details of the land, including attributes such as forest cover, health of forests or exposed rock surfaces, without any individual actually setting foot on it.
- When light is focused on the land terrain and its absorption spectra is recorded, it can be used to extract information about the terrain.
- This is done by comparing the recorded values with reference values of absorbance shown by land with forest cover or exposed rock.
- In fact, the absorbance values vary and depending on the type of the forest, a healthy vegetation will show different values compared to an unhealthy forest cover. It can also provide details of atmospheric composition.
- Both absorption and emission spectroscopy are exact opposites of each other. Since the electronic configurations of elements are different, the spectrum values of these elements will be their 'atomic fingerprint', i.e., it will be unique to each element. It is said that absorption spectrum is the 'photographic negative' of emission spectrum, because the wavelengths that are missing in absorption spectrum are seen in the emission spectrum.

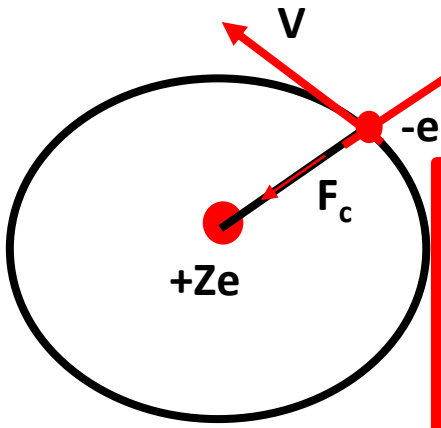
CLASSICAL PLANETARY MODEL OF HYDROGEN ATOM

- Historically, Bohr's model of the hydrogen atom is the very first model of atomic structure that correctly explained the radiation spectra of atomic hydrogen.
- The model has a special place in the history of physics because it introduced an early quantum theory, which brought about new developments in scientific thought and later in the development of quantum mechanics.

- . Niels **Bohr** introduced the atomic **Hydrogen model** in 1913. He described it as a positively charged nucleus, comprised of protons and neutrons, surrounded by a negatively charged electron cloud. In the **model**, electrons orbit the nucleus in atomic shells.
- The movement of electrons between these energy levels produces a spectrum. The Balmer equation is used to describe the four different wavelengths of Hydrogen which are present in the visible light spectrum. These wavelengths are at 656, 486, 434, and 410nm. These correspond to the emission of photons as an electron in an excited state transitions down to energy level n=2. The Rydberg formula, below, generalizes the Balmer series for all energy level transitions. To get the Balmer lines, the Rydberg formula is used with an n_f of 2.

Bohr's Postulates or Bohr's Model of the Hydrogen

- **1)** An atom consistence of a small central part called nucleus in which whole mass of +change of an atom concentrated. The – charged e^- revolve round the nucleus in he closed circular orbit. The centripetal force required for rotation is provided by electrostatic force of attraction between +charge nucleus and – charge e^-
- M=mass of e^-, Z=atomic number
- e=charge on e^- v=orbital velocity
- r=radius of the orbit
- Ze=charge of nucleus
- ε_0 =permittivity of free space =8.885*10^-12f/m



$$F_e = \frac{1}{4\pi\varepsilon_0}\frac{Ze^2}{r^2} ..centripetal...force..F_c = \frac{mv^2}{r}$$

$$\frac{mv^2}{r} = \frac{1}{4\pi\varepsilon_0}\frac{Ze^2}{r^2}$$

Bohr's Second Postulate

- **2)** electrons (e^-) revolve round the nucleus only those orbit whose angular moment is nh/2pi

