1. Details of Module and its structure

Module Detail	
Subject Name	Chemistry
Course Name	Chemistry 01 (Class XI, Semester 01)
Module Name/Title	Structure of Atom: Part 1
Module Id	kech_10201
Pre-requisites	Knowledge of atom, charge
Objectives	After going through this module you will be able to: 1. Know about the discovery of electron, proton and neutron 2. Recall the characteristics of electron, proton and neutron 3. Explain the Rutherford‘s α particle scattering experiment 4. Define atomic number and mass number 5. Understand the drawbacks of Rutherford‘s model of atom
Keywords	Electron, Proton, α Particle, Thomson Model of Atom, Rutherford model of atom

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1. Introduction

In this module, you will study about the sub-atomic particles- electrons, protons and neutrons, a concept very different from that of Dalton and how Rutherford was able to construct a model of atom on the basis of his α particle scattering experiment.

The existence of atoms has been proposed since the time of early Indian and Greek philosophers (400 B.C.) who were of the view that atoms are the fundamental building blocks of matter. According to them, the continued subdivisions of matter would ultimately yield atoms which could not be further divided. The word 'atom' has been derived from the Greek word 'atomos' which means 'uncut-able' or 'non-divisible'. These earlier ideas were based on philosophical thinking and there was no way to test them experimentally.

The atomic theory of matter was first proposed on a firm scientific basis by John Dalton, a British school teacher in 1808. His theory, called **Dalton's atomic theory**, regarded the atom as the ultimate particle of matter.

2. SUB-ATOMIC PARTICLES

Dalton's atomic theory was able to explain the law of conservation of mass, law of constant composition and law of multiple proportion very successfully. However, it failed to explain the results of many experiments, for example, it was known that substances like glass or ebonite when rubbed with silk or fur generate electricity. Many different kinds of sub-atomic particles were discovered in the twentieth century. However, in this section we will talk about only two particles, namely electron and proton.

2.1 Discovery of Electron

In 1830, Michael Faraday showed that if electricity was passed through a solution of an electrolyte, chemical reactions occurred at the electrodes, which resulted in the liberation and deposition of matter at the electrodes. He formulated certain laws which you will study in class XII. These results suggested the particulate nature of electricity.
An insight into the structure of atom was obtained from the experiments on electrical discharge through gases. Before we discuss these results we need to keep in mind a basic rule regarding the behaviour of charged particles: "Like charges repel each other and unlike charges attract each other".
In mid 1850s many scientists, particularly Faraday, began to study electrical discharge in partially evacuated tubes known as **cathode ray discharge tubes**. It is depicted in Fig. 1(a). A cathode ray tube is made of glass containing two thin pieces of metal sealed in it at two ends of the tube. These pieces of metals are known as electrodes. A high electrical potential was applied across the electrodes. The electrical discharge through the gases present in the

tube could be observed only at very low pressures and at very high voltages. The pressure of different gases could be adjusted by evacuation. When sufficiently high voltage was applied across the electrodes, current started flowing through a stream of particles moving in the tube from the negative electrode (cathode) to the positive electrode (anode). These were called **cathode rays or cathode ray particles**.

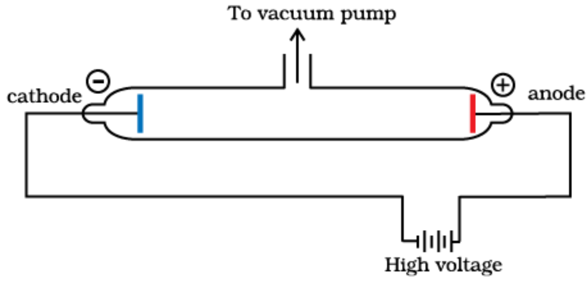


Fig. 1(a) *A cathode ray discharge tube*

Movement of these rays was checked by using a perforated anode and coating the tube behind anode with phosphorescent material zinc sulphide. When these rays, after passing through anode, strike the zinc sulphide coating, a bright spot on the coating is developed (same thing happens in a television set) [Fig. 1(b)].

