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CHAPTER

1

PERIODIC CLASSIFICATION OF ELEMENTS

In 1661, Robert Boyle proposed the definition of an element as "any substance that cannot be decomposed into a simpler substance by a chemical reaction". At that time, only 13 elements were known. Towards the end of the 18th century, another 11 elements have been discovered. When only a few elements were known, it was easy to study and remember their properties. But gradually with the development of science, several new elements were discovered. At present about 114 elements are known to us. Each of these elements show different properties due to different kind of atoms constituting them and they form millions of compounds. A study of the properties of these elements and their compounds individually appear to be a difficult task. So attempts were made, from time to time, to find regularities or trends in the properties of these elements such that they could be classified into a handful of groups having similar properties. The study of the properties of a typical element of a particular group enables scientist to roughly predict the properties of other elements of that group.

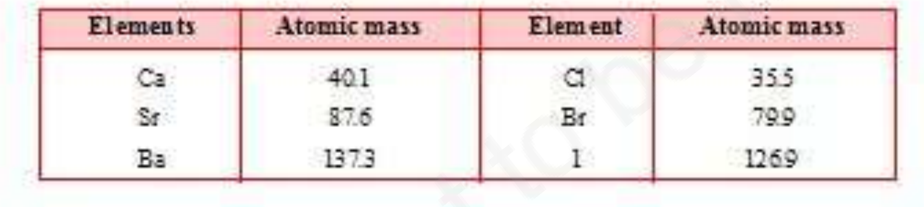
1.1. EARLY ATTEMPTS AT THE CLASSIFICATION OF ELEMENTS

Some elements like copper, silver and gold have characteristic lustre, malleability and ductility. Other elements like graphite (carbon), sulphur and phosphorus do not have lustre, are non-melleable and non-ductile. These important differences led Lavoisier to classify elements into two groups, **metals** and **non-metals**. However, such a grouping was too broad and unbalanced as there are only a few non-metals (about 10%) in one group and a large number of metals (90%) in the other group. Also there was no place for elements with properties resembling those of metals as well as non-metals.

1.1.1. Dobereiner's Triads

In 1803, John Dalton published a table of relative atomic weights (now called atomic masses) of elements. Attempts were then made to find a relationship between the physical and chemical properties of elements and their atomic masses. In 1829, the German Chemist Dobereiner arranged elements having similar chemical properties into **groups of three**. These groups were called **triads.** *Doberiener found that when elements are arranged in a triad in the order of increasing atomic masses, the atomic mass of the middle element was approximately equal to the arithmetic mean of the atomic masses of the other two elements.* This is known as **Dobereiner's law of triads**. For example, the elements calcium, strontium and barium resemble one another in most of their properties. The atomic mass of strontium is almost equal to the average of the atomic masses of calcium and barium. Again, in the case of another three elements, chlorine, bromine and iodine, the atomic mass of bromine is about the average of the atomic masses of chlorine and iodine (Table 1.1)

Table 1.1 Dobereiner's triads



Elements	Atomic mass	Element	Atomic mass
Ca	40.1	Cl	35.5
Sr	87.6	Br	79.9
Ba	137.3	I	126.9

Dobereirner could identify four triads from the elements known at that time. Hence, this system of classification could not become successful because all elements could not be put into triads.

Find out which of the groups in the following table (table 1.2) form Dobereiner's triads?

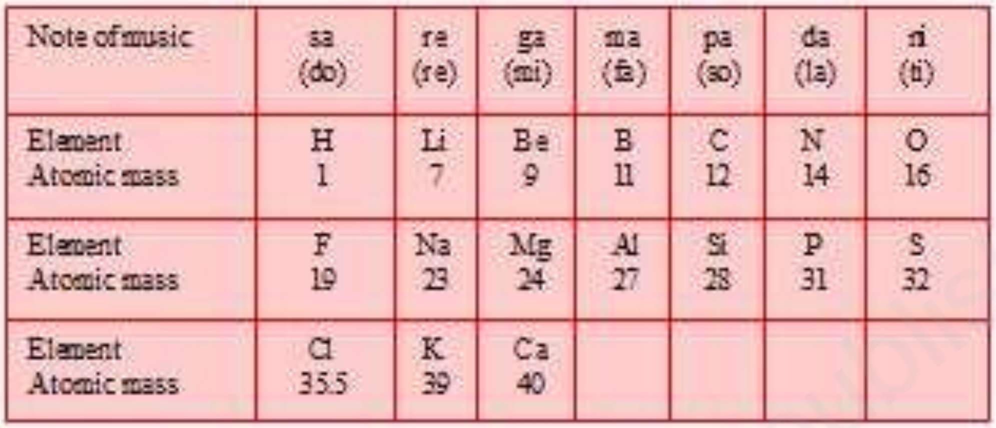
Table 1.2

Group A elements	Atomic mass	Group B	Atomic mass elements
N	14.0	Li	6.9
P	31.0	Na	23.0
As	74.9	K	39.0

1.1.2. Newlands' Octaves

In 1866, John Newlands, an English scientist observed an interesting regularity. He arranged the elements in the increasing order of their atomic masses as follows (Table 1.3).

Table 1.3. Newlands' Octaves



Note of music	sa (do)	re (re)	ga (mi)	ma (fa)	pa (so)	da (la)	ni (ti)
Element Atomic mass	H 1	Li 7	Be 9	B 11	C 12	N 14	O 16
Element Atomic mass	F 19	Na 23	Mg 24	Al 27	Si 28	P 31	S 32
Element Atomic mass	Cl 35.5	K 39	Ca 40				

The properties of the eighth element were found to resemble very closely the properties of the first, wherever one starts in such a grouping. He compared this to the octaves found in musical notes. He suggested a **Law of Octaves** for the classification of elements. For example, starting from hydrogen (H) the eigth element came out to be fluorine (F). The eigth element starting from fluorine was chlorine (Cl). The properties of F and Cl were found similar. Similarly, the properties of Li, Na and K were also found similar.

At the time of Newlands only 62 elements were known. It was found that the law of octaves worked well for lighter elements only upto calcium. But it could not be applied to elements of higher atomic masses, as after calcium every eigth element did not posses properties similar to that of the first.

Let us answer these.

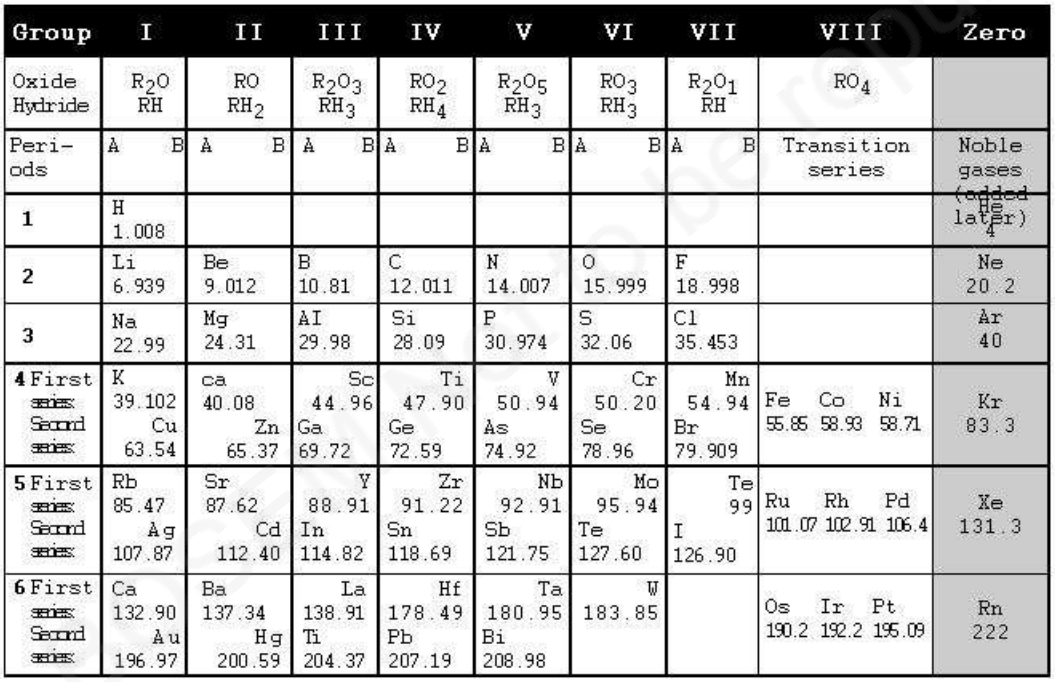
1. What is Dobereiner's law of triads ? Give one example.
2. A, B and C are the elements in Dobereiner's triad. If the atomic mass of A is 40 and that of C is 137, what would be the atomic mass of B ?
3. What is Newlands' Laws of octaves ? Explain with an example.
4. What were the limitations of Newlands' Law of octaves ?

1.1.3. Mendeleev's Periodic Classification

The most notable contribution towards a systematic classification of elements was made by Mendeleev, a Russian chemist in 1869. He studied the physical and chemical properties of the then known 63 elements and their compounds. Among chemical properties, Mendeleev concentrated on the compounds formed by elements with oxygen and hydrogen which are very reactive and formed compounds with most elements.The formula of the oxides and hydrides formed by an element were treated as one of the basic properties of an element for its classification into groups.

He arranged such elements in the increasing order of their atomic masses. He found that elements with similar properties appeared at regular intervals. This discovery led him to formulate a Periodic Law, which states that **"the properties of elements are the periodic functions of their atomic masses"**. This means, when the elements are arranged in order of increasing atomic masses then those with similar properties are repeated at regular intervals. The Mendeleev's periodic table is given in Table 1.4

Table. 1.4. Modified Mendeleev's Periodic Table



Group	I	II	III	IV	V	VI	VII	VIII	Zero
Oxide Hydride	R_2O RH	RO RH_2	R_2O_3 RH_3	RO_2 RH_4	R_2O_5 RH_3	RO_3 RH_3	R_2O_1 RH	RO_4	
Peri- ods	A B	A B	A B	A B	A B	A B	A B	Transition series	Noble gases (added later)
1	H 1.008								He 4
2	Li 6.939	Be 9.012	B 10.81	C 12.011	N 14.007	O 15.999	F 18.998		Ne 20.2
3	Na 22.99	Mg 24.31	AI 29.98	Si 28.09	P 30.974	S 32.06	Cl 35.453		Ar 40
4 First series Second series	K 39.102 Cu 63.54	ca 40.08 Zn 65.37	Sc 44.96 Ga 69.72	Ti 47.90 Ge 72.59	V 50.94 As 74.92	Cr 50.20 Se 78.96	Mn 54.94 Br 79.909	Fe Co Ni 55.85 58.93 58.71	Kr 83.3
5 First series Second series	Rb 85.47 Ag 107.87	Sr 87.62 Cd 112.40	Y 88.91 In 114.82	Zr 91.22 Sn 118.69	Nb 92.91 Sb 121.75	Mo 95.94 Te 127.60	Te 99 I 126.90	Ru Rh Pd 101.07 102.91 106.4	Xe 131.3
6 First series Second series	Ca 132.90 Au 196.97	Ba 137.34 Hg 200.59	La 138.91 Ti 204.37	Hf 178.49 Pb 207.19	Ta 180.95 Bi 208.98	W 183.85		Os Ir Pt 190.2 192.2 195.09	Rn 222

Mendeleev's Periodic Table was published in a German journal in 1872. In the formula for oxides and hydrides at the top of the columns, the letter 'R' is used to represent any of the elements in the group. For example, the hydride of carbon, CH_4*, is written as* RH_4 *and oxide of carbon,* CO_2*, as* RO_2*.*

The periodic table consists of eight vertical columns called **'groups'** and six horizontal rows called **'periods'**. The eight groups are designated as I, II, III, IV, V, VI, VII and VIII. Except group VIII, each group is divided into two sub-groups as A and B. The elements which lie on left hand side of each group constitute sub-group A. These are called **normal or representative elements.** The elements which lie on right hand side of each group constitute sub-group B. These are called **transition elements**. The properties of elements within a sub-group resemble more markedly but they differ from those of the elements of the other subgroup. For example, lithium (Li) , sodium (Na), potassium (K), etc of sub-group I A have close resemblance in properties but they have hardly any resemblance with those of coinage metals (Cu, Ag, and Au) of sub-group I B.

The horizontal rows or periods are numbered 1,2,3, etc.

In arranging the elements in the periodic table, Mendeleev was careful enough to group together elements with similarity in their physical and chemical properties. In doing this he had, in some cases, even to change the order of arranging the elements strictly according to their increasing atomic masses. There was provision for *including unknown elements* as well.