$$Moment..of..electron... = mv..$$

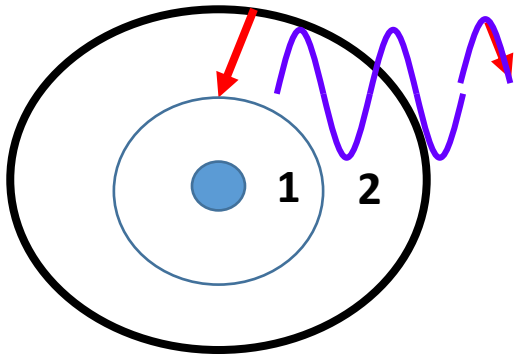
$$angularmomentum..ofelecton.. = mvr$$

$$mvr = \frac{nh}{2\pi}..n..is..positive..\text{int}eger...(n = 1,2,3..)$$

$$principal..quantum..number..$$

Bohr's Third Postulate

- **3)** Electron jump form higher energy orbit to lower energy orbit it emit the photons. Amount of energy emitted is different of energies of electron in the 2 orbit.
- E1=energy of e- in inner or lower orbit,E2=energy of e- in outer or higher orbit
- v= frequency of radiation emitted.



$$h\nu = E_2 - E_1$$

$E_1 =$

Derivation of radius and Energy

For H atom we know $Z = 1$

$$\therefore \frac{mv^2}{r} = \frac{1}{4\pi\varepsilon_0}\frac{e^2}{r^2} \quad ---(1)$$ according. .1. postulates

$\times$ by m and cancelling r we get.

$$m^2v^2 = \frac{1}{4\pi\varepsilon_0}\frac{me^2}{r} \quad ....(2)$$

according. .2. postulates

$$mvr = \frac{nh}{2\pi} \text{ squring both sides}$$

$$m^2v^2r^2 = \frac{n^2h^2}{4\pi^2} .....$$

$$m^2v^2 = \frac{n^2h^2}{4\pi^2r^2} ....(3)$$

equating (2) and (3)

$$\frac{n^2h^2}{4\pi^2r^2} = \frac{1}{4\pi\varepsilon_0}\frac{me^2}{r} .....(4)$$

$$\therefore r = \frac{\varepsilon_0 n^2h^2}{\pi\ me^2} .....(5)$$ using the value of r in (2)

$$m^2v^2 = \frac{1}{4\pi\varepsilon_0}\frac{me^2}{r} ....(2)$$

$$m^2v^2 = \frac{1}{4\pi\varepsilon_0}\frac{me^2}{(\frac{\varepsilon_0 n^2h^2}{\pi\ me^2})} \therefore \frac{1}{4\pi\varepsilon_0}\frac{me^2}{\varepsilon_0 n^2h^2}\pi\ me^2$$

$$m^2v^2 = \frac{1}{4\pi\varepsilon_0}\frac{me^2}{\varepsilon_0 n^2h^2}\pi\ me^2 Or v^2 = \frac{1}{4\varepsilon_0^2}\frac{e^4}{n^2h^2} Or v = \frac{1}{2\varepsilon_0}\frac{e^2}{nh} ...(6)$$

$$\therefore r_n = \frac{\varepsilon_0 n^2h^2}{\pi\ me^2} \text{ radius of } n^{th} \text{ orbit is } \propto \text{ to squre of int egral no}$$

i) $r_n \propto n^2$

ii) velocity of e^- $v = \frac{1}{2\varepsilon_0}\frac{e^2}{nh}$.

iii) **K.E. of n^{th} orbit is K.E** $= \frac{1}{2}.mv^2{}_n = \frac{me^4}{8\varepsilon_0^2 n^2h^2}$.

iv) **P.E** of an electron in n^{th} orbit is due to e^- is at dis tance r from nucleus

work is done moving e^- from dis tance r to ∞

$$\mathbf{P.E} = \int_0^\infty F.dr = \int_0^\infty \frac{1}{4\pi\varepsilon_0}\frac{e^2}{r^2}.dr \text{ Or } \mathbf{P.E} = -\frac{1}{4\pi\varepsilon_0}\frac{e^2}{r}.$$