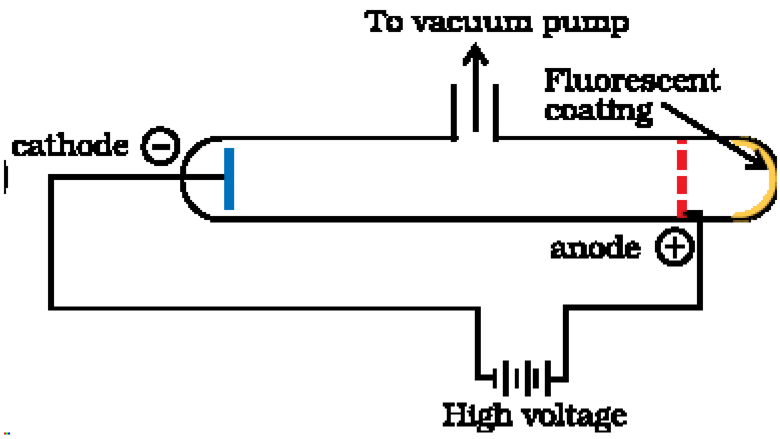


Fig. 1(b) *A cathode ray discharge tube with perforated anode*

The results of these experiments are summarised below.

(i) The cathode rays start from cathode and move towards the anode.

(ii) These rays themselves are not visible but their behaviour can be observed with the help of certain kind of materials (fluorescent or phosphorescent) which glow when hit by them. Television picture tubes are cathode ray tubes and television pictures result due to fluorescence on the television screen coated with certain fluorescent or phosphorescent materials.

(iii) In the absence of electrical or magnetic field, these rays travel in straight lines and strike at point B on the screen (Fig. 2).

(iv) In the presence of electrical or magnetic field, the behaviour of cathode rays are similar to that expected from negatively charged particles, suggesting that the cathode rays consist of negatively charged particles, called **electrons.**

(v) The characteristics of cathode rays (electrons) do not depend upon the material of electrodes and the nature of the gas present in the cathode ray tube.

Thus, we can conclude that electrons are basic constituent of all the atoms.

2.2 Charge to Mass Ratio of Electron

In 1897, British physicist J.J. Thomson measured the ratio of electrical charge (e) to the mass of electron (m_e) by using cathode ray tube and applying electrical and magnetic field perpendicular to each other as well as to the path of electrons (Fig. 2). Thomson argued that the amount of deviation of the particles from their path in the presence of electrical or magnetic field depends upon:

(i) the magnitude of the negative charge on the particle—greater the magnitude of the charge on the particle, greater is the interaction with the electric or magnetic field and thus greater is the deflection.

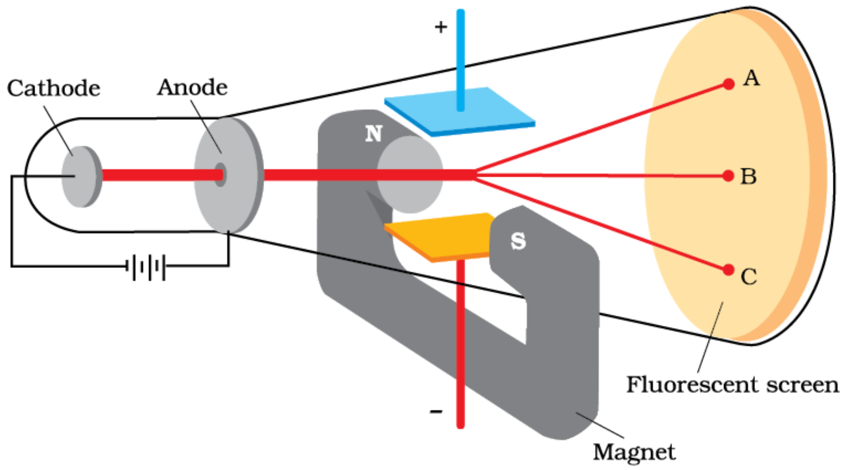


Fig. 2 The apparatus to determine the charge to the mass ratio of electron

(ii) the mass of the particle — lighter the particle, greater the deflection.

(iii) the strength of the electrical or magnetic field — the deflection of electrons from its original path increases with the increase in the voltage across the electrodes, or the strength of the magnetic field.

When only electric field is applied, the electrons deviate from their path and hit the cathode ray tube at point A. Similarly when only magnetic field is applied, electron strikes the cathode ray tube at point C. By carefully balancing the electrical and magnetic field strength, it is possible to bring back the electron to the path followed as in the absence of electric or magnetic field and they hit the screen at point B. Thomson carried out accurate measurements on the extent of deflections observed by the electrons on the electric field strength or magnetic field strength. Based on these measurements, he determined the value of e/m_e as 1.758820×10^{11} C kg^{-1}

Where m_e is the mass of the electron in kg and e is the magnitude of the charge on the electron in coulomb (C). Since electrons are negatively charged, the charge on electron is –e.

