SEMESTER–I (Period–II)

C10CH3

3 TOPIC

Atomic Structure

Learning Objectives

Upon completion of this lesson, learners will be able to:

- Discuss contributors to atomic structures.
- Explain the arrangement of fundamental particles.
- Discuss the concept of atomic theories.
- Discuss atomic number and mass number and their relationship to isotopes.
- Discuss the four quantum numbers.
- Construct electronic configurations for atoms.
- Explain the rules and principles for filling in electrons.

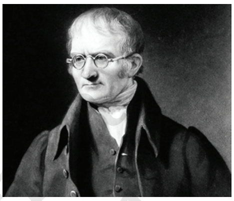
3.1. HISTORY OF ATOMIC STRUCTURE

3.1.1. History of Atomic Chemistry

In the fifth century B.C. the Greek Philosopher Democritus proposed that all matter consists of very small indivisible particles called *atoms* (meaning uncuttable or indivisible). These earlier ideas were mere speculations and did not have any experimental basis. It was in 1808 that an English scientist, John Dalton formulated a precise definition of the indivisible building blocks of matter that are known as atoms. By the end of 19th century it was discovered that atoms consist of subatomic particles such as electrons, protons and neutrons. The **protons** are positively charged particles and are present in the nucleus of the atom. The **electrons** are negatively charged particles and are present in the extra-nuclear part of the atom. The **neutrons** are neutral particles and are present in the nucleus along with the protons. Most of the mass of the atom is concentrated in the **nucleus**.

During the past few years it has been found experimentally that some of the subatomic particles can be further split into two elementary particles, **quarks** and **leptons**.

3.1.2. Dalton's Atomic Theory



John Dalton (1766–1844)

On the basis of laws of chemical combination John Dalton, an English school teacher in Manchester, proposed that behaviour of matter could be explained using an atomic theory. He published his work about atomic theory in 1808. The main points of **Dalton's atomic theory** are:

1. Matter is composed of extremely small particles called atoms.
2. An element consists of only one type of atoms. They have identical properties such as mass, shape, colour, density, chemical properties, etc.
3. Atoms of one element differ from atoms of all other elements in mass, size and chemical properties.
4. Atom is the smallest particle that takes part in chemical reactions.
5. Atoms of different elements can combine in a fixed ratio to form compound.
6. Atoms can neither be created nor destroyed during chemical reactions.

3.1.2.1. Limitations of Dalton's Atomic Theory

Dalton's atomic theory was the first successful attempt which gave us some idea about the inner structure of matter. The main failures of Dalton's atomic theory are:

1. It failed to explain how atoms of different elements differ from each other, i.e., it did not tell anything about internal structure of the atom.
2. It could not explain how and why atoms of different elements combine with each other to form compound-atoms or molecules.
3. It failed to explain the nature of forces that hold together different atoms in a molecule.
4. It did not make any distinction between ultimate particle of an element that takes part in reactions (atom) and ultimate particle that has independent existence (molecule).

Towards the end of 19th century and in the beginning of 20th century new discoveries were made by Sir J.J. Thomson, Neils Bohr, Chadwick and others which revealed the inner structure of atom. In the light of these findings Dalton's atomic theory was suitably modified.

3.1.3. Discovery of Electron

Electron was the first fundamental particle that was discovered. The credit for the discovery of electron goes to J.J. Thomson, a British physicist. Most of the information about electrons is obtained from the study of cathode rays, which were discovered during the experiments with gas discharge tubes. A discharge tube is long glass tube which is fitted with metal electrodes on either end, across which high voltage can be applied. The electrode which is connected to the negative terminal of the power source is called **cathode** while the electrode which is connected to the positive terminal is called **anode**. The tube is also connected to a vacuum pump for controlling the pressure of gas inside the discharge tube.

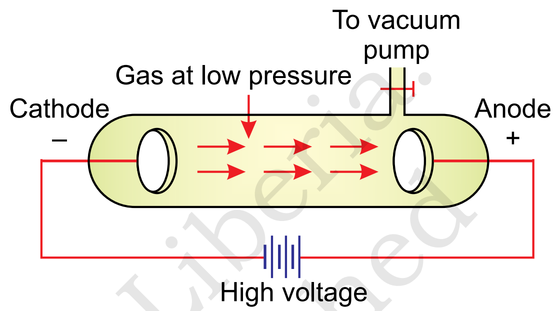


Fig. 3.1. Cathode rays

When the gas pressure inside the discharge tube is one atmosphere, no electric current flows through the tube. If the gas pressure is reduced to about 10^{-2} atmospheres and a potential difference of about 10000 volts is applied to the electrodes, an electric current flows and at the same time light is emitted by the gas. As the gas pressure in the discharge tube is reduced further, a dark space appears in the vicinity of cathode and alternate light and dark bands can be seen between the two electrodes. If the gas pressure is reduced to 10^{-6} atmospheres, the emission of light ceases, instead the end of the glass tube opposite to cathode glows (fluoresces) with a *faint greenish light*. Further investigations revealed that the fluorescence was caused due to the bombardment of the walls of the tube by rays emanating from cathode. These rays were called **cathode rays**. These rays were found to consist of negatively charged material particles, called **electrons**.

3.1.3.1. Properties of Cathode Rays

The cathode rays possess the following properties:

1. ***Cathode rays travel in straight lines.*** An object placed in the path of cathod rays casts a sharp shadow. It shows that cathode rays travel in straight lines.

2. ***Cathode rays consist of material particles.*** This was indicated by the fact that a light paddle wheel placed in the path of cathode rays starts rotating.
3. ***Effect of electric field.*** When electric field is applied to a stream of cathode rays, they get deflected towards positive plate (Fig. 3.2). It showed that cathode rays themselves are negatively charged.

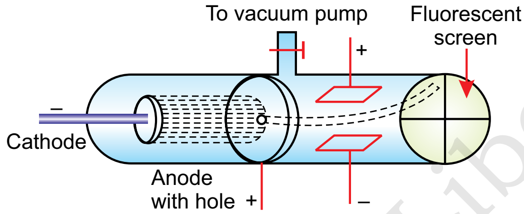


Fig. 3.2. Effect of Electric Field on Cathode Rays

4. ***Effect of magnetic field.*** When magnetic field is applied, the cathode rays get deflected. The direction of deflection again indicates that cathode rays are negatively charged.
5. On striking against walls of the discharge tube *cathode rays produce* ***faint greenish fluorescence***.

The above mentioned properties of cathode rays indicate that the cathode rays consist of a fast-moving stream of negatively charged material particles. These particles were named **electrons**.

3.1.3.2. Characteristics of Electron

Following are the characteristics of electron:

1. The charge to mass ration, e/m was found to be -1.76×10^8 coulombs/g or -1.76×10^{11} coulombs/kg.
2. Charge of electron is 1.60×10^{-19} coulomb.
3. Mass of electron is 9.11×10^{-31} kg, which is nearly equal to $\frac{1}{1840}$th of mass of an atom of hydrogen.

3.1.4. Discovery of the Proton

E. Goldstein, a German scientist, in 1886, discovered the existence of a new type of rays in the discharge tube. He used a perforated cathode (Fig. 3.3) in the discharge tube. The cathode divided the discharge tube in two chambers. On passing the electric discharge at low pressure he observed a new type of rays streaming behind the cathode. The path of these rays became visible due to the glow of the residual gas.

These rays also produced fluorescent glow on striking the walls of the tube behind cathode. These rays were named **anode rays** or **canal rays**. These rays were named canal rays because they passed through '*holes*' or '*canals*' in the cathode.

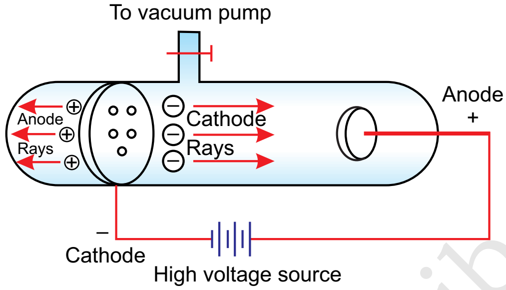


Fig. 3.3. Canal rays

Further investigations of these rays showed that they consist of positively charged material particles.

Some of the **characteristic properties of anode rays** are:

1. Anode rays consist of material particles.
2. Anode rays are deflected by electric field towards negatively charged plate.

   This indicates that *they are positively charged.*
3. The charge to mass ratio of the particles in the anode rays was determined by using Thomson's technique. Charge to mass ratio of the particles in the anode rays depends upon the nature of the gas taken in the discharged tube.