sub value of r we get $\mathbf{P.E} = -\frac{1}{4\pi\varepsilon_0}\frac{e^2}{\frac{\varepsilon_0 n^2h^2}{\pi\ me^2}}$ $Or P.E = -\frac{1}{4\pi\varepsilon_0}\frac{e^2}{\underline{\varepsilon_0 n^2h^2}}\pi\ me^2$

$$\mathbf{P.E} = -\frac{me^4}{4\varepsilon_0^2 n^2h^2}$$

iv) Total energy of electron in n^{th} orbit $\mathbf{E_n = K.E. + P.E}$

$$E_n = \frac{me^4}{8\varepsilon_0^2 n^2h^2} - \frac{me^4}{4\varepsilon_0^2 n^2h^2} Or \mathbf{E_n} = -\frac{\mathbf{me^4}}{\mathbf{8\varepsilon_0^2 n^2h^2}} \therefore \mathbf{E_n} \propto \frac{1}{n^2}$$

GROUND STATE

$$\therefore Ground..state..is..lowest..orbit..of..H..atom.for..which...\Pr inciple..q.No.is1 = n$$

$$..r_n = \frac{\varepsilon_0 n^2 h^2}{\pi\ me^2}...n = 1...r_1 = \frac{\varepsilon_0 h^2}{\pi\ me^2} \therefore r_1 = 0.528..\overset{0}{A}$$

$$r_n = \frac{\varepsilon_0 n^2 h^2}{\pi\ me^2} = r_n = n^2 r_1$$

$$The..energy..ofan..e^{-}.in..ground..state..i.e.1^{st}\,bohr..orbit$$

$$\mathbf{E}_1 = -\frac{\mathbf{me}^4}{\mathbf{8\varepsilon_0}^2\mathbf{n}^2\mathbf{h}^2}..here..n = 1$$

$$\mathbf{E}_1 = -\frac{\mathbf{me}^4}{\mathbf{8\varepsilon_0}^2\mathbf{h}^2} = -13.6eV$$

$$.\mathbf{E}_n = -\frac{\mathbf{me}^4}{\mathbf{8\varepsilon_0}^2\mathbf{n}^2\mathbf{h}^2}....\mathbf{E}_n = \frac{E_1}{\mathbf{n}^2} \therefore E_2 = \frac{E_1}{\mathbf{n}^2} = \frac{-13.6}{4} = -3.4eV, E_3 = \frac{-13.6}{9} = -1.51eV$$

HYDROGEN SPECTRUM

- Electron jump form higher energy orbit to lower energy orbit it emit the photons. Amount of energy emitted is different of energies of electron in the 2 orbit.
- E_1 = energy of e- in inner or lower orbit, E_2 = energy of e- in outer or higher orbit
- v= frequency of radiation emitted.

$\therefore$

$Energy..E_P..is..e..energy..in..lower..P^{Th}orbit..$

$$\mathbf{E}_p = -\frac{\mathbf{me}^4}{\mathbf{8\varepsilon_0}^2 P^2 \mathbf{h}^2}.$$

$Energy..E_n..is..e..energy..in..higher..n^{th}orbit..$

$$\mathbf{E}_n = -\frac{\mathbf{me}^4}{\mathbf{8\varepsilon_0}^2 n^2 \mathbf{h}^2}$$

$when..e^- jump..from..an..higher..E..orbit..to..\mathbf{E}_n..lower...\mathbf{E}_P$

$energy..is..emitted..$

$$h\nu = \mathbf{E}_n - \mathbf{E}_p ... h\nu = -\frac{\mathbf{me}^4}{\mathbf{8\varepsilon_0}^2 n^2 \mathbf{h}^2} - \left(-\frac{\mathbf{me}^4}{\mathbf{8\varepsilon_0}^2 P^2 \mathbf{h}^2}\right)$$

$$h\nu = \frac{\mathbf{me}^4}{\mathbf{8\varepsilon_0}^2 \mathbf{h}^2}\left(\frac{1}{P^2} - \frac{1}{n^2}\right) \therefore \nu = \frac{\mathbf{me}^4}{\mathbf{8\varepsilon_0}^2 \mathbf{h}^3}\left(\frac{1}{P^2} - \frac{1}{n^2}\right)$$

