4 ATOMIC STRUCTURE AND CHEMICAL BONDING

SYLLABUS

(i) Structure of an Atom mass number and atomic number, Isotopes and Octet Rule
- *Definition of an atom;*
- *Constituents of an atom – nucleus (protons, neutrons) with associated electrons; mass number, atomic number.*
- *Electron distribution in the orbits – $2n^2$ rule. Octet rule. Reason for chemical activity of an atom.*
- *Definition and examples of isotopes (hydrogen, carbon, chlorine).*

(ii) Electrovalent and covalent bonding. Structures of various compounds — orbit structure.
- (a) Electrovalent bond : • *Definition* • *Atomic orbit structure for the formation of electrovalent compounds NaCl, $MgCl_2$, CaO*
- (b) Covalent bond : • *Definition* • *Atomic orbit structure for the formation of covalent molecules on the basis of duplet and octet of electrons (examples: hydrogen, chlorine, oxygen, nitrogen, hydrogen chloride, water, ammonia, carbon tetrachloride, methane).*

INTRODUCTION

Many attempts have been made to know as to what is the ultimate particle of matter. The idea of the smallest unit of matter was first given by Maharshi Kanada in the 6th Century B.C. in India. According to him, matter consisted of indestructible minute particles called *paramanus* (*param* means ultimate and *anu* means particle) now called *atoms*". A *paramanu* does not exist in free state, rather it combines with other *paramanus* to form a bigger particle called *anu* (now known as a *molecule*). There are different types of *paramanus*. Each one of them exhibits specific properties.

The Greek philosopher Democritus (460 B.C. – 370 B.C.) called the *paramanu* an "atom", which comes from the Greek word, meaning *indivisible*. But the first scientific theory about the structure of matter was given by **John Dalton** (in 1808). It considered atoms as indivisible particles that are the fundamental building blocks of matter. The existence of different types of matter is considered to be due to different types of atoms constituting them.

The main postulates of Dalton's atomic theory are:

1. Matter consists of very small and indivisible particles called atoms.
2. Atoms can neither be created nor be destroyed.
3. The atoms of an element are alike in all respects but they differ from the atoms of other elements.
4. Atoms of an element combine in small numbers to form molecules.
5. Atoms of one element combine with atoms of another element in a simple ratio to form molecules of compounds.
6. Atoms are the smallest units of matter that can take part in a chemical reaction.

Note : *The latest research on atoms has proved that most of the postulates of Dalton's atomic theory are incorrect except that atoms take part in chemical reactions.*

One of the first indications that atoms consist of subatomic particles, *i.e.* they are made up of charged particles has been put forward by the studies of static electricity* **and the conditions under which electricity is conducted by different substances**. In 1833, Faraday showed that the flow of electricity is due to the flow of charged particles. G.J. Stoney (in 1874) first suggested the name **electron** for this electrical particle. However, it was *J. J. Thomson who showed the existence of these charged particles and thus discovered electrons, negatively charged particles.* **E. Goldstein,** a German scientist **discovered protons**, positively charged particles in atom. Later on, **James Chadwick discovered neutrons**. The electrons, *protons and neutrons are called the fundamental particles of an atom. These particles are the building blocks of all atoms of an element.*

* Electrification of objects by rubbing with each other is called static electricity or frictional electricity.

4.1 DEFINITION OF AN ELEMENT

An element is a substance which is made up of only one type of atoms.

Element is a pure substance that can neither be decomposed into, nor formed from simple substances by ordinary physical or chemical methods.

For example : Carbon is an element because it cannot be split up into two or more simpler substances by ordinary methods like heating, breaking or passing electricity.

Note : Radioactivity* the processes of radioactive decay and high energy nuclear reaction can transfer one elementary substance into another.

4.2 DEFINITION OF AN ATOM

*"An atom is the smallest particle of an element that exhibits all the properties of that element. It **may or may not exist** independently but takes part in every chemical reaction.*

Example : Take a small piece of zinc and grind it into smaller pieces. All these pieces show properties of zinc. On grinding them further they break up into very fine particles which still show the properties of zinc. But, there comes a stage when the particles cannot be further subdivided into particles exhibiting properties of zinc. These indivisible particles are the atoms of zinc.

"In other words an atom is the smallest possible unit of an element".

4.3 CONSTITUENTS OF AN ATOM

Discovery of electrons

William Crooke, a British scientist, noted that gases are ordinarily poor conductors of electricity. However, when a high voltage (10,000 volts) charge from an induction coil is applied to tubes filled with gases at very low pressure (0.01 mm Hg), the gases become good conductors of electricity and begin to flow from cathode to anode in the form of rays. Since these rays originate from the negative plate, *i.e.* the cathode and travel from *the cathode towards the anode,* they are called **cathode rays**.

* The phenomenon due to which certain elements spontaneously emit highly penetrating rays called radiation.

Later **J.J. Thomson** studied the characteristics and the constituents of cathode rays. The apparatus used by him is called a **discharge tube** or **a cathode ray tube** shown in Fig. 4.1.

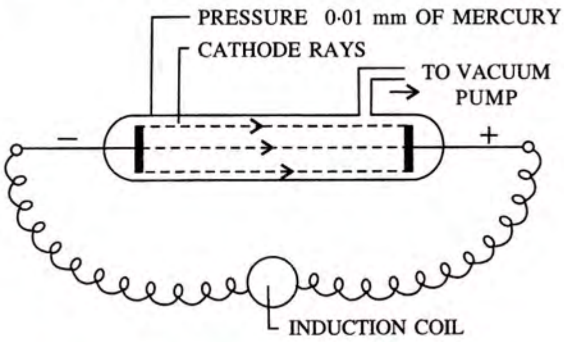


Fig. 4.1 Discharge tube with flowing electrons

A discharge tube is a hard glass tube fitted with two metal plates known as electrodes, one of which is connected to the positive terminal of battery and is called anode (positive electrode) while the other electrode which is connected to the negative terminal of the battery is called cathode (negative electrode). It has a side tube through which gas can be pumped out by using a vacuum pump to create vacuum.

When electrical discharge of 10,000 volts is passed through gases at very low pressures (0·01 mm), cathode rays are produced.

Properties of Cathode Rays

1. They travel from the cathode to the anode in *straight lines*.
2. They cause a *greenish yellow fluorescence* on a soda-glass screen placed in the tube.
3. They are affected by electric field, *i.e.* they are *inflected* towards the positive field and *deflected* from the negative field. This shows that *they carry negative charge*. (Fig. 4.2).

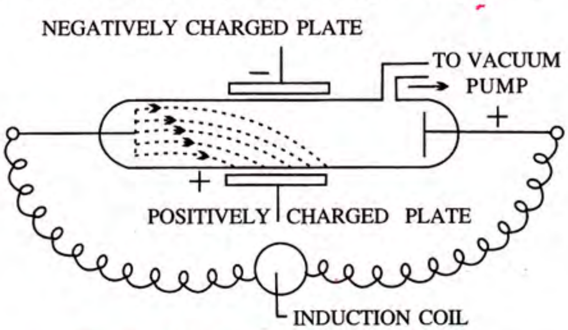


Fig. 4.2 Deflection of cathode rays.

4. When a beam of cathode rays is made to fall upon hard metallic targets like tungsten, X-rays are produced.
5. They penetrate through matter.
6. They cause ionization of the gas through which they pass.
7. The ratio of the charge(e) to mass(m) of the particles constituting cathode rays remains the same ($e/m = 1{\cdot}76 \times 10^{11}$ coulomb/kg) irrespective of the nature of the gas taken and of the metal forming the cathode.
8. They produce the shadow of an opaque object placed in their path and make a light paddle wheel rotate.

Thomson concluded that :

(i) *cathode rays consist of* ***negatively charged*** *particles, now called* **electrons.**

[The name 'electron,' meaning '*atom of negative electricity,*' was given by Johnson Stoney].

(ii) these negatively-charged particles are an integral part of all atoms.

(iii) electrons have both definite mass and definite electric charge, both of which are independent of the nature of the gas in the discharge tube.

Properties of Electrons

1. Electrons from all sources are alike, having identical mass.
2. They are a constituent part of all atoms.
3. The mass of an electron is 1/1837 the mass of a hydrogen atom (or $9{\cdot}108 \times 10^{-31}$ kg).
4. An electron carries unit negative charge of magnitude $-1{\cdot}602 \times 10^{-19}$ coulombs.
5. The electron is extremely small; its radius is less than 1×10^{-15} m.

> An electron may be defined as a subatomic particle having a unit negative charge and a mass equal to 1/1837 of hydrogen atom. It is denoted by the symbol $_{-1}e^{0}$. The superscript 0 represents its mass and subscript –1 represents its electrical charge.

Atoms are found to be electrically neutral, so they must contain, in addition, particles that are positively charged, such that the total negative charge of the electrons is equal to the total positive charge. This realization led to the discovery of positively charged subatomic particles **protons**.

Note : When a high voltage current is passed through a gas taken in the discharge tube, electrical energy breaks up the atoms of the gas into negatively charged particles (electrons) and **positively charged particles (protons) are formed by the removal of one or more electrons from the gaseous atoms.**

4.4. DISCOVERY OF PROTONS

Goldstein noticed another set of rays travelling in a direction opposite to that of the cathode rays, *i.e.* from anode towards cathode, when a perforated cathode was used in the above discharge tube (shown in Fig. 4.3). He called these rays as **canal rays** since these rays passed through holes or 'canals' in the cathode. These rays were named as **positive rays** or **anode rays**.

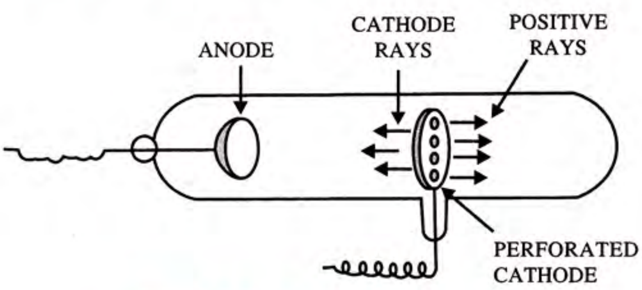


Fig. 4.3 Production of anode rays

Properties of Anode Rays

1. Anode rays travel in a straight line.
2. They consist of minute material particles and hence produce mechanical effects.
3. They are made up of **positively charged** particles.
4. Positive rays are deflected by electric and magnetic fields but in a direction opposite to that of the cathode rays. This means that these rays consist of positively charged particles called the **protons**.
5. These rays produce fluorescence on a zinc sulphide screen.
6. Their e/m *i.e.* charge to mass ratio, differs from gas to gas. Its value is much less than that of an electron and is maximum when hydrogen is taken in the discharge tube.

Proton may be defined as a subatomic particle having mass 1 amu, *i.e.* equal to hydrogen atom and has unit positive charge.

It is denoted as $_{+1}P^1$, the superscript 1 represents its mass 1 amu and the subscript +1 represents one unit positive charge.

Proton is formed by the loss of an electron from a hydrogen atom.

$$\underset{\text{(hydrogen atom)}}{H} - e^- \rightarrow \underset{\text{(proton)}}{H^+}$$

Properties of Protons

1. A proton possesses a unit positive charge **+1** (plus one) of the value $1 \cdot 602 \times 10^{-19}$ coulombs.
2. Its mass is the same as that of a hydrogen atom (**1 a.m.u.**, *i.e.*, 1837 times the mass of an electron, which is $1 \cdot 672 \times 10^{-24}$ g).
3. The proton resides in the central part of an atom, *i.e.* in the nucleus.

After the discovery of electrons and protons **J.J. Thomson** proposed **'plum pudding'** model of the atom.

