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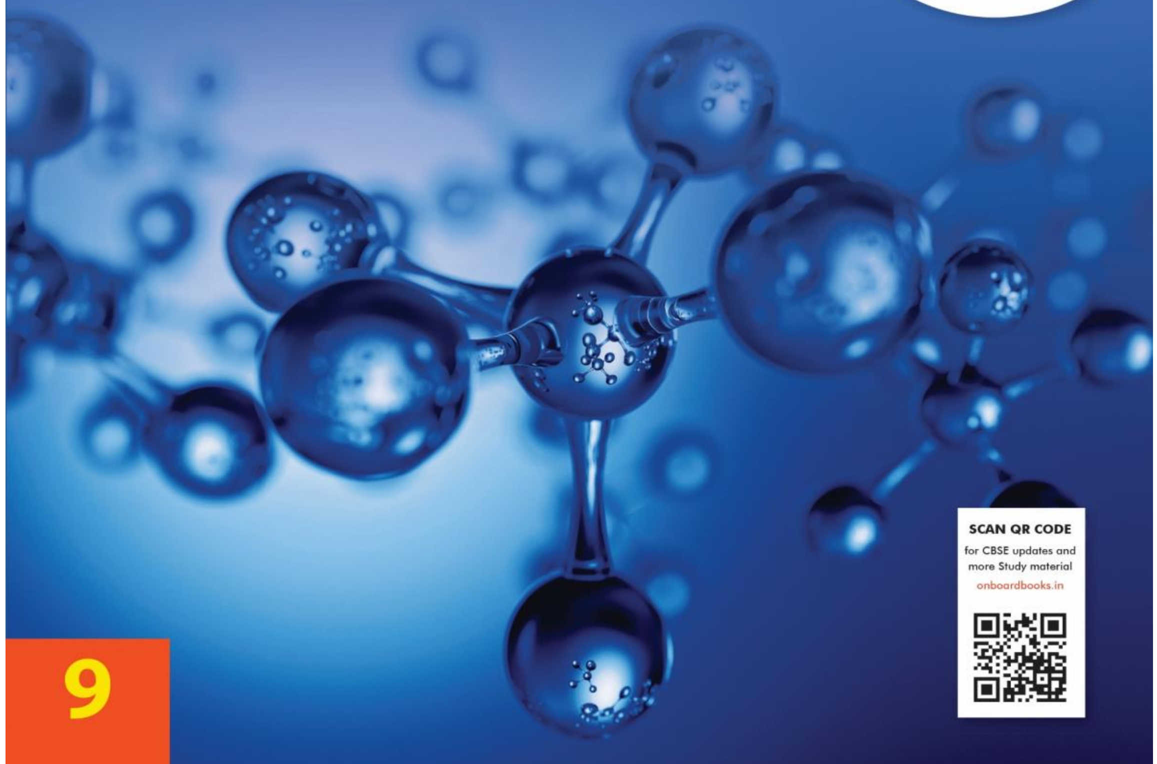

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9

CBSE LIVING SCIENCE CHEMISTRY

CLASS 9

Chapter 4

Structure of the Atom

Learning Objectives

- ❖ **Electron, proton and neutron**
- ❖ **Structure of atom: models of atom**
- ❖ **Isotopes and isobars**

INTRODUCTION

The ancient Greek and Indian philosophers suggested around 400 BC that matter was made up of tiny discrete particles called atoms, but there was no way to verify their idea experimentally. In 1803, John Dalton, a British school teacher, put forward his atomic theory called Dalton's atomic theory. According to this theory, an atom is the smallest, indivisible and indestructible particle of an element that takes part in chemical change. Towards the end of nineteenth century, it was found that an atom was composed of subatomic particles such as electron, proton and neutron. These particles are called elementary fundamental particles because these are essential constituents of atoms.

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ELECTRICAL NATURE OF MATTER

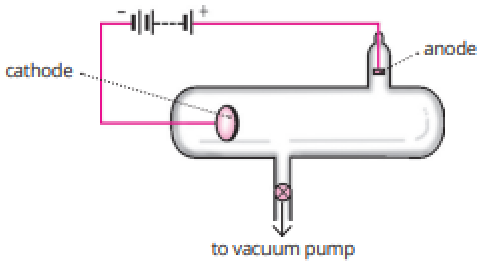
The electrical nature of matter was known to ancient Greeks as early as in 600 BC. They found a hard transparent yellowish-brown substance which could be electrified by friction. William Gilbert (1540–1603) who also observed that there were two types of electrical charges, i.e. positive and negative. The experimental evidence of electrical nature of matter came from the experiment of Michael Faraday who showed in 1833 that the flow of the electricity is due to the flow of the charged particles and when electricity is passed through an electrolyte, the electrolyte decomposes to give the constituent elements. The term electron for the fundamental electrical particle was first suggested in 1874 by G J Stoney who used it for the unit charge on a monovalent negative ion.

The first experimental evidence for the existence of electrons was given by J J Thomson.

DISCOVERY OF ELECTRON – STUDY OF CATHODE RAYS

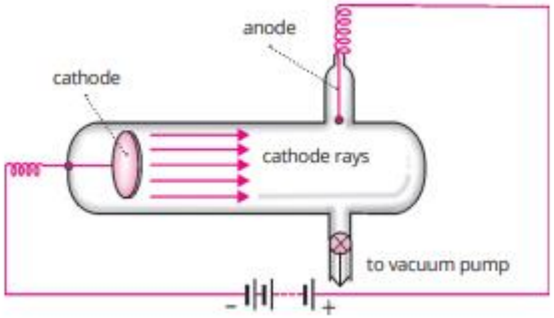
In 1879, William Crookes showed that when an electric discharge at very high voltage is passed through a gas at 0.001 mm of Hg pressure in a discharge tube, the glass wall opposite to the cathode begins to glow.

Discharge tube: A discharge tube is a cylindrical glass tube about 60 cm long, fitted with two metallic electrodes sealed at the two ends. The positive electrode of the discharge tube is called anode and the negative electrode is called cathode. The discharge tube has a side tube fitted with a stopcock (Fig.) through which the air present in the discharge tube can be pumped out by a vacuum pump. When air inside the discharge tube is at 1 atm pressure (760 mm of Hg) and a high voltage (10000–15000 V) is applied, electricity does not flow through the air in the discharge tube. This is because air or any gas at normal pressure is a poor conductor of electricity. But when the pressure in the discharge tube is reduced to 1 mm of Hg, a pink glow is emitted by the air inside the tube.



Emission of pink glow by air inside the discharge tube

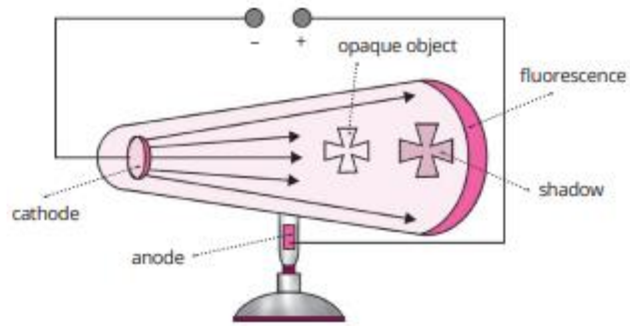
When the pressure of air in the discharge tube is reduced to about 0.001 mm of Hg, the emission of pink light by air stops and the glass walls of the discharge tube opposite to the cathode start to glow with a faint greenish light and inside of the discharge tube turns dark (Fig.). The green glow is due to the fluorescence of the walls produced by the bombardment of the glass by invisible rays coming out from the cathode. These invisible rays coming out from the cathode are called cathode rays. Thus, when electrical discharge is passed through gases at very low pressures, cathode rays are produced.



Emission of cathode rays in the discharge tube opposite to the cathode

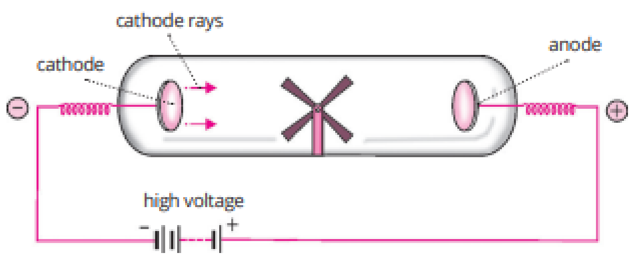
Properties of cathode rays

1. **Cathode rays travel in straight lines**: When an opaque solid object such as a metal cross is placed in the path of cathode rays, a shadow of the metal cross is formed on the wall opposite to the cathode (Fig.). The shadow of the object can be formed only if the cathode rays travel in straight lines.