$\therefore$

$but.c = \nu\lambda..where..\lambda..is..wavelength..and..c..isvelocity..of..light$

$$\therefore \nu = \frac{\mathbf{me}^4}{\mathbf{8\varepsilon_0}^2 \mathbf{h}^3}\left(\frac{1}{P^2} - \frac{1}{n^2}\right)$$

$$\frac{c}{\lambda} = \frac{\mathbf{me}^4}{\mathbf{8\varepsilon_0}^2 \mathbf{h}^3}\left(\frac{1}{P^2} - \frac{1}{n^2}\right) \therefore \frac{1}{\lambda} = \frac{\mathbf{me}^4}{\mathbf{8\varepsilon_0}^2 c\mathbf{h}^3}\left(\frac{1}{P^2} - \frac{1}{n^2}\right)..Or$$

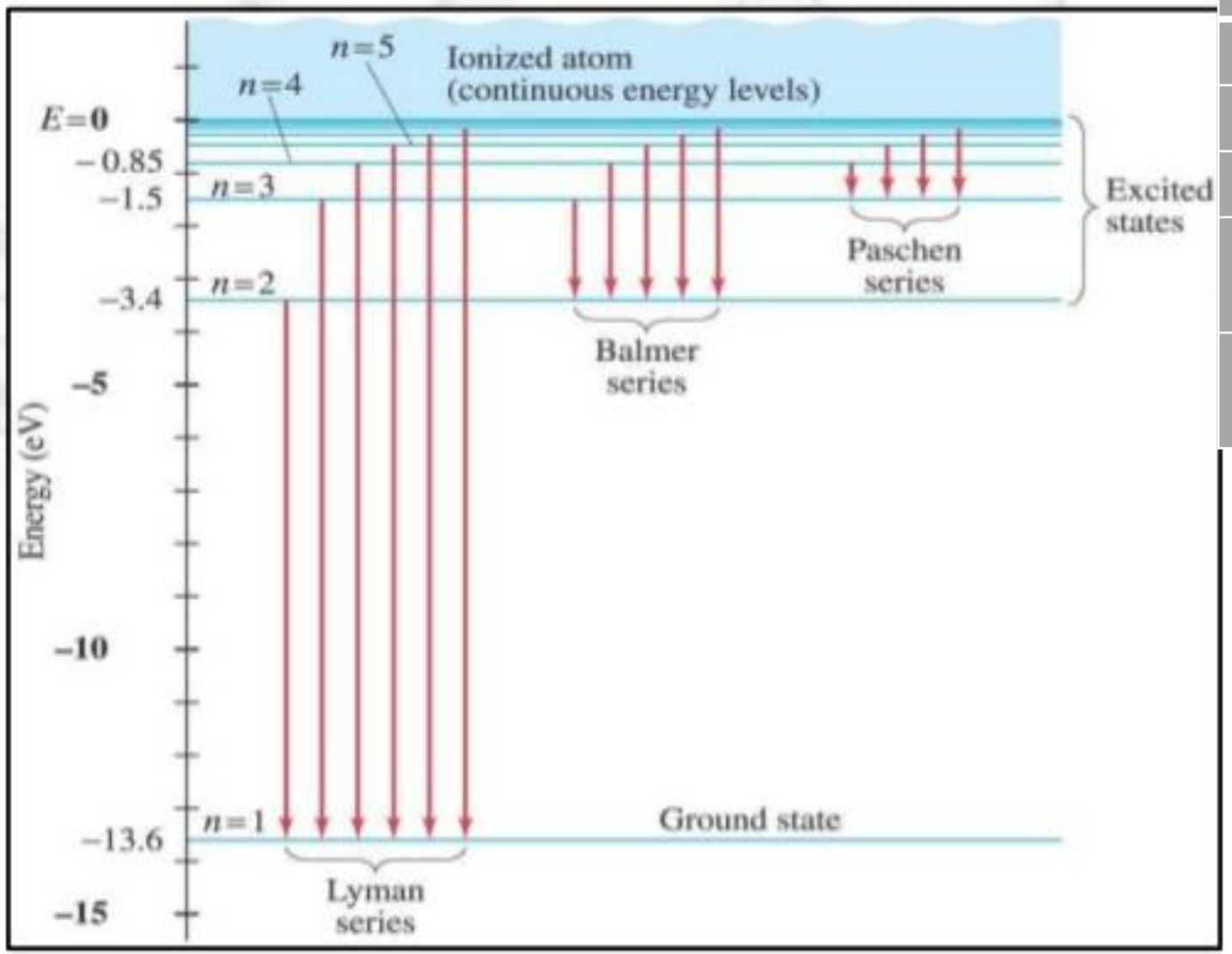
$$\frac{1}{\lambda} = R\left(\frac{1}{P^2} - \frac{1}{n^2}\right)..whwer.R = \frac{\mathbf{me}^4}{\mathbf{8\varepsilon_0}^2 c\mathbf{h}^3}..is..called..Rydberg's..cons\tan t$$