It was observed that e/m ratio was maximum when hydrogen gas was taken in the discharge tube. This indicated that positive ions formed from hydrogen are the lightest. These lightest positively charged particles were named **protons**. The charge and mass of the proton were determined, in the same manner as the one discussed in case of electron. Charge to mass ratio for protons was found to be $\mathbf{9.58 \times 10^{7}}$ **C/kg**. Charge on proton is opposite but equal in magnitude to the charge on the electron i.e , $\mathbf{1.60 \times 10^{-19}}$ **C**. From these two observations mass of a proton works out to be $\mathbf{1.67 \times 10^{-27}}$ **kg**. It is practically the same as the mass of a hydrogen atom and is about 1837 times the mass of an electron.

In general, a proton is represented as p^{+}.

A **proton** is a fundamental particle of atom carrying one unit positive charge and having mass 1.672×10^{-27} kg, which is nearly equal to the mass of an atom of hydrogen.

Thomson's Model of Atom

Thomson was the first to propose a detailed model of the atom. He proposed that an atom consists of a uniform sphere of positive electricity in which the electrons are distributed more or less uniformly. The negative and the positive charge are equal in magnitude. Thus, the atom as a whole is electrically neutral. This model of atom is known as the ***"Plum pudding model"*** because of its similarity to a christmas pudding with currants in it (Fig. 3.4).

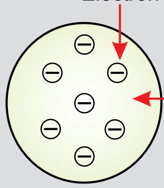
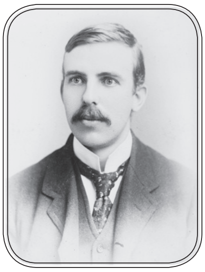


Fig. 3.4. Thomson's model of atom.

3.1.5. Discovery of Nucleus (Rutherford's Gold Foil Experiment)

Ernest Rutherford in 1911 performed an experiment which led to the downfall of Thomson's model. Most of the experimental work was carried out by Geiger and Marsden, two of Rutherford's students at Manchester University. Geiger and Marsden built the apparatus as shown in Fig. 3.5. The experiment involved the bombardment of a thin sheet of gold (thickness 4×10^{-5} cm) by α-particles. The source of α-particles was a piece of radium placed in a lead block. A narrow hole in the lead block allowed the α-particles to travel only in one direction through the evacuated vessel towards a fluorescent zinc sulphide screen.



Ernest Rutherford (1871-1937)

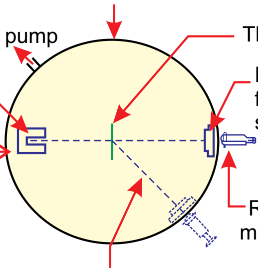


Fig. 3.5. The apparatus used by Geiger and Marsden to investigate the scattering of a-particles by a thin metal foil.

When an α-particle hits the fluorescent screen a tiny flash of light is produced which is observed with the help of a microscope. A thin gold foil was placed between the slit and the screen. The screen and the microscope were rotated to detect scattering of α-particles.

Rutherford observed that:

1. Most of the α-particles (nearly 99%) passed through the gold foil undeflected.
2. Some of the α-particles were deflected by small angles.
3. Very few particles (1 in about 10^4) were either deflected by very large angles or were actually reflected back along their path.

In order to explain the observations of his experiment, Rutherford assumed that the solid gold foil consists of layers of individual atoms which are touching each other so that there is hardly any empty space between them. As such the α-particles striking the gold foil must pass through the atoms. Rutherford explained his observations as follows:

1. Since most of the α-particles pass through the foil undeflected, it indicates that *the most of the space in an atom is empty* (Fig. 3.6).

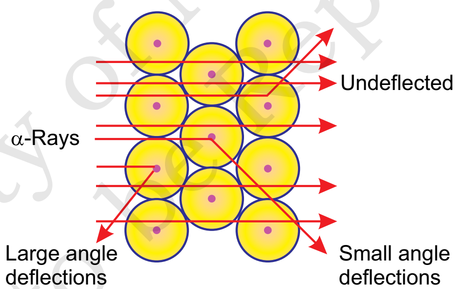


Fig. 3.6. Scattering of a-particles by gold atoms

2. α-particles being positively charged and having considerable mass could be deflected only by some heavy, positively charged centre. The small angle of deflection of α-particles indicates the presence of a *heavy positive centre in the atom.* Rutherford named this positive centre as **nucleus** (Fig. 3.7).

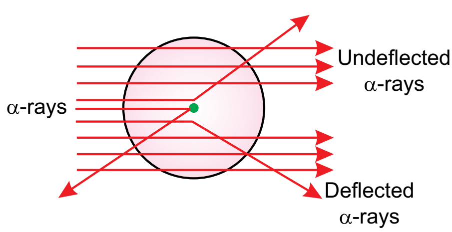


Fig. 3.7.

3. α-particles which make head-on collision with heavy positive centre are deflected through large angles. Since the number of such α-particles is very small, *the space occupied by the heavy positive centre must be very small.*

From the data of scattering experiment, Rutherford was able to calculate the radius of the nucleus. Rutherford calculated that the nucleus of an atom would have radius of about 10^{-14} m. He showed that the radius of the nucleus is about 10^{-4} times the radius of the atom which is about 10^{-10} m.

3.1.5.1. Rutherford's Nuclear Model of Atom

On the basis of scattering experiment Rutherford put forward *nuclear model of atom.* Main points of this model are:

1. *Most of the mass and all the positive charge of an atom is concentrated in a very small region called nucleus. Size of the nucleus is extremely small as compared with the size of the atom.* Radius of the nucleus is of the order of 10^{-15} m, whereas radius of atoms is of the order of 10^{-10} m.
2. *The nucleus is surrounded by electrons which are revolving around it at very high speeds. The electrostatic force of attraction between electrons and the nucleus is balanced by the centrifugal force acting on the revolving electrons.*
3. *Total negative charge on the electrons is equal to the total positive charge on the nucleus so that atom on the whole is electrically neutral.*

Nuclear model of atom can be compared with the solar system. In an atom electrons revolve around the nucleus in just the same way as the planets revolve around the sun. Due to this comparison, revolving electrons are sometimes called ***planetary electrons***.

Drawbacks of Rutherford's Model of Atom

Rutherford model failed in view of electromagnetic theory given by Maxwell. According to this theory a charged particle when accelerated emits energy in the form of electromagnetic radiation. According to Rutherford's model, electrons are revolving around the nucleus. This means, electrons would be in a state of acceleration all the time. Since electrons are charged particles, therefore, electron revolving in an orbit should continuously emit radiations. As a result of this, it would slow down and would no longer be able to withstand the attractive force of the nucleus. Hence, it would move closer and closer to the nucleus and would finally fall in the nucleus by following a spiral path (Fig. 3.8). This means atom should collapse. But actually we know atom is stable. Thus, *Rutherford's model failed to explain stability of atoms.*

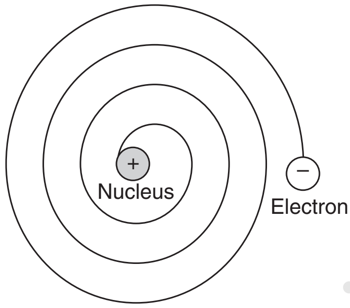


Fig. 3.8. Gradual decrease in the radius of orbit

3.1.6. Discovery of the Neutron

The existence of neutrons in the nucleus was first predicted by **Rutherford** in 1920 to account for the difference in the mass of the atom and the total mass of protons. Neutrons were discovered experimentally by **James Chadwick** in 1932. He bombarded a thin foil of beryllium with fast moving α-particles and observed that highly penetrating rays consisting of neutral particles were produced. These neutral particles were found to have mass 1.675×10^{-27} kg and were named *neutrons*.

In general, a neutron is represented as '***n***'.

> A ***neutron*** is a subatomic particle carrying no charge and having mass 1.675×10^{-27} kg which is almost equal to that of a hydrogen atom.

The mass of the atom is largely due to protons and neutrons in the nucleus of the atom. *Atoms of all elements except ordinary hydrogen contain one or more neutrons in their nuclei.*

The relative masses and relative charges of these three subatomic particles are summarized in Table 3.1.