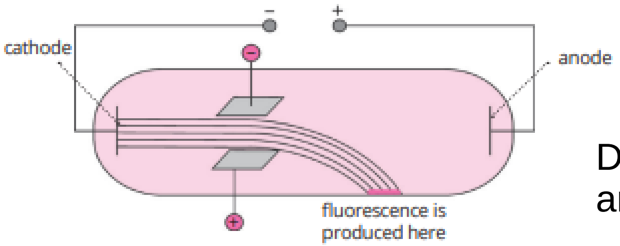
Experiment showing cathode rays casting shadow of an object placed in their path

2. **Cathode rays produce mechanical effects**: Cathode rays produce mechanical motion of a paddle wheel when it is placed in the path of the cathode rays (Fig.). This indicates that the cathode rays consist of material particles having mass and kinetic energy.



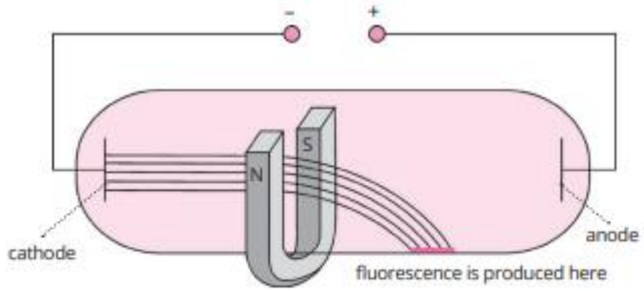
Experiment showing cathode rays producing mechanical motion of paddle wheel placed in their path

3. **Cathode rays are deflected by an electric field**: When cathode rays are passed through a strong electric field formed by placing positively charged and negatively charged plates in their path, they are deflected towards positively charged plate in their path (Fig.) and this shows that they consist of negatively charged particles.



Deflection of cathode rays towards a positively charged plate in an electric field

4. **Cathode rays are deflected by a magnetic field**: When cathode rays are passed through a magnetic field, they are deflected in a direction which shows that they are negatively charged particles. When placed in a magnetic field, the cathode rays are deflected towards the north pole of the magnet. When the cathode rays are passed through electric field and magnetic field in succession arranged perpendicular to each other and the electric and magnetic field strengths are adjusted in such a manner that they balance each other to produce no deflection in the electron beam. Thus, the deflection produced by the magnetic field is cancelled by the deflection produced in the opposite direction by the electric field.



Deflection of cathode rays by a magnetic field

5. **Cathode rays heat thin metal foils to incandescence**: When cathode rays are allowed to strike thin metal foils, a large amount of heat is generated and the metal foil becomes incandescent (white hot).

6. **Cathode rays impart negative charge to the objects in their path**: When cathode rays are allowed to fall on objects, the objects acquire negative charge. This indicates that the cathode rays are energetic negatively charged particles.

7. **Cathode rays cause ionisation of gases**: When gases are exposed to cathode rays, there occurs ionisation of gases.

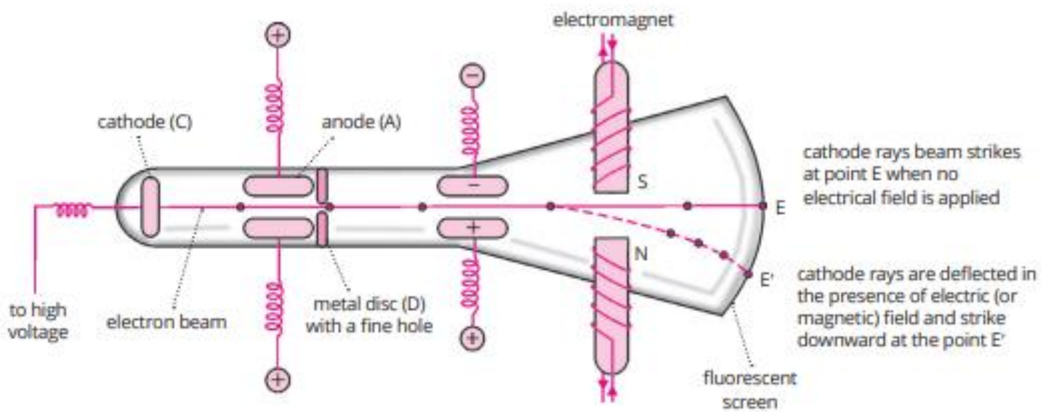
$$X_2(g) + e^- \longrightarrow X_2^+(g) + 2e^-$$

8. Cathode rays expose photographic plates.
9. Cathode rays produce X-rays on striking metal targets.
10. Cathode rays produce fluorescence.

Origin of cathode rays: The gas in the discharge tube is composed of atoms and all atoms contain electrons. On application of high electrical voltage, the electrical energy knocks out some of the electrons from the atoms of the gas taken in the discharge tube. These electrons constitute the cathode rays. Hence, the emission of cathode rays indicates the presence of electrons in atoms

J J THOMSON'S EXPERIMENT

The properties of cathode rays indicate that these consist of negatively charged particles. These particles were called 'corpuscles of negative electricity' by J J Thomson and later G J Stoney used the term 'electrons' for these particles. Actually in his experiment Thomson studied the combined effect of electric and magnetic fields on the cathode rays.



J J Thomson's experiment

When cathode rays are passed through a magnetic field, they are deflected in a direction which shows that they are negatively charged particles. Thomson allowed the cathode rays to pass through magnetic field and electric field arranged perpendicular to each other in succession such that the field produced by the electromagnet was opposite to the electric field. When the cathode rays were passed through electric field only, they were deflected towards the positive plate. Then they were passed through both the electric and the magnetic fields, and the electric and magnetic field strengths were adjusted in such a manner that they balanced each other to produce no deflection in the cathode rays beam. Thus, the deflection produced by the magnetic field was cancelled by the deflection produced in the opposite direction by the electric field (i.e. net deflection = 0).

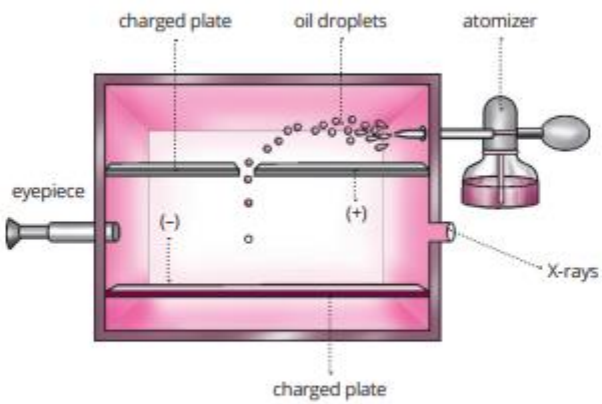
By following this method of balancing the deflections of cathode rays in magnetic and electric fields, Thomson calculated the charge/mass ratio, i.e. *e*/*m* ratio. He found the *e*/*m* ratio of cathode rays to be 1.759×10^{11} C kg^{-1}. It was also found that the nature of cathode rays (i.e. *e*/*m* ratio of the particles) does not depend upon the nature of the gas taken in the discharge tube and the material of the electrodes used in the discharge tube.

$$e/m = 1.759 \times 10^{11}\ \mathrm{C\ kg^{-1}} = \text{constant}$$

It was concluded by Thomson from this experiment that all the atoms contained the same type of negatively charged particles. These negatively charged particles were called electrons. Electrons are essential constituents of all matters. An electron is known as a fundamental particle – a particle found in all atoms.

OIL–DROP EXPERIMENT TO FIND CHARGE ON AN ELECTRON

In 1909, an American chemist, Robert A. Millikan measured the charge (*e*) on an electron by carrying out his famous oil-drop experiment. What Millikan did was to put a charge on a tiny drop of oil, and measure how strong an applied electric field had to be in order to stop the oil-drop from falling (Fig.). Since he was able to work out the mass of the oil-drop and he could calculate the force of gravity on one drop, he could determine the electric charge that the drop must have. By varying the charge on different drops, he noticed that the charge was always a multiple of -1.602×10^{-19} C, the charge on a single electron. This meant that these were electrons carrying this unit charge. This quantity of electric charge was called electronic charge, *e*. This is the minimum value of negative charge and no stable fundamental particle is known, which contains charge less than this value.



Millikan's oil-drop experiment for determination of charge of the electron

This is the reason why the numerical value of one electronic charge represents the fundamental unit of electricity.

