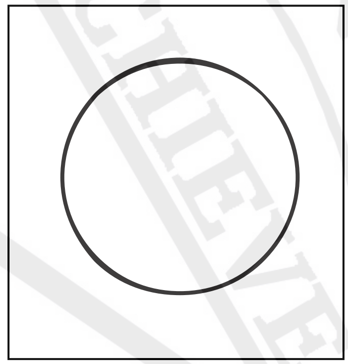
Class XI (Chemistry) Notes: (Structure of Atom)

Properties of fundamental particles of an atom

Name	Symbol	Absolute charge in Coulomb	Relative Charge	Mass IN Kg	Mass In *u*	Approx Mass In *u*
Electron	e	-1.6022×10^{-19}	-1	9.1×10^{-31}	.00054	0
Proton	p	$+1.6022 \times 10^{-19}$	+1	$1.67x \times 10^{-27}$	1.00727	1
Neutron	n	0	0	$1.67x \times 10^{-27}$	1.00867	1

ATOMIC MODELS

A. Thomson Model of Atom



An atom possesses a spherical shape in which

a) The positive charge is uniformly distributed.

b) The electrons are embedded into it in such a manner as to give the most stable electrostatic arrangement

c) Mass of the atom is assumed to be uniformly distributed over the atom.

This model, is also known as, **plum pudding, raisin pudding or watermelon** model.

Successes : This model was able to explain the overall neutrality of the atom,
But the model was not consistent with the results of later experiments.

B. Rutherford's Model of Atom

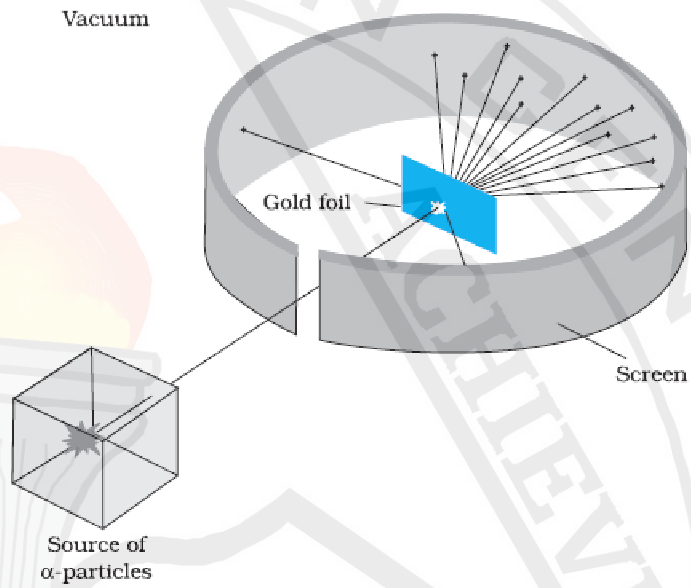
- **The experiment:**

The famous α **particle scattering experiment** is represented in Fig. 2 A stream of high energy α particles from a radioactive source was directed at a thin foil (thickness about 100 nm) of gold metal. The thin gold foil had a circular fluorescent zinc sulphide screen around it, to produce a tiny flash of light whenever α –particles struck the screen.

According to Thomson model of atom, the mass of each gold atom in the foil should have been spread evenly over the entire atom, and α particles had enough energy to pass directly through such a uniform distribution of mass. It was expected that the particles would slow down and change directions only by a small angles as they passed through the foil.

Observations

1. Most of the α particles passed through the gold foil undeflected.
2. A small fraction of the α particles was deflected by small angles.
3. A very few α particles (Less than 1 in 20,000) bounced back, that is, were deflected by nearly 180°.



Conclusions

1. Most of the space in the atom is empty.
2. The deflection of a few positively charged α particles must be due to enormous repulsive force showing that the positive charge of the atom is not spread throughout the atom as Thomson had presumed. The positive charge has to be concentrated in a very small volume that repelled and deflected the positively charged α particles.
3. Calculations by Rutherford showed that the volume occupied by the nucleus is negligibly small as compared to the total volume of the atom. The radius of the atom is about 10^{-10} m, while that of nucleus is 10^{-15} m.

Nuclear Model of the Atom

1. The positive charge and most of the mass of the atom was densely concentrated in extremely small region. This very small portion of the atom was called **nucleus** by Rutherford.
2. The nucleus is surrounded by electrons that move around the nucleus with a very high speed in circular paths called **orbits**.
3. Electrons and the nucleus are held together by electrostatic forces of attraction.