$value..1.097\times10^{-7}m^{-1}$

.

$$wave..number\ \bar{\vartheta} = \frac{1}{\lambda} = R\left(\frac{1}{P^2} - \frac{1}{n^2}\right)$$

LYMAN SERIES, BALMER SERIES & PASCHEN SERIES



SP. Series	Final state	Intial state	SP.region
Lyman	1	2,3,4....	U.V
Blamer	2	3,4,5....	Visible
Paschen	3	4,5,6....	Near I. R
Brackett	4	5,6,7....	Infrared
Pfund	5	7,8,9....	Far. I. R

- In any series Sp.lines.The shortest wavelength corres..to longest transition and longest wavelength corres..to the shortest transition
- balmer.series n=3 to p=2 H alfa

Strengths of Bohr model

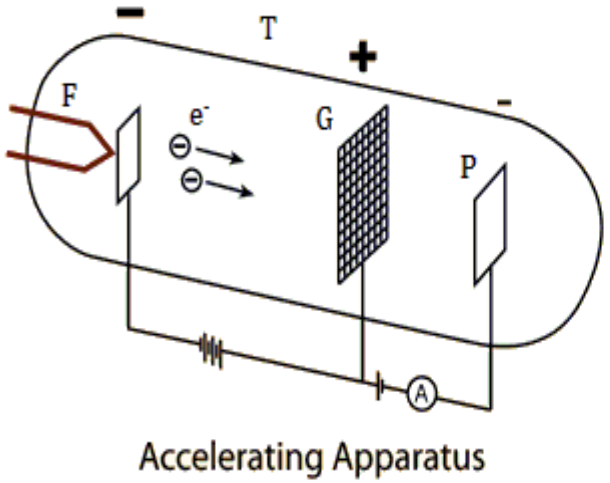
- Bohr atomic model explain the size of H atom and also radius of H atom.
- Bohr atomic model gives the ionization of H atom.
- Bohr atomic model explain formation of spectral lines in the H spectra.
- gives strong scientific basis. for the observation of emission and absorption spectral lines Bal, ly, paschen.
- Bohr found that the energy levels are quantized