Table 3.1. The Relative Masses and Relative Charges of Three Fundamental Particles

Particle	Relative Mass	Mass in kg	Relative Charge	Charge
Proton (p^+)	1	1.673×10^{-27} kg	+ 1 unit	$+ 1.602 \times 10^{-19}$ C
Neutron (n)	1	1.675×10^{-27} kg	0	0
Electron (e^-)	$\frac{1}{1840}$	9.110×10^{-31} kg	– 1 unit	$- 1.602 \times 10^{-19}$ C

3.1.7. Bohr's Model of Hydrogen Atom

In order to overcome the shortcomings of Rutherford's model, Neils Bohr (1913) proposed a new model of hydrogen atom based on radically new concepts. Bohr made a bold suggestion that particles at atomic level behave differently from the macroscopic objects. He proposed that at the atomic level, electron could revolve around the nucleus in stable orbits without continuously radiating energy in the form of electromagnetic radiations. Main postulates of this model are:

1. *The electrons in the hydrogen atom revolve around the nucleus only in certain selected circular orbits. These orbits are associated with definite energies and are called* ***energy shells*** *or* ***energy levels****. These are numbered as 1, 2, 3, 4, ... ,etc., or designated as K, L, M, N, ..., etc. shells (Fig. 3.9).*

   The energy of the electron is minimum in the orbit nearest to the nucleus i.e., K shell. The energy of the electron increases as it moves away from the nucleus (Fig. 3.9).

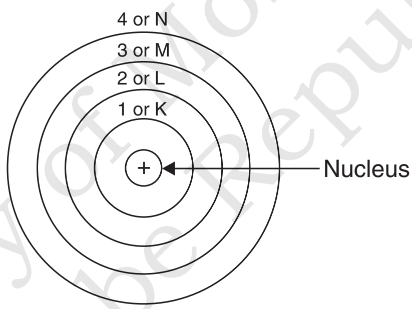


Fig. 3.9. Bohr's orbits

2. *As long as the electron remains in a particular orbit, it does not lose or gain energy. This means that energy of an electron in a particular orbit remains constant. That is why, these orbits are also called* ***stationary states****.*
3. *When energy from some external source is supplied to the electron, it may jump to some higher energy level by absorbing a definite amount of energy (equal to the difference in energy between the two energy levels). When the electron jumps back to the lower energy level it radiates the difference in energy in the form of electromagnetic radiation (Fig. 3.10).*

   The wavelength (λ) of the radiation emitted depends upon the energies of the two levels between which the transition is taking place.

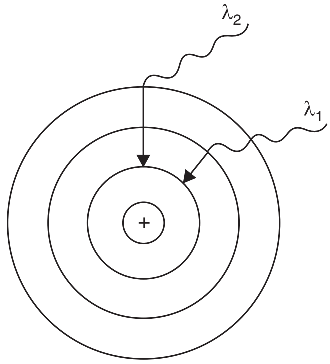


Fig. 3.10.

3.1.7.1. Success of Bohr's Model

Niels Bohr
(1885–1962)

Bohr's model could explain the stability of an atom. According to Bohr's model, an electron revolving in a particular orbit cannot lose energy. The electron can lose energy only if it jumps to some lower energy level. If no lower energy level is vacant then electron will keep on revolving in the same orbit without losing energy and hence it explains the stability of atom.

3.2. FUNDAMENTAL PARTICLES OF AN ATOM

Atoms are made from three fundamental (or sub-atomic)) particles. These are include electrons, protons and neutrons. The atom contain nucleus at its center, which has positively charged protons and neutrons. Electrons are revolving around the nucleus and they carry negative charge.

3.2.1. Arrangement of Particles in an Atom

The fundamental particles are arranged in an atom as follows:

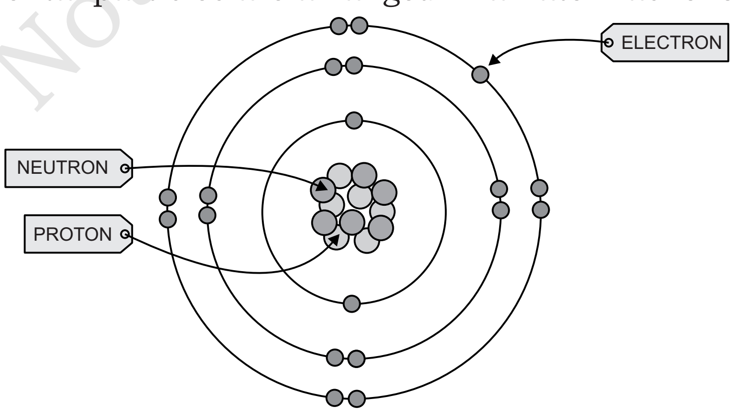


Fig. 3.11. Arrangement of particles in atom

Neutrons and protons are present in the nucleus while electrons revolves around the nucleus.

ACTIVITY 3.1

Aim: ***To Make Model of an Atom***

Materials Required: Some ping-pong balls of different colours (one colour for protons, one for the neutron and other for electrons), glue, a piece of paper and a cardboard.

Procedure:

1. Take a piece of paper and paste it on the cardboard.
2. Draw 4-5 concentric circle on the paper.
3. Now, glue the balls together to represent protons and neutrons. This is the nucleus.
4. Glue the nucleus to the center of the cardboard.
5. Starting from the innermost ring, you can have up to two electrons in the first ring, up to eight in the second ring, up to 18 in the third ring and up to 32 in the fourth ring.
6. Now, glue the electrons on the rings, spacing them evenly. Make sure you don't exceed each ring's maximum number of electrons.

3.3. ISOTOPES

All the atoms of a particular element have same number of protons in their nuclei, however, the number of neutrons may be different. Such atoms have same atomic number but different mass numbers and are known as **isotopes** of the element. Thus:

Isotopes *of an element are the atoms of the element with the same atomic number but different mass numbers.*

For example, hydrogen has three isotopes, protium (H), deuterium (D) and tritium (T). All the three isotopes have atomic number 1, however, their mass numbers are 1, 2 and 3 respectively. The isotopes of other elements do not have special names; they are indicated by giving mass number value on the symbol.

Mass number and atomic number of an atom are generally indicated as follows:

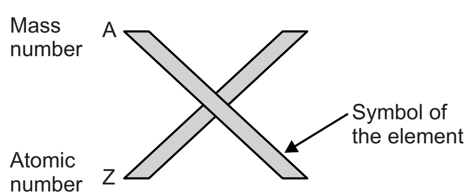


Fig. 3.12. Representation of an isotope

Thus, three isotopes of Hydrogen can be represented as:

$$^{1}_{1}\mathrm{H} \qquad ^{2}_{1}\mathrm{H} \qquad ^{3}_{1}\mathrm{H}$$

Protium **Deuterium** **Tritium**

As isotopes are the atoms with the same atomic number but different mass numbers, then we need to know about atomic number and mass number of the atom.

The two isotopes of chlorine are represented as:

$$^{35}_{17}\mathrm{Cl} \qquad ^{37}_{17}\mathrm{Cl}$$

Chlorine-35 **Chlorine-37**

Another way to describe an isotope is to cite its elemental name and mass number. The two isotopes of chlorine may be represented as chlorine-35 and chlorine-37. Similarly, carbon exists in the form of three isotopes C-12, C-13 and C-14.

3.3.1. Atomic Number

Nucleus carries positive charge due to the presence of protons in it. In 1913, H.G.J. Moseley, a young British physicist, discovered a way to determine nuclear charge accurately. From the charge on the nucleus, the number of protons in it could be determined.

The atomic number of an element is equal to the number of protons in the nucleus of its atom. The atomic number is also known as **proton number** of the element. Further, in an atom, number of protons is equal to the **number of electrons**. Hence, atomic number is also equal to the number of electrons in an atom of the element. Thus,

Atomic number of an element is equal to the number of protons in the nucleus of its atom or the number of extra-nuclear electrons.

Atomic number is denoted by the letter Z.

Atomic Number (Z) = Number of protons = Number of electrons.

For example, an atom of magnesium contains 12 protons, therefore, its atomic numberis 12.

All the atoms of a particular element contain same number of protons in their nuclei.

Therefore, all the atoms of an element have same atomic number. Atoms of different elements contain different number of protons and hence no two elements can have same atomic number. Thus, *each* *element has its characteristic atomic number which can be used to identify the element.*

For example, when we say an element with atomic number 13, we are referring to aluminium. No other element has atomic number 13.

3.3.2. Mass Number

It has already been stated that mass of an atom is mainly concentrated in the nucleus. In the nucleus, there are protons and neutrons. From this it follows that mass of an atom is mainly due to protons and neutrons. Protons and neutrons are collectively called **nucleons**.

The total number of protons and neutrons in the nucleus is called mass number of the atom.

It is generally represented by the letter A.

Mass Number (A) = Number of protons + Number of neutrons
= Number of nucleons

The number of protons in the nucleus is equal to atomic number, Z while *the number of neutrons in the nucleus is sometimes called* ***neutron number***; N.

3.3.3. Calculation of Number of Electrons, Protons and Neutrons

From the knowledge of atomic number and mass number of an element it is possible to calculate number of electrons, protons and neutrons in an atom of the element. For example, atomic number and mass number of aluminium are 13 and 27 respectively.

