3

Atomic Structure

In the previous lesson, you have studied that the atoms are the smallest constituents of matter. But what is the structure of an atom? Why are atoms of different elements different? Let us try to find out the answers to some of these questions in this lesson.

We will start the study of this lesson by recapitulating the postulates of Dalton's atomic theory .At that time, many Greek philosophers believed that the atoms cannot be further subdivided, i.e. they were structure less entities. But as you will study in this lesson, various developments such as the discoveries of sub-atomic particles such as electron, proton etc. led to the failure of this idea. Based on these discoveries, various atomic models were proposed by the scientists. In this lesson, we would discuss how various models for the structure of atom were developed and what were their main features. We would explain the success as well as the shortcomings of these models. These models tell us about the distribution of various sub-atomic particles in the atom. From the knowledge of structure of atom the arrangement of electrons around the nucleus can be obtained. This arrangement is known as electronic configuration. The electronic configurations of some simple elements are discussed in this lesson These electronic configurations would be useful in explaining various properties of the elements. The electronic configuration of an element governs the nature of chemical bonds formed by it. This aspect is dealt in lesson 5 on chemical bonding.

OBJECTIVES

After completing this lesson, you should be able to:

- state the reasons of failure of Dalton's atomic theory;
- name and list the fundamental particles present in the atom;
- recall the developments of various atomic models;
- list the shortcomings of Bohr's atomic model;
- compute the electronic configuration of first 18 elements.

3.1 FAILURE OF DALTON'S ATOMIC THEORY

You have read about Dalton's atomic theory in lesson 2. Dalton's theory explained various laws of chemical combination about which you have read earlier in lesson 2. At that time, the atom was considered to be indivisible. Later, certain experiments showed that an atom is made up of even smaller particles which are called **subatomic particles**. You will now study about the discovery of these subatomic particles namely electrons, protons and neutrons.

3.1.1 Discovery of electron

During 1890s' many scientists performed experiments using cathode ray tubes. A **cathode ray tube** is made of glass from which most of the air has been removed. Such a cathode ray tube has been shown in Fig. 3.1. You can see in the figure that there are two metal electrodes; the negatively charged electrode is called **cathode** whereas the positively charged electrode is called **anode**.

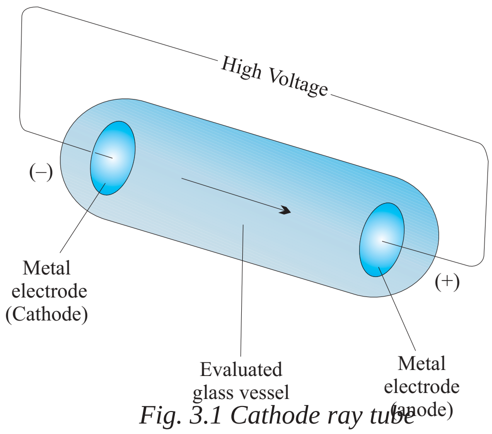

Fig. 3.1 Cathode ray tube

An English physicist J.J. Thomson studied electric discharge through a cathode ray tube. When high voltage was applied across the electrodes, the cathode emitted a stream of negatively charged particles, called **electrons.**

$$\text{mass} = \frac{\text{charge of the electron}}{\text{charge per unit mass of the electron}}$$

$$= \frac{e}{e/m} = \frac{1.60\times10^{-19}\ \text{C}}{1.76\times10^{8}\ \text{C g}^{-1}} = 9.10\times10^{-28}\ \text{g}$$

Since the electrons were released from the cathode irrespective of the metal used for it or irrespective of the gas filled in the cathode ray tube, Thomson concluded that ***all atoms must contain electrons***. **Robert Millikan** (1868-1953) received the Nobel prize in Physics in 1923 for determining the charge of the electron.

The discovery of the electron led to the conclusion that the atom was *no more indivisible* as was believed by Dalton and others. Hence, the idea of indivisibility of atom as suggested by Dalton was proved incorrect. In other words, the atom was found to be *divisible.*

If the atom was divisible, what were are its *constituents*? You have read above that one such particle is an electron. Now, what are the other particles present in an atom? Let us study the next section and find out the answer.