Drawbacks

1. *IT FAILS TO EXPLAIJN THE STABILITY OF AN ATOM*

Nuclear model suggests that electrons should move around the nucleus in well defined orbits. However, when a body is moving in an orbit, it undergoes acceleration (even if the body is moving with a constant speed in an orbit, it must accelerate because of changing direction). So an electron in the nuclear model is under acceleration. According to the electromagnetic theory of Maxwell, charged particles when accelerated should emit electromagnetic radiation. Therefore, an electron in an orbit will emit radiation; the energy carried by radiation comes from electronic motion. The orbit will thus continue to shrink. Calculations show that it should take an electron only 10^{-8} s to spiral into the nucleus. But this does not happen. Thus, the Rutherford model cannot explain the stability of an atom.

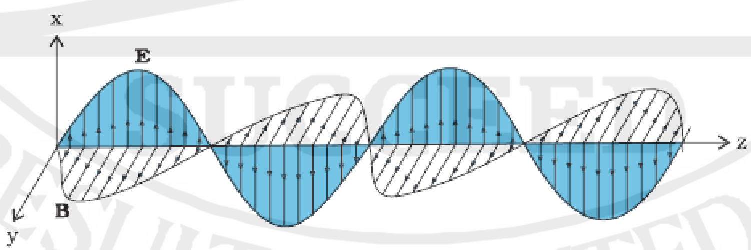
2. *IT SAYS NOTHING ABOUT THE ELECTRONIC STRUCTURE OF ATOMS* i.e., how the electrons are distributed around the nucleus and what are the energies of these electrons.

Question 1: Since the motion of electrons in orbits is leading to the instability of the atom, then why not consider electrons as stationary around the nucleus?

Answer: If the electrons were stationary, electrostatic attraction between the dense nucleus and the electrons would pull the electrons toward the nucleus to form a miniature version of Thomson's model of atom.

Electromagnetic radiation

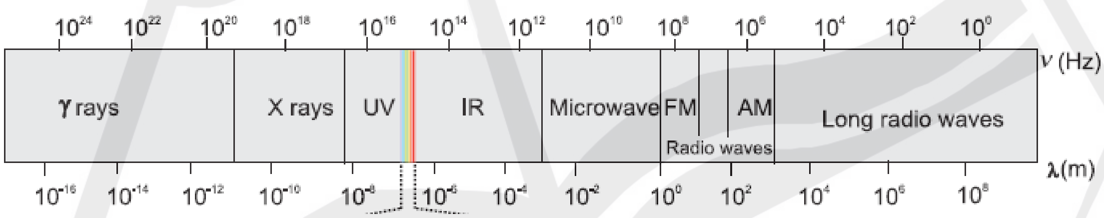
James Maxwell suggested that when electrically charged particle moves under acceleration, alternating electrical and magnetic fields are produced and transmitted. These



fields are transmitted in the forms of waves called **electromagnetic waves** or **electromagnetic radiation**.

Maxwell revealed that light waves are associated with oscillating electric and magnetic fields Its few simple properties are:

- The oscillating electric and magnetic fields produced by oscillating charged particles are perpendicular to each other and both are perpendicular to the direction of propagation of the wave.
- Unlike sound waves or water waves, electromagnetic waves do not require medium and can move in vacuum.
- There are many types of electromagnetic radiations, which differ from one another in wavelength (or frequency). These constitute **electromagnetic spectrum.** Different regions of the spectrum are identified by different names. Some examples are:
  - o Radio frequency region around 10^6 Hz, used for broadcasting
  - o Microwave region around 10^{10} Hz used for RADAR
  - o Infrared region around 10^{13} Hz used for heating
  - o Ultraviolet region around 1016Hz a component of sun's radiation.
  - o The small portion around 1015 Hz, is what is ordinarily called **visible light**. It is only this part which our eyes can see (or detect). Special instruments are



required to detect non-visible radiation.

Characteristics of EMR

a) Frequency(ν)

Number of waves that pass a fixed point per unit time; also, the number of cycles or vibrations undergone in unit time by a body in periodic motion. Frequency *f* is the reciprocal of the time *T* taken to complete one cycle (the period), or 1/*T*. The frequency with which earth rotates is once per 24 hours. Frequency is usually expressed in units called hertz (Hz). One hertz is equal to one cycle per second;

b) Wavelength(λ)

Distance between corresponding points of two consecutive waves. "Corresponding points" refers to two points or particles that have completed identical fractions of their periodic motion. In transverse waves, wavelength is measured from crest to crest or from trough to trough. In longitudinal waves, it is measured from compression to compression or from rarefaction to rarefaction.

$$\lambda f = v$$

c) Wavenumber ($\bar{v}$)

Wavenumber in most physical sciences is a wave property inversely related to wavelength, having SI units of reciprocal meters (m^{-1})

$$\bar{v} = \frac{1}{\lambda}$$

Planck's Quantum Theory

1. Planck suggested that atoms and molecules could emit (or absorb) energy only in discrete quantities and not in a continuous manner Planck gave the name **quantum** to the smallest quantity of energy that can be emitted or absorbed in the form of electromagnetic radiation.
2. The energy (E) of a quantum of radiation is proportional to its frequency (v) and is expressed by the following equation

$$E \alpha v \qquad \Rightarrow E = hv$$

The proportionality constant, 'h' is known as Planck's constant and has the value 6.626×10^{-34} J s.