Number of electrons, protons and neutrons in an atom of it can be calculated as under:

Number of protons = Atomic number = 13
Number of electrons = Atomic number = 13
Number of neutrons = Mass number – Atomic number
= 27 – 13 = 14.

3.4. RELATIVE ATOMIC MASS

Mass of an isotope, relative to the mass of an atom of carbon (C-12 isotope) taken as 12 amu or 12 u, is called the **relative isotopic mass**. The relative masses of isotopes and their relative abundance can be determined using a mass spectrometer.

The relative atomic mass of an element can be calculated from the relative masses and relative abundances of its various isotopes. ***Relative***

atomic mass *of an element is the weighted average of relative masses of its isotopes*. For example, hydrogen exists in the form of three isotopes; protium, deuterium and tritium with relative masses 1.0078, 2.0141 and 3.0161 u respectively. The relative abundance of protium is 99.985%, deuterium is 0.015% and tritium is almost zero. The average atomic mass of hydrogen may be calculated as:

$$\text{Average atomic mass of hydrogen} = \frac{1.0078 \times 99.985 + 2.0141 \times 0.015}{100}$$

$$= \mathbf{1.00794\ u}$$

It may be mentioned here that ***mass number is different from relative atomic mass***. Mass number is always a whole number (as it corresponds to number of protons and neutrons in the nucleus of the isotope) whereas relative atomic mass is generally fractional because it is the weighted average of relative masses of its various isotopes.

Relative atomic mass (RAM) of an element may be defined as the average relative mass of all the isotopes of the element as compared with mass of an atom of carbon (C-12 isotope) taken as 12 u.

Example 3.1: *The nucleus of an atom of a certain isotope contains 6 protons and 7 neutrons. What is the atomic number and the mass number of this isotope? To which element does it belong?*

Solution:

We know that:

$$\begin{aligned} \text{Atomic number} &= \text{Number of protons} \\ \therefore \quad \text{Atomic number} &= 6 \\ \text{Mass number} &= \text{Number of protons + Number of neutrons} \\ &= 6 + 7 = 13. \end{aligned}$$

Since atomic number of the given isotope is 6, it indicates that this isotope is of **carbon** element.

Example 3.2: *The number of protons in the nucleus of an atom of mass number 97 is 41. Find out the number of neutrons in its isotope of mass number 99?*

Solution:

The atomic number of isotopes is same. Therefore, the number of protons in both the isotopes is same and is equal to 41.

$$\begin{aligned} \text{Mass number} &= \text{Number of protons + Number of neutrons} \\ \therefore \quad \text{Number of neutrons} &= \text{Mass number} - \text{Number of protons} \\ &= 99 - 41 = 58. \end{aligned}$$

Example 3.3: *Boron exists in nature in the form of two isotopes B-10 and B-11 with relative abundance 19.91% and 80.09% respectively. The masses of B-10 and B-11 isotopes as determined by mass spectrometer are 10.0129 and 11.0093 respectively. What is the average atomic mass of boron?*

Solution:

$$\% \text{ of B-10} = 19.91\%;\ \ \% \text{ of B-11} = 80.09\%$$

$$\text{Average atomic mass of boron} = \frac{19.91 \times 10.0129 + 80.09 \times 11.0093}{100}$$

$$= \mathbf{10.811\ u}$$

3.5. QUANTUM NUMBERS

In an atom a large number of electron orbitals are permissible. These orbitals are designated by a set of numbers known as **quantum numbers**. In order to specify energy, size, shape and orientation of the electron orbital three quantum numbers are required. These are, principal quantum number, azimuthal quantum number and magnetic quantum number. These quantum numbers arise as a natural consequence during the solution of the Schrodinger wave equation. In order to designate the electron, an additional quantum number called spin quantum number is needed to specify spin of the electron. Thus, each orbital in an atom is designated by a set of three quantum numbers and each electron is designated by a set of four quantum numbers. These quantum numbers are discussed below.

3.5.1. The Principal Quantum Number (n)

This is the most important quantum number as it determines to a large extent the energy of an electron. It also determines the average distance of an electron from the nucleus. It is denoted by the letter n. This quantum number tells us in which principal energy level or shell the electron is present. It can have any whole number value such as 1, 2, 3, 4, ……, etc. The energy levels or energy shells corresponding to these numbers are designated as K, L, M, N, ……, etc. As the value of n increases, the electron gets farther away from the nucleus and its energy increases. The higher the value of n, the higher is the electronic energy. For hydrogen and hydrogen-like species, the energy and size of the orbital are determined by principal quantum number alone.

Energy of the electron in a hydrogen atom is related to principal quantum number by the following relation:

$$E_n = -\,k^2\,\frac{\text{üπ}^2\text{ü}\,m e^4}{n^2h^2} = -\,\frac{\times \qquad ^{-18}}{n^2}\,\text{J}$$

where, m = mass of electron
e = charge on electron
h = Planck's constant
E_n = Energy of the electron in nth principal shell
n = Principal quantum number used to designate the principal shell
k = Coulomb's law constant.

This relation is similar to the expression given by Bohr.

3.5.2. The Orbital Angular Momentum Quantum Number or Azimuthal Quantum Number (*l*)

This quantum number determines angular momentum of the electron. This is denoted by *l*. The value of *l* gives the **sub-level** or **sub-shell** in which the electron is located. *It also determines the shape of the orbital* in which the electron is located. The number of sub-shells within a principal shell is determined by the value of *n* for that principal energy level. *Thus, l may have all possible whole number values from 0 to n—1* for each principal energy level. For a given value of *n*, *l* can have *n* values. The various sub-levels are designated as *s, p, d, f* depending upon the value of *l*, as follows:

Value of *l* →	0	1	2	3	4	5	6
Designation of sub-shell →	*s**	*p*	*d*	*f*	*g*	*h*	*i*

For *n* = 1, *l* can have only one value *i.e.*, 0. It means that an electron in first energy level can be present only in s-sub-shell (l = 0).

So **first energy level** has only **one sub-shell**, *i.e.*, 1*s*.

For *n* = 2, *l* can have values 0 and 1. It means that the electron in second principal energy level may be located either in s-sub-shell (l = 0) or *p*-sub-shell (l = 1).

So **second energy level** has only two **sub-shells**, *i.e.*, 2*s* and 2*p*.

For *n* = 3, possible values of *l* are 0, 1 and 2. This implies that an electron in third principal energy level may be present either in *s*-sub-shell (l = 0) or *p*-sub-shell (l = 1) or *d*-sub-shell (l = 2).

So **third energy level** has **three sub-shells,** *i.e.*, 3*s*, 3*p* and 3*d*. Similarly, **fourth energy level** (n = 4) can have **four sub-shells** 4*s*, 4*p*, 4*d* and 4*f*.

*The names **s, p, d** and **f** are derived from spectroscopic terms sharp, principal, diffuse and fundamental respectively.

The relation between the orbital angular momentum and azimuthal quantum number, l is

Orbital Angular Momentum

$$= \sqrt{l(l+1)}\,\frac{h}{2\pi} = \sqrt{l(l+1)}\,\hbar$$

Azimuthal quantum number is also known as **subsidiary quantum number**.

3.5.3 The Magnetic Quantum Number (m_l)

This quantum number which is denoted by m_l refers to the different orientations of electron cloud in a particular sub-shell. These different orientations are called **orbitals**. The number of orbitals in a particular sub-shell within a principal energy level is given by the number of values allowed to m_l which in turn depends on the values of l. The possible values of m_l range from $+ l$ through 0 to $– l$, thus making a total of $(2l + 1)$ values. Thus, in a subshell, the number of orbitals is equal to $(2l + 1)$.

For l = 0 (i.e., *s*-sub-shell), m_l can have only one value, $m_l = 0$. It means that ***s*-sub-shell** has only **one orbital**.

For l = 1 (i.e., *p*-sub-shell), m_l can have three values, + 1, 0 and – 1. This implies that ***p*-sub-shell** has **three orbitals**.

For l = 2 (i.e., *d*-sub-shell), m_l can have five values, + 2, + 1, 0, – 1, – 2. It means that ***d*-sub-shell** has **five orbitals**.

For l = 3 (i.e., *f*-sub-shell), m_l can have seven values, + 3, + 2, + 1, 0, – 1, – 2, – 3. It means that ***f*-sub-shell** has **seven orbitals**.

The number of orbitals in various types of sub-shells are given below in tabular form:

Sub-shell	*s*	*p*	*d*	*f*	*g*
Value of l	0	1	2	3	4
No. of orbitals ($2l + 1$)	1	3	5	7	9

The relationship between the principal quantum number (n), angular momentum quantum number (l) and magnetic quantum number (m_l) is summed up in Table 3.2.