According to this model :

1. An atom is considered to be a sphere of uniform positive charge and electrons are embedded into it.
2. The total positive charge is equal to the total negative charge so that an atom as a whole is electrically neutral.
3. The mass of an atom is considered to be uniformly distributed.

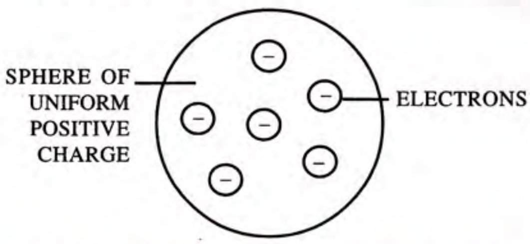


Fig. 4.4 J.J. Thomson's plum-pudding model of the atom.

4.5 DISCOVERY OF NUCLEUS

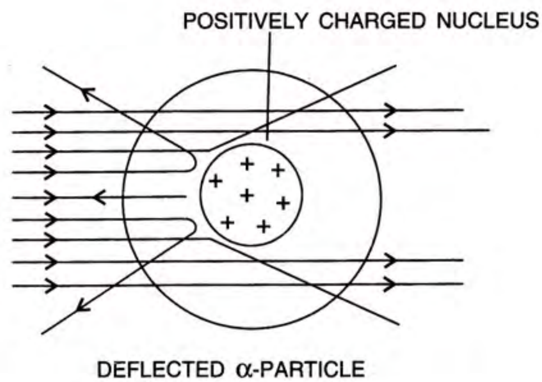
In 1911, Lord Rutherford, a scientist from New Zealand, directed a stream of alpha-particles* towards a very thin (one millionth of a centimetre) gold foil. He selected a gold foil because he wanted as thin a layer as possible and gold is the most malleable metal.

He **observed** that :

(1) Most of the alpha particles passed straight through the foil.

(2) Some alpha particles were slightly deflected from their straight path.

(3) Very few (nearly one in ten thousand) alpha particles were either deflected by very large angles or completely bounced back.



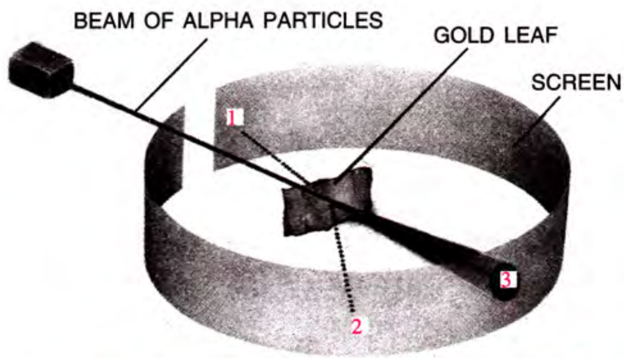


Fig. 4.5 Deflection of α particle by a gold leaf.

1. bounced back; 2. deflected slightly; 3. straight through

Later, Rutherford generalised these results of *alpha particles scattering experiment* and suggested a model of the atom that is known as **Rutherford's Atomic Model.**

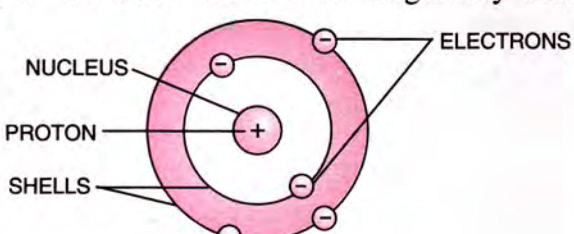
Note : Heavy metals such as platinum will show the same observation with alpha (α) particles as shown by gold foil, but if light nuclei like lithium is used then fast moving α-particles may even push the light nucleus aside and may not be deflected back.

* An **α-particle** is a doubly charged helium ion He^{2+}, containing two protons and two neutrons. It is formed by removing 2 electrons from helium atom.

4.5.1 Rutherford's atomic model

According to this model :

1. *The atom contains a large empty space (extra-nuclear space)*. This is why most of the α-particles pass through the metal foil without deviating from their path.
2. *There is a positively charged mass at the centre of the atom, known as the* **nucleus,** in which the entire mass of the atom is concentrated. Therefore, the nucleus is the densest part of the atom. The electrically positive nature of the nucleus is supported by the fact that the positively charged α-particles approaching the nucleus get deflected. This proved that the nucleus contained protons.
3. *The size of the nucleus is very small* compared to the size of an atom. This is proved by the fact that most of the α-particles pass straight through the gold foil.
4. *The electrons revolve around the nucleus at large distances from each other and from nucleus* (in circular orbits found in the empty space of the atom).
5. *An atom as a whole is electrically neutral*, because the total positive charge of the nucleus is balanced by the total negative charge of the electrons, *i.e.*, the number of protons and electrons in an atom are equal.
6. Electrons revolve around the nucleus in close circular paths (ORBITS). The force of attraction between the negatively charged electrons and the positively charged nucleus is counter balanced by the centrifugal force acquired by the revolving electrons.



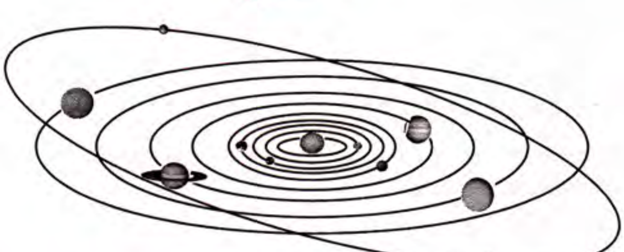


Fig. 4.6 Rutherford's model of the atom was somewhat like that of the solar system

Rutherford's model of atomic structure is similar to the structure of the solar system. Just as in the solar system, the Sun is at the centre (having the maximum mass) and the planets revolve around it, similarly in an atom the nucleus contains the main mass and the electrons revolve around it in orbits or shells.

4.5.2 Drawback of Rutherford's atomic model

The comparison of electrons with planets in the solar system is the main drawback of Rutherford's atomic model. According to the classical laws of mechanics and electrodynamics, if an electrically charged particle is in motion, it inevitably radiates energy. Thus, an electron, when moving round the nucleus continually, should radiate energy, *i.e.* loses energy. As a result, it should be gradually pulled towards the nucleus and end up colliding with it. This should result in the total collapse of the atom (Fig. 4.7).

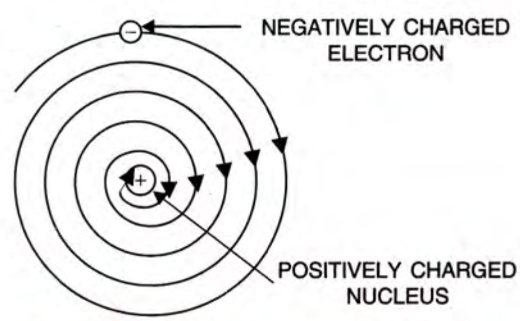


Fig. 4.7 Showing an atom losing energy

If it was so, the atom should be highly unstable and hence matter would not exist in the form that we know. However, we know that an atom is structurally stable.

Thus *Rutherford's model failed to explain the stability of an atom.*

4.6 BOHR'S ATOMIC MODEL

In 1913, Niels Bohr, a Danish physicist, explained the causes of the **stability of the atom** in a different manner.

Main postulates of Bohr's atomic model

(1) The electrons revolving around the nucleus are confined to certain fixed orbits called shells or

energy levels, each of which is associated with a fixed amount of energy.

(2) While it is revolving around the nucleus in an orbit, an electron neither loses nor gains energy.

(3) An electron revolving in a particular orbit, on gaining a certain amount of energy, jumps to the next orbit and vice versa. *Since each orbit is associated with a fixed amount of energy, Bohr called it an **energy level**.*

For convenience, these energy levels are labelled K, L, M, N or I, II, III, IV, etc. The orbit closest to the nucleus is the K shell. It has the least amount of energy and the electrons present in it are called K electrons, and so on with the successive shells and their electrons.

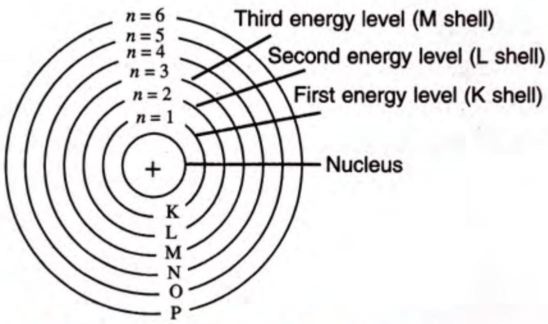


Fig. 4.8 Energy levels around the nucleus of an atom : Bohr's Model.

4.7 DISCOVERY OF NEUTRONS

By now we know that an atom contains electrons and protons, and that the atomic mass of an electron is negligible. Therefore, an atom of helium, which contains 2 protons should have a

$$\text{mass} = 2 \times 1 \text{ a.m.u.} = 2 \text{ a.m.u.}$$

But the atomic mass of a helium atom was found to be approximately 4·0 a.m.u. It was, therefore, proposed that, in the nucleus of an atom, there must be another particle. This particle should not possess any electrical charge and must be equal in mass to the proton.

In 1932 Chadwick discovered these particles by bombarding light nuclei like beryllium with *alpha (α) particles *i.e.* helium nuclei.

$$_4Be^9 + {_2He^4} (\alpha) \rightarrow {_6C^{12}} + \underset{\text{(neutron)}}{_0n^1}$$

These particles are found to be neutral, so named **neutrons**.

A **neutron** is a sub-atomic particle or fundamental particle of an atom with no charge and mass almost equal to the mass of the proton *i.e.* hydrogen atom. Neutron is denoted by $_0n^1$. The superscript 1 represents its mass and subscript 0 represents its electrical charge.

Properties of Neutrons

1. This particle was not found to be deflected by any magnetic or electric field, proving that it is electrically neutral.
2. Its mass is equal to $1{\cdot}676 \times 10^{-24}$ g (**1 amu**).

After the discovery of **electrons, protons** and **neutrons** (subatomic particles) it was found that all atoms have a similar following basic structure.

- Electrons are negatively charged particles that are found outside the nucleus.
- Protons are positively charged particles that are found in the nucleus of an atom.
- Neutrons are electrically neutral particles that are also found in the nucleus. Neutrons are slightly heavier than protons.

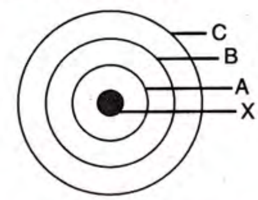
Table 4.1 : Comparison of Dalton's atomic theory with the modern atomic theory

Dalton's atomic theory	***Modern atomic theory***
(i) Atoms are indivisible particles.	(i) Atoms are divisible into sub-atomic particles like protons, electrons and neutrons.
(ii) Atoms can neither be created nor destroyed.	(ii) Atoms can be created and destroyed by nuclear fusion and fission.
(iii) The atoms of an element are alike in all respects, but they differ from the atoms of other elements.	(iii) The atoms of an element may not be alike in all respects, as is seen in the case of isotopes.
(iv) Atoms follow the laws of chemical combination to form compounds.	(iv) In the formation of organic compounds, the laws of chemical combination are not always followed.

* An alpha particle consists of two protons and two neutrons. It is same as Helium atom that has lost its both electrons.