2.3 Charge on the Electron

R.A. Millikan (1868-1953) devised a method known as oil drop experiment, to determine the charge on the electrons. He found the charge on the electron to be -1.6×10^{-19} C. The present accepted value of electrical charge is -1.6022×10^{-19}C. The mass of the electron (m_e) was determined by combining these results with Thomson's value of e/m_e ratio.

$$m_e = \frac{e}{e/m_e} = \frac{1.6022 \times 10^{-19}\,\mathrm{C}}{1.278820 \times 10^{11}\,\mathrm{C\,kg^{-1}}} = 9.1094 \times 10^{-31}\,\mathrm{kg}$$

Millikan's Oil Drop Method

In this method, oil droplets in the form of mist, produced by the atomiser, were allowed to enter through a tiny hole in the upper plate of an electrical condenser. The downward motion of these droplets was viewed through the telescope, equipped with a micrometer eye piece. By measuring the rate of fall of these droplets, Millikan was able to measure the mass of oil droplets. The air inside the chamber was ionized by passing a beam of X-rays through it. The electrical charge on these oil droplets was acquired by collisions with gaseous ions. The fall of these charged oil droplets can be retarded, accelerated or made stationary depending upon the charge on the droplets and the polarity and strength of the voltage applied to the plate. By carefully measuring the effects of electrical field strength on the motion of oil droplets, Millikan concluded that the magnitude of electrical charge, q, on the droplets is always an integral multiple of the electrical charge, e, that is, q = n e, where n = 1, 2, 3... .

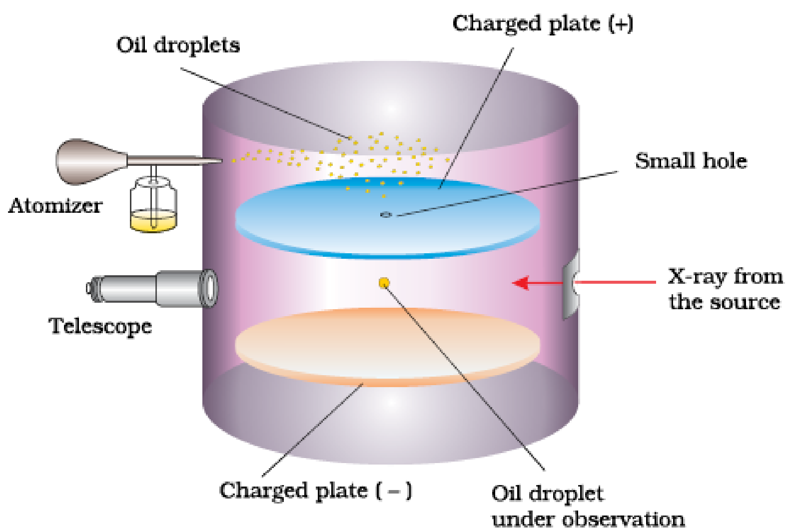


Fig. 3 *The Millikan oil drop apparatus for measuring charge 'e'. In chamber, the forces acting on oil drop are: gravitational, electrostatic due to electrical field and a viscous drag force when the oil drop is moving.*

2.4 Discovery of Protons and Neutrons

Electrical discharge carried out in the modified cathode ray tube led to the discovery of particles carrying positive charge, also known as **canal rays.** The characteristics of these positively charged particles are listed below.

(i) Unlike cathode rays, the positively charged particles depend upon the nature of gas present in the cathode ray tube. These are simply the positively charged gaseous ions.

(ii) The charge to mass ratio of the particles is found to depend on the gas from which these originate.

(iii) Some of the positively charged particles carry a multiple of the fundamental unit of electrical charge.

(iv) The behaviour of these particles in the magnetic or electrical field is opposite to that observed for electron or cathode rays.

The smallest and lightest positive ion was obtained from hydrogen and was called **proton**. This positively charged particle was characterised in 1919. Later, a need was felt for the presence of electrically neutral particle as one of the constituent of atom. Chadwick discovered these particles in 1932. On bombarding a thin sheet of beryllium by α-particles, an electrically neutral particle having a mass slightly greater than that of the protons was emitted. He named these particles as **neutrons**. The important properties of these fundamental particles are given in Table 1.