Table 3.2. Relationship among Values of n, l and m_l

Energy Level	Principal Quantum Numbers (n)	Possible Values of (l)	Designation of Sub-shell	Possible Values of (m_l)	Number of orbitals	
					In a given Sub-shell	In a given Energy Level
K	1	0	1s	0	1	1
L	2	0	2s	0	1	4
		1	2p	+ 1, 0, – 1	3	
M	3	0	3s	0	1	9
		1	3p	+ 1, 0, – 1	3	
		2	3d	+ 2, + 1, 0, – 1, – 2	5	

3.5.4. The Spin Quantum Number (m_s)

This quantum number which is denoted by m_s does not follow from the wave mechanical treatment but arises from the spectral evidence that electron in its motion about the nucleus also rotates or spins about its own axis.

This quantum number determines the orientation of spin angular momentum. Spin angular momentum is quantised and can have two orientations relative to a chosen axis. The spin quantum number can have only two values which are $+\frac{1}{2}$ and $-\frac{1}{2}$. The $+\frac{1}{2}$ value indicates clockwise spin (generally represented by an arrow pointing upwards, i.e., ↑) and the other indicates anti-clockwise spin (generally represented by an arrow pointing downwards i.e., ↓).

Due to its spin, the electron behaves as a tiny magnet. The spin of the electron is responsible for most of the magnetic properties of atoms, molecules or ions. If all the electrons in an atom or molecule are paired, it behaves as a **diamagnetic substance** i.e., it is weakly repelled by the magnetic field. On the otherhand, if atoms or molecules of a substance have one or more unpaired or odd electrons, it behaves as a **paramagnetic substance**, i.e., it is weakly attracted by magnetic field.

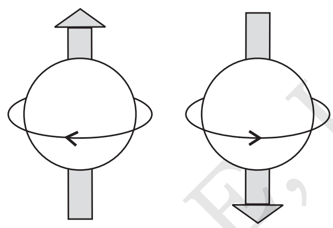


Fig. 3.13. Magnetic field associated with a spinning electron

3.6. ELECTRONIC CONFIGURATION

3.6.1. Dot Notation

A **Lewis electron dot diagram** or **electron dot diagram** is a representation of the valence electrons of an atom that uses dots around the symbol of the element. The number of dots equals the number of valence electrons in the atom. These dots are arranged to the right and left and above and below the symbol, with no more than two dots on a side.

For example, the Lewis electron dot diagram for calcium is simply

·Ca·

3.6.2. Orbital Notation

Each electron present in an atom is present in a subshell. There are totally four subshells present which are labelled as *s, p, d* and *f*. Each subshell can accommodate only a certain number of electrons, i.e. only 2 electrons can be occupied in *s* subshell, 6 electrons in *p* subshell, 10 electrons in *d* subshell and 14 electrons in *f* subshell.

We can write the electronic configuration of an atom by assigning the number of electrons that are present in the atom as the superscript of the subshell. These electrons are filled in the subshells according to the

Aufbau's principle (each added electron occupies the subshell of lowest energy available). Increasing order of energies of various orbitals are:

$1s > 2s > 2p > 3s > 3p > 4s > 3p > 3d > 4p > 5s > 4d > 5p > 6s > 4f > 5d > 6p > 7s > 5f > 6d > 7p$

Electronic configuration of the first 10 elements are mentioned in Table 3.3 given below.

Table 3.3. Electronic Configuration of the First 10 Elements

No. of Atom	Element	Electronic Configuration
1	Hydrogen (H)	$1s^1$
2	Helium (He)	$1s^2$
3	Lithium (Li)	$1s^2 2s^1$
4	Beryllium (Be)	$1s^2 2s^2$
5	Boron (B)	$1s^2 2s^2 2p^1$
6	Carbon (C)	$1s^2 2s^2 2p^2$
7	Nitrogen (N)	$1s^2 2s^2 2p^3$
8	Oxygen (O)	$1s^2 2s^2 2p^4$
9	Fluorine (F)	$1s^2 2s^2 2p^5$
10	Neon (Ne)	$1s^2 2s^2 2p^6$

3.6.2.1. *Electronic Configuration of Transition Elements*

The first series of transition elements (At. Nos. = 21 – 30) which follow calcium are scandium (Sc), titanium (Ti), vanadium (V), chromium (Cr), manganese (Mn), iron (Fe), cobalt (Co), nickel (Ni), copper (Cu) and zinc (Zn). In these elements addition of electrons takes place in the 3d-orbitals. All these elements are known as the **transition elements**. The electronic configuration of scandium (At. No. = 21), the first transition element, for example, is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^1$ while that of zinc (At. No. = 30), the last element of the series, is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}$. Thus, while the filling of $3d$-orbitals begins with scandium, it ends with zinc. The electronic configurations of all these 10 elements are represented in Table 3.4.

Table 3.4. Electronic Configurations of Transition Elements
[Scandium (At. No. 21) to Zinc (At. No. 30)]

Atomic Number	Element	Electronic Configuration			
21	Scandium	$1s^2$	$2s^2\ 2p^6$	$3s^2\ 3p^6\ 3d^1$	$4s^2$
22	Titanium	$1s^2$	$2s^2\ 2p^6$	$3s^2\ 3p^6\ 3d^2$	$4s^2$

23	Vanadium	$1s^2$	$2s^2\ 2p^6$	$3s^2\ 3p^6\ 3d^3$	$4s^2$
24	Chromium	$1s^2$	$2s^2\ 2p^6$	$3s^2\ 3p^6\ 3d^5$	$4s^1$
25	Manganese	$1s^2$	$2s^2\ 2p^6$	$3s^2\ 3p^6\ 3d^5$	$4s^2$
26	Iron	$1s^2$	$2s^2\ 2p^6$	$3s^2\ 3p^6\ 3d^6$	$4s^2$
27	Cobalt	$1s^2$	$2s^2\ 2p^6$	$3s^2\ 3p^6\ 3d^7$	$4s^2$
28	Nickel	$1s^2$	$2s^2\ 2p^6$	$3s^2\ 3p^6\ 3d^8$	$4s^2$
29	Copper	$1s^2$	$2s^2\ 2p^6$	$3s^2\ 3p^6\ 3d^{10}$	$4s^1$
30	Zinc	$1s^2$	$2s^2\ 2p^6$	$3s^2\ 3p^6\ 3d^{10}$	$4s^2$

3.6.3. Orbital Diagram

An orbital diagram is a way of representing the electronic configuration of an atom. Electronic configurations can be determined by applying the Aufbau principle (*each added electron occupies the subshell of lowest energy available*), Pauli Exclusion Principle (*no two electrons can have the same set of four quantum numbers*), and Hund's rule of maximum multiplicity (*whenever possible, electrons retain unpaired spins in degenerate orbitals*).

Orbital diagrams (Orbital box diagrams) of the first 10 elements are mentioned in Table 3.5 given below.

Table 3.5. Electronic Configuration of the First 10 Elements using Orbital Diagram

No. of Atom	Element	Electronic Configuration	Orbital Diagram
1	Hydrogen (H)	$1s^1$	$\underset{1s}{\boxed{\uparrow}}$
2	Helium (He)	$1s^2$	$\underset{1s}{\boxed{\uparrow\downarrow}}$
3	Lithium (Li)	$1s^22s^1$	$\underset{1s}{\boxed{\uparrow\downarrow}}\ \underset{2s}{\boxed{\uparrow}}$
4	Beryllium (Be)	$1s^22s^2$	$\underset{1s}{\boxed{\uparrow\downarrow}}\ \underset{2s}{\boxed{\uparrow\downarrow}}$
5	Boron (B)	$1s^22s^22p^1$	$\underset{1s}{\boxed{\uparrow\downarrow}}\ \underset{2s}{\boxed{\uparrow\downarrow}}\ \underset{2p}{\boxed{\uparrow}\boxed{}\boxed{}}$

6	Carbon (C)	$1s^2 2s^2 2p^2$	1s: ↑↓ 2s: ↑↓ 2p: ↑ ↑ ☐
7	Nitrogen (N)	$1s^2 2s^2 2p^3$	1s: ↑↓ 2s: ↑↓ 2p: ↑ ↑ ↑
8	Oxygen (O)	$1s^2 2s^2 2p^4$	1s: ↑↓ 2s: ↑↓ 2p: ↑↓ ↑ ↑
9	Fluorine (F)	$1s^2 2s^2 2p^5$	1s: ↑↓ 2s: ↑↓ 2p: ↑↓ ↑↓ ↑
10	Neon (Ne)	$1s^2 2s^2 2p^6$	1s: ↑↓ 2s: ↑↓ 2p: ↑↓ ↑↓ ↑↓

3.6.4. Noble Gas Notation

It may be noted that configurations of atoms can also be written in condensed form by taking the configurations of noble gases as the core. The configurations of inert gases representing core are written as $[He]^2$, $[Ne]^{10}$, $[Ar]^{18}$, $[Kr]^{36}$, $[Xe]^{54}$ and $[Rn]^{86}$. For example, electronic configurations of scandium having atomic number 21 may be written as :

$$_{21}Sc : [Ar]^{18}, 3d^1, 4s^2.$$

Electronic configurations of the first ten elements on this pattern are given in Table 3.6.