Thus, an electron has 1 unit of negative charge (–1). Hence, the relative charge of an electron is –1. Therefore, charge on an electron, *e*

$$= -1.602 \times 10^{-19}\ \text{C}$$

Mass of an electron

Mass of an electron is calculated from the values of e/m and e as follows:

$$\text{Mass of electron } (m) = \frac{\text{Charge}}{\text{Charge/Mass}} = \frac{e}{e/m}$$

$$= \frac{1.602 \times 10^{-19}\ \text{C}}{1.759 \times 10^{11}\ \text{C kg}^{-1}}$$

$$= 9.109 \times 10^{-31}\ \text{kg}$$

Let us now compare the mass of an electron with the mass of one hydrogen atom. We know that one gram-atom of hydrogen weighs 1.008 g. Hence, the mass of one hydrogen atom = 1.008 g/Avogadro's number = 1.008 g/6.022 × 10^{23} = 1.673 × 10^{-24} g.

$$\text{Hence,}\quad \frac{\text{Mass of electron}}{\text{Mass of 1 atom of hydrogen}} = \frac{9.109\times10^{-31}\ \text{kg}}{1.673\times10^{-27}\ \text{kg}} = \frac{1}{1837}$$

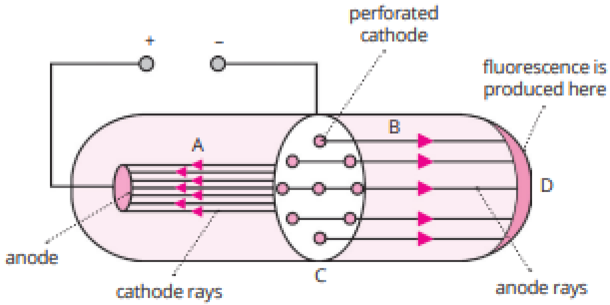
Thus, the mass of an electron is 1/1837 times that of one hydrogen atom. This means that the mass of one hydrogen atom is 1837 times greater than the mass of an electron. Hence, the mass of an electron is much smaller than the mass of one hydrogen atom. We can now calculate the mass of an electron in atomic mass unit (u).

$$\text{Mass of electron } (m) = \frac{1.008}{1837}\,\text{u} = 0.0005487\ \text{u}$$

An electron is defined as a subatomic fundamental particle having mass equal to (1/1837)th of that of a hydrogen atom and carrying one unit of negative charge.

ANODE RAYS

The discovery of electrons in atoms stimulated the search for the discovery of positively charged component of the atom. In 1886, Eugen Goldstein carried out the discharge tube experiments using perforated cathode (Fig). The discharge tube consisted of two chambers A and B, which were separated by a metallic disc C which acted as the perforated cathode.



Goldstein's discharge tube experiment producing anode rays

On lowering the pressure of gas inside the discharge tube and applying high voltage, it was observed that in addition to cathode rays, luminous rays emerged from the anode, which came through the perforations of the cathode. These rays travelled in straight lines in the opposite direction to that of the cathode rays, passed through the perforations (canals) of the cathode and produced fluorescence on a screen placed on the other end D of the second chamber. These rays are also visible due to the glow of the residual gas along their path in the chamber B. Since these rays passed through the canals in the cathode, they are called canal rays or anode rays as they emerged from the anode. They are also called positive rays. The electrical charge on the particles of anode rays was found to be an integral multiple of the magnitude of the charge of an electron (1.602×10^{-19} C) but with a positive sign.

Origin of anode rays

In a discharge tube, when electrical discharge is passed through a gas at a very low pressure (0.001 mm of Hg), cathode rays are formed. The cathode rays are stream of fast moving electrons. When these electrons strike the atoms or molecules of the gas in the discharge tube, one or more electrons are knocked off from the atoms or molecules of the gas and as a result, ionisation of atoms or molecules occurs and positively charged particles are formed. These positively charged particles (ions) of the gas constitute the anode rays.

$$X(g) + e^- \longrightarrow X^+(g) + 2e^-$$

$$X_2(g) + e^- \longrightarrow X_2^+(g) + 2e^-$$

Properties of anode rays

1. Anode rays consist of positively charged particles.

2. Anode rays travel in straight lines.

3. Anode rays are deflected by the electric field and they bend towards the negative plate.

4. Anode rays are deflected by the magnetic field in the direction opposite to that of cathode rays.

5. The nature of anode rays depends upon the nature of the gas taken in the discharge tube. The *e*/*m* ratio for anode rays is very small compared to that for cathode rays or electron.

The *e*/*m* ratio for anode rays is not equal for all gases. The *e*/*m* ratio for anode rays of hydrogen gas is the highest (9.578×10^4 C g^{-1}). Because hydrogen atom is the lightest among all atoms, i.e. *m* is the lowest and thus, *e*/*m* is the highest.

The maximum value of e/m ratio for hydrogen indicates that the positive particles formed from hydrogen has the smallest mass (assuming same charge). Particles in the positive rays are heavier than the particles in the cathode rays.

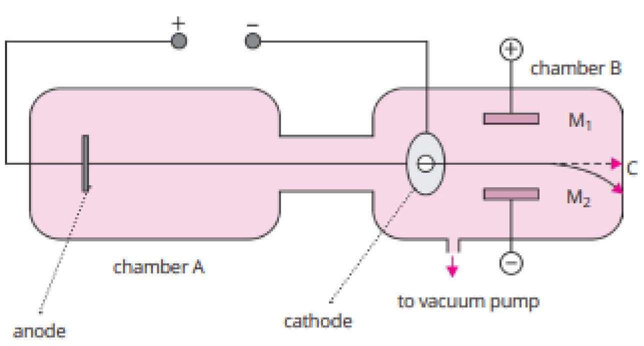
6. Anode rays expose photographic plates and films.

7. Anode rays produce fluorescence in the glass walls of the discharge tube, inside of which, is coated with a fluorescent substance.

8. Anode rays can produce mechanical effect.

Charge and mass of anode rays

Wilhelm Wien studied the deflection of anode rays by electric field. He modified Goldstein's discharge tube and made only one small circular hole at the centre of the cathode. The discharge tube experiment using this system produced only one small fluorescent spot at C (Fig.). Two metallic plates M_1 and M_2 were introduced in the chamber. On application of a strong electric field (2000 V), there occurred a deflection of the fluorescent spot towards the negative electrode. The e/m ratio for anode rays were determined by following Thomson's technique for the determination of *e*/*m* of cathode rays.



Wien's experiment for deflection of anode rays by electric field

DISCOVERY OF PROTON

For the discovery of proton, hydrogen gas was used in the anode ray experiment of Goldstein. The electrons formed at the cathode ionised the hydrogen atoms in the discharge tube.

$$\underset{\text{Hydrogen atom}}{H} \longrightarrow \underset{\text{Proton}}{H^+} + \underset{\text{Electron}}{e^-}$$

The extent of deflection is proportional to the ratio of mass to charge (*m*/*Z*) of hydrogen ion. Since for hydrogen ion, the charge (*Z*) is + 1, *m*/*Z* is simply the mass of hydrogen ion. When other gases were used, it was found that the masses of the atoms were integral multiples of the mass of a hydrogen atom (the smallest atom). The particle in the positive rays when hydrogen gas was used in the discharge tube, was given the name proton (Greek: *protos* = first). Thus, the proton is a hydrogen ion (H^+). Its charge was found to be + 1.602 × 10^{-19} C which is equal in magnitude but opposite in sign to that of an electron. The relative charge of proton is regarded as one unit positive charge (+1).Since the masses of atoms of other gases were integral multiples of mass of one hydrogen atom, the atoms of other gases contain two or more protons. It is concluded that proton is a positively charged particle found in the atoms of all elements. Elements are defined by the number of protons they possess.

Mass of proton

Mass of proton is calculated from the value of e/m and e as follows:

$$\text{Mass of proton } (m) = \frac{\text{Charge}}{\text{Charge/Mass}} = \frac{e}{e/m}$$

$$= \frac{1.602 \times 10^{-19}\ \text{C}}{9.578 \times 10^{4}\ \text{C g}^{-1}} = 1.673 \times 10^{-24}\ \text{g}$$

We know that the mass of an electron is 9.109×10^{-28} g. Hence,

$$\frac{\text{Mass of proton}}{\text{Mass of electron}} = \frac{1.673 \times 10^{-24}\ \text{g}}{9.109 \times 10^{-28}\ \text{g}} = 1837$$

Thus, the mass of a proton is about 1837 times that of an electron. The mass of a proton in the unit of u is 1.0076 u.

A proton is defined as a subatomic particle having mass equal to that of a hydrogen atom (1 u) and carrying one unit of positive charge. Proton is a stable particle and it is stable outside the nucleus of an atom. Proton is considered as a fundamental particle because it is present in all matters.

An atom is composed of two types of particles such as negatively charged electrons and positively charged protons. Since the atom as a whole is neutral, an atom contains equal number of protons and electrons. Hence, the proposal of Dalton that an atom is indivisible is not correct. An atom is not indivisible but consists of smaller fundamental particles.