3. ATOMIC MODELS

Observations obtained from the experiments mentioned in the previous sections have suggested that Dalton's indivisible atom is composed of sub-atomic particles carrying positive and negative charges.

Table 1 Properties of Fundamental Particles

Name	Symbol	Absolute charge/C	Relative charge	Mass/kg	Mass/u	Approx. mass/u
Electron	e	-1.6022×10^{-19}	-1	9.10939×10^{-31}	0.00054	0
Proton	p	$+1.6022\times10^{-19}$	+1	1.67262×10^{-27}	1.00727	1
Neutron	n	0	0	1.67493×10^{-27}	1.00867	1

Different atomic models were proposed to explain the distributions of these charged particles in an atom. Although some of these models were not able to explain the stability of atoms, two of these models, proposed by J. J. Thomson and Ernest Rutherford are discussed below.

3.1 Thomson Model of Atom

J. J. Thomson, in 1898, proposed that an atom possesses a spherical shape (radius approximately 10^{-10}m) in which the positive charge is uniformly distributed. The electrons are embedded into it in such a manner as to give the most stable electrostatic arrangement (Fig. 4). Many different names are given to this model, for example, **plum pudding, raisin pudding or watermelon.** This model can be visualised as a pudding or watermelon of positive charge with plums or seeds (electrons) embedded into it.

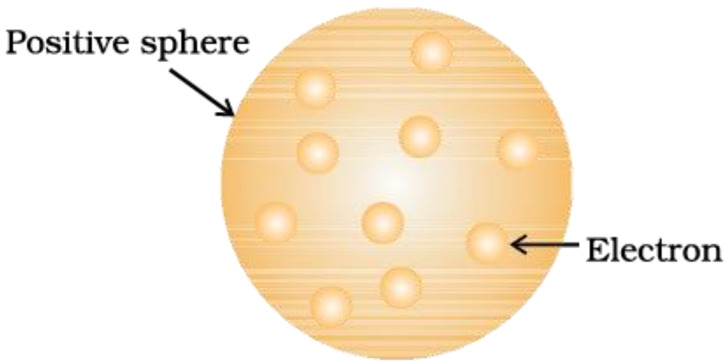
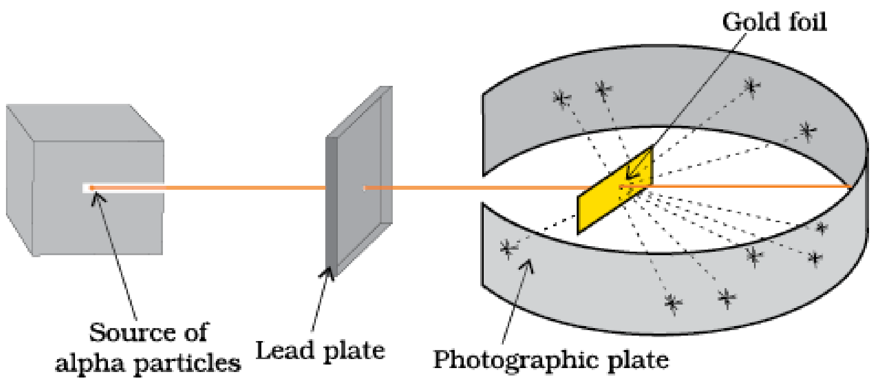


Fig .4 Thomson model of atom

An important feature of this model is that the mass of the atom is assumed to be uniformly distributed over the atom. Although this model was able to explain the overall neutrality of the atom, but was not consistent with the results of later experiments. Thomson was awarded Nobel Prize for physics in 1906, for his theoretical and experimental investigations on the conduction of electricity by gases.