Why is Mendeleev's Classification Appreciated ?

By classifying elements on the basis of their atomic masses, Mendeleev brought some order in the properties exhibited by the elements and their compunds. It made the study of the elements easier. Knowing the properties of one element in a group, the properties of other elements in the group can be easily guessed.

The periodic table could predict errors in the atomic masses of some elements based on their positions in the table. For example, when the table was published the experimental value of the atomic mass of beryllium (Be) was supposed to be 13.65 and its valency, 3. So, the position of Be should have been somewhere else. But Mendeleev placed it at its appropriate position on the basis of its properties and its atomic mass corrected from 13.65 to 9.1 and valency, 2.

While arranging the elements in order of their atomic masses, however, the order had to be ignored at some places to make sure that the elements with similar properties fell in the same group. In doing so, Mendeleev left some vacant places in the table and boldly predicted the existence and properties of some undiscovered elements. For example, **scandium**, **gallium** and **germanium**, discovered later, have properties similar to **eka-boron**, **eka-aluminium** and **eka-silicon** respectively predicted by Mendeleev.

Noble gases like helium (He), neon (Ne) and argon (Ar), were not discovered at the time of Mendeleev. Therefore, there was no group of noble gases in Mendeleev's original periodic table. One of the achievements of Mendeleev's periodic table was that when these gases were discovered they could be placed in a new group called the **zero group** without disturbing the existing order.

Limitations of Mendeleev's Periodic Table

With the discovery of isotopes, the accommodation of the isotopic forms of the same element upset Mendeleev's periodic table. We have learnt that isotopes are the atoms of an element having similar chemical properties (same atomic number), but different atomic masses. If the elements are arranged according to atomic masses, the isotopes have to be placed in different groups of the periodic table. Another problem was that the atomic masses do not increase in a regular manner in going from one element to the next. So it was not possible to predict how many elements could be discovered between two elements especially when we consider the heavier elements.

Let us answer these.

- What was Mendeleev's principle for the classification of elements ?
- Using Mendeleev's Periodic Table, predict the formula for the oxides of the following elements :
  K, C, Al, Si, Ba.
- Besides gallium, which other elements have since been discovered that were left by Mendeleev in his Periodic Table ? (any two).
- Give your reason why the noble gases are placed in a separate group

1.2. Modern Periodic Table

In 1913, Henry Mosely found that the atomic number is a more fundamental property of an element than its atomic mass. This is quite convincing because the atomic number gives us the number of protons in the nucleus and also the number of electrons in an atom. This number increases by one in going from one element to the next higher element. The chemical properties of element depends on the number of electrons in its atom which in turn depends on the atomic number. So, atomic number was adopted as the basis for the classification of elements in the Modern Periodic Table. Thus the modern periodic law states that **“ the properties of elements are the periodic functions of their atomic numbers”**. When elements are arranged in order of increasing atomic number, similar elements occur in regular intervals and this leads us to the classification known as the Modern Periodic Table (Table 1.5).

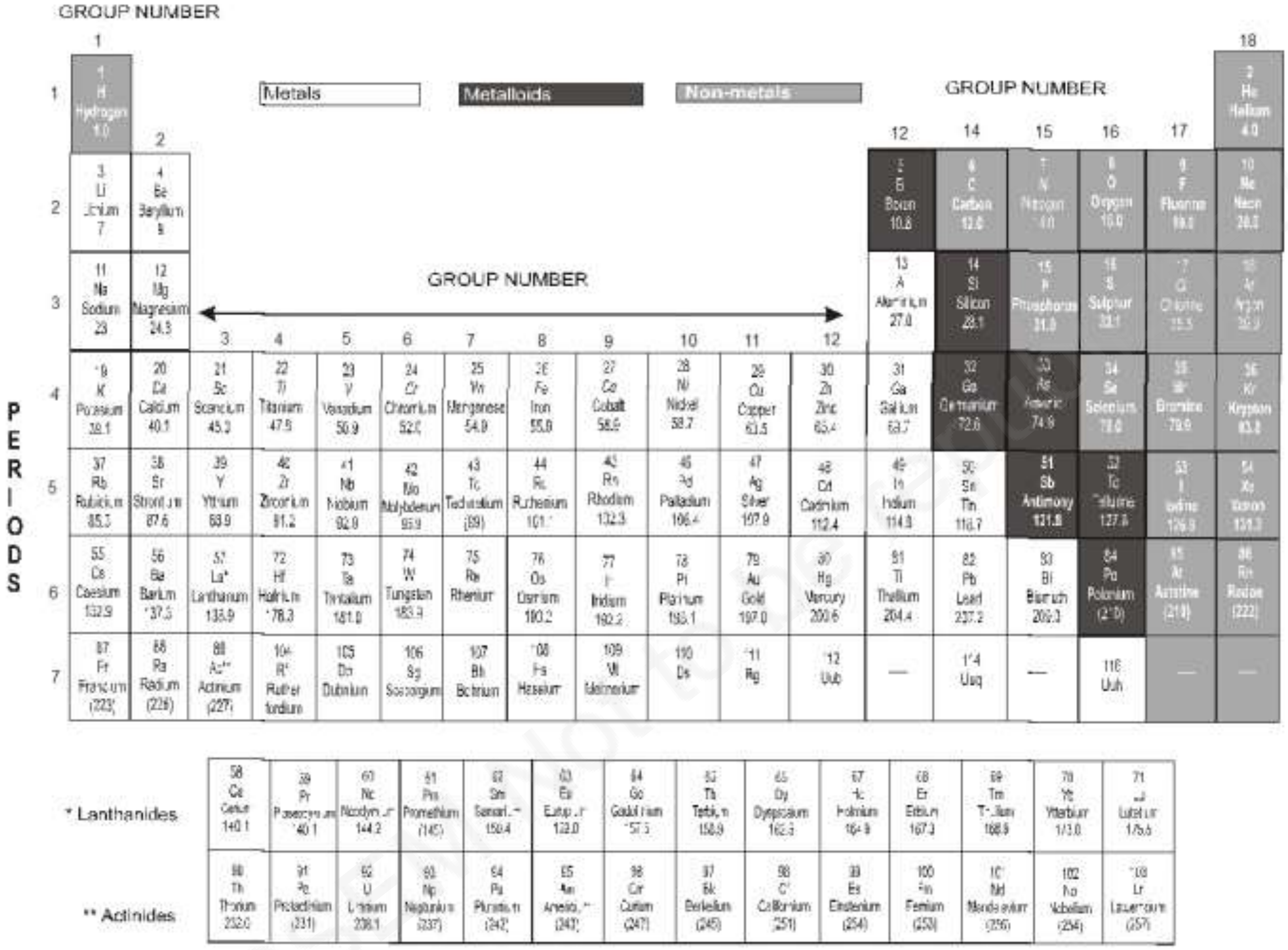
The modern periodic table has 7 horizontal rows called ‘**periods**’ and 18 vertical columns called **‘groups’** or **families**.

Periods:

Look at the periodic table (Table 1.4). We find that the elements in the periods are as follows:

1st period contains 2 elements, H(1) and He (2). It is called very short period.
2nd period contains 8 elements, Li(3) to Ne (10). It is called short period.
3rd period contains 8 elements, Na(11) to Ar(18). It is also a short period.
4th period contains 18 elements, K(19) to Kr (36). It is called long period.
5th period contains 18 elements, Rb(37) to Xe(54). It is also a long period.
6th period contains 32 elements, Cs (55) to Ru(86). It is called very long period.
7th period now contains 25 elements, starting from Fr(87) out of a possiblity of 32 elements. This period is still incomplete.

Two series of elements, 14 each, are placed at the bottom of the main periodic table. These are called **inner-transition elements**. The first series of 14 elements from atomic numbers 58 to 71 (Ce to Lu) is called **Lanthanide series**. The second series of 14 elements from atomic numbers 90 to 103 (Th to Lr) is called **Actinide Series**. They are placed separately at the bottom of the periodic table to avoid undue side-wise expansion of the table and to keep elements with similar properties in a single column.



GROUP NUMBER

Metals | Metalloids | Non-metals

PERIODS

	1	2	3	4	5	6	7	8	9	10	11	12	12	14	15	16	17	18
1	1 H Hydrogen 1.0																	2 He Helium 4.0
2	3 Li Lithium 7	4 Be Beryllium 9											5 B Boron 10.8	6 C Carbon 12.0	7 N Nitrogen 14.0	8 O Oxygen 16.0	9 F Fluorine 19.0	10 Ne Neon 20.2
3	11 Na Sodium 23	12 Mg Magnesium 24.3											13 Al Aluminium 27.0	14 Si Silicon 28.1	15 P Phosphorus 31.0	16 S Sulphur 32.1	17 Cl Chlorine 35.5	18 Ar Argon 39.9
4	19 K Potassium 39.1	20 Ca Calcium 40.1	21 Sc Scandium 45.0	22 Ti Titanium 47.9	23 V Vanadium 50.9	24 Cr Chromium 52.0	25 Mn Manganese 54.9	26 Fe Iron 55.8	27 Co Cobalt 58.9	28 Ni Nickel 58.7	29 Cu Copper 63.5	30 Zn Zinc 65.4	31 Ga Gallium 69.7	32 Ge Germanium 72.6	33 As Arsenic 74.9	34 Se Selenium 79.0	35 Br Bromine 79.9	36 Kr Krypton 83.8
5	37 Rb Rubidium 85.5	38 Sr Strontium 87.6	39 Y Yttrium 88.9	40 Zr Zirconium 91.2	41 Nb Niobium 92.9	42 Mo Molybdenum 95.9	43 Tc Technetium (99)	44 Ru Ruthenium 101.1	45 Rh Rhodium 102.9	46 Pd Palladium 106.4	47 Ag Silver 107.9	48 Cd Cadmium 112.4	49 In Indium 114.8	50 Sn Tin 118.7	51 Sb Antimony 121.8	52 Te Tellurium 127.6	53 I Iodine 126.9	54 Xe Xenon 131.3
6	55 Cs Caesium 132.9	56 Ba Barium 137.3	57 La* Lanthanum 138.9	72 Hf Hafnium 178.5	73 Ta Tantalum 181.0	74 W Tungsten 183.9	75 Re Rhenium	76 Os Osmium 190.2	77 Ir Iridium 192.2	78 Pt Platinum 195.1	79 Au Gold 197.0	80 Hg Mercury 200.6	81 Tl Thallium 204.4	82 Pb Lead 207.2	83 Bi Bismuth 209.0	84 Po Polonium (210)	85 At Astatine (210)	86 Rn Radon (222)
7	87 Fr Francium (223)	88 Ra Radium (226)	89 Ac** Actinium (227)	104 Rf* Rutherfordium	105 Db Dubnium	106 Sg Seaborgium	107 Bh Bohrium	108 Hs Hassium	109 Mt Meitnerium	110 Ds	111 Rg	112 Uub	—	114 Uuq	—	116 Uuh	—	—

* Lanthanides	58 Ce Cerium 140.1	59 Pr Praseodymium 140.1	60 Nd Neodymium 144.2	61 Pm Promethium (145)	62 Sm Samarium 150.4	63 Eu Europium 152.0	64 Gd Gadolinium 157.3	65 Tb Terbium 158.9	66 Dy Dysprosium 162.5	67 Ho Holmium 164.9	68 Er Erbium 167.3	69 Tm Thulium 168.9	70 Yb Ytterbium 173.0	71 Lu Lutetium 175.0
** Actinides	90 Th Thorium 232.0	91 Pa Protactinium (231)	92 U Uranium 238.1	93 Np Neptunium (237)	94 Pu Plutonium (242)	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (245)	98 Cf Californium (251)	99 Es Einsteinium (254)	100 Fm Fermium (253)	101 Md Mendelevium (256)	102 No Nobelium (254)	103 Lr Lawrencium (257)

Groups : The 18 vertical columns (groups or families) are numbered from 1 to 18. Group 18 is also known as **zero group** which consists of noble gases. Elements of the groups 1,2 and 13 to 18 are called **normal** or **representative elements** while those of groups 3 to 12 are called **transition** elements. At present we shall study only the first twenty elements of the eight groups (groups 1, 2 and 13 to 18) of the normal or representative elements.

Activity 1.1.