DISCOVERY OF NEUTRON

In 1920, Rutherford observed that the atomic masses of all elements (except hydrogen) are much higher than that expected on the basis of the presence of only protons and electrons in an atom. Rutherford predicted the presence of another type of particle which must be electrically neutral and has a mass almost equal to that of a proton. When a sheet of lighter metal such as beryllium, lithium or boron was bombarded with high energy α-particles, there was emission of a highly penetrating radiation. When this radiation is allowed to strike a paraffin block, protons having very high velocity are released. In 1932, Chadwick showed that the highly penetrating radiation, produced by lighter metals, was due to neutral particles of approximately unit mass but little heavier than proton. Chadwick named these particles as neutrons (*n*). Since neutron has charge 0 and mass 1 u, it is represented as $^{1}_{0}n$. The discovery of neutron became rather late (1932) because neutron is chargeless (Italian: *neutro* = neutral). The neutron is generated by the following nuclear reaction:

$$^{9}_{4}\mathrm{Be} + {}^{4}_{2}\mathrm{He} \longrightarrow {}^{12}_{6}\mathrm{C} + {}^{1}_{0}n$$

Neutron is defined as a subatomic particle having mass equal to that of a hydrogen atom (1 u) carrying zero charge.

Characteristics of a neutron

1. Neutron is an electrically neutral particle having no electrical charge.
2. Neutron has a mass of 1.676×10^{-24} g which is slightly higher than that of proton (1.673×10^{-24} g). The mass of neutron in the atomic mass unit is 1.0089 u.

Neutron is a stable particle inside the nucleus but it is unstable outside the nucleus.

Other particles in the nucleus: Nucleus consists of several other unstable particles such as mesons, positrons, neutrinos, antineutrinos, gluons, quarks, antiquarks and Higgs Bosons. Protons and neutrons are made up of various combinations of smaller elementary particles called quarks having fractional charges and properties such as colour and 'flavours' (six varieties such as up, down, strange, charm, top and bottom). Quarks are about 1000 times smaller than proton. Quarks are always found in pairs. Quarks are bound by gluons. Another elementary particle called Higgs Boson (God particle) has recently been discovered (2012).

Comparison among proton, neutron and electron

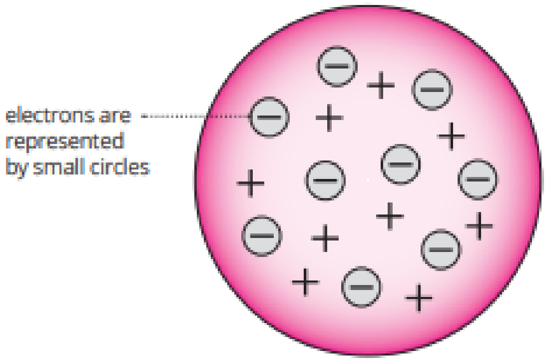
Fundamental particle	Mass in g	Mass in u	Relative charge	Location in the atom
Proton	1.673×10^{-24} g	1 u	+ 1	Nucleus
Neutron	1.676×10^{-24} g	1 u	0	Nucleus
Electron	9.109×10^{-28} g	$\frac{1}{1837}$ u	– 1	Outside nucleus

STRUCTURE OF THE ATOM

J J Thomson's model of an atom

In 1898, Thomson proposed a model for the structure of an atom on the basis of his discovery of the composition of the atom. This atomic model is called Thomson's plum-pudding model of atom. According to this model, an atom is considered to be a sphere (radius = 10^{-8} cm) of uniform positive charge into which the negatively charged electrons are embedded just like raisins are embedded in a plum-pudding (Fig.). An important feature of the plum-pudding model is that the mass of atom is considered to be evenly spread over the atom.

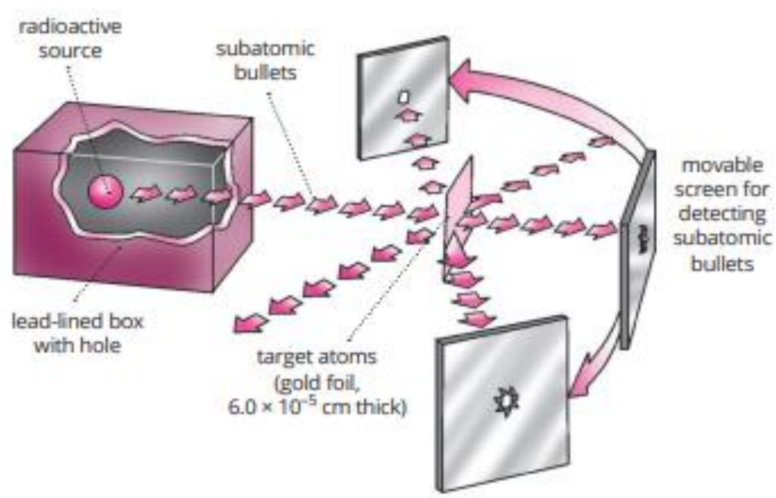
Thomson's model could not explain the results of a-particle scattering experiment carried out by Rutherford. Since both Dalton's atomic theory and Thomson's plum-pudding model of atom could not explain many experimental facts, these have been abandoned.



Thomson's plum-pudding model of atom

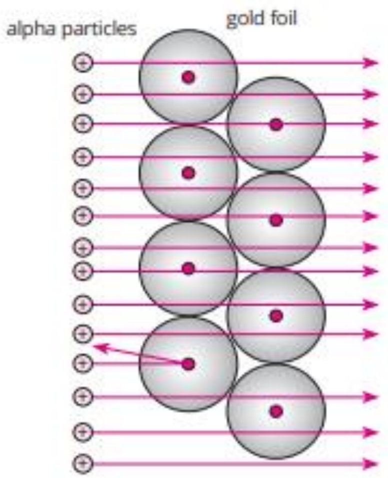
Rutherford's alpha-particle scattering experiment

In 1911, Rutherford performed an experiment which is now called Rutherford's alpha-particle scattering experiment. In this experiment, a stream of high energy a-particles from a radioactive source (radium, an α-particle emitter) was allowed to bombard a thin gold foil (thickness = 6.0×10^{-5} cm) (Fig.). A circular fluorescent screen coated with zinc(II) sulphide was set up around the gold foil. A tiny flash of light called scintillation was produced whenever an α-particle struck the zinc(II) sulphide-coated screen. The scintillations were observed with the help of a movable microscope.



Rutherford's gold foil experiment

The following observations were made from the alpha-particle scattering experiment:
1. Most of the a-particles passed through the gold foil without any deflection (Fig.).
2. A small fraction of the α-particles was deflected through small angles.
3. Only a very few α-particles (1 in 10000) after head on collision with a positive body returned back suffering a deflection of 180°.



Scattering of α-particles by a thin gold foil

In the light of the above experimental results, Rutherford drew the following conclusions:
1. Since most of the α-particles passed through the gold foil without any deflection, the major part of space in an atom is empty.
2. Since very few α-particles are deflected by small angles, the deflection is due to enormous repulsive force between positively charged α-particles and some positive body present within the atom. The α-particles coming close to this positive body got deflected by small angles. This positively charged body inside the atom was named as nucleus.
3. Since some α-particles were deflected back at 180° after head-on collision with this positive body and α-particles are heavy particles, the positive body must be heavy.

4. Since only a very few a-particles got deflected back, the heavy positive body occupies a very small volume as compared to the total volume of the atom. The radius of the nucleus is around 10^{-13} cm while that of the atom is 10^{-8} cm. That means if the radius of the nucleus is 1 cm, then the radius of the atom would be about 1 km.
5. Since the a-particles have appreciable mass and they are deflected by the nucleus, the nucleus of an atom has an appreciable mass.

Based on these conclusions Rutherford gave the following model of an atom:
1. An atom consists of a tiny positively charged nucleus surrounded by negatively charged electrons. The positive charge of the nucleus is due to protons.
2. The nucleus and electrons are held together by coulombic force of attraction.
3. The volume of the nucleus is extremely small as compared to the total volume of atom.
4. Most of the mass of atom is concentrated in the nucleus. The mass of the nucleus is due to protons and neutral particles (called later as neutrons by Chadwick and Rutherford in 1932) having mass almost equal to the mass of proton.
5. The number of electrons in an atom is equal to the number of protons in it.

Drawbacks of Rutherford’s model of atom

1. It does not say anything about the distribution of electrons around the nucleus and the energy of electrons.
2. Due to change in direction, a moving body undergoes acceleration. A charged particle on acceleration should emit electromagnetic radiation. Hence, an electron in an orbit is expected to emit radiation, the energy carried by radiation coming from motion of electron and the orbit is expected to shrink continuously. Thus, the electron should follow a spiral path and ultimately fall into the nucleus. But this does not happen. Thus, the stability of an atom is not explained.