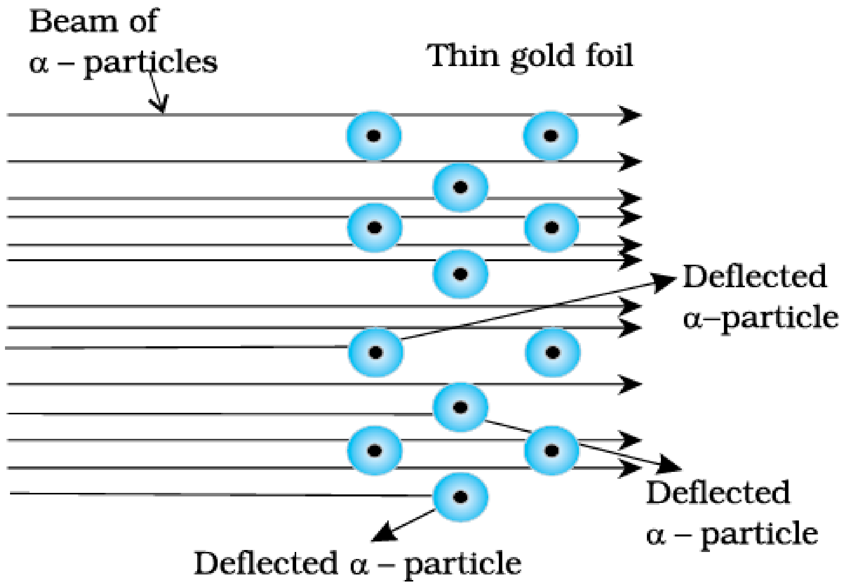
3.2 Rutherford's Nuclear Model of Atom

Rutherford and his students (Hans Geiger and Ernest Marsden) bombarded a very thin gold foil with α–particles. Rutherford's famous α–**particle scattering experiment** is represented in Fig. 5. A stream of high energy α–particles from a radioactive source was directed at a thin foil (thickness ~100 nm) of gold metal.



A. *Rutherford's scattering experiment*

B.



B. Schematic molecular view of the gold foil

Fig.5 *Schematic view of Rutherford's scattering experiment. When a beam of alpha (α) particles is "shot" at a thin gold foil, most of them pass through without much effect. Some, however, are deflected.*

The thin gold foil had a circular fluorescent zinc sulphide screen around it. Whenever α–particles struck the screen, a tiny flash of light was produced at that point.

The results of scattering experiment were quite unexpected. 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. However, it was observed that :

(i) Most of the α– particles passed through the gold foil undeflected.

(ii) A small fraction of the α–particles was deflected by small angles.

(iii) Very few α– particles (~1 in 20,000) bounced back, that is, were deflected by nearly 180°.

On the basis of these observations, Rutherford drew the following conclusions regarding the structure of atom:

(i) Most of the space in the atom is empty as most of the α–particles passed through the foil undeflected.

(ii) A few positively charged α– particles were deflected. The deflection 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.

(iii) 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. One can appreciate this difference in size by realising that if a cricket ball represents a nucleus, then the radius of atom would be about 5 km.

On the basis of above observations and conclusions, Rutherford proposed the nuclear model of atom (after the discovery of protons). According to this model:

(i) The positive charge and most of the mass of the atom is densely concentrated in extremely small region. This very small portion of the atom was called **nucleus** by Rutherford.

(ii) The nucleus is surrounded by electrons that move around the nucleus with a very high speed in circular paths called **orbits**. Thus, Rutherford's model of atom resembles the solar system in which the nucleus plays the role of sun and the electrons that of revolving planets.

(iii) Electrons and the nucleus are held together by electrostatic forces of attraction.

3.3 Atomic Number and Mass Number

The presence of positive charge on the nucleus is due to the protons in the nucleus. As established earlier, the charge on the proton is equal but opposite to that of electron. The number of protons present in the nucleus is equal to atomic number (Z). For example, the number of protons in the hydrogen nucleus is 1, in sodium atom it is 11, therefore their atomic numbers are 1 and 11 respectively. In order to keep the electrical neutrality, the number of electrons in an atom is equal to the number of protons (atomic number, Z). For example, number of electrons in hydrogen atom and sodium atom are 1 and 11 respectively.

Atomic number (Z) = number of protons in the nucleus of an atom

= number of electrons in a neutral atom (1)

While the positive charge of the nucleus is due to protons, the mass of the nucleus is due to protons and neutrons. The protons and the neutrons present in the nucleus are collectively known as **nucleons**. The total number of nucleons is termed as **mass number (A)** of the atom.

Mass number (A) = number of protons (Z) + number of neutrons (n) (2)

3.4 Isobars and Isotopes

The composition of any atom can be represented by using the normal element symbol (X) with super-script on the left hand side as the atomic mass number (A) and subscript (Z) on the left hand side as the atomic number (i.e.,$^{A}_{Z}X$).