Table 3.6. Electronic Configurations of Elements using Noble Gas Notation

Atomic Number	Symbol of Element	Electron Configuration	Atomic Number	Symbol of Element	Electron Configuration
1	H	$1s^1$	11	Na	$[Ne]3s^1$
2	He	$1s^2$	12	Mg	$[Ne]3s^2$
3	Li	$[He]2s^1$	13	Al	$[Ne]3s^2 3p^1$
4	Be	$[He]2s^2$	14	Si	$[Ne]3s^2 3p^2$
5	B	$[He]2s^2 2p^1$	15	P	$[Ne]3s^2 3p^3$
6	C	$[He]2s^2 2p^2$	16	S	$[Ne]3s^2 3p^4$
7	N	$[He]2s^2 2p^3$	17	Cl	$[Ne]3s^2 3p^5$
8	O	$[He]2s^2 2p^4$	18	Ar	$[Ne]3s^2 3p^6$
9	F	$[He]2s^2 2p^5$	19	K	$[Ar]4s^1$
10	Ne	$[He]2s^2 2p^6$	20	Ca	$[Ar]4s^2$

3.6.5. KLMNOPQ Notation

The electrons can be distributed in the KLMN based electron shell. The K shell is the first shell or energy level, L is the second shell, M is third, and so on. The KLMN notations indicate the total number of electrons with each principal quantum number which is '*n*'.

The total number of electrons accommodated by the energy shell is given by $2n^2$, where '*n*' is the shell number. The values of shell and the principal quantum number is tabulated as:

Shell and '*n*' value	Max.number of electron
K shell, $n = 1$	$2(1)^2 = 2$
L shell, $n = 2$	$2(2)^2 = 8$
M shell, $n = 3$	$2(3)^2 = 18$
N shell, $n = 4$	$2(4)^2 = 32$

According to **Bohr** and **Bury**:

1. Electrons fill up the lowest energy level first.
2. The maximum number of electrons in any orbit is given by the formula $2n^2$ where n is the number of the orbit, i.e., K = 2, L = 8, M = 18, N = 32, etc.
3. The outermost orbit in a stable atom cannot have more than 8 electrons even if it can accommodate more electrons according to rule 2.
4. The penultimate energy shell, i.e., the energy shell preceding the outermost shell, cannot have more than 18 electrons.

Let us apply the above rules to write electronic configuration of some elements.

I. CASE OF HYDROGEN ATOM

Atomic number of hydrogen is 1. There is only one electron in hydrogen atom which goes into the lowest energy shell i.e., K shell. Thus, the electronic configuration of hydrogen atom is:

$$\begin{array}{ccc} & & \textbf{K} \\ \textbf{H} & : & \textbf{1} \end{array}$$

Atomic diagram of hydrogen atom is shown in Fig. 3.14.

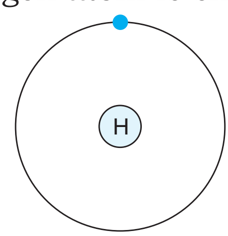


Fig. 3.14. Atomic diagram of hydrogen.

Here, circle in the centre represents the nucleus and the ring drawn around the nucleus represents the first energy shell. The dot in the ring represents the electron.

II. CASE OF OXYGEN ATOM

Now consider an atom of oxygen. Atomic number of oxygen is 8. Thus, there are eight electrons in an atom of oxygen. Two electrons are accommodated in the first shell (K-shell) and the remaining six electrons are accommodated in second shell (L-shell).

	K	L
O :	**2**	**6**

The atomic diagram of oxygen is shown in Fig. 3.15.

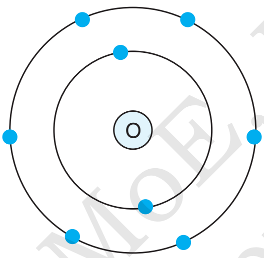


Fig. 3.15. Atomic diagram of oxygen

III. CASE OF CALCIUM ATOM

Atomic number of calcium is 20. An atom of calcium contains 20 electrons. 2 Electrons go in the first shell (K-shell), 8 electrons go to the second shell (L-shell), next 8 electrons go to the third shell (M-shell) and the remaining 2 electrons go to the fourth shell (N-shell). Although the third shell can accommodate a maximum of 18 electrons, all the 10 electrons cannot go into it because that would make 10 electrons in the outermost shell whereas the outermost shell cannot have more than 8 electrons. After the third level acquires 8 electrons, the fourth level begins to fill. The next 2 electrons go to the fourth energy shell (N-shell). Distribution of electrons of various elements has been shown in Table 3.7.

	K	L	M	N
Ca :	**2**	**8**	**8**	**2**

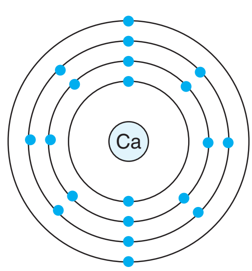
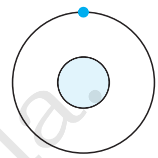
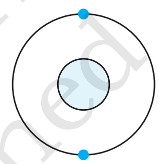
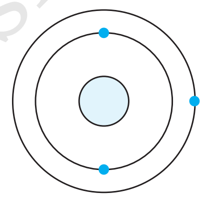
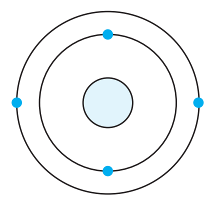
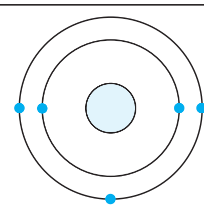
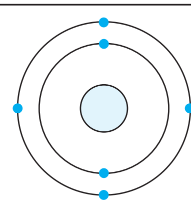
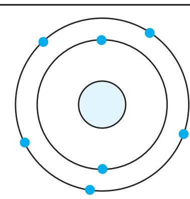
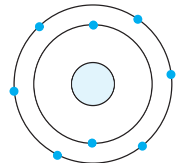
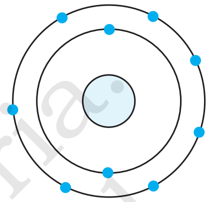
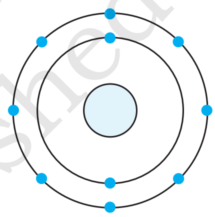
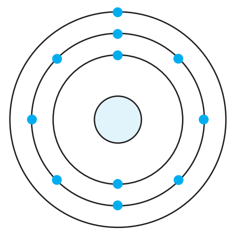
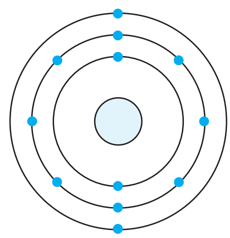
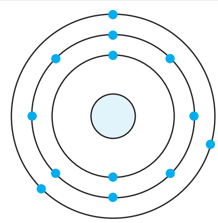
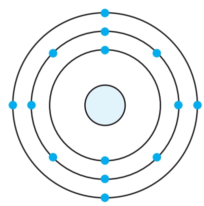
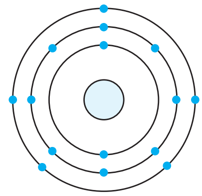
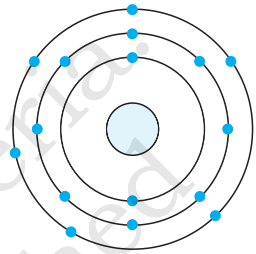
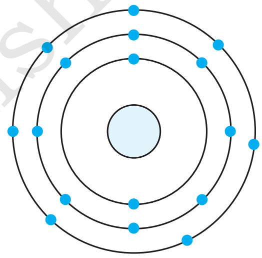
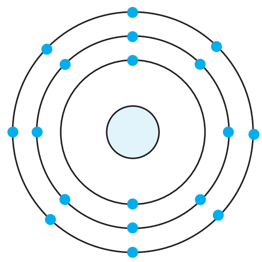
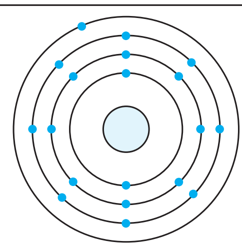
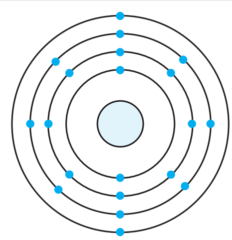