EXCERCISE 4(A)

1. What is the contribution of the following in Atomic structure.
   (a) Maharshi Kanada (b) Democritus
2. State Dalton's atomic theory.
3. What is an α (alpha) particle ?
4. What are cathode rays? How are these rays formed?
5. (a) What is the nature of charge on cathode rays ?
   (b) State the properties of cathode rays.
6. How are X-rays produced ?
7. Why were anode rays also called as 'canal rays' ?
8. How does cathode rays differ from anode rays ?
9. State one observation which shows that atom is not indivisible.
10. (a) Name an element which does not contain neutron.
    (b) If an atom contains one electron and one proton, will it carry any charge or not ?
11. On the basis of Thomson's model of an atom, explain how an atom as a whole is neutral.
12. Which sub-atomic particle was discovered by
    (a) Thomson (b) Goldstein (c) Chadwick.
13. Name the sub-atomic particle whose charge is :
    (a) +1 (b) –1 (c) 0.
14. (a) Which metal did Rutherford selected for his α particle scattering experiment and why ?
    (b) What do you think would be the observation of α-particle scattering experiment if carried out on (i) heavy nucleus like platinum (ii) light nuclei like lithium.
15. On the basis of Rutherford's model of an atom, which subatomic particle is present in the nucleus of an atom ?
16. Which part of atom was discovered by Rutherford.
17. How was it shown that atom has empty space ?
18. State one major drawback of Rutherford's model.
19. In the figure given alongside :



    (a) Name the shells denoted by A, B and C. Which shell has least energy ?
    (b) Name X and state the charge on it.
    (c) The above sketch is of .... model of an atom.
20. Give the postulates of Bohr's atomic model.

4.8 ATOM — ITS STRUCTURE

The atom is built up of a number of sub-atomic particles*. The three sub-atomic particles of great importance in understanding the structure of an atom are *electrons, protons and neutrons*, the properties of which are given in Table 4.2.

Table 4.2 Properties of sub-atomic particles

Particle	*Symbol*	*Charge (electronic unit)*	*mass (amu)*	*mass (grams)*
1. Electron	${}_{-1}e^0$ or e^-	–1	1/1840	9.1×10^{-28} g
2. Proton	${}_1^1H$ or p^+	+1	1	1.6×10^{-24} g
3. Neutron	${}_0^1n$ or n	0	1	1.6×10^{-24} g

There are two structural parts of an atom, the nucleus and the empty space in which there are imaginary paths called **orbits**.

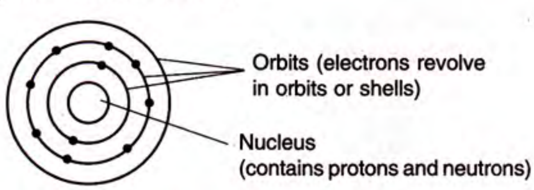


Fig. 4.9 Showing structure of an atom

Nucleus : The protons and neutrons [collectively called **nucleons**] are found in the central part or nucleus of the atom.

Orbits : It is the imaginary path where electrons revolve around the nucleus of the atom.

4.9 ATOMIC NUMBER [Z]

- An atom of an element has its own characteristic number of protons in its nucleus, which distinguishes it from the atoms of other elements.
- This characteristic number (number of protons) is called the *atomic number* of the element.
- Atomic number is denoted by Z.
- The atomic number is, therefore, equal to the *number of electrons* in the atom of an element as well, since in a neutral atom,

  no. of protons = no. of electrons.
- Number of protons gives the total positive charge present in the nucleus of an atom.

The atomic number of an element is the number of :
- protons present in the nucleus of its atom.
- electrons present in its neutral atom.
- positive charge in the nucleus of its atom.

* Modern theories about the atom include the existence of over 200 subatomic particles such as positron, neutrino, etc.

Symbolic representation of Element	*Atomic Number* [Z]	*Mass Number* [A]
$^{A}_{Z}X$ X denotes the symbol of the element Superscript A = Mass number Subscript Z = Atomic number	Atomic number [Z] = no. of protons [p] = no. of electrons [e] $\therefore$ Z = p = e	Mass number [A] = no. of neutrons [n] + no. of protons [p] $\therefore$ A = n + p [$\because$ p = Z] $\therefore$ no. of neutrons [n] = A – Z
Example : $^{35}_{17}Cl$	Atomic number [Z] = p = e = 17 Mass number [A] = n + p = 35 No. of neutrons [n] = A – Z = 18	

Mass Number [A]

- Mass number of an element is the *total number of protons and neutrons* present in its nucleus. Mass number is denoted by A.
- Mass number, however, is a *whole number approximation* of the atomic mass calculated in atomic mass units.

4.10 DISTRIBUTION OF ELECTRONS IN THE ORBITS — BOHR-BURY SCHEME

Distribution of electrons in different shells is called **electronic configuration**.

Distribution of electrons into different shells (orbits) of an atom was suggested by Bohr and Bury.

The following rules are followed for writing the number of electrons in different energy levels or shells.

1. The **maximum possible number of electrons** in a particular shell is given by the formula **$2n^2$**, where *n* denotes the serial number of that shell.

   First or **K shell** has (2×1^2) = 2 electrons
   Second or **L shell** has (2×2^2) = 8 electrons
   Third or **M shell** has (2×3^2) = 18 electrons
   Fourth or **N shell** has (2×4^2) = 32 electrons
   and so on.

2. Electrons are not accomodated in a given shell, unless the inner shells are filled. In other words the **shells are filled in stepwise manner**.

3. *The outermost shell of an atom cannot accommodate more than 8 electrons*, even if it has the capacity to accommodate more electrons. For example electronic arrangement in Calcium having 20 electrons is

K	L	M	N
2	8	8	2

   The maximum capacity of *M* shell is 18 electrons but only 8 electrons are filled in it. This is due to the fact that the elements become stable only after acquiring 8 electrons in the outermost shell.

4. The outermost shell of a chemically stable atom can accommodate a maximum of 8 electrons, *except Helium atom, which has only one shell, and can thus accommodate a maximum of 2 electrons only.*

Structure of Magnesium atom

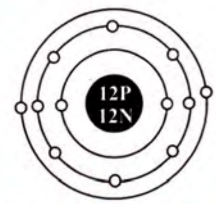


Fig. 4.10 The structure of the magnesium atom

Magnesium has atomic number 12 and atomic mass 24. Hence its nucleus has 12 protons and (24 – 12) = 12 neutrons. It is surrounded by 12 electrons that are allotted to different shells as follows.

K shell or I shell = 2 electrons
L shell or II shell = 8 electrons
M shell or III shell = 2 electrons

So its electronic configuration is 2, 8, 2.

In lighter elements, up to argon (atomic number 18), each inner shell is completely filled before any electron can occupy an outer shell. However, in elements heavier than argon, the situation changes. Although the third shell can accommodate up to 18 electrons, yet the fourth shell begins to be filled after it has only 8 electrons.

For example, potassium (atomic number 19) has 19 electrons. Its electronic configuration is :

K	L	M	N
2	8	8	1

ORBITAL DIAGRAMS OF ATOMS OF THE FIRST TWENTY ELEMENTS

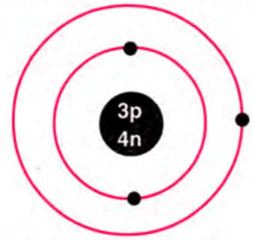
Atom of Hydrogen $^{1}_{1}H$

Electronic Config. 1
Valency = 1

Atom of Helium $^{4}_{2}He$

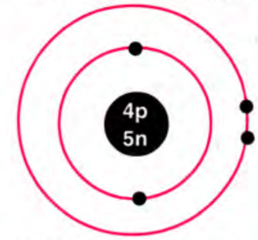
Electronic Config. 2
Valency = 0

Atom of Lithium $^{7}_{3}Li$



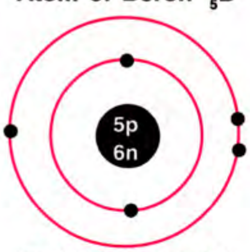
Electronic Config. 2,1
Valency = 1

Atom of Beryllium $^{9}_{4}Be$



Electronic Config. 2,2
Valency = 2

Atom of Boron $^{11}_{5}B$

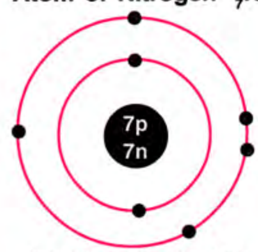


Electronic Config. 2,3
Valency = 3

Atom of Carbon $^{12}_{6}C$

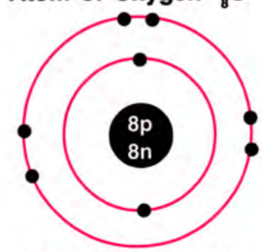
Electronic Config. 2,4
Valency = 4

Atom of Nitrogen $^{14}_{7}N$



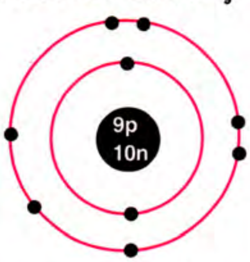
Electronic Config. 2,5
Valency = –3

Atom of Oxygen $^{16}_{8}O$



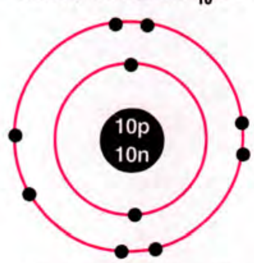
Electronic Config. 2,6
Valency = –2

Atom of Fluorine $^{19}_{9}F$



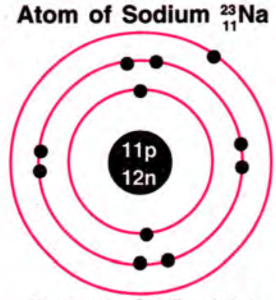
Electronic Config. 2,7
Valency = –1

Atom of Neon $^{20}_{10}Ne$



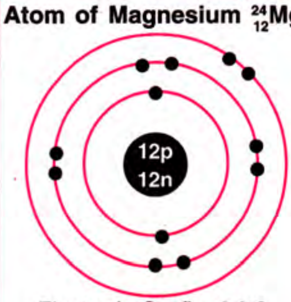
Electronic Config. 2,8
Valency = 0

Atom of Sodium $^{23}_{11}Na$



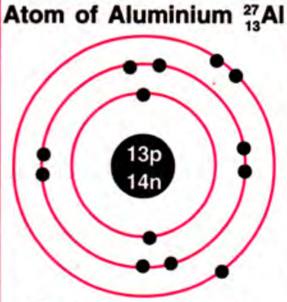
Electronic Config. 2,8,1
Valency = 1

Atom of Magnesium $^{24}_{12}Mg$



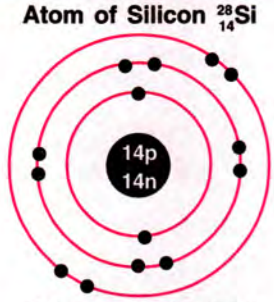
Electronic Config. 2,8,2
Valency = 2

Atom of Aluminium $^{27}_{13}Al$



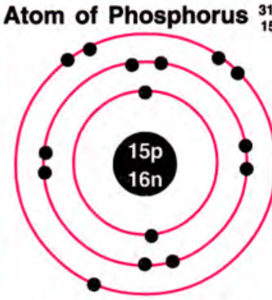
Electronic Config. 2,8,3
Valency = 3

Atom of Silicon $^{28}_{14}Si$



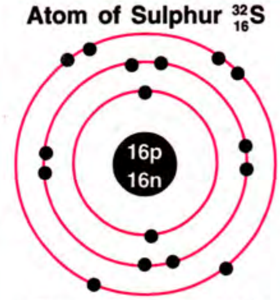
Electronic Config. 2,8,4
Valency = 4

Atom of Phosphorus $^{31}_{15}P$



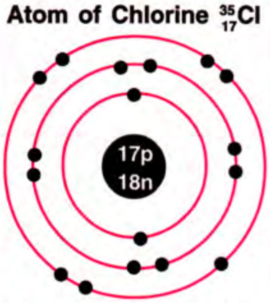
Electronic Config. 2,8,5
Valency = –3

Atom of Sulphur $^{32}_{16}S$



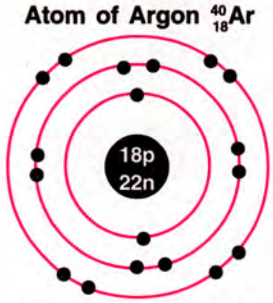
Electronic Config. 2,8,6
Valency = –2

Atom of Chlorine $^{35}_{17}Cl$



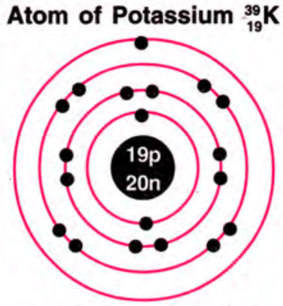
Electronic Config. 2,8,7
Valency = –1

Atom of Argon $^{40}_{18}Ar$



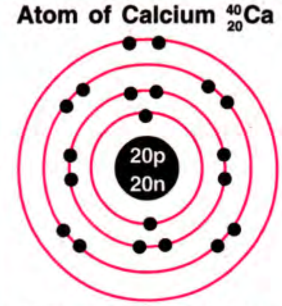
Electronic Config. 2,8,8
Valency = 0

Atom of Potassium $^{39}_{19}K$



Electronic Config. 2,8,8,1
Valency = 1

Atom of Calcium $^{40}_{20}Ca$



Electronic Config. 2,8,8,2
Valency = 2

Main features of the structure of atoms :