- Look at the periodic table. Find out the elements present in Group 1 of the table.
- Write down electronic configuration of the first four elements of the group.
- What similarity do you find in their electronic configurations ?
- How many valence electrons are present in these elements ?
- Repeat the above activities with those elements of group 2 and 17 elements.
- Compare your results with Table 1.6

Table 1.6. Electronic configuration of some elements of group 1,2 and 17

Group-1 elements	Atomic no.	Electronic configuration K, L, M, N,	Group-2 elements	Atomic no.	Electronic configuration K, L, M, N,	Group-17 elements	Atomic no.	Electronic configuration K, L, M, N
H	1	1						
Li	3	2, 1	Be	4	2, 2	F	9	2, 7
Na	11	2, 8, 1	Mg	12	2, 8, 2	Cl	17	2, 8, 7
K	19	2, 8, 8, 1	Ca	20	2, 8, 8, 2			

We find that the elements of group 1 have the same number of valence electrons i.e. 1, and those of group 2 and 17 have 2 and 7 valence electrons respectively.

Similarly, you will find that elements present in any one group have the same number of valence electrons. Hence, we can say that groups in the periodic table signify identical outer shell electronic configuration. On the other hand, the number of shells increases as we go down the group.

The electronic configuration of the first twenty elements in the Periodic Table are shown in Table 1.7.

Table 1.7. Electronic configuration of first twenty elements in the Periodic Table

Period	**Group1**	**--------------GROUP NUMBER-------------------**						**Group 18**
1st period	H 1	2	13	14	15	16	17	2 He 2
2nd period	3 Li 2,1	4 Be 2,2	5 B 2,3	6 C 2,4	7 N 2,5	8 O 2,6	9 F 2,7	10 Ne 2,8
3rd period	11 Na 2,8,1	12 Mg 2,8,2	13 Al 2,8,3	14 Si 2,8,4	15 P 2,8,5	16 S 2,8,6	17 Cl 2,8,7	18 Ar 2,8,8
4th period	19 K 2,8,8,1	20 Ca 2,8,8,2						
Valence Electrons	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**

The chart shows the arrangement of first twenty elements in the periodic table on the basis of their electronic configuration. The numbers given above the symbols of the elements are atomic numbers while the numbers given below the symbols are electronic configurations.

The atomic number of Hydrogen and Helium are 1 and 2 respectively. They have 1 and 2 electrons respectively in the K-shell. The atomic number of Lithium is 3, its electronic configuration is 2, 1. Thus Lithium has 1 electron in its outermost shell. With the increase in atomic number, the number of electrons in the outermost shell increases from 1(in Li) to 8 (in Ne). When the atomic number is 11 (in sodium), the electronic configuration is 2, 8, 1. Thus Li and Na have the same number of electrons in the outermost shell (they belong to group 1). They also have similar properties. Thus with the repetition of electronic configuration, the properties of elements are also repeated. In the same way, the group 17 elements having 7 electrons in their outermost shell show similar chemical properties.

Activity 1.2

☞ Find out the elements present in the 2nd and 3rd period of the periodic table.
☞ Write down their electronic configuration.

- ☞ Do these elements contain the same number of valence electrons ?
- ☞ Do they contain the same number of shells ?

We have observed that the elements in a period do not have the same number of valence electrons as we move from left to right, though they contain the same number of shells. Or we can say that atoms of different elements with the same number of occupied shells are placed in the same period. H and He belong to the first period, since the electrons in the atoms are filled in K shell. Na, Mg, Al, Si, P, S, Cl and Ar belong to the third period as the electrons of these elements are filled in K, L and M shells.

The progression of 2, 8, 8, 18, 18, 32, 32 elements in the seven periods are in accordance with the electronic configuration of elements. We have already learnt in class IX that the maximum number of electrons that can be accommodated in any shell is given by the formula, $2n^2$ where 'n' is the number of the given shell. For example,

K – shell : $2\times(1)^2 = 2$, hence first period has two elements, i.e. , H and He.

L – shell : $2\times(2)^2 = 8$, hence second period has eight elements, from Li to Ne.

M – shell : $2\times(3)^2 = 18$, but the outermost shell cannot have more than 8 electrons. Therefore, third period also has eight elements from Na to Ar.

The maximum number of elements which can be accommodated in 4th, 5th, 6th and 7th period are 18, 18, 32 and 32 respectively. You will learn about these in higher classes.

The position of an element in the periodic table tells us about its chemical reactivity. As you have learnt, the valence electrons determine the chemical characteristic of the element. Differences in electronic arrangement results in variation of the properties of the elements. In other words, elements having similar electronic configuration of the valence shell have similar chemical properties. On arranging the elements in the increasing order of their atomic numbers, the elements with the same valence electrons appear at regular intervals. Hence, the chemical properties of elements show similarities.

1.3. TRENDS ON THE BASIS OF MODERN PERIODIC TABLE

From the study of the modern periodic table, we find that there is a regular periodic repetition of the electronic configuration of elements as the atomic number increases. Consequently, in some properties of elements there is a periodic trend or the rythmic behaviour of the atoms as a function of the atomic number.

1.3.1. Valency

We have learnt that the combining capacity or the valency of an element is determined by the number of valence electrons present in the outermost shell of its atom.

Activity 1.3.

- ☞ Write down the electronic configuration of sodium (atomic number 11), magnesium (at. no. 12) and oxygen (at. no. 8). What are their valence electrons? What are their valencies?
- ☞ Similarly, find out the valencies of the first twenty elements.
- ☞ Find out how does the valency vary in a period on going from left to right ?
- ☞ How does the valency vary in a group?

We find that on moving from left to right in a period, the number of valence electrons in elements increases from 1 to 8 (Table 1.8). In the first period, the increase is from 1 to 2.

Table 1.8 Valency of the Elements of 2nd and 3rd period

Elements of 2nd Period

Elements	Li	Be	B	C	N	O	F	Ne
Atomic number	3	4	5	6	7	8	9	10
Electronic configuration.	2,1	2,2	2,3	2,4	2,5	2,6	2,7	2,8
Valence electrons	1	2	3	4	5	6	7	8
Valency	1	2	3	4	3	2	1	0

Elements of 3rd Period

Elements	Na	Mg	Al	Si	P	S	Cl	Ar
Atomic number	11	12	13	14	16	16	17	18
Electronic Configuration	2,8,1	2,8,2	2,8,3	2,8,4	2,8,5	2,8,6	2,8,7	2,8,8
Valence electrons	1	2	3	4	5	6	7	8
Valency	1	2	3	4	3	2	1	0

Let us recall what we have learnt in class IX, that atoms of elements having a comletely filled outermost shell i.e. 8 electrons show little chemical activty. An outermost shell , which has eight electrons was said to possess an octet in the outermost shell. In other words, the valency of the atoms having octet in their valence shells are zero. These elements are known as inert or noble ***gas elements.*** *They are all gases and belong to group 18 of the periodic table. Of these inert elements, helium atom has two electrons in its outermost shell i.e. K-shell for which it represents full or stable structure.*

The combining capacity (valency) of the atoms of an element is the tendency to react or form molecules with other atoms of the same or different elements. It was thus explained as the tendency to attain a fully filled outermost shell or stable structure, i.e. octet of electrons in the outermost shell (except Helium structure which has two electrons in its outermost shell i.e. K-shell).

The number of electrons gained, lost or shared, so as to make the octet of electrons in the outermost shell, gives us directly the valency of the element. In other words, in a period, the valency of an element is either equal to the number of electrons in the valence shell or eight minus the number of electrons in the valence shell.

Thus, we find that on moving left to right in each of the short periods, the valency of elements increases from 1 to 4 and then decreases to 0 (zero).

The elements in a given group of the periodic table have the same valence electron (s). Hence, in a group, the elements have the same valency (Table 1.9)

Table 1. 9 . Valency of elements

Group I Elements				**Group 17 Elements**			
Symbol	At.no.	Electronic configuration	Valency	Symbol	At.no.	Electronic configuration	Valency
Li	3	2,1	1	F	9	2,7	1
Na	11	2,8,1	1	Cl	17	2,8,7	1
K	19	2, 8, 8, 1	1	Br	35	2,8, 18, 7	1

1.3.2. Atomic Size

The term atomic size refers to the radius of an atom. It may be visualised as the distance between the centre of the nucleus and the outermost shell of an isolated atom. The atomic radius is generally expressed in ‘picometre’ units whose symbol is ‘pm’.

1 picometre = m.
or 1 pm = 10^{-12} m.

The atomic radii of hydrogen and helium are 37 pm and 93 pm respectively.

The following table (Table1.10) gives the atomic radii of the second and third period elements.

Table 1.10. Atomic radii of 2nd and 3rd period elements in picometres (pm)

Elements								
2nd period	Li	Be	B	C	N	O	F	Ne
Atomic radius(pm)	123	90	80	77	75	74	72	112
3rd period	Na	Mg	Al	Si	P	S	Cl	Ar
Atomic radius (pm)	186	160	143	118	110	104	99	154

It is observed that the *atomic radius decreases in moving from left to right along a period for normal elements.* The elements of the groups 1, 2, and 13 to 18 are called normal or representative elements. In a period, the atomic number increases from left to right. This means that the number of protons (positive charges of nucleus) and electrons in the atoms increases. The electrons are added to the same shell while the charge on the nucleus increases steadily. The increase in nuclear charge tends to pull the electrons closer to the nucleus and reduces the size of the atom. The atomic radii of noble gases (group 18 elements) have comparatively larger atomic radii than their preceding groups. This is due to the increase of electronic repulsions between completely filled outer shells.

How does atomic radius vary as we go down a group ? Table 1.11gives the atomic radii of group 1 and group 17 elments.

Table 1.11 Increase in atomic radii in a group

Group 1 Elements	Atomic radius (pm)	Group 17 Elements	Atomic radius (pm)
Li	123	F	72
Na	186	Cl	99
K	231	Br	114
Rb	244	I	133
Cs	262		

We can see that the atomic size increases down the group. This is because electrons are added to new shells as we go down the group. This increases the distance between the outermost electrons and the nucleus so that the atomic radius increases in spite of the increase in nuclear charge.

1.3.3. Metallic and non-metallic character

Activity 1.4

- Examine the elements of the second and the third period. Which of them are metals and which are non-metals ?
- On which side of the periodic table do you find the metals ?
- On which side of the periodic table do you find the non-metals ?

We can see that metals like lithium, sodium, magnesium etc. are found towards the left-hand side of the periodic table while the non metals like oxygen, sulphur, chlorine etc. are found on the right hand side. Towards the middle, we find boron and silicon, which are classified as semi-metals or metalloids because they exhibit some properties of both metals and non - metals.

Thus, we find that **on moving from left to right in a period, the metallic character of elements decreases i.e, it changes from metallic nature to metalloid, then to non-metallic and finally the last element in a period is inert in nature (Table 1.12).**

Table 1. 12. Metallic and non-metallic character of elements in 2nd & 3rd period

Elements of 2nd Period	Li	Be	B	C	N	O	F	Ne
Metallic/non-metallic character	metal	metal	meta-lloid	non metal	non metal	non metal	non metal	Noble gas
Elements of 3rd period	Na	Mg	Al	Si	P	S	Cl	Ar
Character	metal	metal	metal	meta-lloid	non metal	non metal	non metal	Noble gas

→ *Metallic or Electropositive character decreases* →

In the modern periodic table, a zig-zag line separate metals from non-metals. The borderline elements - boron, silicon, germanium, arsenic, antimony, tellurium and polonium are intermediate in properties and are called **semi-metals** or **metalloids**. We shall learn in the next chapter that metals lose electrons and form positive ions, so metals are electro positive in nature. The tendency of an element to lose electrons and form positive ions (cations) is called **Electropositive character.** Since metal have such a tendency the electropositive character is also known as **metallic character**.

Thus, metals like Na, K and Mg are electropositive, because they have high tendency to lose electrons.

$Na \longrightarrow Na^{+} + e^{-}$
2, 8, 1 → 2, 8 (electron)
(atom) (sodium ion)

$K \longrightarrow K^{+} + e^{-}$
2, 8, 8, 1 → 2, 8, 8 (electron)
(atom) (potassium ion)

$Mg \longrightarrow Mg^{2+} + 2e^{-}$
2, 8, 2 → 2, 8 (electron)
(atom) (magnesium ion)

The valency of Na and K is 1 and that of Mg is 2.