3.1.2 Discovery of proton

In 1886, **Eugen Goldstein** observed that rays flowing in a direction opposite to that of the cathode rays were positively charged. Such rays were named as **canal rays** because they passed through the holes or the canals present in the perforated cathode. In 1898, Wilhelm

Wien, a German physicist, measured *e*/*m* for canal rays. It was found that the particles constituting the canal rays are much heavier than electrons. Also unlike cathode rays, the nature and the type of these particles varied depending upon the gas present in the cathode ray tube. The canal rays had positive charges which were whole number multiples of the amount of charge present on the electron. The positive nature of the canal rays was explained as follows:

In a cathode ray tube, the electrons emitted from the cathode collide with the atoms of the gas present in the tube and knock out one or more electrons present in them. This leaves behind positive ions which travel towards the cathode. If the cathode has holes in it ,then these positive ions can pass through these holes or canals. Hence, they are called the *canal rays*. The canal rays are shown in Fig. 3.2.

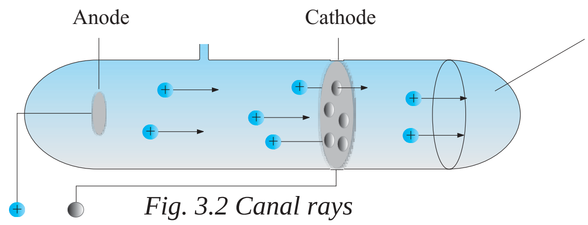

Fig. 3.2 Canal rays

When the cathode ray tube contained hydrogen gas, the particles of the canal rays obtained were the lightest and their *e/m* ratio was the highest. Rutherford showed that these particles were identical to the *hydrogen ion* (hydrogen atom from which one electron has been removed). These particles were named as ***protons*** and were shown to be present in all matter.

Now it is the time to check your understanding. For this, take a pause and solve the following questions:

CHECK YOUR PROGRESS 3.1

1. Name the extremely small particles which constitute matter.
2. What do we call the negatively charged particles emitted from the cathode?
3. What is a cathode ray tube?
4. What is an anode?
5. Why the canal rays obtained by using different gases have different *e/m* values?

3.2 EARLIER MODELS OF ATOM

Based on the experimental observations, different models were proposed for the structure of the atom. In this section, we will discuss two such models namely Thomson model and Rutherford model.

3.2.1 Thomson model

All matter is made of atoms and all the atoms are electrically neutral. We have just seen that all atoms contain the electrons. Based on these facts, *Thomson concluded that there must be an equal amount of positive charge present in the atom.* He proposed that *an atom could be considered as a sphere of uniform positive charge in which electrons are embedded.* This is shown below in Fig.3.3.

This model is similar to a water-melon according to which an atom can be thought of as a sphere of positive charge in which the electrons are embedded like seeds. This model is also called **plum pudding model** or **raisin pudding model** because the electrons resembled the raisins dispersed in a pudding (an English dessert).

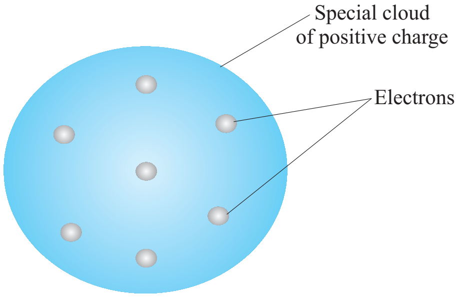

Fig. 3.3 Thomson model of atom

During this period only, the phenomenon of **radioactivity** was also being studied by the scientists. This phenomenon of spontaneous emission of rays from atoms of certain elements also proved that the atom was divisible and it contained sub –atomic particles. **Ernest Rutherford** and his coworkers were also carrying out experiments which revealed that the radiation could be of three types: α*(alpha)*, β*(beta) and* γ*(gamma)*. You will study more about them in lesson 14.

In 1910, Rutherford and his co-workers performed an experiment which led to the downfall of the Thomson model. Let us now study about the contribution of Rutherford.

3.2.2 Rutherford's model

Rutherford who was a student of J.J Thomson was studying the effect of alpha (a) particles on matter. The alpha particles are helium nuclei. They are obtained by the removal of two electrons from the helium atom. Hans Geiger (Rutherford's technician) and Ernest Marsden (Rutherford's student) directed α particles from α radioactive source on a thin piece of gold foil (about 0.00004 cm thick). This is shown below in Fig. 3.4.