(i) *Atoms of all the elements (except hydrogen) are made up of three fundamental (sub-atomic) particles: electrons, protons and neutrons. (Hydrogen atom lacks neutrons).*

(ii) *The nucleus is located at the centre of the atom. It contains protons and neutrons, which account for the total mass of that atom.*

(iii) *The nucleus is positively charged due to the presence of protons in it.*

(iv) *The electrons are outside the nucleus and have negligible mass.*

(v) *The number of electrons in an atom is equal to the number of protons in it, hence the atom is electrically neutral.*

(vi) *The electrons revolve rapidly in fixed circular paths around the nucleus. These circular paths are called energy levels or shells or orbits.*

(vii) *The atoms of different elements contain different numbers of electrons, protons and neutrons.*

Table 4.3 Arrangement of electrons in atoms of elements having atomic number from 1 to 20

Elements	*Symbol*	*Atomic No. (Z)* No. of protons/ No. of electrons	*Mass No. (A)* No. of protons + neutrons	No. of neutrons (A – Z)	*Distribution of Electrons* 1st or K-shell	2nd or L-shell	3rd or M-shell	4th or N-shell	*Valency*	*Metal / non-metal/ noble gas*
Hydrogen	H	1	1	—	1				1	Non–metal
Helium	He	2	4	2	2				0	Noble gas
Lithium	Li	3	7	4	2	1			1	Metal
Beryllium	Be	4	9	5	2	2			2	Metal
Boron	B	5	11	6	2	3			3	Non–metal (solid)
Carbon	C	6	12	6	2	4			4	Non–metal (solid)
Nitrogen	N	7	14	7	2	5			3	Non–metal (gas)
Oxygen	O	8	16	8	2	6			2	Non–metal (gas)
Fluorine	F	9	19	10	2	7			1	Non–metal (gas)
Neon	Ne	10	20	10	2	8			0	Noble gas
Sodium	Na	11	23	12	2	8	1		1	Metal (solid)
Magnesium	Mg	12	24	12	2	8	2		2	Metal (solid)
Aluminium	Al	13	27	14	2	8	3		3	Metal (solid)
Silicon	Si	14	28	14	2	8	4		4	Non–metal (solid)
Phosphorus	P	15	31	16	2	8	5		3	Non–metal (solid)
Sulphur	S	16	32	16	2	8	6		2	Non–metal (solid)
Chlorine	Cl	17	35, 37	18, 20	2	8	7		1	Non–metal (gas)
Argon	Ar	18	40	22	2	8	8		0	Noble gas
Potassium	K	19	39	20	2	8	8	1	1	Metal (solid)
Calcium	Ca	20	40	20	2	8	8	2	2	Metal (solid)

EXCERCISE 4(B)

1. (a) Name the three fundamental particles of an atom. Give the symbol and charge of each particle.
   (b) Draw the orbital diagram of $^{40}_{20}Ca^{2+}$ ion and state the number of three fundamental particles present in it.
2. Complete the table given below by identifying **P**, **Q**, **R** and **S**.

Element	*Symbol*	*No. of Protons*	*No. of Neutrons*	*No. of Electrons*
Sodium	$^{23}_{11}Na$	11	**P**	11
Chlorine	$^{35}_{17}Cl$	**Q**	18	17
Uranium	**R**	92	146	92
S	$^{19}_{9}F$	9	10	9

3. The atom of an element is made up of 4 protons, 5 neutrons and 4 electrons. What are its atomic number and mass number ?
4. The atomic number and mass number of sodium are 11 and 23 respectively. What information is conveyed by this statement?
5. Write down the names of the particles represented by the following symbols and explain the meaning of superscript and subscript numbers attached.

$${}_1p^1,\ {}_0n^1,\ {}_{-1}e^0$$

6. From the symbol $^{24}_{12}Mg$, state the mass number, the atomic number and electronic configuration of magnesium.
7. Sulphur has an atomic number 16 and a mass of 32.

   State the number of protons and neutrons in the nucleus of sulphur. Give a simple diagram to show the arrangement of electrons in an atom of sulphur.
8. Explain the rule according to which electrons are filled in various energy levels.
9. Write down the electronic configuration of the following :

   (a) $^{27}_{13}X$ (b) $^{35}_{17}Y$

   Write down the number of electrons in X and neutrons in Y.

4.11 VALENCE ELECTRONS

The outermost shell of an atom is called its ***valence shell,*** *and the electrons present in the valence shell are known as* **valence electrons.**

Hydrogen has only one electron in its valence shell, hence it has one valence electron. Similarly, carbon has 4 electrons in the valence shell and so it has 4 valence electrons.

The chemical properties of elements are decided by these valence electrons, since they are the ones that take part in chemical reactions.

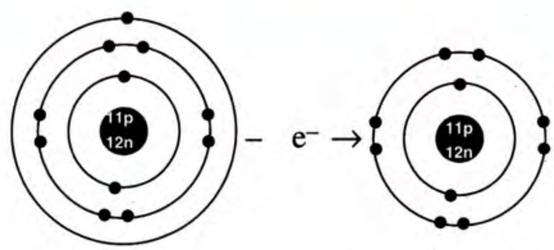
The elements with the same number of electrons in the valence shell show similar properties; those with different number of valence electrons show different chemical properties.

Elements, which have valence electrons 1 or 2 or 3 (except hydrogen), are **metals.**

These elements can lose electrons to form ions* which are positively charged and are called CATIONS.

e.g. Na – e^- → Na^+

2, 8, 1 → 2, 8



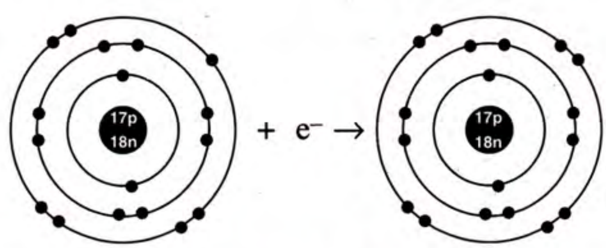
They are good reducing agents.

Elements with 4 or 5 or 6 or 7 electrons in their valence shells are **non-metals.**

These elements can gain electrons to form ions which are negatively charged and are called ANIONS.

e.g. Cl + e^- → Cl^-

2, 8, 7 → 2, 8, 8



They are good oxidising agents.

* An ion is an electrically charged atom or group of atoms formed by loss or gain of electrons.

Elements with the valence shell having 8 electrons, and helium with 2 electrons, cannot gain or loose electrons so they are chemically inactive and are called **inert gases** or noble gases.

Valency of an element is the combining capacity of that element with other elements and is equal to the number of electrons that take part in a chemical reaction.

Valency of the elements having 1, 2 or 3 valence electrons is 1, 2 or 3 respectively while valency of elements with 4, 5, 6 and 7 valence electrons is 4, 3, 2 and 1 (8 minus valence electrons) respectively.

4.12 REASON FOR CHEMICAL ACTIVITY OF AN ATOM

Atoms of all noble gases (except Helium) have eight electrons in their outermost shell. This arrangement is known as OCTET (Table 4.4).

Table 4.4. Electronic configuration of the inert or rare gases.

Inert gas	*Atomic No.*	*Electronic configuration*					
		K	L	M	N	O	P
He	2	2					
Ne	10	2	8				
Ar	18	2	8	8			
Kr	36	2	8	18	8		
Xe	54	2	8	18	18	8	
Rn	86	2	8	18	32	18	8

Among all the elements, only six inert gases (also called ***noble*** or ***rare gases***), namely, Helium (He), Neon (Ne), Argon (Ar), Krypton (Kr), Xenon (Xe) and Radon (Rn), are non-reactive, *i.e.* inert. They do not combine with other atoms to form molecules, since their outermost shells are already filled. So their combining capacity or valency is **zero**.

Helium has 2 electrons **[duplet]** in its shell (K shell), which can occupy a maximum of 2 electrons (according to the $2n^2$ rule). So it is a stable element, even though the 8-electron rule does not apply in its case.

Thus the chemical activity of an atom depends upon the number of electrons in the valence shell of its atom. **Chemically active atoms have an incomplete octet**, *i.e.* less than 8 electrons or less than 2 electrons (in case of hydrogen) in their valence shell. The ***valence electrons* of atoms contribute to the formation of molecules.**

Why do atoms combine ?

Kossel and Lewis, in 1916, independently put forward the **octet theory of valency** or *the electronic theory of valency.* According to them, the rare gases have a stable electronic configuration.

The atoms of all elements, other than the inert gases, combine because they have incomplete valence shells and tend to attain a stable configuration.

Elements tend to combine with one another to attain the stable electronic configuration of the nearest inert gas, *i.e.* 8 electrons in its valence shell. This is known as OCTET RULE. However, in the case of helium, valence shell has 2 electrons and this is known as DUPLET RULE.

This is achieved either by gaining electrons or losing electrons or sharing of electrons with other atoms of same element or another element.

The number of electrons gained, lost or shared to attain the octet in the outermost shell, gives the combining capacity of the element that is, its **valency**.

As a result of this redistribution of electrons, a force of attraction develops between atoms, which binds them together to form molecules. This force is known as a **chemical bond.**

*A **chemical bond** may be defined as the force of attraction between the two atoms that binds them together as a unit called molecule.*

The chemical combination of atoms involves redistribution of electrons so as to leave each atom with a stable electronic configuration.

4.13 ISOTOPES

It is noticed that two atoms of the same element contain the same number of protons but they may differ in the number of neutrons. However, electrons, protons and neutrons are arranged in the same way and give the atoms identical chemical properties. An element showing similar chemical properties but possessing different mass is said to represent ISOTOPY and the varieties of the atom are called ISOTOPES.