Fig. 3.16. Atomic diagram of calcium

Table 3.7. Distribution of Electrons in Various Energy Shells in the First 20 Elements

Element	Number of Electrons	Electronic Configuration	Atomic Diagram
Hydrogen	1	K 1	
Helium	2	K 2	
Lithium	3	K L 2, 1	
Beryllium	4	K L 2, 2	
Boron	5	K L 2, 3	
Carbon	6	K L 2, 4	
Nitrogen	7	K L 2, 5	

Oxygen	8	K L 2, 6	
Fluorine	9	K L 2, 7	
Neon	10	K L 2, 8	
Sodium	11	K L M 2, 8, 1	
Magnesium	12	K L M 2, 8, 2	
Aluminium	13	K L M 2, 8, 3	
Silicon	14	K L M 2, 8, 4	

Phosphorus	15	K L M 2, 8, 5	
Sulphur	16	K L M 2, 8, 6	
Chlorine	17	K L M 2, 8, 7	
Argon	18	K L M 2, 8, 8	
Potassium	19	K L M N 2, 8, 8, 1	
Calcium	20	K L M N 2, 8, 8, 2	

After the fourth shell acquires 2 electrons, the third shell begins to fill to its capacity of 18. Thus, in the atoms from atomic numbers 21–30, the electrons beyond 20th electron go to the third shell. For example, the electronic configuration of iron (Z = 26) is

		K	L	M	N
Fe	**:**	**2**	**8**	**14**	**2**

3.7. RULES FOR FILLING ORBITALS IN AN ATOM

An atom in its lowest energy state is said to be in the normal state or the *ground state*. The ground state is the most stable state for the atom. The filling of orbitals in the ground state is determined by the following *rules*:

1. AUFBAU RULE

According to Aufbau rule, ***the electrons are added progressively to the various orbitals in their order of increasing energies, starting with the orbital of lowest energy.***

Increasing order of energies of various orbitals is :

1*s*, 2*s*, 2*p*, 3*s*, 3*p*, 4*s*, 3*d*, 4*p*, 5*s*, 4*d*, 5*p*, 6*s*, 4*f*, 5*d*, 6*p*, 7*s*, ...

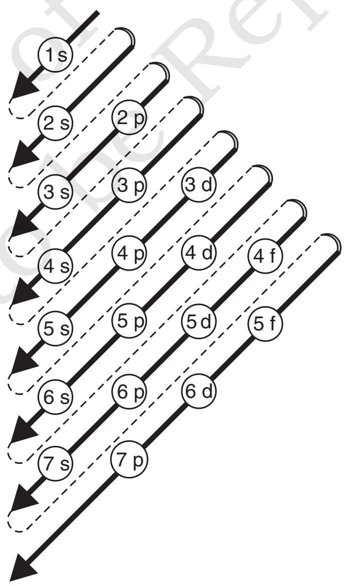


Fig. 3.17. Memory aid for remembering order of energies of various orbitals

Figure 3.17 shows a simple memory aid for remembering the increasing order of energies of various orbitals.

2. PAULI'S EXCLUSION PRINCIPLE

Pauli's exclusion principle states that ***no two electrons in an atom can have same set of all the four quantum numbers***.

From this it follows that **an orbital cannot have more than two electrons**. Moreover, if an orbital has two electrons then **they must have opposite spins**.

3. HUND'S RULE OF MAXIMUM MULTIPLICITY

Hund's rule states that ***the pairing of electrons in the orbitals of a particular sub-shell (p, d or f) does not take place until all the orbitals of the sub-shell are singly occupied. Moreover, the singly occupied orbitals must have the electrons with parallel spins***.

The basis of this rule is that two electrons in a particular orbital feel greater repulsion and hence while filling orbitals of equal energy pairing of electrons is avoided as long as it is possible. Moreover, the singly occupied orbitals should have electrons with parallel spin because this corresponds to state of lower energy. This can be explained in terms of magnetic effects of electron spin. This rule helps us in writing the ground state configurations of those atoms which have partially filled *p*, *d* or *f* sub-shells in them. The application of these rules has been illustrated in the following electronic configurations.

ADDITIONAL ACTIVITIES/EXPERIMENTS

The flame photometer

Flame photometry is one of the branches of atomic absorption spectroscopy. It is also known as flame emission spectroscopy. Currently, it has become a necessary tool in the field of analytical chemistry. Flame photometer can be used to determine the concentration of certain metal ions like sodium, potassium, lithium, calcium and cesium etc. In flame photometer spectra the metal ions are used in the form of atoms. The International Union of Pure and Applied Chemistry (IUPAC) Committee on Spectroscopic Nomenclature has named this technique as flame atomic emission spectrometry (FAES).

In this Experiment, you will calibrate a flame photometer using standard sodium and potassium solutions then measure the Na^+ and K^+concentrations in a redissolved oral rehydration sachet.

EXPERIMENT 3.1

To determine concentration of Na^+ and K^+ in solution by flame photometry. (This activity can be conducted under the supervision of your teacher.)

Reagents: oral rehydration sachet

NaCl standards: 0.25, 0.5, 1.0, 2.0, 4.0 and 5.0 mM

KCl standards: 0.1, 0.2, 0.5, 1.0, 1.5, 2.0 mM

Procedure:

1. Carefully open the oral rehydration sachet and empty the contents into a clean 250 *ml* beaker. Add about 150 *ml* distilled water and gently churn the contents until dissolved.
2. Pour the solution into a 200 *ml* volumetric flask and rinse out the beaker with small amounts of distilled water, adding the washings to the flask. Finally, make up the flask to exactly 200 *ml* and mix thoroughly.
3. Make a 1/50 dilution of the redissolved sachet solution by accurately pipetting 2 *ml* of the solution into a 100 *ml* volumetric flask and making up to 100 *ml* with distilled water.
4. Ensure that the photometer drain is leading into a sink and that the instrument is connected to gas, air and electricity supplies. Ensure the mains supply gas tap is off.
5. Turn the "Sensitivity" and instrument "Gas" controls control fully counterclockwise(towards you).
6. Insert the sodium optical filter.
7. Switch on the instrument and unclamp the galvanometer by turning counterclockwise.
8. Open the mica window, turn on the mains gas supply, light the gas and close the window.

   (**CAUTION:** Do not lean over the instrument or you will set your hair alight.)
9. Turn on the air supply control and adjust the air pressure to 10 lb/in^2. Leave for 1-2 minutes to stabilise.
10. Place a beaker of distilled water into position at the left hand side of the instrument and insert the narrow draw tube into it to allow water to pass through the photometer.

11. Adjust the gas control to give a flame with a large central blue cone then, with water passing through the instrument, slowly close the gas control until ten separate blue cones just form.
12. Set the galvanometer to zero using the “Set zero” control.
13. Replace the distilled water with the 5 mM NaCl standard and adjust the “Sensitivity” control till the galvanometer reads 100.

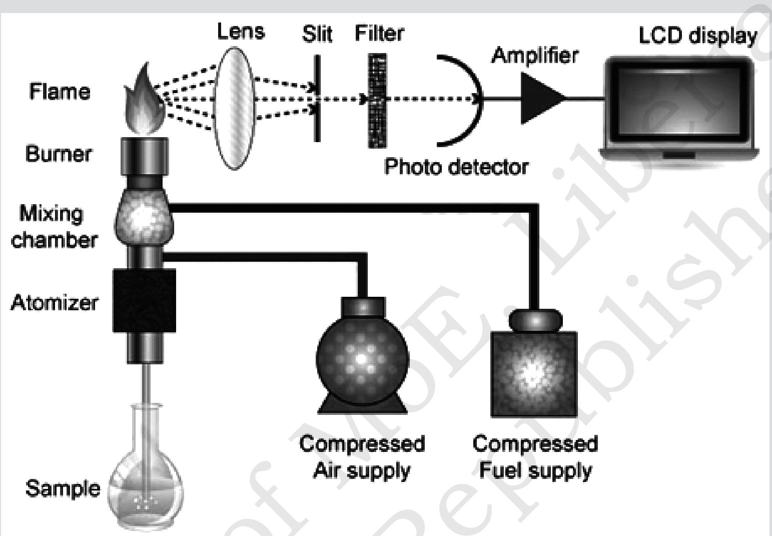


Fig. 3.18. Flame photometer

14. Quickly but carefully, replace the 5 mM NaCl standard with standards of decreasing concentration from 4 mM to 0.25 mM and note the readings in the Table below.
15. Run water through the instrument again for 1-2 min then place the draw tube into a beaker containing the 1 in 50 diluted rehydration sachet solution and note the galvanometer reading.
16. Run water through the instrument again and replace the sodium with the potassium filter.
17. Repeat the above procedure with the KCl standards, setting to 100 with 2.0 mM KCl, then reading the others in reverse order. Then read the 1 in 50 diluted rehydration sachet solution.
18. Finally, run water through the instrument until the flame appears free of colour again.