Present concept of atom – Bohr's model of atom

In 1913, Niels Bohr, a Danish physicist, proposed a model to explain the structure of an atom. The postulates of Bohr's atomic model are:

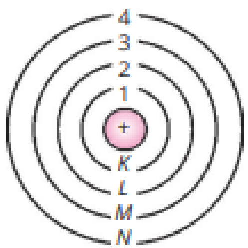
1. In an atom, the electrons revolve around the nucleus in certain definite circular paths called orbits. Such orbits differ from each other in their radii. Only certain discrete orbits of electrons are allowed inside the atom.

2. Each orbit has a definite energy. These orbits are known as energy levels or energy shells. The orbits or energy shells are numbered as 1, 2, 3, 4, … or *K*, *L*, *M*, *N*, ... shells starting from the nucleus. The integers 1, 2, 3, 4, …are called principal quantum numbers (*n*). The energy shell nearest to the nucleus has minimum energy and the energy shell farthest from the nucleus has maximum energy.

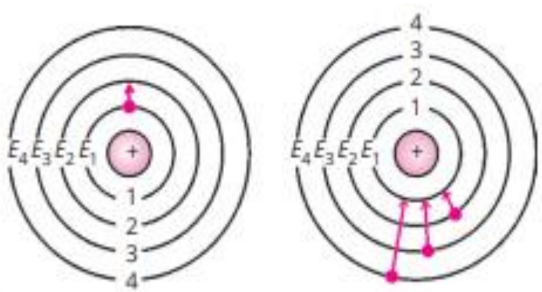
3. As long as an electron remains in a particular orbit, it does not radiate energy.

4. An electron loses energy when it jumps from an orbit of higher energy (E_2) to an orbit of lower energy (E_1) and energy equal to $E_2 – E_1$ is given out in the form of electromagnetic radiation. An electron gains energy from outside when it jumps from an orbit of lower energy (E_1) to an orbit of higher energy (E_2). The change in energy, ΔE is given by: $\Delta E = E_2 – E_1 = h\nu$, where h is Planck's constant and ν is the frequency of radiation emitted or absorbed.

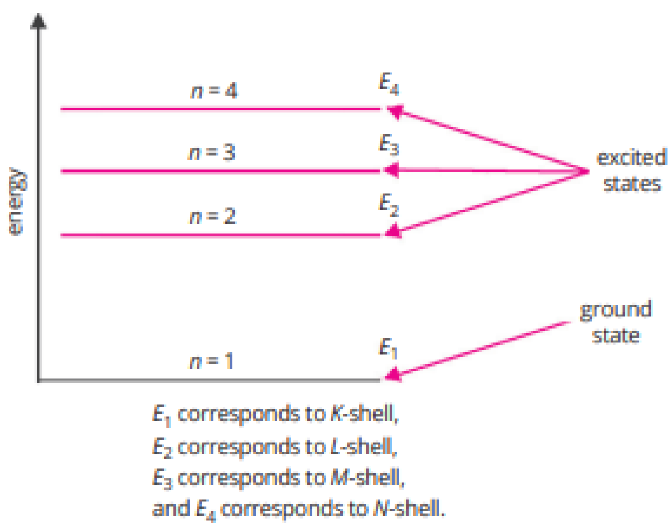
The lowest energy state (E_1) is called ground state and the higher energy states (E_2, E_3, E_4, …) are called excited states.



A few energy shells around the nucleus of an atom



Change in the energy of an electron when electron jumps from one energy shell to other ($E_1 < E_2 < E_3$) [$v_1 = (E_3 – E_2) < v_2 = (E_2 – E_1)$]



Energy of energy shells of hydrogen atom ($E_1 < E_2 < E_3 < E_4$)

Arrangement of electrons in different shells: Bohr–Bury Scheme

The arrangement of electrons in various shells of an atom of an element is known as electronic configuration of the element. We should know the following rules for writing down the electronic configuration of an element:

1. Number of electrons present in one atom of the element.
2. Maximum number of electrons which can be accommodated in different energy shells (electron orbits) of the atom.

The distribution of electrons in different energy shells of an atom is governed by a scheme known as Bohr–Bury Scheme. According to this scheme:

a. The electron orbits are designated by the number n where n = 1, 2, 3, 4, ... The maximum number of electrons which can be accommodated in an energy shell (orbit) is given by $2n^2$ where n stands for the number of the orbit.

i. For first energy shell, $n = 1$. The maximum number of electrons in the first energy shell $= 2n^2 = 2 \times (1)^2 = 2$.

ii. For second energy shell, $n = 2$. The maximum number of electrons in the second energy shell $= 2n^2 = 2 \times (2)^2 = 8$.

iii. For third energy shell, $n = 3$. The maximum number of electrons in the third energy shell $= 2n^2 = 2 \times (3)^2 = 18$.

iv. For fourth energy shell, $n = 4$. The maximum number of electrons in the fourth energy shell $= 2n^2 = 2 \times (4)^2 = 32$.

Orbit number (n)	Orbit designation	Maximum number of electrons ($2n^2$)
1	K	$2 \times 1^2 = 2$
2	L	$2 \times 2^2 = 8$
3	M	$2 \times 3^2 = 18$
4	N	$2 \times 4^2 = 32$

Electron orbits and maximum number of electrons in various electron orbits

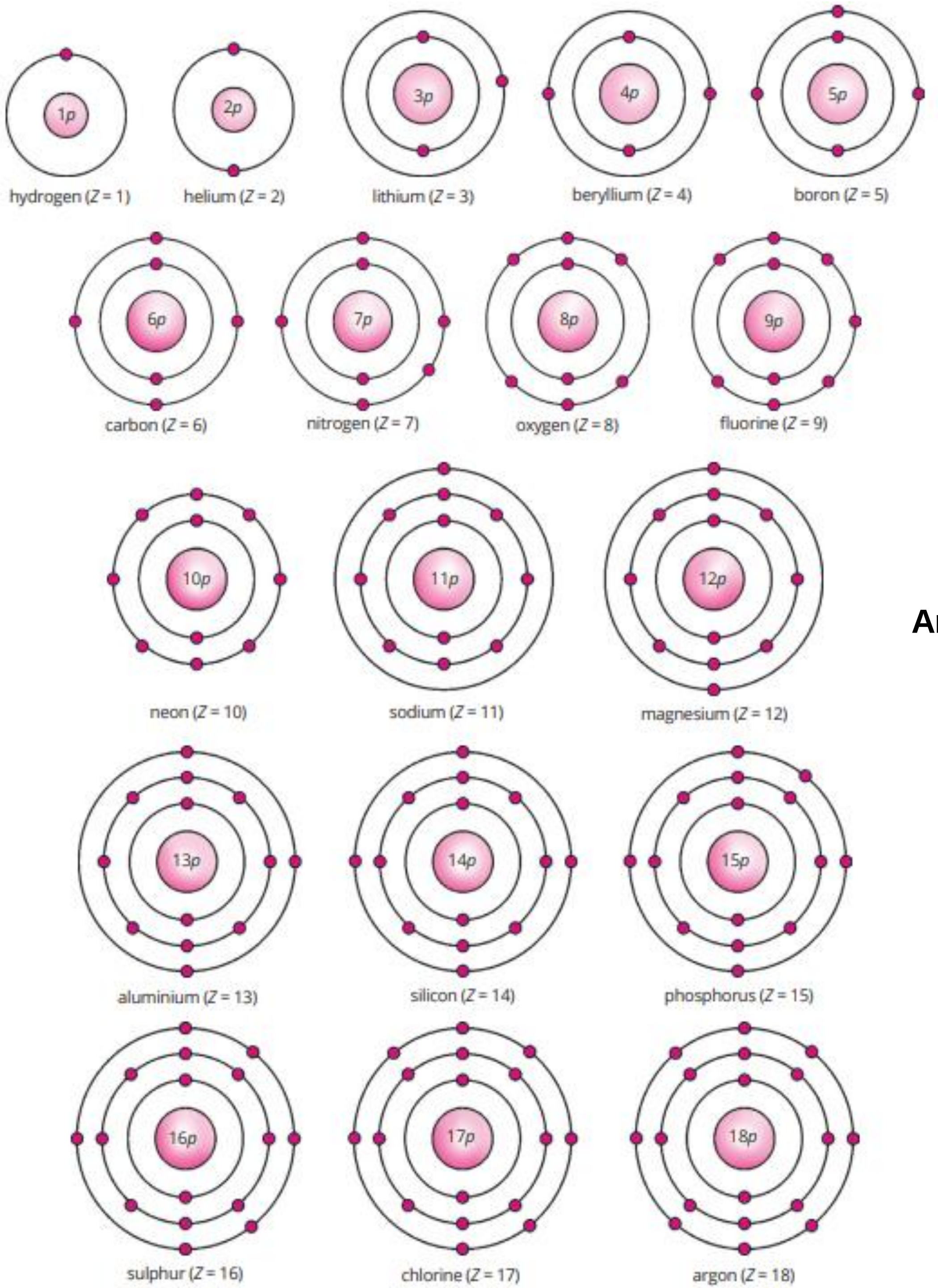
b. The outermost shell of an atom cannot accommodate more than eight electrons, even if it has the capacity to accommodate more electrons. For example, when *N* shell is the outermost shell of an atom, it can accommodate maximum of eight electrons only, although *N* shell has the capacity to accommodate a maximum of 32 electrons.
This is because the presence of eight electrons in the outermost shell makes the atom very stable. Note that when the outermost shell is *K* shell, the maximum number of electrons which can be accommodated is only two.