Isotopes may be defined *as atoms of the same* *element having the same atomic number but different mass numbers.*

Examples of Isotopes

Element	*Number of isotopes*	
1. Hydrogen	Three	$^{1}_{1}H$, $^{2}_{1}H$, $^{3}_{1}H$
2. Carbon	Three	$^{12}_{6}C$, $^{13}_{6}C$, $^{14}_{6}C$
3. Chlorine	Two	$^{35}_{17}Cl$, $^{37}_{17}Cl$
4. Oxygen	Two	$^{16}_{8}O$, $^{18}_{8}O$
5. Potassium	Two	$^{39}_{19}K$, $^{41}_{19}K$

Structures of the isotopes

1. Hydrogen

Hydrogen has three isotopes, namely *protium*, *deuterium* (or heavy hydrogen) and *tritium,* each having the same atomic number, *i.e.* 1 but different mass numbers, *i.e.* 1, 2 and 3 respectively. In other words, **isotopes** differ only in their number of neutrons. Thus, protium has no neutron, deuterium has one and tritium has two neutrons (Fig. 4.11). They are represented as $^{1}_{1}H$ and $^{2}_{1}H$ and $^{3}_{1}H$ respectively, in which *the subscript denotes the atomic number and the superscript the mass number.*

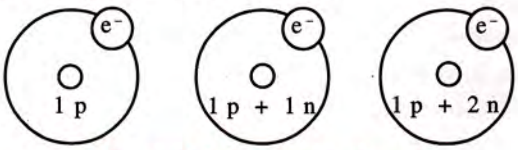
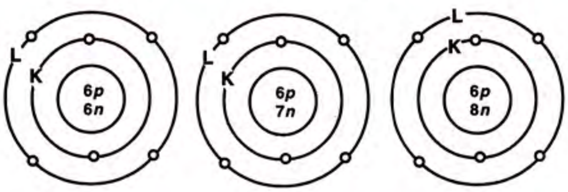


Fig. 4.11 Isotopes of hydrogen

No. of particles	Protium	Deuterium	Tritium
No. of electrons	1	1	1
No. of protons	1	1	1
Atomic number	1	1	1
No. of neutrons	0	1	2
Mass number	1	2	3

2. Carbon

Carbon (atomic number 6) has three isotopes, of mass numbers 12, 13 and 14 respectively. These are represented by the symbols $^{12}_{6}C$, $^{13}_{6}C$ and $^{14}_{6}C$ respectively. (Fig. 4.12)

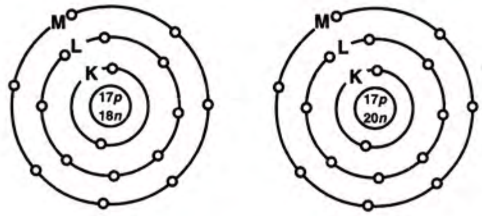


$^{12}_{6}C$ $^{13}_{6}C$ $^{14}_{6}C$

Fig. 4.12 Structures of isotopes of carbon atoms.

3. Chlorine

The two isotopes of chlorine, of mass numbers 35 and 37 respectively, have the same atomic number (17), *i.e.* the same number of protons and electrons. However, they differ only with respect to the number of neutrons contained in their respective nuclei. They are represented as $^{35}_{17}Cl$ and $^{37}_{17}Cl$. Isotope $^{35}_{17}Cl$ has 35 – 17 = 18 neutrons, and isotope $^{37}_{17}Cl$ has 37 – 17 = 20 neutrons, (See Fig. 4.13).



$^{35}_{17}Cl$ $^{37}_{17}Cl$

Fig. 4.13 Structures of isotopes of chlorine atoms.

Properties of Isotopes

Isotopes have similar chemical properties.

Isotopes have the same atomic number (Z), *i.e.* the same number of electrons and protons and hence the same electronic configuration. Chemical properties are determined by the electronic configuration of an atom. Thus, *isotopes of an element are chemically alike. For example : The chlorine* atom of mass number 37 undergoes the same reactions as the chlorine atom of mass number 35.

Isotopes differ in few physical properties such as density, boiling point etc.

Physical constants or physical properties depend on the atomic mass (mass number) and isotopes have different mass numbers (A) due to different number of neutrons.

For example, the boiling point of protium (H^{1}_{1}) is 20·38 K and that of deuterium ($^{2}_{1}H$) is 23·5 K. Thus, *the higher the mass of the isotope of an element, the higher is the boiling point.*

Isotopes like tritium, carbon-14 illustrate another difference in physical properties that can occur between isotopes as they are ***radioactive***. The extra neutrons in their nuclei cause them to be unstable so the nuclei break up spontaneously emitting certain types of radiation. Thus carbon-14 and tritium ($^{3}_{1}H$) are known as ***radio-isotopes***.

The spontaneous decay of unstable radioisotopes is known as radioactivity.

Uses of isotopes

1. Isotope of cobalt $^{60}_{27}Co$ is used in radiotherapy for treating cancer and other diseases.
2. $^{14}_{6}C$ is used for determining the age of historical and geological material.
3. An isotope of Uranium $^{235}_{92}U$ is used as a fuel in nuclear reactors.
4. An isotope of Iodine $^{131}_{53}I$ is used in the treatment of goitre.

Fractional atomic masses

An atom can have a fractional atomic mass. *For example*, the atomic mass of chlorine is 35·5. The reason for fractional atomic mass is that the atomic mass of an element represents the weighted average of all the naturally occurring isotopes of that element.

For example, any natural sample of chlorine contains 75% of $^{35}_{17}Cl$ and 25% $^{37}_{17}Cl$, *i.e.* these isotopes are in the ratio of 3 : 1 respectively. Thus, average atomic mass of chlorine is calculated as $\frac{35\times3+37\times1}{4} = 35{\cdot}5$. Therefore atomic mass of chlorine atom is 35.5.

ISOBARS

They are atoms of different elements with the same mass number, but different atomic numbers. Since the properties of elements depend upon atomic number, so they have both different physical and chemical properties.

Examples :

(i) $^{40}_{18}Ar$ and $^{40}_{20}Ca$

(ii) $^{54}_{24}Cr$ and $^{54}_{26}Fe$.

EXCERCISE 4(C)

1. How does the Modern atomic theory contradict and correlate with Dalton's atomic theory?
2. (a) What are inert elements ?
   (b) Why do they exist as monoatoms in molecules?
   (c) What are valence electrons ?
3. In what respects do the three isotopes of hydrogen differ ? Give their structures.
4. Match the atomic numbers 4, 14, 8, 15, and 19 with each of the following :
   (a) A solid non-metal of valency 3.
   (b) A gas of valency 2.
   (c) A metal of valency 1.
   (d) A non-metal of valency 4.
5. Draw diagrams representing the atomic structures of the following :
   (a) Sodium atom (b) Chlorine ion
   (c) Carbon atom (d) Oxygen ion.
6. What is the significance of the number of protons found in the atoms of different elements ?
7. Elements X, Y and Z have atomic numbers 6, 9 and 12 respectively. Which one :
   (a) forms an anion (b) forms a cation
   (c) has four electrons in its valence shell ?
8. Element X has electronic configuration 2, 8, 18, 8, 1. Without identifying X,
   (a) predict the sign and charge on a simple ion of X.
   (b) write if X will be an oxidising agent or a reducing agent. Why ?
9. Define the terms :
   (a) mass number (b) ion (c) cation
   (d) anion (e) element (f) orbit.
10. From the symbol $_{2}He^{4}$ for the element helium, write down the mass number and the atomic number of the element.
11. Five atoms are labelled A to E.

Atoms	Mass No	Atomic No.
A	40	20
B	19	9
C	7	3
D	16	8
E	14	7

   (a) Which one of these atoms :
      (i) contains 7 protons
      (ii) has electronic configuration 2, 7.
   (b) Write down the formula of the compound formed between C and D.
   (c) Predict : (i) metals (ii) non-metals.
12. An atom of an element has two electrons in the M shell.

What is the (a) atomic number (b) number of protons in this element ?

13. $^{24}_{12}Mg$ and $^{26}_{12}Mg$ are symbols of two isotopes of magnesium.
   (a) Compare the atoms of these isotopes with respect to :
      (i) the composition of their nuclei.
      (ii) their electronic configurations.
   (b) Give reasons why the two isotopes of magnesium have different mass numbers.

14. What are nucleons? How many nucleons are present in phosphorus ? Draw its structure.

15. What are isotopes ? With reference to which fundamental particle do isotopes differ ? Give two uses of isotopes.

16. Why do $^{35}_{17}Cl$ and $^{37}_{17}Cl$ have the same chemical properties ? In what respect do these atoms differ ?

17. Explain fractional atomic mass. What is the fractional mass of chlorine ?

18. (a) What is meant by "atomic number of an element"?
   (b) Complete the table given below :

	No. of Protons	*No. of Electrons*	*No. of Neutrons*	*Atomic number*	*Mass number*
$^{35}_{17}Cl$					
$^{37}_{17}Cl$					

   (c) Write down the electronic configuration of (i) chlorine atom (ii) chlorine ion.

19. Name the following :
   (a) The element which does not contain any neutron in its nucleus.
   (b) An element having valency '*zero*'.
   (c) Metal with valency 2.
   (d) Two atoms having the same number of protons and electrons but different number of neutrons.
   (e) The shell closest to the nucleus of an atom.

20. Give reasons
   (a) Physical properties of isotopes are different
   (b) Argon does not react.
   (c) Actual atomic mass is greater than mass number.
   (d) $^{35}_{17}Cl$ and $^{37}_{17}Cl$ do not differ in their chemical reactions.

21. An element A atomic number 7 mass number 14
   B electronic configiration 2, 8, 8
   C electrons 13, neutrons 14
   D Protons 18 neutrons 22
   E electronic configuration 2, 8, 8, 1

   State (i) valency of each element (ii) which one is a metal (iii) which is a non-metal (iv) which is an inert gas.

22. Choose the correct option
   (a) Rutherford's alpha-particle scattering experiment discovered
      A. Electron **B.** Proton
      C. Atomic nucleus **D.** Neutron
   (b) Number of valence electrons in O^{2-} is :
      A. 6 **B.** 8
      C. 10 **D.** 4
   (c) Which of the following is the correct electronic configuration of potassium ?
      A. 2, 8, 9 **B.** 8, 2, 9
      C. 2, 8, 8, 1 **D.** 1, 2, 8, 8

23. Explain :
   (a) octet rule for formation of sodium chloride
   (b) duplet rule for formation of hydrogen.

24. Complete the following table relating to the atomic structure of some elements.

Element Symbol	Atomic Number	Mass Number	Number of Neutrons	Number of Electrons	Number of Protons
Li	3	6			
Cl	17		20		
Na			12		11
Al		27			13
S		32	16		

4.14 ELECTROVALENT (OR IONIC) BOND

Atoms of metallic elements that have 1, 2 or 3 valence electrons can lose electron(s) to atoms of non-metallic elements, which have 5, 6 or 7 electrons respectively in their outermost shell and thereby forming an electrovalent compound.

After the transfer of electron(s), both the combining atoms acquire the electronic configuration of the nearest inert gas.

A metallic atom, which loses electron(s), becomes a positively charged ion and is known as a **cation** and a non-metallic atom, which gains electron(s), becomes a negatively charged ion and is known as an **anion.**

An **ion** is a charged particle which is formed due to the gain or the loss of one or more electrons by an atom.

A metallic element, whose one atom readily loses electron(s) to form a positively charged ion, is an ***electropositive element***.

$$Na - e^- \rightarrow Na^+ \text{ (cation)}$$

A non-metallic element, whose atom readily accepts electron(s) to form a negatively charged ion, is an ***electronegative element.***

$$Cl + e^- \rightarrow Cl^- \text{ (anion)}$$

The **cation** and the **anion** being oppositely charged attract each other and form a chemical bond. Since this chemical bond formation is due to the electrostatic force of attraction between a cation and an anion, it is called an ***electrovalent*** (or an ***ionic***) ***bond.***

Electrovalent (or ionic) compounds : *The chemical compounds formed as a result of the transfer of electrons from one atom of an element to one atom of another element are called* ***ionic*** *(or electrovalent)* ***compounds****.*
Electrovalency : *The number of electrons that an atom of an element loses or gains to form a electrovalent bond is called its* ***electrovalency****.*

The **metals** of group numbers 1, 2 and 13 have a tendency to lose their valence electrons. So they combine with the **non-metals** of Groups 15, 16 and 17, which have a tendency to gain electron(s) and form ionic bonds.