Isobars are the atoms with same mass number but different atomic number for example, $^{14}_{6}C$ and $^{14}_{7}N$. On the other hand, atoms with identical atomic number but different atomic mass number are known as **Isotopes**. According to equation 2, it is evident that difference between the isotopes is due to the presence of different number of neutrons present in the nucleus. For example, considering hydrogen atom again, 99.985% of hydrogen atoms contain only one proton. This isotope is called **protium ($^{1}_{1}H$)**. Rest of the percentage of hydrogen atom contains two other isotopes. The one containing 1 proton and 1 neutron is called **deuterium** ($^{2}_{1}D$, 0.015%) and the other one possessing 1 proton and 2 neutrons is called **tritium ($^{3}_{1}T$)**. The latter isotope is found in trace amounts on the earth. Other examples of commonly occuring isotopes are: carbon atoms containing 6, 7 and 8 neutrons besides 6 protons ($^{12}_{6}C$,$^{13}_{6}C$,$^{14}_{6}C$); chlorine atoms containing 18 and 20 neutrons besides 17 protons ($^{35}_{17}Cl$,$^{37}_{17}Cl$).

Lastly an important point to mention regarding the isotopes is that the chemical properties of atoms are controlled by the number of electrons, which are determined by the number of protons in the nucleus. Number of neutrons present in the nucleus have very little effect on the chemical properties of an element. Therefore, all the isotopes of a given element show same chemical behaviour.

Problem 1: Calculate the number of protons, neutrons and electrons in $^{80}_{35}\text{Br}$.
Solution
In this case, $^{80}_{35}\text{Br}$, Z = 35, A = 80, species is neutral
Number of protons = number of electrons = Z = 35
Number of neutrons = 80 – 35 = 45, (equation 2)
Problem 2: The number of electrons, protons and neutrons in a species are equal to 18, 16 and 16 respectively. Assign the proper symbol to the species.
Solution
The atomic number is equal to number of protons = 16.
The element is sulphur (S).
Atomic mass number = number of protons + number of neutrons
= 16 + 16 = 32
Species is not neutral as the number of protons is not equal to electrons. It is anion (negatively charged) with charge equal to excess electrons = 18 – 16 = 2. Symbol is $^{32}_{16}\text{S}^{-2}$.
Note: Before using the notation $^{A}_{Z}\text{X}$, find out whether the species is a neutral atom, a cation or an anion. If it is a neutral atom, equation (1) is valid, i.e., number of protons = number of electrons = atomic number. If the species is an ion, determine whether the number of protons are larger (cation, positive ion) or smaller (anion, negative ion) than the number of electrons. Number of neutrons is always given by A–Z, whether the species is neutral or ion.

3.5 Drawbacks of Rutherford Model

Rutherford nuclear model of an atom is like a small scale solar system with the nucleus playing the role of the massive sun and the electrons being similar to the lighter planets. Further, the coulomb force$\left(\frac{\text{k}q_1q_2}{r^2}\right)$where q_1 and q_2 are the charges, r is the distance of separation of the charges and k is the proportionality constant) between electron and the nucleus is mathematically similar to the gravitational force $\left(\frac{\text{G}m_1m_2}{r^2}\right)$where$m_1$ and m_2 are the masses, r is the distance of separation of the masses and G is the gravitational constant. When classical mechanics (*a theoretical science based on Newton's laws of motion*)is applied to the solar system, it shows that the planets describe well-defined orbits around the sun. The theory can also calculate precisely the planetary orbits and these are in agreement with the experimental measurements. The similarity between the solar system and 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 describing planet like orbits is under acceleration. According to the electromagnetic theory of Maxwell, charged particles when accelerated should emit electromagnetic radiation. (This feature does not exist for planets since they are uncharged). The energy carried by radiation comes from electronic motion. Consequently, the electron

will keep on losing energy in the form of radiation and 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.
If the motion of an electron is described on the basis of the classical mechanics and electromagnetic theory, you may ask that 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. 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.
Another serious drawback of the Rutherford model is that 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.

4. Summary

- Atoms are the building blocks of elements.
- They are the smallest parts of an element that react chemically.
- Dalton regarded atom as invisible.
- Later it was found that atoms are divisible and consist of three fundamental particles viz. electrons, protons and neutrons.
- Thomson proposed that an atom consists of uniform sphere of positive charge with electrons embedded into it.
- Rutherford proposed that an atom is made of a tiny positively charged nucleus at its centre, with electrons revolving around it in circular orbits.
- Main drawback of Rutherford's model of atom was that it could not explain the stability of the atom i.e. why the electrons do not fall into the nucleus.