Ernest Rutherford, (1871-1937) who received the Nobel Prize in Chemistry in 1908 for proposing the nuclear model of the atom.

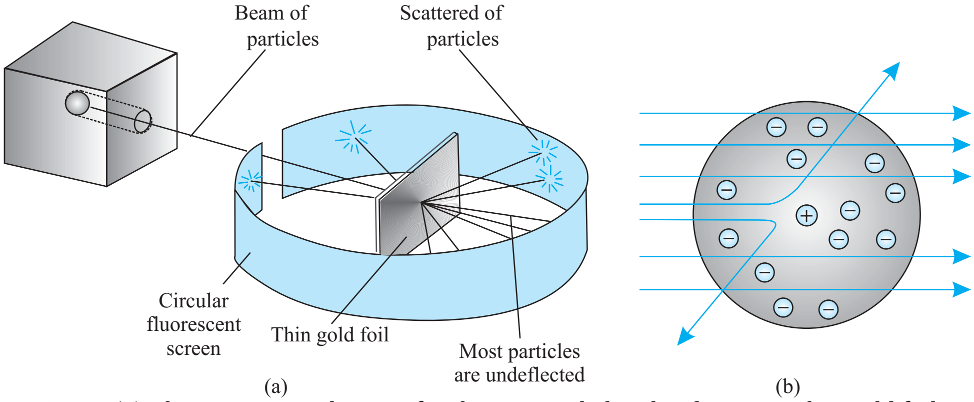

Fig.3.4 (a) The experimental set-up for the α particle bombardment on thin gold foil, (b)Scattering of α particles

If Thomson model was correct,then most of the a particles should pass through the gold foil and their path should only be deflected by a small amount. They were surprised to find out that although the majority of the a particles passed through the gold foil undeflected (or were deflected with minor angles), some of them were deflected by a large angles and a few even bounced back. This is shown in Fig. 3.4(b). In 1911, Rutherford explained the above observation by proposing another model of the atom. He suggested that :

(i) *Most* of the mass of atom and *all* of its positive charge reside in a very small region of space at the centre of the atom, called the **nucleus**.

(ii) The electrons revolve around the nucleus in circular paths.

This model is also known as **Rutherford's nuclear model** of the atom and is shown in Fig. 3.5.

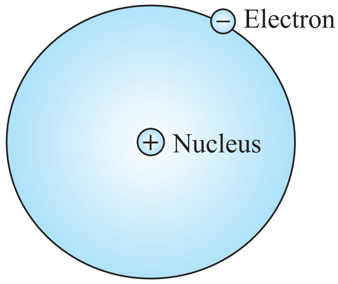

Fig. 3.5 Rutherford's nuclear model of atom

This model resembeled the solar system in which the nucleus was similar to the Sun and the electrons were similar to the planets. Ruthurford was able to predict the size of the nucleus by carefully measuring the fraction of α particles deflected. He estimated that the radius of the nucleus was atleast 1/10000 times smaller than that of the radius of the atom. We can imagine the size of the nucleus with the following similarity. If the size of the atom is that of a cricket stadium then the nucleus would have the size of a fly at the centre of the stadium.

Thus, most of the space in the atom is empty through which the majority of the α-particles could pass. When the α- particles come close to the nucleus, they are repelled by its positive charge and hence they show a large deflection. Wherefrom this positive charge comes in the nucleus?

The nucleus was supposed to contain positively charged particles, called **protons**. The positive charge on a proton was equal but opposite in nature to that on an electron. This quantity of charge, i.e. 1.602×10^{-19} C is called the electronic charge and is expressed as a **unit charge**, i.e., the charge of an electron is –1 whereas that of a proton is +1.

CHECK YOUR PROGRESS 3.2

1. Who proposed the nuclear model for the structure of atom?
2. Define nucleus.
3. What is a proton?

3.3 DISCOVERY OF NEUTRON

Although Rutherford's model of the atom could explain the electrical neutrality and the results of scattering experiment but a major problem regarding the atomic masses remained unsolved.