On the other hand, non-metals accept electrons and form negative ions, so non-metals are called electronegative elements. The tendency of an element to accept electrons to form an anion is called **electronegative character** or **non-metallic character.** For example, the elements fluorine (F), Chlorine (Cl) and Oxygen (O) are electronegative in nature, because they have high tendency to gain electrons.

$$\underset{\substack{2,\,7\\(\text{atom})}}{\mathrm{F}} \quad + \quad \underset{(\text{electron})}{e^{-}} \quad \longrightarrow \quad \underset{\substack{2,\,8\\(\text{fluoride ion})}}{\mathrm{F}^{-}}$$

$$\underset{2,\,8,\,7}{\mathrm{Cl}} \quad + \quad e^{-} \quad \longrightarrow \quad \underset{\substack{2,\,8,\,8\\(\text{chloride ion})}}{\mathrm{Cl}^{-}}$$

$$\underset{2,\,6}{\mathrm{O}} \quad + \quad 2\,e^{-} \quad \longrightarrow \quad \underset{\substack{2,\,8\\(\text{oxide ion})}}{\mathrm{O}^{2-}}$$

As we move from left to right in a period, the effective nuclear charge (positive charge on nucleus) increases due to gradual increase in the number of protons. Due to this, the valence electrons are pulled in more strongly by the nucleus and the tendency to lose electrons will decrease. Down the group, the effective nuclear charge experienced by valence electrons is decreasing because the outermost electrons are farther away from the nucleus due to the addition of new shells. Therefore, these can be lost easily. Hence **metallic character (or electropositive nature) decreases across a period and increases down a group**. *The reverse will be the trend in electronegative character of the elements.*

1.3.4. Nature of Oxides of the elements

We know that oxides of metals are basic and that of non-metals are acidic in general. With the knowledge of the trends in metallic and non-metallic properties of elements in the periodic table, can you predict the nature of the oxides formed by the elements ?

As one moves from left to right in a period, the basic nature of the oxides gradually decreases and acidic nature of the oxides gradually increases. For example, in the third period of the periodic table, sodium oxide is highly basic and magnesium oxide is comparatively less basic. The aluminium and silicon oxides are amphoteric while phosphorous oxides are acidic. Sulphur oxides are more acidic whereas chlorine oxides are highly acidic in nature (Table 1.12).

Table 1.12. Nature of oxides of third period Elements

Elements of 3rd period	Na	Mg	Al	Si	P	S	Cl	Ar
Oxides	Na_2O	MgO	Al_2O_3	SiO_2	P_2O_5	SO_2SO_3	Cl_2O_7	
Nature	Strongly basic	Basic	Amphoteric	Amphoteric	Acidic	Acidic	Highly Acidic	

Can you predict what will be the basic character down a group? As the electropositive character increases down a group , so also is the basic character. For example, calcium oxide is more basic than magnesium oxide and strontium oxide is more basic than calcium oxide.

The modern periodic table is a very satisfactory attempt to the classification of elements. It summarises the vast knowledge of the chemical elements and helps us to see the relationship between their physical and chemical properties and their electronic configuration to a very large measure. However it is still not free from certain defects, the details of which you will learn in higher classes. Chemists are still trying to evolve a still better periodic table.

Let us answer these.

1. What name has been given to the horizontal arrangement of elements in the periodic table ? Is there any gradation in the properties of elements across these rows? Give two examples to illustrate.
2. How do the properties of group 1 elements differ from those of group 17 elements? Illustrate the differences in at least two properties ?
3. From the standpoint of electronic configuration, determine which element will be the first and which will be the last in a period of the periodic table ?

4. Accommodation of isotopic form of the same element was an upset in Mendeleev’s periodic table. How could the Modern Periodic table remove the anomaly ?

5. Name two elements you would expect to show chemical reactions similar to lithium. What is the basis of your choice ?

6. **Name**

(i) three elements having one electron in their shells,

(ii) two elements having four electrons in their outermost outer most shells,

(iii) three elements with completely filled shells.

7. Fluorine, chlorine and bromine belong to group 17 of the periodic table.What similarity is there in the atoms of these elements ?

8. By considering their position in the periodic table, which one of the following elements is expected to have maximum metallic character ?

In, Sn, Sb, Te, I.

9. How would be the tendency to gain electrons change as we go from left to right across a period ?

10. How would be the tendency to gain electrons change as we go down a group ?

11. Why is non-metals are found on the right-hand side of the periodic table towards the top?

SUMMARY

- Elements are classified on the basis of similarities in their properties.
- The earliest attempt to classify the elements resulted in grouping the then known elements into two groups, metals and non-metals.
- Dobereiner classified elements having similar properties into triads and stated that when elements are arranged in the order of increasing atomic masses in a triad, the atomic mass of the middle element was approximately equal to the arithmetic mean of the atomic masses of the other two elements.

- Newlands found that when elements are arranged in the order of increasing atomic masses, the properties of the eight element were found to resemble the properties of the starting element.
- Mendeleev arranged the elements in increasing order of their atomic masses and according to their chemical properties. He stated that the properties of elements are periodic functions of their atomic masses.
- Mendeleev's periodic table could predict the properties of several elements on the basis of their positions in the periodic table. He even predicted the existence of some yet to be discovered elements on the basis of gaps in the periodic table.
- In the modern periodic table, the basis for the classification of elements is the atomic number which is a more fundamental property of an element. It could remove the anomalies in the arrangement of elements based on increasing atomic mass.
- Elements in the modern periodic table are arranged in seven horizontal rows called periods and 18 vertical columns called groups.
- Elements thus arranged show periodicity in their properties.
- Trends like atomic size, valency, metallic and non-metallic character of elements could be predicted from their position in the periodic table.

EXERCISES

1. Why is there a need to classify the elements ? What difficulties are likely to arise if the classification is not made ?
2. Name the period and group to which the element with atomic number 15 belongs.
3. State one reason for keeping boron and aluminium in the same group of the periodic table.
4. Write the electronic configurations of the elements Na, Al, Ca, O and Si and justify their positions in the periodic table.
5. Why is atomic number more important than atomic mass to know the properties of the elements ?

6. State how atomic size varies from left to right in a period and from top to bottom in a group.

7. **The atomic radii of the elements of second period are given below:**

2nd period elements :	B	Be	O	N	Li	C
Atomic radii (pm):	80	90	74	75	123	77

(a) Arrange the elements in the decreasing order of their atomic radii, keeping the element with the largest atomic radius first.

(b) Find out whether the arragement is in the pattern of a period or not?

(c) From this data, infer how the atomic size of the elements changed as one goes from left to right in a period.

8. By considering their position in the periodic table, which one of the following elements would you expect to have the maximum metallic character ?

Ga, Ge, As, Se, Br.

9. **The position of three elements A, B and C in the periodic table are shown below.**

Group16	Group 17
—	—
—	A
—	—
B	C

(a) State whether A is a metal or non-metal ?

(b) State whether C will be more reactive or less reactive than A

(c) Will C be larger or smaller in size than B ?

(d) What type of ion (cation or anion) will be formed by A ?

10. Element X forms a chloride with the formula XCl_2. With which of the following elements, X would most likely be in the same group of the periodic table ?

Na, Mg, Al, C.

11. Name two elements you would expect to show chemical properties similar to Li. What is the basis of your choice ?

12. What properties do all elements in the same group of boron in the periodic table have in common ?

13. **An atom has electronic configuration 2, 8, 6.**

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

(b) Which of the following elements would it be similar in chemical properties?

N(7), F(9), P(15), O(8), Ar(18)

14. Nitrogen (at.no. 7) and Phosphorous (at. no. 15) belong to group 15 of the periodic table. Which of them will be more electronegative and why?

15. In the modern Periodic table, calcium (at. no. 20) is surrounded by elements with atomic numbers 12, 19, 21 and 38. Which of them have physical and chemical properties resembling calcium ?

★★★★★★

CHAPTER

2

CHEMICAL BONDING

We have learnt in previous classes that the smallest particle of an element or compound that is capable of an independent existence and shows all the properties of that substance is called a **molecule**. Molecules are made up of atoms. Molecule of an element is composed of atoms of the same kind. For example, a molecule of hydrogen consists of two atoms of hydrogen, two atoms of oxygen unite to form a molecule of oxygen etc. Atoms of different elements combine together in definite proportions to form molecules of compounds. For example, hydrogen chloride is a compound whose molecules are made up of one atom of hydrogen and one atom of chlorine. Similarly, a molecule of water contains two atoms of hydrogen and one atom of oxygen. When atoms of elements combine to form molecules, a force of attraction is developed between the combining atoms which holds them together. This force is called **chemical bond**. Thus, a **chemical bond is one which holds the atoms together in a molecule.**

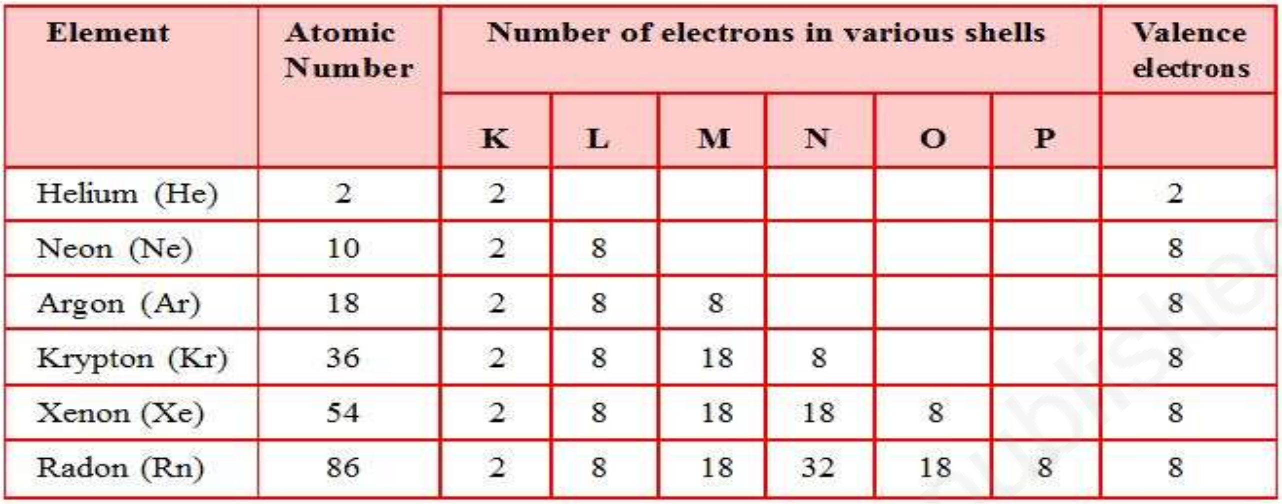
2.1. ATTAINMENT OF A STABLE ELECTRONIC CONFIGURATION

How do atoms combine together to form molecules and why do atoms form bonds ? The molecule formed must be more stable than the individual atoms or molecules would not be formed. It means that the process must be energetically favourable and lead to a minimum energy.

In the previous class, we have studied the structure of atom. We have also learnt that the chemical properties of elements depend upon the number of electrons present in the outermost shell of the atom or **valence electrons.** It is a well known fact that atoms of the noble gases do not normally react with any other atoms and their molecules are mono-atomic, that is, they contain only one atom. The lack of reactivity indicates that these atoms are extremely stable and their energy is so favourable that it cannot be further minimised by compound formation. They must be having a complete outer shell of electrons, so that we conclude that this is a very stable electronic arrangement or configuration.

The electronic configuration of these elements are given below (Table 2.1).

Table 2.1. Electronic configuration of noble gas atoms



Element	Atomic Number	Number of electrons in various shells						Valence electrons
		K	L	M	N	O	P	
Helium (He)	2	2						2
Neon (Ne)	10	2	8					8
Argon (Ar)	18	2	8	8				8
Krypton (Kr)	36	2	8	18	8			8
Xenon (Xe)	54	2	8	18	18	8		8
Radon (Rn)	86	2	8	18	32	18	8	8

We can see from the above table that all the atoms of noble gases (with the exception of helium) have eight electrons in the valence shell. Helium, on the other hand, has two electrons in its first and the only energy level (K-shell) which cannot hold more than two electrons. *Two electrons in the first (K-shell) and the only shell of helium (He) and eight electrons in the outermost shells of other noble gases give a stable configuration of electrons.* We can therefore, conclude that atoms having eight electrons in their outermost shells (two in case there is K-shell) are very stable and unreactive. Eight electrons in the outermost shell of an atom is called **octet** of electrons. Two electrons in the outermost shell of helium is called **duplet** of electrons. Thus, eight electrons in the outermost shell of an atom makes the atom stable. When there is only K-shell, in an atom, two electrons in this shell make the atom stable.

All atoms other than noble gases have less than eight electrons in their outermost shells. In other words, the outermost shells of the atoms of these elements do not have stable configurations.

Thererefore, **atoms combine with each other or with other atoms to acquire stable nearest noble gas electronic configurations. The principle of attaining eight electrons in the valence shell of atoms is called octet rule**.

However, it may be noted that for hydrogen and lithium atoms, the stable configuration means two (duplet) electrons which corresponds to the nearest noble gas helium. This is called **duplet rule**.