Note : Group 1 elements are most electropositive, *i.e.*, they are metallic in nature. Their metallic nature increases down the group, so **caesium** is most electropositive.

Group 17 elements are most electronegative. Fluorine is the most electronegative element. Thus, *caesium fluoride CsF is the most ionic compound*.

Bonds formed between metals and non-metals are ionic or electrovalent bonds.

Why are Ionic Compounds Stable?

Ionic compounds are formed by ions but there also exists a repulsive force between ions for like charges. Since the electrostatic force of attraction between opposite charges is much higher, it makes the ionic compounds stable.

Examples of electrovalent (ionic) compounds.

$NaCl$	(Sodium chloride)
$MgCl_2$	(Magnesium chloride)
CaO	(Calcium oxide)
KBr	(Potassium Bromide)
$CaCl_2$	(Calcium chloride)

4.14.1 Structures of some electrovalent compounds

1. Sodium chloride (NaCl)

The electronic configuration of a **sodium atom is 2, 8, 1.** It has one electron in excess of the stable electronic configuration of the nearest noble gas, neon, (2, 8). Therefore, an atom of sodium shows a tendency to give up the electron from its outermost shell, so as to acquire a stable electronic configuration of neon.

$$\underset{\substack{(2,\,8,\,1)\\ \text{atom}}}{Na} \rightarrow \underset{\substack{(2,\,8)\\ \text{cation}}}{Na^+} + 1e^+$$

However, *after giving up one electron, the sodium atom is no more electrically neutral. It has eleven protons in its nucleus but only ten electrons revolving around it. Therefore, it has a net positive charge of +1. This positively charged atom is called sodium ion and is written as* Na^+ and its electronic configuration resembles that of the noble gas neon.

The electronic configuration of **chlorine is 2, 8, 7.** It has an electronic configuration with one electron less than that of the nearest noble gas, argon (2, 8, 8). Therefore, the chlorine atom shows a tendency to acquire an electron to attain octet in its outermost shell.

$$\underset{\substack{(2,\,8,\,7)\\ \text{atom}}}{Cl} + 1e^- \rightarrow \underset{\substack{(2,\,8,\,8)\\ \text{anion}}}{Cl^-}$$

An atom of chlorine is electrically neutral, as it contains 17 protons in its nucleus and 17 electrons revolving around it. But, *after acquiring an electron* *from the sodium atom, the chlorine atom does not remain electrically neutral. It has one electron more than the number of protons in its nucleus and therefore has charge of –1 represented as* Cl^- *i.e.* **chloride ion.**

Chloride ion has an octet of electrons in its outermost shell, and its electronic configuration resembles that of the noble gas argon (Fig. 4.14).

Thus, when an atom of sodium combines with an atom of chlorine (electronic configuration 2, 8, 7), one electron is transferred from the sodium atom to the chlorine atom, resulting in the formation of a sodium chloride molecule.

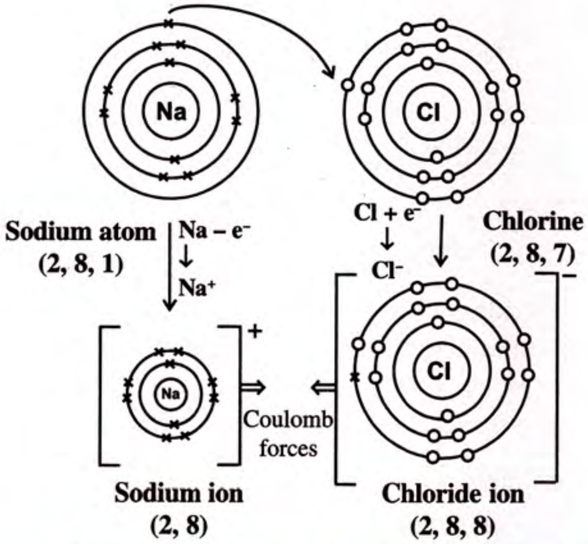


Fig. 4.14 Orbit structure of electrovalent bonding in sodium chloride.

The cation Na^+ and anion Cl^- are attracted towards each other, due to electrical charge or coulomb force existing between them and form an ionic compound.

Electron dot symbol (Lewis symbol)

The electron dot symbol for an atom consists of the symbol of the element surrounded by dots representing the outermost shell electrons. The paired electrons are represented by a pair of dots, whereas the unpaired electron in the outermost orbit is represented by a single dot.

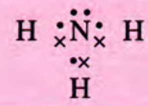
Example : Electron dot symbol of Hydrogen is H•

and of Oxygen is :O:

Symbols other than dots, such as circles and crosses can be used to distinguish between the electrons of different atoms in a molecule,

for example :

Ammonia (NH_3) can be represented as



Electron dot structure of NaCl

$Na\bullet + \bullet\ddot{\underset{\cdot\cdot}{Cl}}: \rightarrow Na^+ + Cl^-$ or NaCl

2. Magnesium chloride ($MgCl_2$)

The number of valence electrons of magnesium (atomic number 12) is 2 and that of chlorine (atomic number 17) is 7. Magnesium atom acquires a stable configuration of 8 electrons by losing two electrons from its outermost shell (one each to each atom of chlorine) and thus becomes a positive magnesium ion, Mg^{2+}.

Mg	–	$2e^-$	→	Mg^{2+}
(2, 8, 2)				(2, 8)
atom				cation

However, each chlorine atom, which contains 7 electrons in its outermost shell, can accept only 1 of the 2 electrons donated by a magnesium atom. Therefore, for each magnesium atom forming a magnesium ion, there must be two chlorine atoms to form two chloride ions.

2Cl	+	$2e^-$	→	$2Cl^-$
(2, 8, 7)				(2, 8, 8)
2 Cl atoms				2 chloride anions

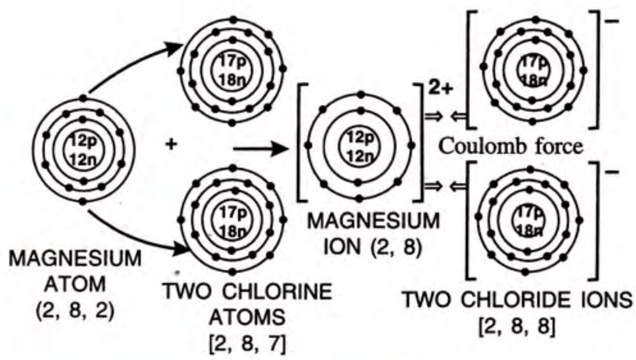
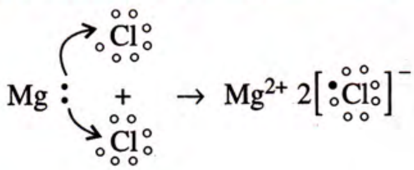


Fig. 4.15 Orbit structure of formation of magnesium chloride ($MgCl_2$) molecules.

Thus, the ratio of magnesium to chloride ions in magnesium chloride must be 1 : 2, so the molecular formula of the compound magnesium chloride is $MgCl_2$ (Fig. 4.14).

Electron dot structure of magnesium chloride



3. Calcium oxide (CaO)

The number of valence electrons of a calcium atom (atomic number 20) is 2, and that of an oxygen atom is 6, *i.e.*, oxygen requires 2 electrons to attain octet. In the presence of oxygen, each calcium atom loses its 2 valence electrons to one oxygen atom (Fig. 4.15). As a result, the calcium atom forms a calcium ion with charge +2 (Ca^{2+}), and the oxygen atom forms an oxide ion with charge −2 (O^{2-}). Since only one oxygen atom is needed to accept the 2 valence electrons donated by a calcium atom, the formula of calcium oxide is CaO and not Ca_2O_2.

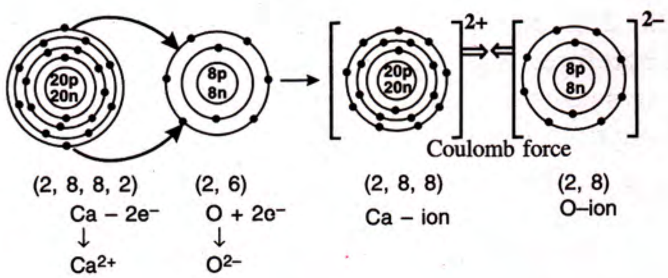
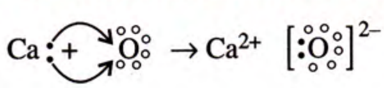


Fig. 4.15 Orbit structure of formation of calcium oxide (CaO).

$$\underset{\text{atom}}{\underset{(2,\,8,\,8,\,2)}{Ca}} - 2e^- \rightarrow \underset{\text{cation}}{\underset{(2,\,8,\,8)}{Ca^{2+}}}$$

$$\underset{\text{atom}}{\underset{(2,\,6)}{O}} + 2e^- \rightarrow \underset{\text{anion}}{\underset{(2,\,8)}{O^{2-}}}$$

Electron dot structure of calcium oxide



In the formation of an electrovalent bond, the transfer of electron(s) is involved. The **electropositive atom** undergoes **oxidation**, while the ***electronegative atom*** undergoes ***reduction***. This is known as REDOX PROCESS.

For example :

Formation of sodium chloride : Sodium chloride is formed by the combination of sodium and chlorine.

$$2Na + Cl_2 \rightarrow 2Na^+ + 2Cl^- \text{ (or 2 NaCl)}$$

The reaction can be written as two half reactions :

$$2Na \rightarrow 2Na^+ + 2e^- \text{ (Oxidation)}$$

$$Cl_2 + 2e^- \rightarrow 2Cl^- \text{ (Reduction)}$$

Oxidation

$$2Na + Cl_2 \rightarrow 2Na^+ + 2Cl^- \text{ (Redox Reaction)}$$

Reduction

Oxidation and reduction always occur simultaneously because the electron(s) lost by the reducing agent must be gained by the oxidising agent.

For example :

Oxidised

$$\underset{\text{Oxidising agent}}{CuO} + \underset{\text{Reducing agent}}{H_2} \rightarrow Cu + H_2O$$

Reduced

In this reaction, hydrogen acting as a reducing agent reduces Cu(II) oxide to copper. This is a reduction reaction.

$$Cu^{2+} + 2e^- \rightarrow Cu \text{ (Reduction)}$$

At the same time, hydrogen is oxidised to water by the oxidising agent Cu(II) oxide, and this is an oxidation reaction.

$$2H - 2e^- \rightarrow 2H^+ \text{ (Oxidation)}$$

Thus, the net reaction is a redox reaction.

It can be inferred from the above example that an oxidising agent is an acceptor of electron(s) and a reducing agent is a donor of electron(s).