The mass of helium atom (which contains 2 protons) should be double than that of a hydrogen atom (which contains only one proton). [The electron being very light weight particle as compared to that of a proton, its contribution to the atomic mass can be ignored]. Actual ratio of helium and hydrogen masses is 4:1. Rutherford and others, thus, suggested that there must be one more type of subatomic particle present in the nucleus which may be neutral but must have mass. Later in 1932, James Chadwick showed the existence of this third type of subatomic particle. This was named as **neutron.** *The neutron was found to have a mass slightly higher than that of a proton electrically neutral.* Thus, if the helium atom contained 2 protons and 2 neutrons in the nucleus, its mass ratio to hydrogen as 4:1 could be explained. The characteristics of these three particles, called as **fundamental particles** are given in Table 3.1.

James Chadwick (1891-1972) was a British physicist. He received the Nobel prize in 1935 for showing the existence of neutron in the nucleus of an atom.

Table 3:1 Characteristics of the subatomic particles.

Particle	Symbol	Mass(kg)	Charge	Coulomb (C) in multiple units
Electron	*e*	9.10939×10^{-31}	-1.6022×10^{-19}	-1
Proton	*p*	1.67262×10^{-27}	$+1.6022 \times 10^{-19}$	+1
Neutron	*n*	1.67493×10^{-27}	0	0

CHECK YOUR PROGRESS 3.3

1. What is a neutron?
2. How many neutrons are present in the α-particle?
3. How will you distinguish between an electron and a proton?

3.4 ATOMIC NUMBER, MASS NUMBER AND ISOTOPES

Why do the atoms of different elements differ from each other? The numbers of protons present in the atom of an element are different from those present in the atom of another element. Thus, *the number of protons present in the atom of each element is fixed and is a characteristic property of that element as you have already learnt in lesson 2.* This number is called the **atomic number** and is denoted by *Z* .Hydrogen has one proton in its nucleus and therefore, its atomic number is 1. Similarly, two protons are present in the nucleus of helium atom and hence its atomic number is 2. What about the number of electrons present in hydrogen and helium? Since the atom is electrically neutral, the number of electrons present in these atoms is 1 and 2 respectively.In addition to the protons, the helium atom also has neutrons present in its nucleus. *The total number of protons and neutrons present in the nucleus of an atom of an element is called its* **mass number**. It is denoted by *A*. Helium nucleus contains 2 protons and 2 neutrons; hence, its mass number is 4.The atomic number and the mass number of an element (X) can be denoted as follows:

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

Thus, helium can be represented as $^{4}_{2}H$

Similarly, $^{12}_{6}C$ means that the carbon atom has 6 protons and hence 12–6 = 6 neutrons. But some carbon atoms can have 7 or 8 neutrons also. The mass number of these carbon atoms would be 6+7=13 or 6+8=14.Such *atoms which have the same atomic number but have different mass number are called* **isotopes.** Thus, carbon has three isotopes.

These isotopes can be represented as shown below:

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

CHECK YOUR PROGRESS 3.4

1. How is atomic number related to the number of protons present in the atom?
2. What is the mass number of an atom which has 7 protons and 8 neutrons?
3. Calculate the number of neutrons present in the following isotopes of hydrogen.

$$^{1}_{1}H,\ ^{2}_{1}H,\ ^{3}_{1}H$$

3.5 DRAWBACKS OF RUTHERFORD'S MODEL

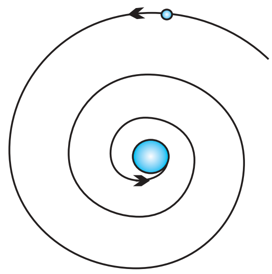

Fig. 3.6 Spiral path of an electron

As you have studied in section 3.3, Rutherford's model could not solve the problem of atomic mass. The existence of the neutron thus accounted for the mass of the atom. Another drawback which Rutherford's model suffered was that it could not explain the stability of the atom. According to the electromagnetic theory of radiation, a moving charged particle, such as the electron which is constantly accelerating because of change in directions of motion, should emit radiation. The energy of the radiation would come from the motion of the electron. Thus, the electron would emit radiation and follow a spiral path as shown in Fig.3.6.

The energy of the electron would keep on decreasing (as the electron would keep on emitting radiation) till the electron finally falls into the nucleus. But actually it does not happen. The electron does not collapse into the nucleus. Thus, Rutherford's model needed the improvements which were later on suggested by Bohr. Bohr's model now will be discussed in the next section.