In the light of octet rule, Kossel and Lewis, in 1916 developed an important theory of chemical combination between atoms known as **Electronic Theory of Valency**. According to this theory, atoms combine to acquire the nearest noble gas electronic configuration by losing, gaining or sharing electrons. This leads to the formation of chemical bonds between the combining atoms.

There are three common types of bonds. These are

(i) Electrovalent or ionic bond

(ii) Covalent bond

(iii) Co-ordinate bond.

2.1.1. Electrovalent or Ionic Bond

In the previous chapter, we have seen that elements can be divided into (a) metals or electropositive elements, whose atoms give up one or more electrons fairly readily, and (b) non-metals or electronegative elements which take up electrons.

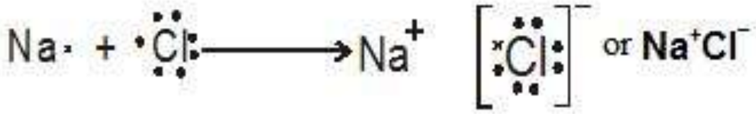
An ionic or electrovalent bond is formed when an electropositive element combines with an electronegative element. Let us consider the sodium atom (atomic number 11), which has the electronic configuration 2, 8, 1. The first and second shells of electrons are full, but the third shell contains only one electron. When this atom reacts, it will do so in such a way that a stable electron configuration is attained. The nearest noble gas neon has the configuration 2,8. If the sodium atom can lose one electron from its outer shell, it will attain this configuration except that sodium atom becomes sodium ion with a net charge of +1. This is because there are eleven positive charges on the nucleus and only ten electrons. When supplied with energy sodium atoms tend to do this, so sodium is an electropositive element.

$Na \longrightarrow Na_+$ + electron
2,8,1 2,8
Sodium Sodium
atom ion

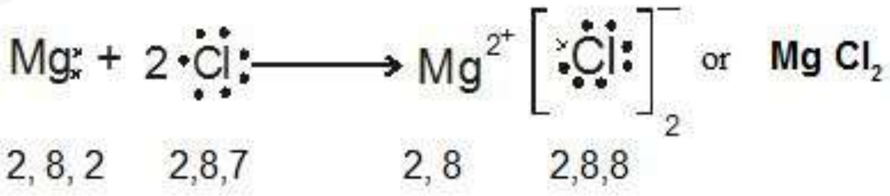
Chlorine atom has the electronic configuration 2, 8, 7. It is only one electron short of the stable inert gas configuration of argon, 2, 8, 8. Chlorine atoms tend to react and gain electrons that is they are electronegative. By gaining one electron, the electrically neutral chlorine atom becomes a chloride ion with a net charge of 1–.

atom

When sodium and chlorine react together, the outer electron of the sodium atom is transferred to the chlorine atom to produce sodium ion Na^+ and chloride ion Cl^-. The two ions (Na^+ and Cl^-), being oppositely charged, are now held together by the electrostatic force of attraction. Since both sorts of atoms attain a very stable inert gas configuration, the process is energetically favourable and sodium chloride, Na^+Cl^- is formed readily. This may be illustrated diagrammatically using dots and crosses to represent the valence electrons only as follows:

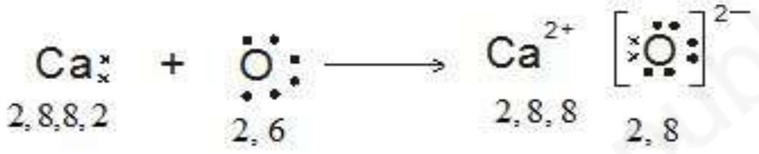


In a similar way, an atom of magnesium (Mg) can combine with two atoms of chlorine (Cl) to form magnesium chloride $MgCl_2$. The electronic configuration of Mg (atomic number 12) is 2, 8, 2. It has 2 electrons in its outermost shell. Thus the Mg atom transfers its two valence electrons to two Cl atoms and changes into a stable magnesium ion (Mg^{2+}), with a stable octet of electrons in its outer shell. The positively charged magnesium ion (Mg^{2+}) and negatively charged chloride ions (Cl^-) are now held together by the electrostatic force of attraction.



It should be noted that two chlorine atoms are needed to take up the two electrons coming from a single magnesium atom.

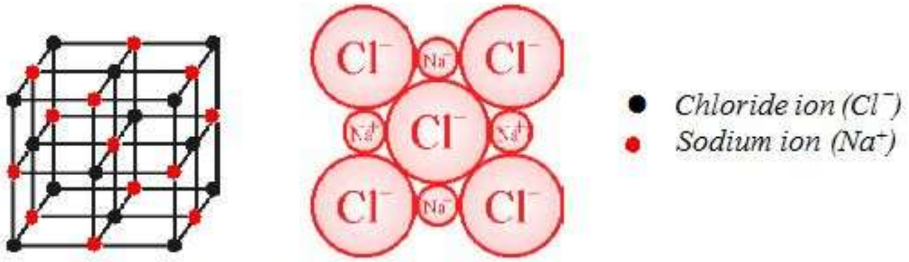
Combination of calcium (Ca) and oxygen (O) atoms to form calcium oxide (CaO) can be explained in a similar way. The atomic numbers of Ca and O are 20 and 8 respectively. So, their electronic configurations are 2, 8, 8, 2 and 2, 6, respectively, so that , Ca atom has two valence electrons, whereas the O atom has 6. When the two atoms combine, the two valence electrons of Ca are transferred to the valence shell of oxygen. In doing so, calcium, as well as oxygen atom, achieve a stable octet of electrons and thus attain stable noble gas configuration. Ca atom acquires two positive charges and becomes calcium ion (Ca^{2+}) and oxygen atom, two negative charges and becomes oxide (O^{2-}) ion. The two oppositely charged ions (Ca^{2+} and O^{2-}) are now held together by electrostatic force of attraction forming calcium oxide $Ca^{2+}O^{2-}$ or CaO.



In the above examples, we find that the stable structure like that of the noble gases is attained by combining atoms by the transfer of one or more electrons from one atom to another. **The type of combination between atoms which depend upon the transfer of electrons and in which oppositely charged ions are held together by electrostatic force of attraction, is called electrovalent or ionic bond**. *The number of electrons lost or gained by an atom to attain the stable electronic configuration is known as the electrovalency of the atom.* The compounds formed by electrovalent or ionic bonds are called electrovalent or ionic compounds.

Characteristics of Ionic or Electrovalent Compounds

(a) **Ionic compounds are generally crystalline in nature**. X-rays studies of these compounds have shown that they do not exist as single molecules but as ions which are arranged in a well defined geometric pattern. This is because of the fact that the forces of attraction are not restricted to a single unit, such as one Na^+ and one Cl^-, but due to uniform electric field around an ion, each ion is attracted to a large number of other ions. The ionic bond is **non-directional**. This results into a three dimensional aggregate of the oppositely charged ions arranged in a regular manner. For example, the arrangement of ions in sodium chloride is such that each Na^+ ions is surrounded by six Cl^- ions and each Cl^- ion is surrounded by six Na^+ ions.



The arrangement of ions in sodium chloride crystal.
The right hand side diagram shows the relative size of ion.
The left hand side diagram shows only the arrangement of ion, but not their relative sizes.

Fig. 2.1. Formation of aggregates in ionic crystals.

(b) **Ionic compounds have high melting and boiling points**. This is due to the strong electrostatic force of attraction between the oppositely charged ions. As a result of this, large amount of energy is needed to break the bond and hence melting and boiling points are high.

(c) **They conduct electricity in molten state and in their aqueous solution.** Solid ionic compounds do not conduct electricity as the ions are fixed at particular positions and are not free to move. However, when these are melted or dissolved in water, the ions become free and, therefore, conduct electricity.

(d) **Ionic compounds are usually soluble in water and other polar solvents** but insoluble in organic solvents such as benzene, ether, carbon tetrachloride, kerosene etc.

Let us answer these.

1. Which electronic configuration favour inert or noble behaviour of elements ?
2. Why do most of the elements form ions ?
3. What kind of elements form cation ? Support your answer by two examples?
4. What kind of elements form anions ? Support you answer by two examples.
5. Sodium atom reacts vigorously with water, but sodium ion does not. Why ?
6. What is an ionic bond ? State at least three conditions for its formation.
7. A metal A (atomic number 19) burns in chlorine to produce a white solid chloride ACl. By means of diagrams, show the arrangement of electrons in A before and after the reaction.
8. Give reasons of the following : (a) Ionic compounds have high melting and boiling points. (b) Ionic compounds are good conductors of electricity in fused state or aqueous solution.

2.1.2. Covalent Bond

The formation of ionic bond cannot explain the chemical combination between identical atoms or atoms of similar chemical nature to form molecules. Thus it cannot account for the formation of molecules like H_2, O_2, Cl_2, CO_2; and also for the organic molecules like methane (CH_4), carbon tetrachloride (CCl_4) etc. In such cases, the combining atoms attain the stable noble gas configuration by sharing electrons. The sharing always occurs in pairs and in each pair of shared electrons, one electron is contributed by each atom. **The bond formed by mutual sharing of electrons between the combining atoms is called covalent bond**. Sharing of one pair of electrons form a **single bond**, two pairs a **double bond** and three pairs a **triple bond**.

Let us consider the formation of hydrogen molecule (H_2) from two atoms of hydrogen. The nucleus of hydrogen has one proton and the only one electron revolves in the K-shell which is also the valence shell. In the formation of hydrogen molecule, H_2, when the two hydrogen atoms approach each other, the single electron of both the atoms form a shared pair. This electron pair revolve around the two nuclei of H-atoms i.e., the two H-nuclei are bound together by the electron pair. It is contained in the electron shells of both the atoms and each hydrogen atom attain the electronic configuration of the nearest inert gas helium. In the conventional system of writing the structural formula of molecules, a single covalent bond is represented by a dash (–) between the combining atoms. This may be represented as follows.

$$\underset{\substack{1\\ \text{atom}}}{H\cdot} \quad + \underset{\substack{1\\ \text{atom}}}{\cdot H} \longrightarrow H{:}H \text{ or } H-H \text{ or } H_2$$

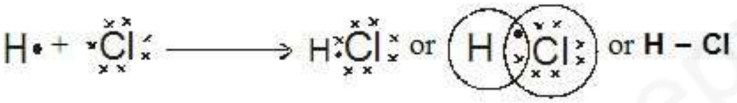
Pictorially hydrogen molecule may be represented as . Once the bond is formed, both the atoms have a stable configuration of the noble gas helium.

Formation of Chlorine molecule (Cl_2)

The atomic number of chlorine is 17. The electronic configuration of the Cl atom is 2, 8, 7. Thus, a Cl atom needs one electron to complete its octet. The two combining atoms of chlorine gives one of its electrons to be shared between them and constitute a shared pair of electrons. The two shared electrons are counted for both the atoms. Thus, each Cl atom in a chlorine molecule (Cl_2) acquires a stable configuration. Showing the outer (valence) electrons only, this may be represented as follows:

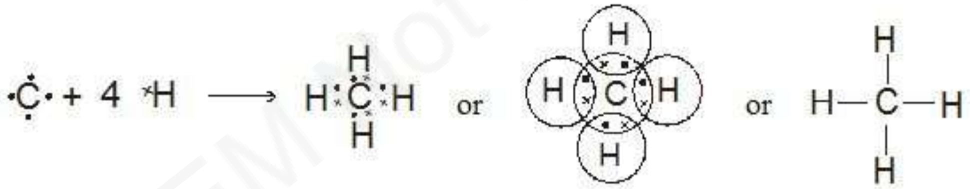
Formation of hydrogen chloride (HCl) molecule

Hydrogen atom has only one electron and chlorine atom has seven electrons in its valence shell. Hence, by mutual sharing of electron pair between them, both the atoms acquire nearest noble gas configurations. Hydrogen atoms acquire stable electronic configuration (duplet) of helium whereas chlorine atoms acquire stable electronic configuration (octet) of argon.



Formation of methane (CH_4) molecule

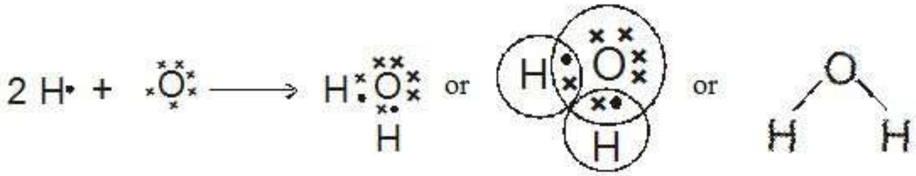
A carbon atom has four electrons in its valence shell. It shares its valence electrons with those of four H atoms. Thus, an atom of carbon forms four single covalent bonds with four H atoms.