c. It is not always necessary to fill up an orbit completely before starting the next higher orbit. For filling of electrons in various shells, we have to also remember the following:

The filling of electrons in the second shell (n = 2) begins only after the first shell is filled with two electrons. The filling of electrons in the third shell (n = 3) begins only after the second shell is filled with eight electrons. The filling of the fourth shell (n = 4) begins even before the third shell is completely filled.

Electronic configurations of first twenty elements

Element	Symbol	Atomic number	Electronic distribution			
			Orbit No. 1	2	3	4
			Shell *K*	*L*	*M*	*N*
Hydrogen	H	1	1			
Helium	He	2	2			
Lithium	Li	3	2	1		
Beryllium	Be	4	2	2		
Boron	B	5	2	3		
Carbon	C	6	2	4		
Nitrogen	N	7	2	5		
Oxygen	O	8	2	6		
Fluorine	F	9	2	7		
Neon	Ne	10	2	8		
Sodium	Na	11	2	8	1	
Magnesium	Mg	12	2	8	2	
Aluminium	Al	13	2	8	3	
Silicon	Si	14	2	8	4	
Phosphorus	P	15	2	8	5	
Sulphur	S	16	2	8	6	
Chlorine	Cl	17	2	8	7	
Argon	Ar	18	2	8	8	
Potassium	K	19	2	8	8	1
Calcium	Ca	20	2	8	8	2



Arrangements of electrons in the first 18 elements

ATOMIC NUMBER AND MASS NUMBER

Atomic number

The atomic number of an element is equal to the number of protons present in the nucleus of an atom of that element. It is represented by *Z*.

Atomic number of an element = Number of protons in the nucleus of its atom

Since an atom is electrically neutral, the total number of electrons in an atom is also equal to its atomic number. In a chemical reaction, the electrons take part while the protons do not take part. Hence, during a chemical reaction, the number of protons remains the same while the number of electrons changes. Thus, the atomic number of an element does not change during a chemical reaction.

Mass number

The sum of the number of protons and neutrons in the atom of an element is called its mass number. It is represented by *A*. Hence,

$$\text{Mass number } (A) = \text{Number of protons } (p) + \text{Number of neutrons } (n)$$

or $$A = p + n$$

But the atomic number (Z) of an element is equal to the number of protons in the atom of the element.

Hence, $$A = Z + n$$

$$\text{Mass number} = \text{Atomic number} + \text{Number of neutrons}$$

Hence,

Number of neutrons (n)

$$= \text{Mass number } (A) - \text{Atomic number } (Z) = A - Z$$

An atom of an element X having mass number *A* and atomic number *Z* is represented as

$$^{A}_{Z}X$$

The protons and neutrons present in the nucleus of an atom are collectively called nucleons.

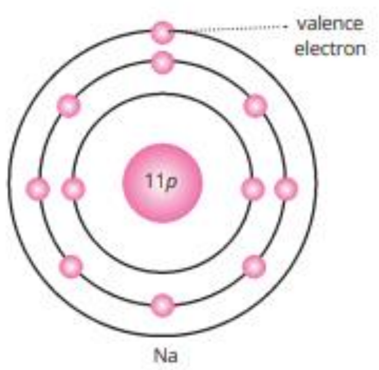
First 20 elements – atomic number, mass number, number of protons, neutrons and electrons

Name	Symbol	Atomic number (*Z*)	Mass number (*A*)	Number of protons (*P*)	Number of neutrons (*N*)	Number of electrons (*E*)
Hydrogen	$^{1}_{1}H$	1	1	1	0	1
Helium	$^{4}_{2}He$	2	4	2	2	2
Lithium	$^{7}_{3}Li$	3	7	3	4	3
Beryllium	$^{9}_{4}Be$	4	9	4	5	4
Boron	$^{11}_{5}B$	5	11	5	6	5
Carbon	$^{12}_{6}C$	6	12	6	6	6
Nitrogen	$^{14}_{7}N$	7	14	7	7	7
Oxygen	$^{16}_{8}O$	8	16	8	8	8
Fluorine	$^{19}_{9}F$	9	19	9	10	9
Neon	$^{20}_{10}Ne$	10	20	10	10	10
Sodium	$^{23}_{11}Na$	11	23	11	12	11
Magnesium	$^{24}_{12}Mg$	12	24	12	12	12
Aluminium	$^{27}_{13}Al$	13	27	13	14	13
Silicon	$^{28}_{14}Si$	14	28	14	14	14
Phosphorus	$^{31}_{15}P$	15	31	15	16	15
Sulphur	$^{32}_{16}S$	16	32	16	16	16
Chlorine	$^{35}_{17}Cl$	17	35	17	18	17
Argon	$^{40}_{18}Ar$	18	40	18	22	18
Potassium	$^{39}_{19}K$	19	39	19	20	19
Calcium	$^{40}_{20}Ca$	20	40	20	20	20

VALENCE ELECTRONS AND VALENCY

Valence electrons

The electrons present in the outermost shell of an atom are called valence electrons. The valence electrons are also responsible for chemical reactions of an element. The electronic configuration of sodium atom is 2, 8, 1 (Fig.). Only one electron is present in the outermost shell (*M*-shell) or the valence shell of sodium atom. Hence, the number of valence electrons in sodium is 1.



Electronic configuration of sodium atom (Z = 11)

Valency

The number of valence electrons of an element which actually takes part in chemical reactions is called the valency of that element. Valency of an element may also be defined as follows:

The number of hydrogen atoms or chlorine atoms or twice the number of oxygen atoms which combine with one atom of an element is called its valency

For example, in forming water, one atom of oxygen combines with two atoms of hydrogen. Hence, the valency of oxygen (O) in water (H_2O) is two. The valency of aluminium in its oxide Al_2O_3 is found out as follows:

Number of oxygen atoms combining with two atoms of Al = 3

Number of oxygen atoms combining with one atom of Al = $\frac{3}{2}$

Hence, the valency of Al (as per definition it is twice the number of oxygen atoms) = $2 \times \frac{3}{2} = 3$.

Element	Atomic number	Electronic configuration			No. of valence electrons	Valency
		K	*L*	*M*		
Li	3	2	1		1	1
Be	4	2	2		2	2
B	5	2	3		3	3
C	6	2	4		4	4
N	7	2	5		5	3
O	8	2	6		6	2
F	9	2	7		7	1
Na	11	2	8	1	1	1
Mg	12	2	8	2	2	2
Al	13	2	8	3	3	3
Si	14	2	8	4	4	4
P	15	2	8	5	5	3, 5
S	16	2	8	6	6	2
Cl	17	2	8	7	7	1
Ar	18	2	8	8	8	0

Number of valence electrons and valency of some common elements

In general, the valency of an element is equal to the number of valence electrons or is equal to eight minus the number of valence electrons. For example, hydrogen, sodium, potassium, rubidium and caesium contain one electron each in their outermost shell (valence shell) and hence, they exhibit the valency of 1. Similarly, magnesium, aluminium and carbon having 2, 3 and 4 valence electrons, respectively exhibit the valency of 2, 3 and 4, respectively. When the number of valence electrons in an atom is close to its maximum capacity (8, octet), the valency of the element is obtained by subtracting the number of valence electrons from the octet.

For example, number of valence electrons in fluorine (F) is 7 and hence, the valency of fluorine is 8 – 7 = 1. The elements having the completely filled outermost shell exhibit the valency of zero. Thus, the elements having the completely filled outermost shell are chemically inert and they exhibit little chemical reactivity. For example, helium, neon and argon have completely filled outermost shell and they are chemically inert. Since these elements do not exhibit chemical reactivity, they are called noble gases (earlier name inert gases).