CHECK YOUR PROGRESS 3.5

1. What were the two drawbacks of Rutherford's model?

3.6 BOHR'S MODEL OF ATOM

In 1913, Niels Bohr proposed a model which was an improvement over Rutherford's nuclear model. Bohr proposed that an electron moves around the nucleus in a well defined circular path. He set down following two main postulates to explain the stability of atom particularly hydrogen atom

(i) An electron can have only a definite circular path around the nucleus with specific energy values. This circular path he called **orbit** or **energy level**

(ii) Electron may go to next higher energy level (orbit) when given a definite amount of energy. In other words, an electron absorbs energy when it goes to higher energy level from a lower energy level.

Contrary to this, electron will emits out a definite amount of energy when it comes from a higher energy level to lower energy level. If E_2 is energy of an electron in higher energy level and E_1 is energy of electron in lower energy level, then energy released ΔE will be expressed as,

$$\Delta E = E_2 - E_1$$

If the electron remians in the same orbit, the energy would neither be released nor absorbed. These orbits will, therefore, were called **stationary orbits** or **stationary states**.

Niels Bohr (1885-1962). He was a Danish physicist He was awarded the Nobel Prize in Physics in 1922.

Although Bohr model could explain a number of aspects related to hydrogen atom but it could not explain stability of atoms having more than one electron. After the nature of electron was studied in detail, it was found that an electron cannot remain in a fixed circular orbit as envisaged by Bohr. Bohr model was rejected on this ground.

Based on the nature of electron, concept of circular orbit was modified and a three dimensional shell with definite energy came into existence. These shells are similar to circular path/energy levels given by Bohr. These shells are represented by letters *K*, *L*, *M*, *N* etc. Each shell is associated with a definite energy. The energies of these shells go on increasing as we move away from the nucleus. The maximum number of electrons which can be accommodated in each shell is given by $2n^2$ where n can take values 1, 2, 3….etc. Thus, the first shell can have two electrons whereas the second shell can have 8 electrons. Similarly the maximum number of electrons present in third and fourth shells would be 18 and 32, respectively. Each shell could be further sub-divided into various sublevels of energy called **subshells.** These subshells are denoted by letters *s, p, d, f,* etc about which you would study in your higher classes.

CHECK YOUR PROGRESS 3.6

1. What are stationary states?
2. What will happen to the energy of electron when it goes from an orbit of higher energy to that of a lower energy?
3. What is a shell?
4. How many electrons can be present in a *L*-shell?

3.7 ELECTRONIC CONFIGURATION OF ELEMENTS

From the above discussions, you are aware that shells of different energies exist in an atom. The electrons occupy these shells according to the increasing order of their energy. You also know that the first shell can have two electrons whereas the second shell can accommodate eight electrons. Keeping these points in mind, let us now study the filling of electrons in various shells of atoms of different elements.

Hydrogen atom has only one electron. Thus electronic configuration of hydrogen can be represented as 1.

The next element helium (He) has two electrons in its atom. Since the first shell can accommodate two electrons; hence, this second electron can also be placed in first shell. The electronic configuration of helium can be represented as 2.

The third element, Lithium (Li) has three electrons. Now the two electrons occupy the first shell whereas the third electron goes to the next shell of higher energy level, i.e. second shell. Thus, the electronic configuration of Li is 2, 1.

Similarly, the electronic configurations of beryllium (Be) and boron (B) having four and five electrons respectively can be written as follows:

Be	-	4 electrons	Electronic configuration - 2, 2.
B	-	5 electrons	Electronic configuration - 2, 3.

The next element carbon (C) has 6 electrons. Now the sixth electron also goes to the second shell which can accommodate eight electrons. Hence, the electronic configuration of carbon can be represented as 2, 4. Similarly, the next element nitrogen having 7 electrons has the electronic configuration 2, 5.

The electronic configuration of other elements can be given on the same lines. The electronic configuration of first twenty elements is given in Table 3.2 and depicted in Fig. 3.7.