In a molecule of methane, the carbon atom and each of the four H atoms acquire the stable electronic configurations of neon and helium respectively. In methane, we have four carbon-hydrogen (C–H) single bonds. Thus, a methane molecule has four pairs of shared electrons.

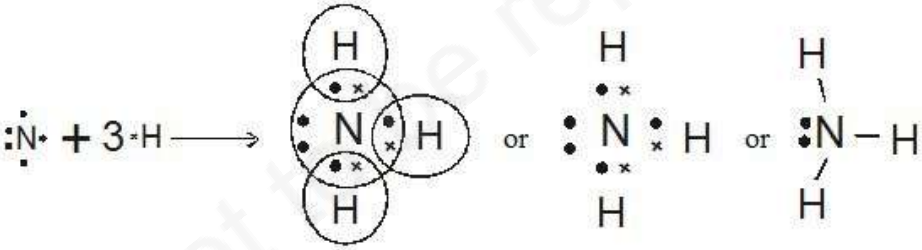
Formation of water (H_2O) molecule

A molecule of water is made up of one oxygen atom and two hydrogen atoms. An oxygen atom is two electrons short and hydrogen atoms are one electron short of a noble gas structure. Thus one atom of oxygen shares two of its electrons with two hydrogen atoms, thus attaining a stable configuration.



Formation of ammonia (NH_3) molecule

A molecule of ammonia is formed by the combination of one atom of nitrogen with three atoms of hydrogen. There are five electrons in the valence shell of nitrogen atom. So, it has three electrons short of to complete its octet and to be stable. When it shares three of its electrons with three H atoms, the stable electronic configuration of neon is achieved. The three H atoms acquire the stable helium configuration.



Multiple Covalent Bonding

We have seen that electron pair shared between two atoms is responsible for the formation of a covalent bond. This shared pair is also called **bonding pair of electrons**. The pair of electrons which do not take part in bonding are called **lone pairs**. The bonding atoms may share more than one pair of electrons also depending upon their requirement to complete the stable elctronic configuration or octet.

If the two atoms share one electrons pair, the bond is known as **single covalent bond** and is represented by one dash (–).

If the two atoms share two electron pairs, bond is known as **double covalent bond** and is represented by two dashes $(=)$.

If the two atoms share three electron pairs, bond is known as **triple covalent bond** and is represented by three dashes $(\equiv)$.

Formation of Double Bond

Two oxygen atoms combine to form an oxygen molecule(O_2) by sharing two electron pairs. We know that each oxygen atom has six electrons in the valence shell and requires two electrons to complete the octet. Therefore, in the formation of oxygen molecule, both the atoms contribute two electrons each for sharing resulting in two electron pairs. Hence there is double bond between the two oxygen atoms.

A molecule of carbon dioxide (CO_2) is formed when an atom of carbon combines with two atoms of oxygen. The C atom has four electrons in its valence shell and thus requires four more electrons to complete its octet while oxygen atom requires two electrons to complete its octet. This is achieved when the carbon atom shares its four electrons with two oxygen atoms to form a CO_2 molecule.

Formation of Triple Bond

In the formation of nitrogen molecule (N_2), each of the two nitrogen atoms having five valence electrons, provides three electrons to form three electron pairs for sharing. Thus a triple bond is formed between the two atoms.

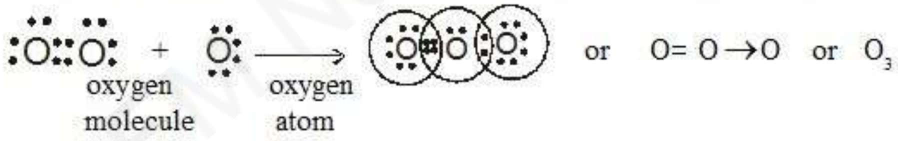
We have seen from the above examples that in the formation of covalent molecules, an atom shares electrons with other atoms as needed to acquire nearest noble gas configurations (octet or duplet).

The type of combination between atoms which results from the mutual sharing of pair of electrons between the combining atoms is called covalent bond and the molecules formed are called covalent molecules. The number of electrons contributed by an atom for sharing in the formation of a covalent molecule is known as **covalency**. Thus, the covalency of hydrogen, oxygen and nitrogen in the above examples is one, two and three respectively.

2.1.3. Co-ordinate or Dative Bond

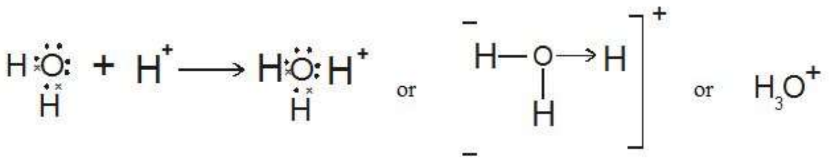
A covalent bond results from the sharing of an electron pair, where each atom contributes one electron to the bond, as seen above. It is also possible to have an electron pair bond in which both the electrons of the shared pair came originally from one of the atoms. Such type of bond is formed between atoms, one of which is deficient in at least two electrons while the other has already acquired a stable configuration (or octet). The atom which contributes the electron pair is called the *donor* while the other which only shares the electron pair is known as *acceptor*. Such bonds are called **co-ordinate covalent or dative bonds**. The usual representation of co-ordinate bond is to point an arrow (→) from donor to the acceptor.

Let us consider the formation of ozone (O_3) molecule. A molecule of oxygen (O_2) contains two atoms of oxygen joined by covalent double bond so that each atom has stable octet configuration. In O_2 molecule, each oxygen atom has two pairs of unused valence electrons. Such pairs of electrons which are not involved in bond formation are called **lone pairs**. Now, if an atom of oxygen having six valence electrons comes close to O_2 molecule, the new atom may share a lone pair of electrons of one of the oxygen atoms of O_2 molecule. It gives rise to the formation of a co-ordinate covalent bond.



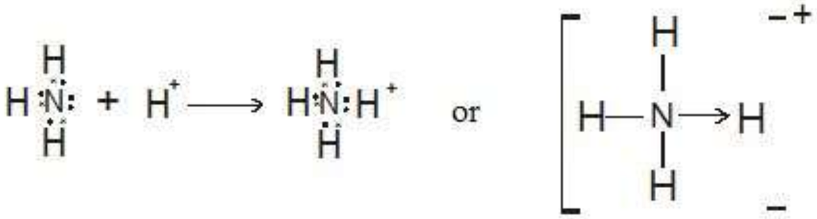
Formation of hydronium ion (H_3O^+) and ammonium (NH_4^+) ion

Water, H_2O has lone pair of electrons on oxygen atom while hydrogen ion (H^+) has empty K-shell thereby can accommodate two electrons. Thus, oxygen in H_2O can easily provide two electrons while H^+ ion can accept these



Hydronium ion

Similarly ammonium ion (NH_4^+) is formed by co-ordinate bond between ammonia (NH_3) molecule and hydrogen ion (H^+).



Ammonium ion

Characteristics of Covalent substances

Covalent substances show the following characteristic properties.

(a) **Physical state :** The covalent compounds do not exist as ions but they exist as molecules. There are weak inter molecular forces between the molecules. Hence, generally they exist as liquids or gases at room temperature.

(b) **Low melting and boiling points**: Since the forces between the molecules are weak, a comparatively small amount of heat energy is required to overcome these forces. Hence covalent compounds generally have low melting and boiling points.

(c) **Solubility:** In general, they are sparingly soluble in polar solvents like water but soluble in non-polar solvents like benzene, carbon tetrachloride, kerosene, etc.

(d) **Conductivity :** They do not conduct electricity. This is because they are made up of neutral molecules, not ions.

2.1.4. Do covalent bonds show partial ionic character ?

If we consider a covalent bond between similar atoms, the shared pair of electrons is equally attracted by the two atoms. As a result of this, it is situated exactly in between the two identical nuclei. The common examples are H_2, F_2, Cl_2 etc.

H_2

Cl_2

F_2

These are examples of pure covalent bonds.

However, if the atoms forming bonds are different, the electron attracting tendency of the two atoms may differ. *The relative tendency of an atom to attract shared electrons to itself when combined in a compound is termed as the electronegativity of the atom.* As a result of the difference in electronegativity, the shared electron pair will be displaced towards the atom with more electron attracting tendency or electronegativity. For example, the compound hydrogen chloride is formed by sharing a pair of electrons between the hydrogen and the chlorine atom.

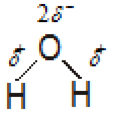
Of the two nuclei, the attraction exerted on the shared pair of electrons by the chlorine nucleus is more than that by the hydrogen nucleus. In the hydrogen chloride molecule, therefore, the shared electron pair would be attracted more towards the chlorine end of the molecule. So the chlorine end of the molecule will be slightly negative and the hydrogen end slightly positive, these are represented as $\delta^+$and δ^- charges.

or $H^{\delta+}—Cl^{\delta-}$

In other words, the hydrogen chloride molecule develops a partial ionic character even though it is a covalent molecule. Such molecules are said to be **partially polar**. Almost all covalent bonds between unlike atoms are partially polar. They are thus **polar covalent compounds**.

Similarly, the molecules of HF and H_2O are polar covalent compounds.

$H^{\delta+}—F^{\delta-}$



Unlike the covalent bond in the HF and HCl molecule, the water (H_2O) molecule contains two covalent bonds between the oxygen and hydrogen atoms. These two bonds are not symmetrically arranged, and H_2O molecule is V-shaped. This results in the polar character of water. Hence water is a polar solvent.

Let us answer these.

1. Using dots (•) and crosses (×) to represent valence electrons, show the formation of covalent bonds in CO_2 and CCl_4.
2. Why does carbon tetrachloride not conduct electricity ?
3. How does a co-ordinate bond differ from a normal covalent bond ?
4. Why does H_2 molecule exist while He_2 does not exist ?
5. Why is HF is polar while H_2 and F_2 are not polar.

2.3. BONDING IN METALS

We know that metals are generally hard solids and they are made up of atoms. It has been established that the atoms in metals are closely packed together.

What force holds the atoms in metals closely together ? What type of bonding exist between the atoms of metals ? We have learnt that metal atoms can easily lose electrons to form positively charged ions and they are good conductors of electricity. This implies that the valence electrons of metal atoms are loosely held by the nucleus and can move freely.

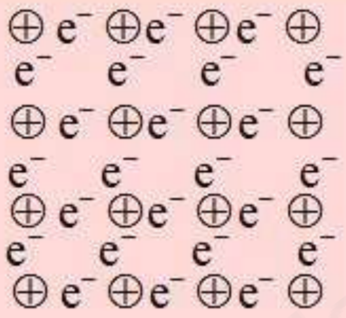


Fig. 2.1. Electron sea model of metal.

To account for the bonding in metals, Lorentz has proposed a model known as *electron sea model*. According to this model, a metal lattice consists of positively charged **kernels** arranged in a regular way surrounded by loosely held valence electrons. Kernels are metal nuclei and other electrons except valence electrons. Alternatively, it can be assumed that there is a sea of valence electrons in which positively charged **kernels** are immersed. It is for this reason that it is called **electron sea model**. *The simultaneous force of attraction between the positively charged kernel and the mobile electrons which binds the metal atoms together is called metallic bond.*

Let us consider lithium metal. The electronic configuration of lithium is 2, 1; it has one valence electron. Without the valence electron, the kernel is Li^+. The arrangement of kernel of lithium atom, ⊕ and valence electrons (e^-) in its metal lattice is shown in Fig.2.1.

SUMMARY

The interaction between two or more atoms or groups of atoms holding the atoms together is called chemical bond.

- Noble gas atoms have stable electronic configurations.
- The type of combination which depends upon the transfer of electrons and in which oppositely charged ions are held together by electrostatic force of attraction is called electrovalent or ionic bond.
- Atoms which can lose electrons and form positive ions (cations) are said to be electropositive and the elements are called electropositive elements. Metals are generally electropositive.
- Atoms which can gain electrons and form negative ions (anions) are said to be electronegative and the elements are called electronegative elements. Non-metals are generally electronegative.
- The number of electrons gained or lost by an atom in the formation of ionic bonds is equal to its electrovalency.
- Covalent bonds are formed by mutual sharing of electrons between the combining atoms.
- Compounds formed by covalent bonds are called covalent compounds.
- The number of electrons contributed by an atom for sharing in the formation of covalent bonds to attain stable configuration is known as its covalency.
- A co-ordinate covalent bond is formed by sharing of electrons between two atoms where both the electrons of the shared pair come originally from one atom and none from the other .
- The force that holds the atoms closely together by simultaneous force of attraction by between the positively charged kernel and the mobile valence electrons in a metal is known as the metallic bond.