To attain stability, atoms of elements gain, lose or share electrons with other atoms so as to complete their octet. Hence, valency is defined as the number of electrons gained, lost or shared by the atom of an element so as to complete the octet of electrons in the outermost shell. The elements having one or seven electrons in their valence shell exhibit greater chemical reactivity.

ISOTOPES AND ISOBARS

Isotopes

Isotopes were discovered by Soddy in 1913. Atoms of the same element which have the same atomic number (hence same number of protons) but different mass numbers (because of presence of different number of neutrons) are called isotopes [Greek: *isos* = equal, *topos* = place (i.e. in the periodic table of elements)].

These may also be defined as follows:
Atoms of the same element which have same number of protons but different number of neutrons inside their nuclei are called isotopes.

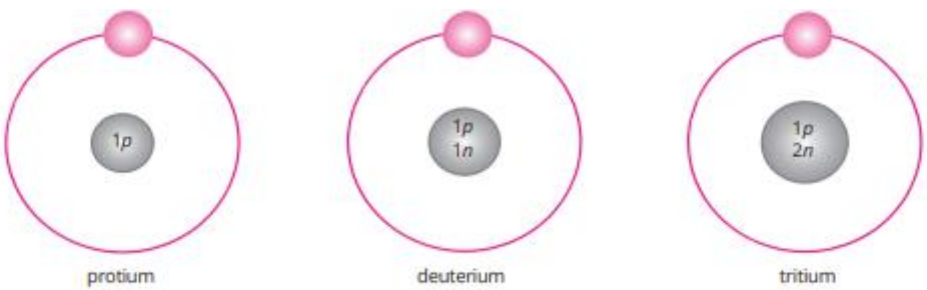
Since the chemical properties of an element are dependent on its number of electrons, and the isotopes have same number of electrons because of same atomic number, the isotopes have similar chemical properties. Since the mass numbers of isotopes are different because of the presence of different number of neutrons, the physical properties like density, melting point and boiling point, etc. which depend on the atomic mass, show variation.

Isotopes of hydrogen

There are three isotopes of hydrogen: protium, deuterium and tritium (Fig.). They are represented as:

$^{1}_{1}H$	$^{2}_{1}H$	$^{3}_{1}H$
Protium	Deuterium	Tritium

Isotopes of hydrogen



Isotopes of hydrogen

Name of the isotope	Symbol of the isotope	Mass number (*A*)	Number of protons	Number of neutrons	Number of electrons
Hydrogen (Protium)	$^{1}_{1}H$ (H)	1	1	0	1
Deuterium	$^{2}_{1}H$ (D)	2	1	1	1
Tritium	$^{3}_{1}H$ (T)	3	1	2	1

Isotopes of carbon

There are three isotopes of carbon. They are represented as:

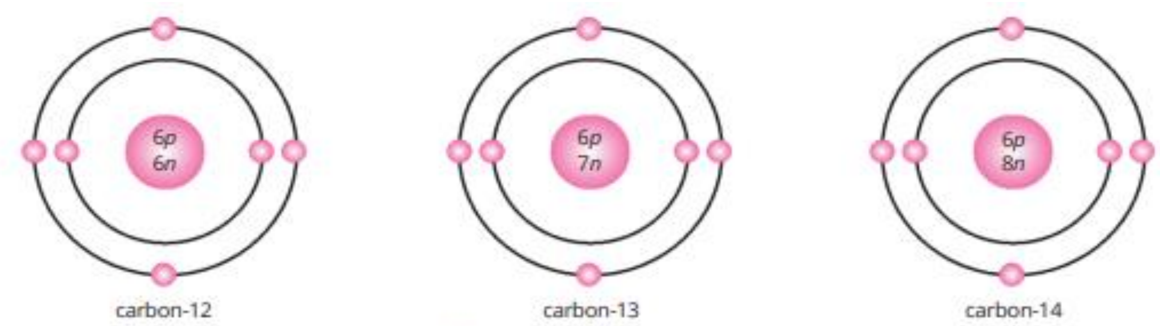
$$^{12}_{6}C,\ ^{13}_{6}C,\ ^{14}_{6}C$$

The relative natural abundance of isotopes of carbon is: carbon-12: 98.892%, carbon-13: 1.108%, carbon-14: negligible.

Isotopes of carbon

Name of the isotope	Symbol of the isotope	Mass number (*A*)	Number of protons	Number of neutrons	Number of electrons
Carbon-12	$^{12}_{6}C$	12	6	6	6
Carbon-13	$^{13}_{6}C$	13	6	7	6
Carbon-14	$^{14}_{6}C$	14	6	8	6

Isotopes of carbon



Characteristics of isotopes

1. The isotopes of an element have the same number of protons and electrons but different number of neutrons.
2. The isotopes of an element have different mass numbers and therefore, different masses.
3. The isotopes of an element possess the same electronic configuration and the same number of valence electrons and exhibit the same chemical properties.
4. The isotopes of an element exhibit different physical properties such as density, melting point, boiling point, etc.

Applications of isotopes

1. **Medicinal use:** Isotopes are used for the treatment of various diseases. For example,
   a. Isotopes $^{60}_{27}Co$, $^{225}_{88}Ra$ and $^{198}_{79}Au$ are used in the treatment of cancer.
   b. Isotope $^{32}_{15}P$ is used for locating cancer. It is used for the treatment of blood cancer (leukaemia). It is also used in patients suffering from bone fracture to find absorption of phosphorus in their bones.
   c. Isotopes $^{73}_{33}As$ and $^{131}_{53}I$ are used for locating brain tumor.
   d. Isotope $^{131}_{53}I$ is used for the detection of thyroid disorder and its treatment.
   e. Isotope $^{24}_{11}Na$ is used for detection of blood clots in the circulatory system.
   f. Isotope $^{59}_{26}Fe$ is used for the diagnosis of anaemia.

Tracer technique: During the photosynthesis in the green leaves of plants, CO_2 and H_2O are consumed and glucose ($C_6H_{12}O_6$) and O_2 are formed. The oxygen produced may originate from H_2O or CO_2 or both. By exchanging the ^{16}O atoms of CO_2 with ^{18}O, it has been established that O_2 produced in the reaction comes from H_2O and not from CO_2, since O_2 produced does not contain ^{18}O. Calvin was awarded 1961 Nobel Prize in Chemistry for his work on this mechanism of photosynthesis.

Radiocarbon dating: The method of estimating the age of dead objects containing carbon such as fossils and pieces of wood by measuring the amount of the $^{14}_{6}C$ isotope in the dead object relative to that of the $^{14}_{6}C$ isotope in a living object is called **radiocarbon dating**. The method was discovered by Libby and he was awarded the Nobel Prize in Chemistry for his work in developing the technique of radiocarbon dating. The $^{14}_{6}C$ isotope formed in the upper atmosphere is converted into $^{14}CO_2$ which is incorporated into plants during photosynthesis and is then assimilated by animals which eat plants. When a tree or animal dies, the uptake of $^{14}CO_2$ by it ceases and the level of $^{14}_{6}C$ in the dead tree decreases progressively due to decay of $^{14}_{6}C$ isotope as per the reaction below:

$$^{14}_{6}C \longrightarrow \, ^{14}_{7}N + \underset{(\beta \text{ particle})}{^{0}_{-1}e^{-}}$$

Hence, an old sample of dead tree has less radioactivity than that in a living tree. The age of dead tree is found out by comparing its radioactivity with the radioactivity present in the living tree. For example, if the dead tree has only 50% carbon-14 radioactivity as compared to that in a living tree, then its age is 5760 years which is the half-life period of $^{14}_{6}C$ isotope [This means that in 5760 years, the concentration of $^{14}_{6}C$ is lowered to half (50%) of its initial concentration and after another 5760 years, the concentration gets lowered to 25% (50% of 50%) of the initial concentration.]. If the dead tree has only 25% carbon-14 radioactivity as compared to that in a living tree, its age is $5760 \times 2 = 11520$ years (please note that by age here we mean the number of years ago when the plant should have died). This method is extensively used for finding the age of archaeological samples and objects of historical importance. This technique is also used to determine the age of wine, glaciers and snowfields in which the radioactive level of tritium ($^{3}_{1}H$) is measured.

Agricultural use: The radioisotope $^{32}_{15}P$ is used to monitor the uptake of phosphorus by plants. This is done by incorporating $^{32}_{15}P$ in fertilisers and measuring the radioactivity in the plants. This has led to the improvement in the preparation of fertilisers. The radiations from radioisotopes are used to develop disease free, high yielding seeds. The radiations from radioisotopes are used for preservation of foodstuffs such as fruits, potatoes, onions, fish, meat, etc. and for disinfecting food grains.