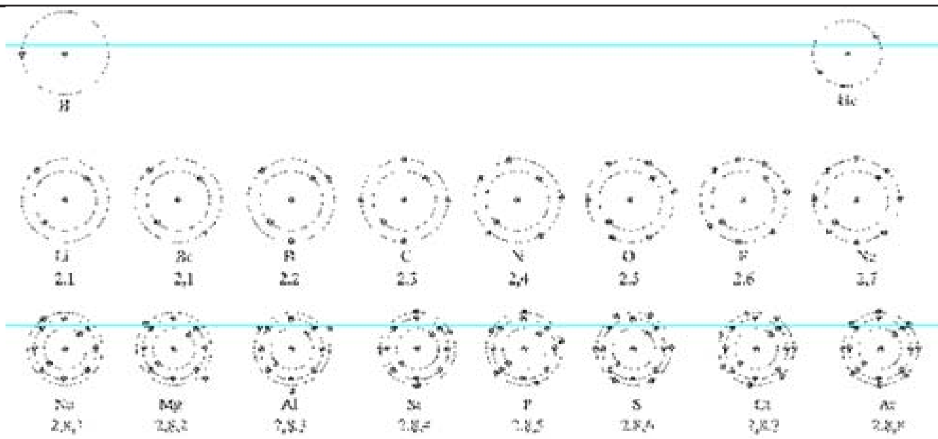

Fig. 3.7 Electronic configuration of some elements

Table 3.2: Electronic distribution in shells of first twenty elements

Element/symbol	No. of electrons	Arrangement of electrons in shells	Electrons distribution in shells	Common valency
Hydrogen, H	1	1 in first shell	1	1
Helium, He	2	2 in first shell	2	0
Lithium, Li	3	2 in first shell + 1 in second shell	2,1	1
Beryllium, Be	4	2 in first shell + 2 in second shell	2,2	2
Boron, B	5	2 in first shell + 3 in second shell	2,3	3
Carbon, C	6	2 in first shell + 4 in second shell	2,4	4
Nitrogen, N	7	2 in first shell + 5 in second shell	2,5	3
Oxygen, O	8	2 in first shell + 6 in second shell	2,6	2
Fluorine, F	9	2 in first shell + 7 in second shell	2,7	1
Neon, Ne	10	2 in first shell + 8 in second shell	2,8	0
Sodium, Na	11	2 in first shell + 8 in second shell + 1 in third shell	2,8,1	1
Magnesium, Mg	12	2 in first shell + 8 in second shell + 2 in third shell	2,8,2	2
Aluminium, Al	13	2 in first shell + 8 in second shell + 3 in third shell	2,8,3	3
Silicon, Si	14	2 in first shell + 8 in second shell + 4 in third shell	2,8,4	4
Phosphorus, P	15	2 in first shell + 8 in second shell + 5 in third shell	2,8,5	3,5
Sulphur, S	16	2 in first shell + 8 in second shell + 6 in third shell	2,8,6	2
Chlorine, Cl	17	2 in first shell + 8 in second shell + 7 in third shell	2,8,7	1
Argon, Ar	18	2 in first shell + 8 in second shell + 8 in third shell	2,8,8	0
Potassium, K	19	2 in first shell + 8 in second shell + 8 in third shell + 1 in fourth shell	2,8,8,1	1
Calcium, Ca	20	2 in first shell + 8 in second shell + 8 in third shell + 2 in fourth shell	2,8,8,2	2

3.7.1 Valence electron and valency

We have just discussed the electronic configuration of first 20 elements. We can see from the table 3.2 that electrons are located in different shells around the nucleus. The electrons in the last shell (popularly known as valence shell) govern the chemical properties of the atoms. These electrons are known as valence electrons. Valency or combining capacity of an atom of an element depends on the number of these electrons as mentioned in lesson 2. Valency of 20 elements along with their electronic configuration is also provided in Table 3.2.

In next lesson, you would study how these electronic configurations are useful in understanding the periodic arrangement of elements. These electronic configurations are also helpful in studying the nature of bonding between various elements which will be dealt in lesson 5.

CHECK YOUR PROGRESS 3.7

1. How many shells are present in the nitrogen atom?
2. Name the element which has the completely filled first shell.
3. The electronic configuration of an element having atomic number11 is______________

LET US REVISE

- Electrons are present in all the atoms.
- Thomson proposed the plum-pudding model of the structure of atom.
- Rutherford's model of the structure of atom suggested that most of the mass and all of positive charge of an atom is concentrated in its nucleus and the electrons revolve around it in.
- The neutrons are neutral particles present in the nucleus.
- Atomic number is the number of protons present in the nucleus of an atom.
- Mass number gives the number of protons and neutrons present in an atom
- Isotopes have same atomic number but different mass numbers.
- Bohr's model gave the idea of definite orbits or stationary states.
- The electrons occupy various shells in an atom in the increasing order of their energy. The maximum number of electrons which can be accommodated in a shell is $2n^2$.