With this theory, Planck was able to explain the distribution of intensity in the radiation from black body as a function of frequency or wavelength at different temperatures.

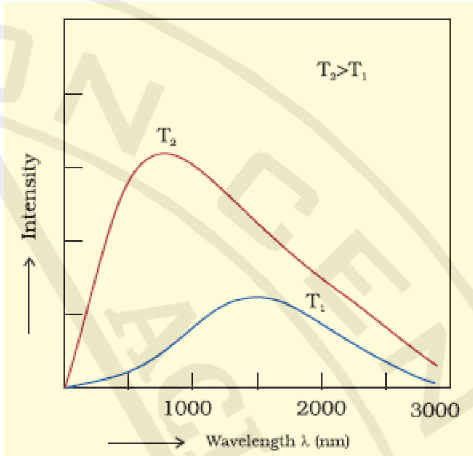
Successes of Quantum Theory

Some of the experimental phenomenon such as diffraction and interference can be explained by the wave nature of the electromagnetic radiation. However, following are some of the observations which could be explained with the help Quantum theory

- The nature of emission of radiation from hot bodies (black -body radiation)
- Ejection of electrons from metal surface when radiation strikes it (photoelectric effect)
- Variation of heat capacity of solids as a function of temperature

- Line spectra of atoms with special reference to hydrogen.

Black body radiation

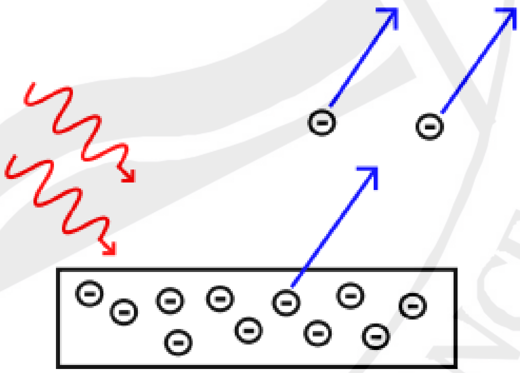


When solids are heated they emit radiation over a wide range of wavelengths. For example, when an iron rod is heated in a furnace, it first turns to dull red and then progressively becomes more and more red as the temperature increases. As this is heated further, the radiation emitted becomes white and then becomes blue as the temperature becomes very high.

In terms of frequency, it means that the radiation emitted goes from a lower frequency to a higher frequency as the temperature increases. The red colour lies in the lower frequency region while blue colour belongs to the higher frequency region of the electromagnetic spectrum.

The ideal body, which emits and absorbs all frequencies, is called a black body and the radiation emitted by such a body is called black body radiation.

The exact frequency distribution of the emitted radiation (i.e., intensity versus frequency curve of the radiation) from a black body depends only on its temperature. At a given temperature, intensity of radiation emitted increases with decrease of wavelength, reaches a maximum value at a given wavelength and then starts decreasing with further decrease of wavelength, as shown in the graph above.



Photoelectric effect

The **photoelectric effect** is a phenomenon in which electrons are emitted from matter (metals and non-metallic solids, liquids, or gases) after the absorption of energy from electromagnetic radiation such as X-rays or visible light. The emitted electrons can be referred to as **photoelectrons**

Following are the results of photoelectric effect

1. The electrons are ejected from the metal surface as soon as the beam of light strikes the surface, i.e., there is no time lag between the striking of light beam and the ejection of electrons from the metal surface.
2. The number of electrons ejected is proportional to the intensity or brightness of light.
3. For each metal, there is a characteristic minimum frequency,v0 (also known as **threshold frequency**) below which photoelectric effect is not observed.
4. At a frequency $\nu > \nu_0$, the ejected electrons come out with certain kinetic energy. The kinetic energies of these electrons increase with the increase of frequency of the light used.

All the above results could not be explained on the basis of laws of classical physics. .

According to the classical physics, the energy content of the beam of light depends upon the brightness of the light. Thus number of electrons ejected and kinetic energy associated with them should depend on the brightness of light. It has been observed that though the number of electrons

ejected does depend upon the brightness of light, the kinetic energy of the ejected electrons does not.