[Also refer chapter 6, article 6.2 (viii) & (ix)]

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EXCERCISE 4(D)

1. How do atoms attain noble gas configuration ?
2. Define :
   (a) a chemical bond, (b) an electrovalent bond,
   (c) a covalent bond.
3. What are the conditions necessary for the formation of an electrovalent bond ?
4. An atom X has three electrons more than the noble gas configuration. What type of ion will it form ? Write the formula of its (i) sulphate (ii) nitrate (iii) phosphate (iv) carbonate (v) hydroxide.
5. Mention the basic tendency of an atom which makes it to combine with other atoms.
6. What type of compounds are usually formed between metals and non-metals and why ?
7. In the formation of the compound XY_2, an atom X gives one electron to each Y atom. What is the nature of bond in XY_2 ? Draw the electron dot structure of this compound.
8. An atom X has 2,8,7 electrons in its shell. It combines with Y having 1 electron in its outermost shell.
   (a) What type of bond will be formed between X and Y ?
   (b) Write the formula of the compound formed.
9. Draw orbit structure and electron dot diagrams of NaCl, $MgCl_2$ and CaO.
10. Compare :
    (a) sodium atom and sodium ion
    (b) chlorine atom and chloride ion, with respect to
    (i) atomic structure, (ii) electrical state,
    (iii) chemical action, (iv) toxicity.
11. The electronic configuration of fluoride ion is the same as that of a neon atom. What is the difference between the two ?
12. (a) What do you understand by redox reactions? Explain oxidation and reduction in terms of loss or gain of electrons.
    (b) Divide the following redox reactions into oxidation and reduction half reactions.
    (i) $Zn + Pb^{2+} \rightarrow Zn^{2+} + Pb$
    (ii) $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$
    (iii) $Cl_2 + 2Br^- \rightarrow Br_2 + 2Cl^-$
    (iv) $Sn^{2+} + 2Hg^{2+} \rightarrow Sn^{4+} + Hg_2^{2+}$
    (v) $2Cu^+ \rightarrow Cu + Cu^{2+}$
    (c) Potassium (at No. 19) and chlorine (at No. 17) react to form a compound. Explain on the basis of electronic concept :
    (i) oxidation (ii) reduction
    (iii) oxidising agent (iv) reducing agent.

4.15 COVALENT (MOLECULAR) BOND

*The chemical bond that is formed between two combining atoms by mutual sharing of one or more pairs of electrons is called a **covalent** (or a molecular) **bond** and the compound formed due to this bond is called a **covalent compound**.*

The molecule formed due to the sharing of electrons (covalent bond) is called a **covalent molecule.**

The atoms of non-metals usually have 5, 6 or 7 electrons in their outermost shell (except carbon which has 4 and hydrogen which has just 1 electron in the outermost shell). *The atoms of such elements do not favour the loss of its electrons due to energy considerations and thus the transfer of electrons is not possible*. Therefore, these atoms can complete their octet only by mutually sharing one or more pairs of electrons. **Each atom contributes equal number of electron(s).** So, whenever a non-metal combines with another non-metal (to attain stable configuration), the sharing of electrons takes place between their atoms and a covalent bond is formed.

For example, hydrogen is a non-metal and chlorine is also a non-metal. When hydrogen combines with chlorine to form hydrogen chloride (HCl), the sharing of electrons takes place between hydrogen and chlorine atoms and a covalent bond is formed. It should be noted that a covalent bond can also be formed between two atoms of the same non-metal. For example, two chlorine atoms combine together by the sharing of an electron each to form a chlorine molecule (Cl_2) and a covalent bond is formed between the two chlorine atoms.

Covalent bonds are of three types :

1. Single covalent bond
2. Double covalent bond
3. Triple covalent bond

A **single covalent** bond is formed by the sharing of one pair of electrons between the atoms, each atom contributing one electron.

A single covalent bond is denoted by putting a short line (—) between the two atoms. So, a hydrogen molecule can also be written as H—H.

Likewise, molecules of chlorine, hydrogen chloride, water, ammonia, methane and carbon tetrachloride are examples of single covalent bonds.

Cl—Cl	H—Cl	H—O—H
Chlorine	Hydrogen chloride	Water

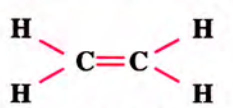
```
              H                  Cl
              |                  |
H—N—H       H—C—H            Cl—C—Cl
  |           |                  |
  H           H                  Cl
Ammonia     Methane       Carbon tetrachloride
```

Similarly, **a double bond** is formed by the sharing of two pairs of electrons between two atoms.

A double bond is actually a combination of two single bonds, so it is represented by putting two short lines (=) between the two atoms. *For example*, oxygen molecule, O_2, contains a double bond between two atoms and it can be written as O=O.

A triple bond is a combination of three single bonds, so it is represented by putting three short lines (≡) between the two atoms. Nitrogen molecule, N_2, contains a triple bond, so it is written as N ≡ N.

Some molecules have a combination of single bond as well as a double or a triple bond. *For example*, ethene (C_2H_4) molecule has one double covalent bond and four single covalent bonds.



Ethene (C_2H_4)

Similarly, **ethyne** (C_2H_2) molecule has one triple covalent bond and two single covalent bonds.

Ethyne (C_2H_2) H – C ≡ C – H

***The covalency** of an atom is the number of its electrons taking part in the formation of shared pairs.* Thus, the covalency of hydrogen is 1, oxygen 2, nitrogen 3 and carbon 4.

Non-polar and polar covalent compounds

Covalent compounds are non polar when shared pair of electron(s) are equally distributed between the two atoms. No charge separation takes place. The molecule is symmetrical and electrically neutral.

If the two covalently bonded atoms are identical the shared electron pair(s) is at equal distance from the combining atoms *i.e.*, the shared electron pair(s) is equally attracted by the nuclei of the two types of charge, such molecules are non-polar. For example hydrogen (H_2), chlorine (Cl_2), oxygen (O_2), *etc*., are perfectly non-polar compounds.

***The covalent compounds are said to be polar when the shared pair of electrons are not at equal distance between the two atoms.** This results in the development of fractional positive and negative charges on them. They ionise in water. For example, hydrogen chloride.*

Covalent bond is formed when both atoms have *four* or more electrons in their outermost shells, *i.e.*, non-metals (exceptions are H, Be, B, Al, *etc*.).

Hydrogen can combine with all non-metals of Group IVA to VIIA with the help of covalent bonds.

Note : As we move across a period, the electrovalent character of chlorides, oxides, etc., decreases. An example of the decreasing electrovalent character of chlorides and oxides of the elements of 3rd period is shown below :

Group →	Group I	Group II	Group III	Group IV	Group V	Group VI	Group VII
Chloride →	NaCl	$MgCl_2$	$AlCl_3$	$SiCl_4$	PCl_3/PCl_5	S_2Cl_2	
Bonding →	Ionic solid	Ionic solid	Partially ionic-partially covalent solid	Covalent liquid	Covalent liquid/solid	Covalent liquid	

Oxide →	Na_2O	MgO	Al_2O_3	SiO_2	P_2O_5	SO_2/SO_3	Cl_2O_7
Bonding →	Ionic Solid	Ionic Solid	Ionic Solid	Covalent Solid	Covalent Solid	Covalent Gas/Solid	Covalent Gas

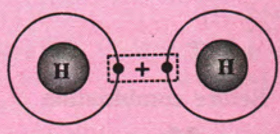
4.15.2 Some covalent molecules and their structures

1. Hydrogen molecule (Non-polar compound)

A hydrogen atom has one electron in its only shell. It needs one more electron to attain **duplet**. To meet this need, hydrogen atom shares its electron with another hydrogen atom. Thus, one electron each is contributed by the two atoms of hydrogen, and the resulting pair of electrons is mutually shared by both the atoms to form a hydrogen molecule.

Formation of a hydrogen molecule

Before combination



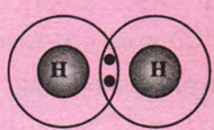
H atom   H atom

Electron dot structure

H· + ·H

H-atom   H-atom

After combination



Hydrogen molecule

H:H or H—H → H_2

One shared pair of electrons   **Single covalent bond**

[A hydrogen molecule contains two atoms of hydrogen]

2. Chlorine molecule (Non-polar compound)

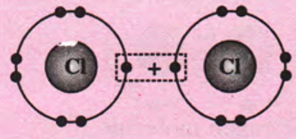
Electronic configuration	*Nearest noble gas*
$_{17}Cl$ [2, 8, 7]	Argon ($_{18}Ar$) [2, 8, 8]

To attain the stable electronic configuration of the nearest noble gas, chlorine needs one electron.

When two chlorine atoms come closer, each contributes one electron and form ***one shared pair of electrons*** between them. Both the atoms of chlorine thus attain an octet. A single covalent bond [Cl — Cl] is formed between them.

Formation of a chlorine molecule

Before combination



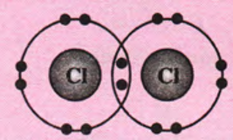
Cl-atom   Cl-atom

Electron dot structure

:C̤̈l· + ·C̤̈l:

Cl-atom   Cl-atom

After combination



Chlorine molecule

:C̤̈l : C̤̈l: or [Cl—Cl] → Cl_2

One shared pair of electrons   **Single covalent bond**

[A chlorine molecule contains two atoms of chlorine]

3. Nitrogen molecule (Non-polar compound)

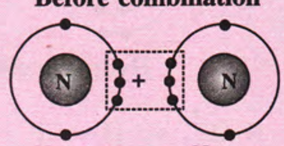
Electronic configuration	*Nearest noble gas*
Nitrogen ($_7N$) [2, 5]	Neon ($_{10}Ne$) [2, 8]

To attain the stable electronic configuration of the nearest noble gas, nitrogen needs three electrons.

When two nitrogen atoms come closer, each contributes three electrons and so they have ***three shared pairs of electrons*** between them. Both atoms attain an octet, resulting in the formation of a triple covalent bond [N ≡ N] between them.

Formation of a nitrogen molecule

Before combination



N-atom   N-atom

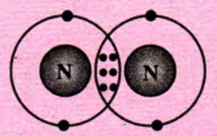
Electron dot structure

:N⁝ + ⁝N:

N-atom   N-atom

After combination

Mutual sharing of three pairs of elements



Nitrogen molecule

:N⋮ ⋮N: or [N≡N] → N_2

Three shared pairs of electrons — **Triple covalent bond**

[A nitrogen molecule contains two atoms of nitrogen]

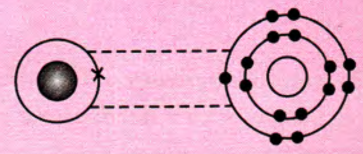
4. Hydrogen chloride (polar compound)

Atoms involved	*Electronic configuration*	*Nearest noble gas*
Hydrogen Chlorine	$_1H$ [1] $_{17}Cl$ [2, 8, 7]	Helium [2] Argon [2, 8, 8]

To attain the stable electronic configuration of nearest noble gas, hydrogen needs one electron and chlorine also needs one electron.

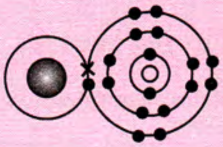
In the case of hydrogen chloride molecule hydrogen atom shares an electron pair with chlorine atom, such that hydrogen acquires a duplet configuration and chlorine an octet, resulting in the formation of a single covalent bond.

Before combination



H× :Cl:

After combination



Hydrogen chloride molecule

H ×Cl: or H — Cl

[One molecule of hydrogen chloride contains a total of two atoms *i.e.* one atom of hydrogen and one atom of chlorine]

5. Water molecule (Polar compound)

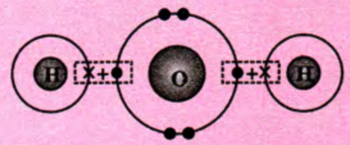
Atoms involved	*Electronic configuration*	*Nearest noble gas*
Hydrogen Oxygen	$_1H$ [1] $_8O$ [2, 6]	Helium [2] Neon [2, 8]

To attain the stable electronic configuration of the nearest noble gas, hydrogen needs one electron and oxygen needs two electrons.

In the case of a water molecule, *each of the two hydrogen atoms share an electron pair with the oxygen atom* such that hydrogen acquires a duplet configuration and oxygen an octet, resulting in the formation of two single covalent bonds.