Drawbacks of Bohr model

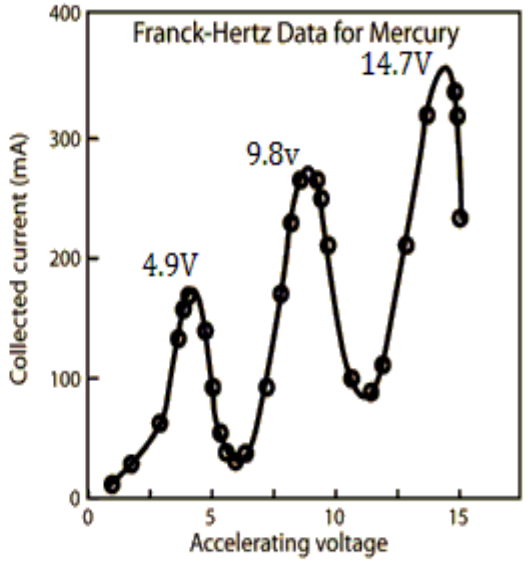
- It is valid only for one electron atoms, eg: H
- Orbits were taken as circular but according to Somerfield these are elliptical.
- Intensity of spectral lies could not be explained.
- Nucleus was taken as stationary but it also rotates on its own axis.
- It could not be explained the minute structure in spectrum line.
- This does not explain the Zeeman effect (splitting up of spectral lines in magnetic filed) and Stark effect (splitting up in electric filed). This does not explain the doublets in the spectrum of some of the atoms.

Franck-Hertz Experiment

- In 1914, James Franck and Gustav Hertz performed an experiment which demonstrated the existence of excited states in mercury atoms, and confirm the quantum theory which predicted that electrons occupied only discrete, quantized energy states.
- Electrons were accelerated by a voltage toward a positively charged grid in a glass envelope filled with mercury vapor.
- Past the grid was a collection plate held at a small negative voltage with respect to the grid. The values of accelerating voltage where the current dropped gave a measure of the energy necessary to force an electron to an excited state.

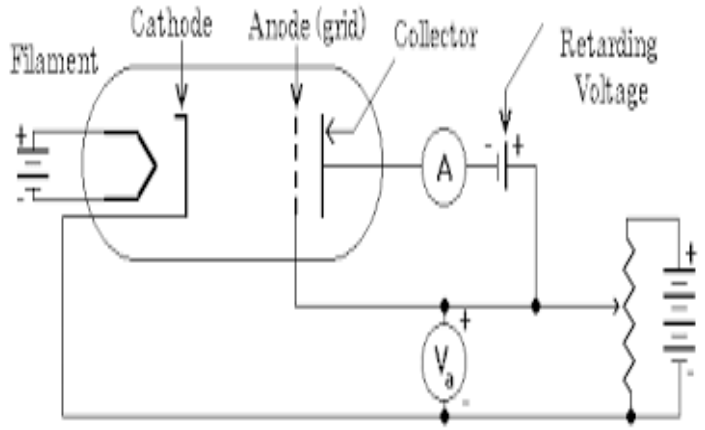


Accelerating Apparatus



Franck-Hertz Experiment

- The tube filled low pressure gas (like Hg is used) and fitted with 3 electrodes.
- When accelerating voltage applied Between F. and G. is order of 0 to 40v and distance between F. and G is 1or2 cm and that of G to plate collector less than 1mm.
- Retarding potential applied Between G. and collector plate (P)
- filament is heated e- are emitted by means of low tension battery. Electrons are emitted These emitted Electrons are accelerated toward grid. and F to G
- If these e have K.E grater than retarding potential they can reach the collector plate. once they reach collector plate indicated by (A)current



- When accelerating voltage is slowly increase more and more e-are reaching to plate if K.E grater than retarding potential

Franck-Hertz Experiment

- On the way when e passing through gas and these e collide with Hg atoms, Hg atoms are heavy as compared to e and elastic collision takes place and very small amount K. E. of e transfer to Hg atom and Retain K.E. reach the collector plate.
- It is found that When accelerating voltage reaches to critical value there is sudden drop in current indicates that most of the e not reaching to collector plate. because of losing the K.E. while collide with Hg atoms.
- When accelerating voltage reaches to critical value there is inelastic collision between e and Hg atom K. E. of e are absorbed by Hg atom.
- Father increase in accelerating voltage current increase after one elastic collision