TERMINAL EXERCISES

A. Fill in the blanks.

1. The nucleus consists of ———————and ————————
2. The model which resembled the solar system was proposed by——
3. Anode rays travel towards——————————
4. An electron has ——————— charge.

B. Classify the following statements as true or false.

1. The plum pudding model was proposed by Rutherford.
2. Cathode is the negatively charged electrode.
3. Neutrons are constituents of atoms of all elements.
4. The number of electrons present in a neutral atom is always equal to the number of protons.

C. Multiple choice type questions.

1. An α-particle has

   (a) 2 protons only. (b) 2 neutrons only

   (c) 2 protons and 2 neutrons (d) 2 neutrons

2. Isotopes have

   (a) same mass number (b) same atomic number

   (c) different atomic number (d) same mass as well as atomic number

3. The mass of a neutron

   (a) is less than that of a proton. (b) is greater than that of a proton.

   (c) is equal to that of a proton. (d) zero

4. The filling of second shell starts with

   (a) He (b) Li

   (c) C (d) N

5. The electronic configuration of Cl is

   (a) 2, 8 (b) 2, 8, 4

   (c) 2, 8, 6 (d) 2, 8, 7

6. Which of the following elements has completely filled shells?

   (a) H (b) O

   (c) Ne (d) Mg

D. Descriptive type questions.

1. How can you say that electrons are present in all types of matter?
2. Define an orbit.
3. Calculate the number of neutrons present in $^{16}_{8}O$ and $^{19}_{9}F$
4. The mass number of iron is 56. If 30 neutrons present in its atom, what is its atomic number?
5. Which of the following are isotopes? $^{12}_{6}C$, $^{14}_{6}C$, $^{14}_{7}N$

ANSWERS TO CHECK YOUR PROGRESS

3.1

1. atoms
2. electrons
3. A glass tube from which most of the air has been removed. It has two electrodes.
4. It is a positively charged electrode.
5. because the positive ions resulting from the different gases have different masses.

3.2

1. J.J. Thompson
2. The small region of space at the centre of the atom where most of the mass and all of the positive charge is located.
3. An alpha particle is the helium nucleus which is obtained by the removal of two electrons from the helium atom.

3.3

1. A neutron is a neutral subatomic particle having mass slightly higher than proton.
2. 2
3. (i) An electron has negative charge whereas a proton has a positive charge.

(ii) An electron is present outside the nucleus whereas a proton is present in the nucleus.

(iii) The electron has very less mass as compared to a proton.

3.4

1. Atomic number is equal to the number of protons present in the nucleus of the atom.
2. 15
3. 0, 1, 2.

3.5

1. It could not explain the correct atomic masses and the stability of atoms.

3.6

1. Stationary states are energy levels of definite energy. When an electron is present in a stationary state, its energy does not change.
2. Its energy would decrease.
3. A shell is a group of energy levels having similar energy.
4. 8 electrons.

3.7

1. 2
2. He
3. 2, 8, 1

GLOSSARY

Alpha particles: Positively charged particles ejected at high speeds from certain radioactive substances;

Atom: The smallest particle of an element that retains the chemical properties of that elemen.

Atomic nucleus: The tiny central core of an atom that contains neutrons and protons.

Atomic number: The number of protons in the nucleus of an atom of an element.

Electron: A negatively charged subatomic particle found in the space about the nucleus.

Electron shell: The collectio of orbitals with same principal quantum number.

Electronic configuration: The complete description of the orbitals occupied by all the electrons in an atom on ion.

Isotopes: Forms of an element composed of atoms with same atomic number but different mass number owing to a difference in a number of neutrons.

Mass number: The number of proton plus neutrons in the nucleus of an atom of an element.

Neutrons: An electrically neutral subatomic particle found in the nucleus.

Orbital: Regions occupied by electrons in S, P, d, f, subshells, represened by three dimensional boundary surface diagram..

Proton: A positively charged subatomic particle found in the nucleus.