EXERCISES

1. Define the terms ‘valence shell’ and ‘valence electrons’.
2. Why do elements combine chemically ?
3. What is an electrovalent bond and how is it formed ?
4. What is a covalent bond and how is it formed ?
5. How does an ionic bond differ from a covalent bond ? Illustrate your

Answer with two examples of each type of bond.

6. What is a polar molecule ? Give example.

7. Give electron dot representation for the following molecules.

F_2, HF and H_2O.

8. State four properties each of electrovalent and covalent compounds.

9. The elements *w*, *x*, *y* and *z* have atomic numbers 7, 9, 10 and 11 respectively. Write the formula of the molecules that would be formed between the following pairs of elements and indicate the type of the bonding present using dots (**.**) and crosses (×) to represent valence electrons.

(a) *w* and *x* **(b)** *x* and *x* **(c)** *w* and *z* **(d)** *y* and *y*.

★★★★★★

CHAPTER

3

ACIDS, BASES AND SALT

You know the tastes of various fruits and vegetables. Some fruits and vegetables such as lemon, tamarind, green mango, oxalis, Indian goose berry (amla), unripe grapes, curd are sour to taste. Do you know the taste of sodium bicarbonate ? It is bitter in taste. Add some sodium bicarbonate on a half-cut lemon. What do you see ? There will be gas bubbles escaping from the lemon surface.

Those substances which have sour taste are said to contain acids. Different types of acids are present in different kinds of food materials. Acids present in food materials are generally called **organic acids**. Oxalic acid, lactic acid, tartaric acid, citric acid are some common organic acids found in food materials. Other types of acids are hydrochloric acid, sulphuric acid, nitric acid, phosphoric acid etc. which are called **mineral acids**. They are generally prepared from minerals. The liquid present in lead storage battery is dilute sulphuric acid. It is sour to taste.

Substances like potassium hydroxide, sodium hydroxide, calcium hydroxide etc. which are bitter to taste, are called **bases**. Sodium carbonate and sodium bicarbonate also shows basic character in their solutions. Bases react with acids to form compounds called **salts**. You know the salty taste of common salt.

Acids change the blue colour of litmus to red. Bases turn the colour of red litmus to blue. The colour of litmus generally remains unaffected in salt solutions. You must have seen the change of curry stain (turmeric stain) in soap solution. Turmeric and litmus are examples of natural indicators. They are called so because they differentiate acidic and basic solutions by changing their colours. Here we will discuss the reaction of acids and bases. We will also study how acids and bases interact with each other in reactions known as **neutralisation reactions.**

Litmus solution is a water soluble purple dye extracted from Lichen. Lichen is a plant belonging to Thallophyta family. Acid turns the purple color of litmus red while base turns it into blue. There are many natural substances like turmeric, red cabbage leaves, colour petals of Petunia, Geranium, Hydrangea etc. which can indicate the presence of acid or base in a solution. They are generally called acid-base indicators or simply indicators .

The teacher can ask the students to identify the solutions by using red and blue litmus papers.

3.1. Chemical Properties of Acids and Bases

3.1.1. Acids and Bases in the Laboratory

Activity 3.1

- Find the following solution samples from your laboratory : sulphuric acid (H_2SO_4) hydrochloric acid (HCl), acetic acid (CH_3COOH), oxalic acid ($H_2C_2O_4$), sodium hydroxide (NaOH), sodium carbonate (Na_2CO_3), ammonium hydroxide (NH_4OH).
- With the help of a dropper or glass tube, take out a few drops of each sample solution on a watch glass or a small test tube. Add a drop of the following indicators as given in the Table 3.1.
- Observe the change in the colour of each solution taken and note them in the Table 3.1.

Table 3.1

Sample Solution	Phenolphthalein	Blue litmus	Red litmus	Methyl orange
Hydrochloric acid				
Sulphuric acid				
Acetic acid				
Oxalic acid				
Sodium hydroxide				
Sodium carbonate				
Ammonium hydroxide				

Activity 3.2

- Prepare an aqueous solution of turmeric powder and filter. Collect the filtrate.
- Collect samples of lime juice, passion fruit juice, unripe tomato juice, soap solution, detergent solution and common salt solution. Filter, if necessary.
- Take 10 drops or about 1mL of each solution in separate test tubes and add 1 or 2 drops of the indicators and mix thoroughly as shown in the Table 3.2.
- Observe the change in colour of each solution taken and note them in the table.

Table 3.2

Sample solution	Turmeric solution	Phenolphthalein
Lime juice		
Passion fruit juice		
Unripe Tomato juice		
Soap solution (Toilet Soap)		
Detergent (washing powder) solution		
Common salt (NaCl) solution		

Let us answer these

1. From the observed facts draw your conclusion comparing with the results of the activity 3.1.
2. Suggest the usefulness of turmeric powder as an indicator.
3. Is sodium chloride solution acidic or basic ?

Indicators tell us the nature of a solution – acidic or basic by a change in their colour in different solutions.

3.1.2. Reaction of acids and bases with metals

Activity 3.3

Under the guidance of your teacher perform this activity

- Set up the apparatus as shown in Fig. 3.1
- Take 2-3 pieces of zinc granules in a double necked Woulf's bottle, fitted with a thistle funnel and a delivery tube.

- ☞ Add some dilute sulphuric acid through the funnel to cover the zinc granules. See that the end of the Thistle funnel is under the acid.
- ☞ Observe the surface of zinc granules. Do you see the bubbles formed on it?
- ☞ Allow the gas formed to pass through the soap solution.

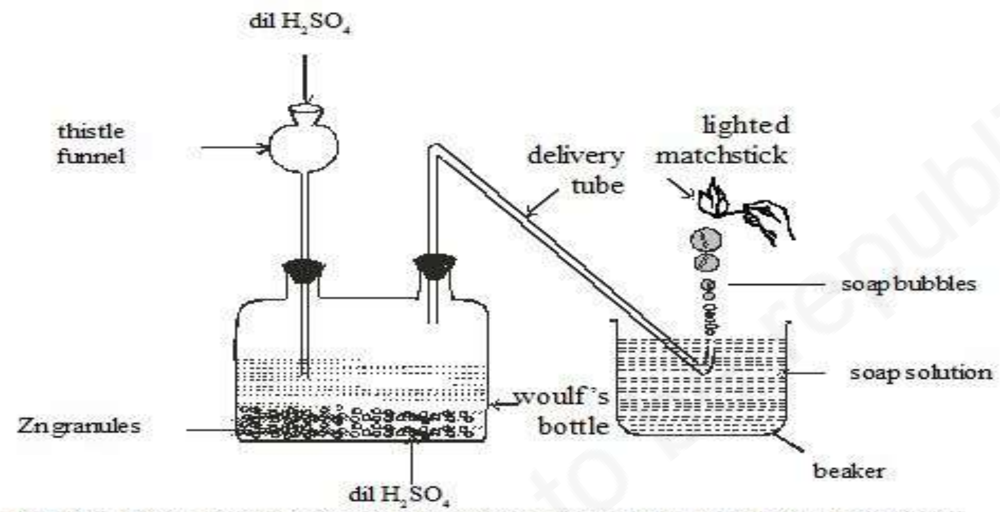


Fig 3.1. Zinc granules react with dilute sulphuric acid to give hydrogen gas. Hydrogen burns in air with a pop sound.

Let us answer these observations:

- ⇨ Why are soap bubbles formed?
- ⇨ Why do the bubbles go up in the air ?
  Compare with the hydrogen balloons you have seen in the market place.
  Did the ballons go up in the air when released ? Why ?
- ⇨ Take a lighted match stick near the rising soap bubbles.
  What do you observe ?
- ⇨ Repeat this activity using acetic acid solution in place of dilute suphuric acid solution.
- ⇨ Are there any differences in your observations ?

Note carefully the composition of acids–sulphuric acid (H_2SO_4), hydrochloric acid (HCl), nitric acid (HNO_3), phosphoric acid (H_3PO_4), acetic acid (CH_3COOH) etc. All of them contain hydrogen and another part e.g., Cl^- and SO_4^{2-} in HCl and H_2SO_4 respectively.

In the above activity, the hydrogen (H) of the acid is replaced by the metal releasing hydrogen gas (H_2). Then the metal combines with the other part of the acid to form a compound which is called salt. In short, we can write

Acid + Metal   Salt + Hydrogen gas

For example,

Sulphuric acid + Zinc   Zinc sulphate + Hydrogen gas

H_2SO_4 + Zn   Zn SO_4 + H_2

Can you write the equation for the reaction between zinc and acetic acid ?

Activity 3.4.

Repeat the activity 3.3 using pieces of copper metal instead of zinc granules

- ☞ Do you see any gas formed on the metal surface ?
- ☞ Can you draw a conclusion from this observation ?
- ☞ Does this verify the general statement that
  Acid + Metal   Salt + Hydrogen gas ?
- ☞ Are there exceptions to the above statement ?

Activity 3.5.

Repeat the activity 3.3 using dil sodium hydroxide (NaOH) in place of dilute sulphuric acid.

- ☞ You can repeat the activity using fine pieces of aluminium metal with dilute sodium hydroxide.
- ☞ Can you draw a conclusion from these observations ?
- ☞ Repeat all the steps in activity 3.3 and record all your observations.

Here, the metal zinc reacts with the alkali sodium hydroxide to release hydrogen gas and forms a salt. The reaction can be written as

Sodium hydroxide + Zinc   Sodium zincate (salt) + Hydrogen gas

2 NaOH + Zn Na_2ZnO_2 + H_2

Aluminium can also release hydrogen in a similar way. Thus,

Sodium hydroxide + Aluminium +Water Sodium meta aluminate + Hydrogen gas

$$2NaOH + 2Al + 2H_2O \quad 2NaAlO_2 + 3H_2$$

Alluminium powder and NaOH solution is used to prepare hydrogen for gas balloon sold in the market. Here, you find that hydrogen gas is formed in the reaction. However, such reactions do not occur with all metals. Besides zinc and aluminium, other elements which react with sodium hydroxide to liberate hydrogen are tin and silicon.

Lime and lime water are also bases. A suspension of lime (CaO) in water is called milk of lime but the clear supernatant liquid is called lime water. Chemically, it is calcium hydroxide $Ca(OH)_2$. Lime water readily absorbs carbon dioxide forming calcium carbonate which is insoluble in water.

Lime is used in the treatment of acidic soil. The acids in soil are released by the decay of organic matter. Lime neutralises these acids. Nitrifying bacteria (those bacteria which can increase the nitrogen content of the soil) which cannot function in acid soils become active again when the soil is treated with lime.

3.1.3. Reactions of acids with metal carbonates and metal hydrogen carbonates.

Activity 3.6

- Take about 0.5g of a metal carbonate like zinc carbonate, magnesium carbonate, sodium carbonate etc. in a test tube as shown in the Fig 3.2.
- Add about 3ml of dilute hydrochloric acid into the test tube.
- Note the evolution of a colourless gas.
- Pass the gas into a test tube containing clear lime water as in the figure.
- Record the changes you have seen.

Now repeat the above steps, taking sodium bicarbonate or baking soda in place of metal carbonate and record your observations.

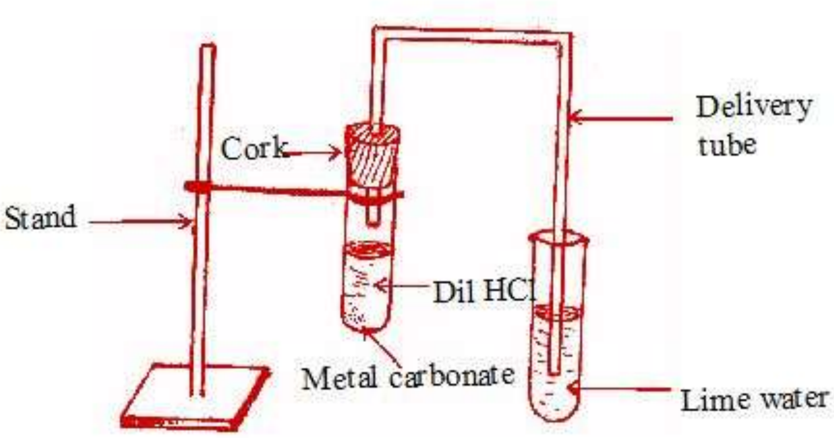


Fig. 3.2. Action of acids on metal carbonates and bicarbonates.