19. When the instrument is no longer required, switch off in the following sequence:
    (i) Turn off the gas control and the mains gas supply.
    (ii) Wait for the flame to die out.
    (iii) Turn off the air supply.
    (iv) Switch off the electricity.
    (v) Clamp the galvanometer.

$[Na^+]$ (mM)	**5.0**	**4.0**	**2.0**	**1.0**	**0.5**	**0.25**	**0**
Galvo. reading	100						
$[K^+]$ (mM)	**2.0**	**1.5**	**1.0**	**0.5**	**0.2**	**0.1**	**0**
Galvo. reading	100						

20. Plot the galvanometer readings against Na^+ and K^+ concentrations on the graph paper provided (separate graph for each ion) and from these calibration curves determine the Na^+ and K^+ concentrations in the diluted sachet solution. Finally, calculate the Na^+ and K^+ concentrations in the undiluted sachet solution.

	Galvanometer reading	**Diluted concentration (mM)**	**Undiluted concentration (mM)**
Sodium ion			
Potassium ion			

EXPERIMENT 3.2

Aim: *To compare thermal conductivity of metals and non-metals.*

(*This activity can be conducted under the supervision of your teacher.*)

Materials required: Copper or aluminium wire, carbon or graphite rod, spirit lamp or Bunsen burner, wax, laboratory stand and needles.

Procedure:

1. Take the copper or aluminium wire and fix the wire to a stand. Attach a pin to one of the free ends of the wire using wax, as shown in the figure.

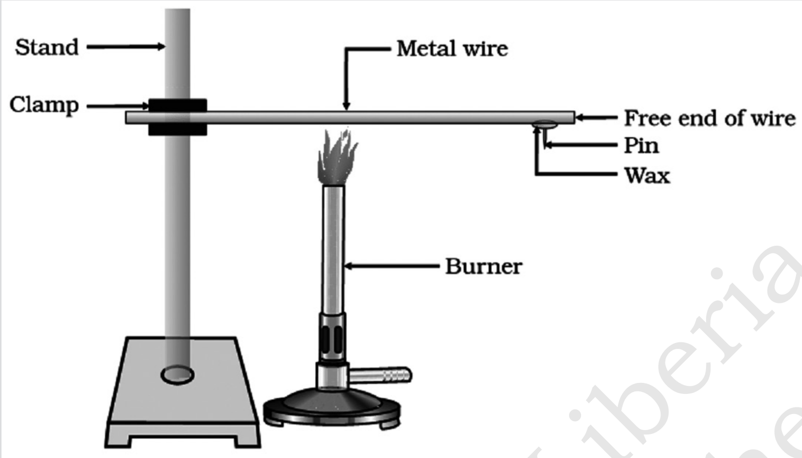


Fig. 3.19. Experimental Set-up for Comparing Thermal Conductivity of Metals and Non-metals

2. Start heating the wire from its other free end using a spirit lamp or a Bunsen burner.
3. Observe whether the wax melts or not and the pin falls or not.
4. Repeat the same procedure with a carbon or graphite rod and record your observations.

Observation:

S.No.	Sample	Observation
1	Copper or aluminium wire	Pin drops
2	Carbon or graphite rod	Pin doesn't drop

Explanation: Metals are a good conductor of heat. Here, once we start heating, heat transfers to the area of wax. It melts the wax. So, the pin drop on the table.

Conclusion: This experiment demonstrates that metals are a good conductor of heat.

GLOSSARY

- **Alpha Particles:** He^{2+} ions or helium nuclei.
- **Atomic Number:** The number of protons present in the nucleus of an atom.
- **Electron:** A sub-atomic particle carrying one unit negative charge and having negligible mass (9.1×10^{-31} kg).

- **Isotopes:** Atom of an element having different mass numbers.
- **Isotopic Mass:** Mass of an isotope relative to the mass of an atom of C-12 taken as 12 u.
- **Mass Number:** The total number of protons and neutrons present in the nucleus an atom.
- **Mass Spectrometer:** An instrument used to find isotopic masses, relative isotopic abundance and relative molecular masses.
- **Neutron:** A neutral sub-atomic particle having mass almost equal to that of a hydrogen atom.
- **Nucleus:** Heavy positively charged centre of an atom. It contains protons and neutrons.
- **Proton:** A sub-atomic particle carrying one unit positive charge and having mass nearly equal to the mass of an atom of hydrogen.
- **Relative Atomic Mass (RAM):** It is the weighted average of isotopic masses of different isotopes of the element.
- **Specific Charge:** e/m ratio for a charged particle.

SUMMARY

- In the fifth century B.C. the Greek Philosopher Democritus proposed that all matter consists of very small indivisible particles called atoms (meaning uncuttable or indivisible).
- The protons are positively charged particles and are present in the nucleus of the atom.
- The electrons are negatively charged particles and are present in the extra-nuclear part of the atom.
- The neutrons are neutral particles and are present in the nucleus along with the protons. Most of the mass of the atom is concentrated in the nucleus.
- Cathode rays consist of negatively charged material particles called electrons.
- Electrons are the fundamental subatomic particles carrying negative charge (-1.602×10^{-19} C) and having mass 9.1×10^{-31} kg. Discovered by J.J. Thomson.
- Charge to mass (e/m) ratio for electrons is 1.76×10^{8} C/g or 1.76×10^{11} C/kg.
- Alpha Particles are He^{2+} ions or helium nuclei.

- Rutherford's Experiment led to the discovery of nucleus. Radius of nucleus ($\sim 10^{-14}$ m) is very small as compared with radius of atom ($\sim 10^{-10}$ m).
- Atomic Number (Z) (Proton Number) the number of protons present in the nucleus of an atom.
- Mass Number (A) (Nucleon Number) the total number of protons and neutrons present in the nucleus of an atom.
- Isotopes are the atoms of an element having different mass numbers.
- Relative Atomic Mass (RAM) of an element is the weighted average of isotopic masses of different isotopes of the element.
- An orbital diagram is a way of representing the electronic configuration of an atom.

EVALUATION

I. Multiple Choice Questions

1. Which of these are negatively charged particles?

(a) Electrons (b) Protons

(c) Neutrons (d) None of these

2. Name the subatomic particles that are present in the nucleus of an atom.

(a) Electrons and protons (b) Electrons and neutrons

(c) Protons and neutrons (d) None of these

3. Who discovered electrons?

(a) John Dalton (b) J.J. Thomson

(c) James Chadwick (d) Ernest Rutherford

4. Who determined the charge to mass ratio of an electron?

(a) Robert Milikan (b) Robert Boyle

(c) J.J. Thomson (d) None of these

5. Who discovered nucleus?

(a) John Dalton (b) J.J. Thomson

(c) James Chadwick (d) Ernest Rutherford

6. Who discovered proton?

(a) J.J. Thomson (b) James Chadwick

(c) E. Goldstein (d) Ernest Rutherford

7. This is an isotope of hydrogen having mass number 2.

(a) Protium (b) Deuterium

(c) Tritium (d) None of these

8. The total number of protons and neutrons in the nucleus is called _______ .

(a) atomic number (b) mass number

(c) Both (a) and (b) (d) None of these

II. State True or False

1. Protons are negatively charged particles.

2. Dalton's atomic theory explains the law of chemical combination by mass.

3. Mass of an electron is nearly equal to $\frac{1}{1840}$th of mass of an atom of hydrogen.

4. An atom as a whole is electrically neutral.

5. Neutrons are experimentally discovered by James Chadwick.

6. Atomic number of an element is equal to the number of neutron in the nucleus of its atom.

7. Mass number is different from relative atomic mass.

III. Answer the Following Questions

1. Write down the brief history of atomic chemistry.

2. Write down the main points of Dalton's atomic theory.

3. Briefly explain the discovery of electrons.

4. Briefly explain the discovery of nucleus.

5. Briefly explain the Bohr's model of hydrogen atom.

6. What do you mean by isotopes? Give two examples of isotopes.

7. What do you mean by quantum number? What are the four kinds of quantum numbers?

8. Explain the rules and principles for filling in electrons.