Einstein's Explanation

- In Einstein's picture, the basic elementary process involved in photoelectric effect is the absorption of a light quantum by an electron. This process is instantaneous.
- When a photon of sufficient energy strikes an electron in the atom of the metal, it transfers its energy instantaneously to the electron during the collision and the electron is ejected without any time lag or delay.
- Greater the energy possessed by the photon, greater will be transfer of energy to the electron and greater the kinetic energy of the ejected electron. In other words, kinetic energy of the ejected electron is proportional to the frequency of the electromagnetic radiation.
- Since the striking photon has energy equal to $h\nu$ and the minimum energy required to eject the electron is $h\nu_0$ (also called work function, w_0), then the difference in energy $(h\nu - h\nu_0)$ is transferred as the kinetic energy of the photoelectron.

- Following the conservation of energy principle, the kinetic energy of the ejected electron is given by the equation

$$h\nu = h\nu_0 + \frac{1}{2}m_e v^2{}_{\max}$$

where m_e is the mass of the electron and $v_{\max}$ v is the velocity associated with the ejected electron.

- Lastly, a more intense beam of light consists of larger number of photons, consequently the number of electrons ejected is also larger as compared to that in an experiment in which a beam of weaker intensity of light is employed.
- Thus, whatever may be the intensity i.e., the number of quanta of radiation per unit area per unit time, photoelectric emission is instantaneous. Low intensity does not mean delay in emission, since the basic elementary process is the same.
- Intensity only determines how many electrons are able to participate in the elementary process (absorption of a light quantum by a single electron) and, therefore, the photoelectric current.

Isotopes and Isobar

- A specific type of atom is designated by using its **chemical symbol**, which is an abbreviation of its name in German, Latin, or English, with the **A**, the mass number, placed in the upper left and **Z**, the atomic number, placed in the lower left corner. For example, $^{23}Na_{11}$, has a mass number of 23 and an atomic number of 11.
- A
- **Atomic number (Z):** The atomic number represents the **number of unit positive charges** on the nucleus and is equal to the **number of protons** within the nucleus, since each proton carries unit positive charge. In electrically neutral atoms, it also represents the number of electrons, which carry unit negative charge.

- **Mass number (A):** The mass number is equal to the **total number** of nucleons, which is the sum of the number of protons and neutrons. **A** does not equal the total mass of the atom; rather, it represents a whole number approximation of the mass, as expressed in amu.

- The **number of neutrons** is simply defined as the **A - Z.**
- **Isotopes** are atoms of the same element that differ in **mass**. For example, ^{87}Sr and ^{86}Sr or ^{238}U and ^{235}U. Isotopes have **similar chemical characteristics** and are studied using a **mass spectrograph** or **spectrometer**. Most elements have at least two naturally occurring isotopes.
- **Isobars** are nuclides that have the **same mass number** but **different atomic numbers**. For example, ^{36}S and ^{36}Ar are isobars; they both contain a total of 36 nucleons (protons plus neutrons), but the sulfur isotope has 16 protons and 20 neutrons, while the argon isotope has 18 protons and 18 neutrons. Isobars **do not have** similar chemical characteristics!

Atomic Spectra

- Since ordinary white light consists of waves with all the wavelengths in the visible range, a ray of white light is spread out into a series of coloured bands called **spectrum**.
- The spectrum of white light, that we can see, ranges from violet at 7.50×10^{14} Hz to red at 4×10^{14} Hz. Such a spectrum is called **continuous spectrum.** Continuous because violet merges into blue, blue into green and so on
- When electromagnetic radiation interacts with matter, atoms and molecules may absorb energy and reach to a higher energy state. With higher energy, these are in an unstable state. For returning to their normal (more stable, lower energy states) energy state, the atoms and molecules emit radiations in various regions of the electromagnetic spectrum.
- The spectrum of radiation emitted by a substance that has absorbed energy is called an **emission spectrum**. Atoms, molecules or ions that have absorbed radiation are said to be **"excited"**. To produce an emission spectrum, energy is supplied to a sample by heating it or irradiating it and the wavelength (or frequency) of the radiation emitted, as the sample gives up the absorbed energy, is recorded.
- An absorption spectrum is like the photographic negative of an emission spectrum. A continuum of radiation is passed through a sample which absorbs radiation of certain wavelengths. The missing wavelength which corresponds to the radiation absorbed by the matter, leave dark spaces in the bright continuous spectrum.

- The emission spectra of atoms in the gas phase, on the other hand, do not show a continuous spread of wavelength from red to violet, rather they emit light only at specific wavelengths with dark spaces between them. Such spectra are called **line spectra** or **atomic spectra** because the emitted radiation is identified by the appearance of bright lines in the spectra