Generation of cheaper electricity in nuclear reactors: The isotopes $^{235}_{92}U$ and $^{239}_{94}Pu$ are used as fuel in nuclear reactors for generation of cheaper electricity by a cleaner method using controlled nuclear reactions. The large amount of heat energy generated in these reactions is used to convert water into steam which drives the turbines of a generator to generate electricity.

Industrial use: Isotope $^{32}_{15}P$ is used in the manufacturing of steel from cast iron in order to find out the complete removal of phosphorus from steel. For this purpose the cast iron containing a small quantity of $^{32}_{15}P$ isotope is used. The disappearance of radioactivity in the molten steel indicates the complete removal of phosphorus from the steel. Radioactive isotopes are used to detect minor cracks in the underground gas-pipelines, oil-pipelines and water-pipelines. A solution of a radioactive substance is introduced in the pipeline. If there is minor crack in the pipeline, a high level of radiations will be detected at the place of the crack. The wear and tear of engines are found out by incorporating a radioactive isotope of a metal in the piston and measuring the radioactivity of the lubricating oil at various intervals of time. The appearance of radioactivity in the lubricating oil gives an indication that the wear and tear of the piston have started.

Reason for the fractional atomic masses of elements

Most elements have fractional atomic masses and not whole numbers. Even atoms such as H (1.0079), C (12.011), N (14.0067), O (16.9994), F (18.9984) have fractional atomic masses. But deviation of atomic mass of these elements from whole number is very small. In some elements, the deviation of atomic mass from whole number is quite significant. For example, Cl (35.5) and Cu (63.5). The fractional atomic masses of elements are due to the existence of their various isotopes having different masses. The atomic mass of an element is the average of the relative masses of all the naturally occurring isotopes of the element.

Isobars

The atoms of different elements with different number of protons (i.e. different atomic numbers) but equal sum of the number of protons and neutrons (i.e. same mass number) are called isobars (Greek: *isos* = equal, *baros* = weight). Isobars have different physical and chemical properties. Example of isobars is:

$$^{14}_{6}C,\ ^{14}_{7}N.$$

On Board! BOOKS

SUMMARY

1. A discharge tube is a cylindrical glass tube fitted with an anode and a cathode and containing a gas at a very low pressure through which electric current of a very high voltage (10000–15000 V) is passed.
2. **Cathode rays:** When a high voltage is passed through a discharge tube containing air at a very low pressure (0.001 mm of Hg), discharge tube turns dark and the glass wall of the discharge tube opposite to the cathode starts to glow with a faint greenish light. The green glow is due to invisible rays coming out from the cathode. These rays are called **cathode rays**.
3. Properties of cathode rays: **a.** Cathode rays travel in straight lines. **b.** Cathode rays produce mechanical effect. **c.** Cathode rays are deflected by electric and magnetic fields. **d.** Cathode rays consist of negatively charged electrons. **e.** Cathode rays expose photographic plates.
4. The *e/m* ratio of cathode rays is 1.759×10^{11} C kg^{-1}.
5. The electrical charge on an electron is -1.602×10^{-19} C. The mass of electron is 9.109×10^{-31} kg. The mass of electron is 1/1837 times that of one hydrogen atom. The mass of electron in atomic mass unit is 0.0005487 u.
6. **Electron:** Electron is an essential constituent of atoms of all elements and matter.
7. An electron is defined as a subatomic fundamental particle having mass equal to (1/1837)th of that of a hydrogen atom and carrying one unit of negative charge. Energy of electron is negative. Electron in stable outside an atom. Electron was discovered by JJ Thomson.
8. **Anode rays:** In the discharge tube, when a perforated cathode was used, positively charged rays called **anode rays** emerged from the anode.
9. Anode rays are formed due to the ionisation of atoms or molecules of the gas taken in the discharge tube.
10. The positively charged particles in the anode rays are heavier than electron.
11. The anode rays get deflected in electric and magnetic fields.
12. The nature of anode rays depends upon the nature of the gas taken in the discharge tube.
13. The electrical charge on the particles of anode rays is an integral multiple of the magnitude of the charge on an electron (1.602×10^{-19} C) but with a positive sign.
14. **Proton:** Proton is a subatomic fundamental particle having mass equal to that of hydrogen atom and carrying one unit of positive charge. Proton is stable inside and outside the nucleus. Proton was discovered by E Goldstein.
15. **Neutron:** Neutron is a subatomic electrically neutral fundamental particle having mass equal to that of a proton. When a sheet of lighter metals such as Be, Li or B is bombarded with high energy α-particles, there occurs emission of a highly penetrating radiation containing electrically neutral particles called **neutrons**. Neutron was discovered by Chadwick. Neutron is unstable outside the nucleus.

18. **Bohr's model of atom:** Electrons are distributed in different shells with discrete energy around the nucleus, while revolving in discrete orbits the electrons do not radiate energy.

19. **Orbits:** In an atom, the electrons revolve around the nucleus in certain definite circular paths called **orbits**. Orbits are also known as **energy levels** or **energy shells**. The orbits or energy shells are numbered as 1, 2, 3, 4, ... or *K*, *L*, *M*, *N*, ... shells starting from the nucleus.

20. **Atomic number (*Z*):** Atomic number of an element is equal to the number of protons present in the nucleus of an atom.

21. **Mass number (*A*):** The sum of the number of protons and neutrons in the atom of an element is called its **mass number**.

    Mass number = Number of protons + Number of neutrons.

22. **Nucleons:** The protons and neutrons present in the nucleus of an atom are collectively called **nucleons**.

23. **Valence electrons:** The electrons present in the outermost shell of an atom are called **valence electrons**. The valence electrons dictate the valency and chemical reactions of an element.

24. **Valency:** The number of hydrogen atoms or chlorine atoms or twice the number of oxygen atoms which combines with one atom of an element is called its **valency**.

25. **Noble gases:** The elements having completely filled outermost shell, which exhibit little chemical reactivity are called **noble gases**.

26. **Isotopes:** The atoms of an element which have the same number of protons but different number of neutrons are called **isotopes**. Isotopes have the same atomic number but different mass numbers.

27. **Isobars:** The atoms of different elements which have the same mass number but different atomic numbers are called **isobars**.

28. **Radiocarbon dating:** The method of estimating the age of old carbon containing dead objects such as fossils and pieces of wood by measuring the radioactivity of the $^{14}_{6}C$ isotope in the dead object relative to that of the $^{14}_{6}C$ isotope in a living object is called **radiocarbon dating**.

29. Most elements have fractional atomic mass due to the existence of isotopes.

MIND MAP

FUNDAMENTAL (SUB-ATOMIC) PARTICLES OF AN ATOM

Fundamental particle	Mass in g	Mass in u	Relative charge	Location in the atom
Proton	1.673×10^{-24} g	1 u	+ 1	Nucleus
Neutron	1.676×10^{-24} g	1 u	0	Nucleus
Electron	9.109×10^{-28} g	$\frac{1}{1837}$ u	– 1	Outside nucleus

STRUCTURE OF AN ATOM

Thomson's Model of Atom

According to this model :

- an atom is considered to be a sphere of uniform positive charge in which electrons are embedded
- an atom is electrically neutral

Rutherford's Model of Atom

According to this model :

- most of the mass of an atom resides in the nucleus (positively charged centre in atom)
- electrons revolve around the nucleus in circular paths
- size of the nucleus is very small as compared to the size of the atom

Bohr's Model of Atom

According to this model :

- only discrete orbits of electrons are allowed inside the atom
- while revolving in discrete orbits the electrons do not radiate energy

MIND MAP

BOHR-BURY SCHEME

Distribution of electrons in different orbits :

- maximum number of electrons in a shell is given by the formula $2n^2$
- outermost orbit (shell) cannot accommodate more than 8 electrons
- shells are filled in a step-wise manner only

VALENCE ELECTRONS AND VALENCY

- electrons present in the outermost shell of an atom are valence electrons
- a definite combining capacity of the atom of an element is its valency

ATOMIC NUMBER AND MASS NUMBER

- Atomic number (*Z*) of an element = Number of protons in the nucleus of its atom
- Mass number (*A*) = Number of protons (*p*) + Number of neutrons (*n*)
- $^{A}_{Z}X$ — A → Mass number; X → Symbol of an element; Z → Atomic number

ISOTOPES AND ISOBARS

- atoms of the same element, having same atomic number but different mass numbers are isotopes, e.g., $^{35}_{17}Cl$, $^{37}_{17}Cl$
- atoms of different elements with different atomic numbers but same mass number are isobars, e.g., $^{40}_{18}Ar$, $^{40}_{20}Ca$