Formation of a water molecule

Before combination

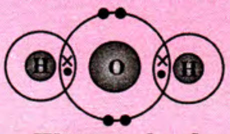


Electron dot structure

H× + :O: + ×H

H-atom O-atom H-atom

After combination



Water molecule

H ×O× H or H—O—H → H_2O

One shared pair of electrons — **Two single covalent bond**

[One molecule of water contains a total of three atoms, *i.e.* one atom of oxygen and two atoms of hydrogen]

6. Ammonia molecule (Polar compound)

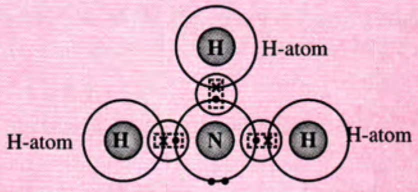
Atoms involved	*Electronic configuration*	*Nearest noble gas*
Nitrogen Hydrogen	$_7N$ [2, 5] $_1H$ [1]	Neon [2, 8] Helium [2]

To attain the electronic configuration of the nearest noble gas, nitrogen needs three electrons and hydrogen needs one electron.

When a molecule of ammonia is to be formed, one atom of nitrogen *shares three electron pairs, one with each of the three atoms of hydrogen.*

Formation of an ammonia molecule

Before combination



Electron dot structure

H• + •N• + •H (with H above, joined by +)

Ammonia molecule

After combination

H
H:N:H

One shared pair of electrons with each hydrogen atom

OR

H
|
H—N—H

Three single covalent bonds

[The ammonia molecule contains a total of four atoms, *i.e.*, one atom of nitrogen and three atoms of hydrogen]

7. Carbon tetrachloride molecule (Non-polar compound)

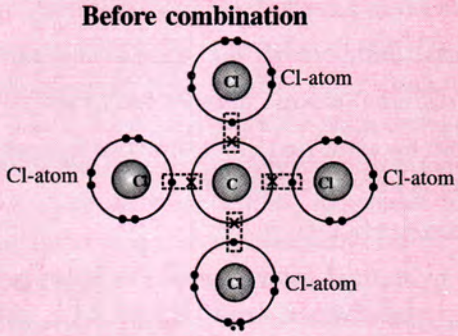
Atoms involved	*Electronic configuration*	*Nearest noble gas*
Carbon	$_6C$ [2, 4]	Neon [2, 8]
Chlorine	$_{17}Cl$ [2, 8, 7]	Argon [2, 8, 8]

To attain the stable electronic configuration of the nearest noble gas, carbon needs four electrons and chlorine needs one electron.

When a molecule of carbon tetrachloride is to be formed, one atom of carbon *shares four electron pairs, one with each of the four atoms of chlorine.*

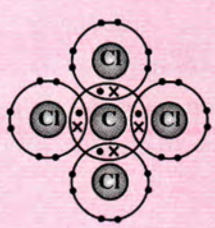
Formation of a carbon tetrachloride molecule

Before combination



Electron dot structure

:Cl: + ·C· + ·Cl: + :Cl: (with :Cl: above and below, joined by +)



After combination

:Cl:
:Cl ×C× Cl:
:Cl:

One shared pair of electrons with each chlorine atom

Carbon tetrachloride molecule

or

Cl
|
Cl—C—Cl → CCl_4
|
Cl

Four single covalent bonds

[One molecule of carbon tetrachloride contains five atoms in all, *i.e.* one atom of carbon and four atoms of chlorine]

8. Methane molecule (Non-polar compound)

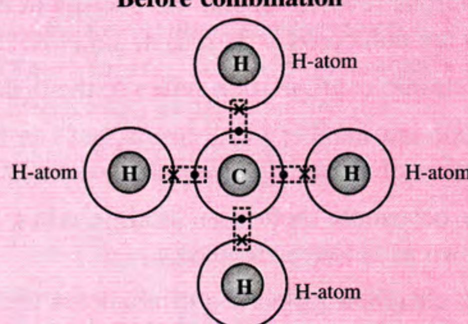
Atoms involved	*Electronic configuration*	*Nearest noble gas*
Carbon Hydrogen	$_6C$ [2, 4] $_1H$[1]	Neon [2, 8] Helium [2]

To attain the stable electronic configuration of the nearest noble gas, carbon needs four electrons and hydrogen needs one electron.

When a molecule of methane is to be formed, one atom of carbon *shares four electron pairs, one with each of the four atoms of hydrogen.*

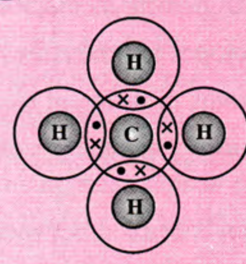
Formation of a methane molecule

Before combination



Electron dot diagram

H×+•C•+×H (with H× + above and + ×H below)



Methane molecule

After combination

H
H ×C× H
H

or

H
|
H—C—H
|
H

Four single covalent bonds

[A methane molecule contains a total of five atoms, *i.e.* one atom of carbon and four atoms of hydrogen]

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CHAPTER AT A GLANCE

- The idea of the smallest unit of matter *i.e.* PARMANU which combines to give ANU was given by the Indian scientist Maharshi Kanada.
- Greek philosopher Democritus called the paramanu an 'ATOM'.
- John Dalton gave the first systematic idea of structure of atom.
- Gases conduct electricity at very low pressure 0·01 mm of mercury and high potential at 10,000 volts emitting rays known as cathode rays. These rays consist of negatively charged particles called **electrons**.
- Electrons are present in every atom. Its mass is $\frac{1}{1837}$ the mass of hydrogen atom. Its charge is $1{\cdot}602 \times 10^{-19}$ coulombs.
- Goldstein showed the presence of positively charged particles, the nature of which depends upon the gas used in the discharge tube. These positively charged subatomic particles are **protons**.
- Chadwick discovered neutron in 1932. Neutron is an electrically neutral subatomic particle with mass equal to that of a proton.
- Rutherford showed the presence of a small heavy, positively charged centre — **nucleus**.
- Electrons, protons and neutrons are the fundamental particles or subatomic particles. Protons and neutrons are present in the central part of the atom *i.e.* in nucleus and collectively known as **nucleons.** Electrons are outside the nucleus in orbits.
- Atomic number (Z) of an element is equal to the number of protons and mass number (A) of an atom is equal to the sum of the number of protons and neutrons present in the nucleus.
- Electrons revolve around the nucleus in definite circular paths which are known as **orbits or energy shells** and are numbered as 1, 2, 3, 4, ... *i.e.* K, L, M and N, ...
- Electronic configuration refers to the distribution of electrons in different shells.
- Maximum number of electrons that can be filled in these shells is given by $2n^2$ rule, where n denotes the number of shells. However, the outermost orbit cannot have more than 8 electrons.
- The outermost shell of an atom is called its **valence shell** and the electrons present in the valence shell are known as **valence electrons**.
- The chemical properties of elements are decided by these valence electrons, since they are the ones that take part in chemical reactions.
- Isotopes are the atoms of the same element possessing different mass numbers but the same atomic number.
- Atoms of different elements that have the same mass number but different atomic number are called isobars.
- During chemical reactions, atoms of all the elements become stable by acquiring electronic configuration of the nearest inert gas, *i.e.* eight electrons in the outermost orbit, this is known as octet rule.
- Bonds formed between metals and non-metals are ionic or electrovalent bonds.
- The chemical bond that is formed between two combining atoms by mutual sharing of one or more pairs of electrons is called a covalent (or a molecular) bond and the compound formed due to this bond is called a covalent compound.

EXCERCISE 4(E)

1. Explain the following :
   (a) Electrovalent compounds conduct electricity in molten or aqueous state.
   (b) Electrovalent compounds have a high melting point and boiling point while covalent compounds have low melting and boiling points.
   (c) Electrovalent compounds dissolve in water whereas covalent compounds do not.
   (d) Electrovalent compounds are usually hard crystals yet brittle.
   (e) Polar covalent compounds conduct electricity.
2. A solid is crystalline, has a high melting point and is water soluble. Describe the nature of the solid.
3. Match the atomic number 4, 8, 14, 15 and 19 with each of the following :
   (a) A solid non-metal of valency 3.
   (b) A gas of valency 2.
   (c) A metal of valency 1.
   (d) A non-metal of valency 4.
4. (a) Elements X, Y and Z have atomic numbers 6, 9 and 12 respectively. Which one :
   (i) forms an anion , (ii) forms a cation,
   (b) State the type of bond between Y and Z and give its molecular formula.
5. Taking $MgCl_2$ as an electrovalent compound, CCl_4 as a covalent compound, give four differences between electrovalent and covalent compounds.
6. Potassium chloride is an electrovalent compound, while hydrogen chloride is a covalent compound. But, both conduct electricity in their aqueous solutions. Explain.
7. (a) Name two compounds that are covalent when taken pure but produce ions when dissolved in water.
   (b) For each compound mentioned above, give the formulae of ions formed in aqueous solutions.
8. An element M burns in oxygen to form an ionic compound MO. Write the formula of the compounds formed if this element is made to combine with chlorine and sulphur separately.
9. Give electron dot diagram of the following :
   (a) magnesium chloride, (b) nitrogen,
   (c) methane (d) hydrogen chloride
10. State the type of bonding in the following molecules.
    (a) water, (b) calcium oxide,
    (c) hydrogen chloride
11. Element M forms a chloride with the formula MCl_2 which is a solid with high melting point. M would most likely be in the group in which .... is placed.
    [(a) Na (b) Mg (c) Al (d) Si.]
12. (a) Compound X consists of molecules.
    Choose the letter corresponding to the correct answer from the options A, B, C and D given below :
    (i) The type of bonding in X will be :
    A. ionic B. electrovalent
    C. covalent D. molecular
    (ii) X is likely to have a :
    A. low melting point and high boiling point,
    B. high melting point and low boiling point,
    C. low melting point and low boiling point,
    D. high melting point and high boiling point.
    (iii) In the liquid state, X will :
    A. become ionic, B. be an electrolyte,
    C. conduct electricity, D. not conduct electricity.
    (b) If electrons are getting added to an element Y; then
    (i) is Y getting oxidized or reduced ?
    (ii) what charge will Y migrate to during the process of electrolysis ?
13. Choose the correct answer from the choices A, B, C and D:
    (i) The property which is characteristic of an electrovalent compound is that :
    A. it is easily vaporized,
    B. it has a high melting point,
    C. it is a weak electrolyte,
    D. it often exists as a liquid.
    (ii) When a metal atom becomes an ion :
    A. it loses electrons and is oxidized,
    B. it gains electrons and is reduced,
    C. it gains electrons and is oxidized,
    D. it loses electrons and is reduced.
14. Identify the following reactions as either oxidation or reduction :
    (i) $O + 2e^- \rightarrow O^{2-}$ (ii) $K - e^- \rightarrow K^+$
    (iii) $Fe^{3+} + e^- \rightarrow Fe^{2+}$
15. (a) Name the charged particles which attract one another to form electrovalent compounds.
    (b) In the formation of electrovalent compounds, electrons are transferred from one element to another. How are electrons involved in the formation of a covalent compound ?

(c) The electronic configuration of nitrogen is (2, 5). How many electrons in the outer shell of a nitrogen atom are not involved in the formation of a nitrogen molecule ?

(d) In the formation of magnesium chloride (by direct combination between magnesium and chlorine), name the substance that is oxidized and the substance that is reduced.

16. (a) Which of the following is **not** a common characteristic of an electrovalent compound ?

A. High melting point.

B. Conducts electricity when molten.

C. Consists of oppositely charged ions.

D. Ionizes when dissolved in water.

(b) What is the term defined below ?

A bond formed by a shared pair of electrons, each bonding atom contributing one electron to the pair.