The reaction taking place in the above can be written as

Metal carbonate + Hydrochloric acid Metal chloride + Water + Carbon
(salt)   dioxide

e.g., $ZnCO_3$ (s) + 2 HCl(aq) $ZnCl_2$ (aq) + H_2O (l) + CO_2 (g)

zinc carbonate   zinc chloride

Na_2CO_3 (s) + 2 HCl (aq) NaCl(aq) + H_2O(l) + CO_2(g)

sodium carbonate   sodium chloride

With sodium bicarbonate (sodium hydrogen carbonate) the reaction is

$NaHCO_3$ (s) + HCl (aq) NaCl(aq) + H_2O(l) + CO_2(g)

Sodium hydrogen carbonate

When the evolved carbon dioxide is passed into lime water, the clear solution becomes milky due to the formation of insoluble calcium carbonate. The reaction is

Ca $(OH)_2$(aq) + CO_2(g) $CaCO_3$(s) + H_2O (l)
Lime water   Calcium carbonate
(insoluble)

When carbon dioxide is passed into lime water for a longer period, the milkyness of the solution will start disappearing. The solution will again become clear. This is due to the conversion of insoluble calcium carbonate into soluble calcium hydrogen carbonate. The reaction is

$CaCO_3$ (s) + H_2O(l) + CO_2(g) $\longrightarrow$ Ca $(HCO_3)_2$ (aq)
(insoluble)   (soluble)

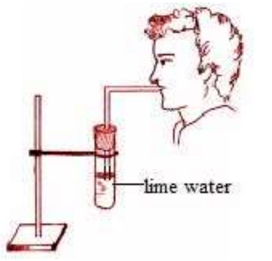


Fig. 3.3. Passing breath into lime water

Activity 3.7.

- Take about 5ml of lime water in a clean test tube.
- Exhale your breath into it as shown in Fig 3.3.
- Observe the change in lime water solution.
- What inference can you derive from the observation you made ?

3.1.4. Neutralisation Reactions

How do Acids and Bases React with each other ?

Different acids and different bases have different strengths. Generally dilute solutions of these substances are prepared by dissolving them in water. Such solutions generally have similar strengths. They are usually available in the laboratory. If different strengths are required for a specific purpose, they are further diluted with fixed amounts of water.

In acid-base reactions, solutions with similar strengths are usually taken. For example, samples of dilute HCl and dil NaOH solutions in the laboratory have nearly the same strength.

Activity 3.8

- Take about 5ml of dilute NaOH solution in a test tube and add two drops of phenolphthalein indicator.

  (Note that volume of a test tube is about 20 mL).
- Note the colour of the solution
- Add dilute HCl into the above solution drop by drop and mix well.
- Do you see any change in the colour of the mixture ?

- Continue to add dil HCl in the above manner till the solution becomes almost colourless.
- Now add a few drops of dilute NaOH to the above colourless solution.
- What is the colour of this solution ?
- Why is this happened ?

The effect of a base can be annulled by an acid. Similarly, we can find out that the effect of an acid can also be annulled by a base. (For this, repeat the steps of **activity 3.8** taking 5 mL of dilute HCl in the test tube first and adding dilute NaOH dropwise to it after adding two drops of phenolphthalein).

The reaction between the acid and the base is

Acid + Base Salt + Water

e.g. $HCl\ (aq) + NaOH\ (aq)$ $NaCl\ (aq) + H_2O\ (l)$

Such a reaction between an acid and a base producing salt and water is known as **neutralisation reaction**.

The neutralisation reaction between dilute sulphuric acid and sodium carbonate solution can be studied using methyl orange indicator also. For this follow the similar steps in activity 3.8.

Consider again the reaction between dilute HCl and dilute NaOH solutions to form sodium chloride solution and water.

$$HCl\ (aq) + Na\,OH\ (aq) \quad NaCl\ (aq) + H_2O\ (l)$$

Here, 1 mole of HCl react with 1 mole of NaOH to form 1 mole of NaCl and 1 mole of water.

Molecular mass of HCl = 1 + 35.5 = 36.5
Molecular mass of NaOH = 23 + 16 + 1= 40
Molecular mass of NaCl= 23 +35.5 = 58.5
Molecular mass H_2O = 1 ×2 + 16 = 18

Therefore ,

$$HCl\ (aq) + NaOH\ (aq) \quad NaCl\ (aq) + H_2O(l)$$

36.5 g 40 g 58.5 g 18 g

36.5 g of HCl reacts with 40 g of NaOH to produce 58.5 g of NaCl and 18 g of H_2O.

Let us answer these.

1. Dissolve a small amount of $NaHCO_3$ in water. Add 1 drop of methyl orange into it. What is the colour of the mixture ?
   Add lime juice in the above solution till the evolution of the gas stops. What is the colour of the mixture now ?
2. How many grams of sulphuric acid (H_2SO_4) will be needed to react completely with a solution containing 10.6g of sodium carbonate (Na_2CO_3)?

3.1.4. Reaction of Acids with metal Oxides

Activity 3.9

- Collect samples of zinc oxide, magnesium oxide, copper oxide, mangnese dioxide, ferric oxide etc. and note their colours.
- Take a small amount of an oxide in a test tube.
- Add dilute hydrochloric acid into it and warm on a spirit lamp.
- What happens to the oxide ?
- Note the colour of the solution

In the case of colourless oxides, like that of zinc and magnesium , you will get a colourless solution when hydrochloric acid dissolve the oxides. The reaction in case of zinc oxide is

$$ZnO(s) \quad + \quad 2HCl\,(aq) \longrightarrow ZnCl_2\,(aq) \quad + \quad H_2O\,(l)$$

Zinc oxide (white) **Zinc (II) chloride(colourless)**

If you take the black copper oxide you will get a bluish green solution. The reaction is

$$CuO(s) + 2HCl\,(aq) \longrightarrow CuCl_2\,(aq) + H_2O\,(l)$$

Copper (II) oxide **Copper (II) chloride**

(Black) (Bluish green)

Thus the general reaction between a metal oxide and an acid can be written as

Metal oxide + Acid ⟶ Salt + Water

Let us answer these.

1. What is the nature of metal oxides – acidic or basic ?
2. Two metallic compounds A and B react separately with acid to form salt and water. The compound A also produces an effervescence. What could be A and B ?
3. When milk becomes curd lactic acid is produced as one of the products. Give reason why is it not advisable to keep curd in copper and brass vessels?
4. Carbon dioxide reacts with lime water, $Ca(OH)_2$ to form salt, $CaCO_3$ and water as

   $Ca(OH)_2 + CO_2 \; CaCO_3 \quad + \quad H_2O$

   base salt water

   Indicate the nature of CO_2 . What will be the general nature of non-metallic oxides?
5. Write the reaction between sodium hydroxide and sulphur dioxide (SO_2).
6. What will happen when CO_2 gas is passed into water ? Write the chemical equation.
7. Write the reaction that takes place between water and sulphur dioxide.
8. Give definition of base.
9. Define acids.
10. Give definition of acidic oxides.

3.2. What Is Common In All Acids ?

You have seen that most metals react with acids to liberate hydrogen gas (Activity 3.3). You have also come across that metals replace all or some of the hydrogen (H) atoms present in acids and these replaced hydrogens came out as hydrogen (H_2) gas. Let us see the chemical formulas of some common acids :

Hydrochloric acid	HCl
Sulphuric acid	H_2SO_4
Sulphurous acid	H_2SO_3
Nitric acid	HNO_3
Nitrous acid	HNO_2
Phosphoric acid	H_3PO_4

Boric acid	H_3BO_3
Hydrofluoric acid	HF
Hydrobromic acid	HBr

All these acids known as **inorganic acids** (mineral acids), contain at least one or more hydrogen atoms which can be replaced by metals.

Let us see some of the acids found naturally in plants, fruits and insects which have acidic properties.

Plants and animal sources	**Name of the acid present**	**Chemical formula**
Lemon and oranges etc (citrus fruits)	Citric acid	$C_3H_4(OH)(COOH)_3$
Apple, Peach etc	Malic acid	$C_2H_3(OH)_2(COOH)_2$
Tamarind and grapes	Tartaric acid	$C_2H_2(OH)_2(COOH)_2$
Oxalis	Oxalic acid	$(COOH)_2$
Sour Milk (Curd)	Lactic acid	$C_2H_4(OH)COOH$
Vinegar	Acetic acid	CH_3COOH
Ants, Bees, Wasp	Formic acid	HCOOH
Cabbage, green leafy vegetables etc	Ascorbic acid (vitamin C)	$C_5H_5(OH)_4COOH$

If you observe the chemical formulas of these natural acids carefully it can be easily noted that all them are made up of carbon, hydrogen and oxygen elements. Some of the hydrogens are present as –OH, –COOH and some of the hydrogens are attached to carbon. In other words, there are different hydrogens in these acids. These acids are known as **organic acids**. Those hydrogen atoms (H) present only in the form –COOH can be replaced. For example, acetic acid reacts with sodium metal to form sodium acetate liberating hydrogen gas. The reaction is

$$\underset{\text{acetic acid}}{2\,CH_3COOH\,(aq)} + \underset{\text{sodium}}{2\,Na\,(s)} \longrightarrow \underset{\text{sodium acetate (salt)}}{2\,CH_3COONa\,(aq)} + H_2\,(g)$$

Organic acids contain hydrogen atoms which are not acidic i.e., which are not replaceable by metals. In the case of CH_3COOH, the three hydrogen atoms attached to carbon as $–CH_3$ are not replaced by metals i.e., these three hydrogens are not acidic.

3.2.1. What Constitute Bases ?

From the activities 3.5, 3.6 and 3.8 you have learnt that bases may be metal hydroxide, metal oxides. Metal carbonates and metal hydrogen carbonates also show basic character. All of them combine with acid forming a salt and water. For example, sodium hydroxide, NaOH, sodium oxide, Na_2O and calcium hydroxide, $Ca(OH)_2$, are bases. Sodium carbonate, Na_2CO_3 and Sodium hydrogen Carbonate , Na HCO_3 also have basic properties.

Water soluble hydroxide and metal oxide, which dissolve in water to form metal hydroxides are known as alkalis. However, the alkali NH_4OH is not a metal hydroxide.

Thus, sodium hydroxide is an alkali,

$$\underset{\text{sodium oxide}}{Na_2O} + H_2O \quad \underset{\text{sodium hydroxide}}{2Na\,OH}$$

There are metal hydroxides (bases) like aluminium hydroxide $Al(OH)_3$, ferric hydroxide $Fe(OH)_3$, etc. which are not soluble in water. Such hydroxides are not alkalis.

3.2.2. What happens to an acid when dissolved in water ? Will dry HCl show the acidic reactions ?

Activity 3.10

☞ Take a clean dry hard test tube and set up the apparatus as shown in the Fig. 3.4.

☞ Add about 1g of NaCl into it.

☞ Add about 5 mL of conc. H_2SO_4 into the test tube.

☞ Quickly cover the test tube with the cork fitted with delivery tube protected by a calcium chloride guard tube.

☞ Test the issuing gas with a dry blue litmus paper.

☞ Is there any change in the colour of litmus paper ?

☞ Now dip the blue litmus paper in water and test the gas with it as before ?

☞ Do you see any change in the colour of litmus paper ?

Dry litmus paper is unaffected by the dry HCl.

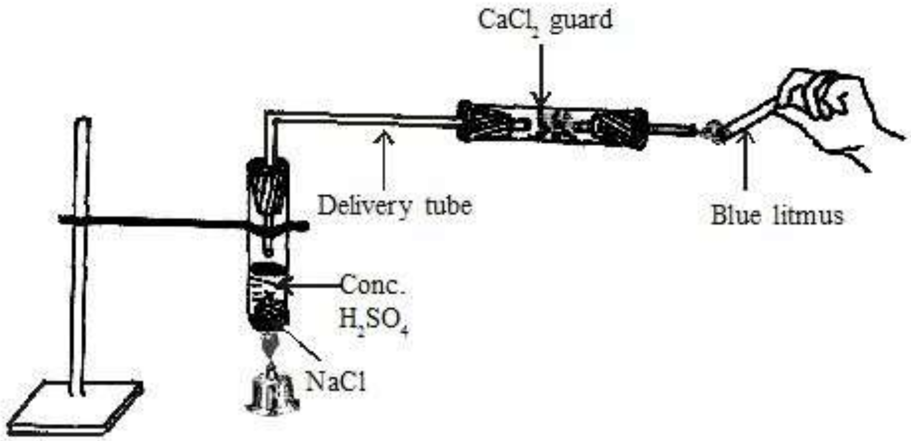


Fig. 3.4. Reaction between conc. H_2SO_4 and NaCl to form HCl

$$\text{Conc } H_2SO_4\,(l) + 2NaCl\,(s) \longrightarrow 2HCl\,(g) + Na_2SO_4\,(aq)$$

However colour of the blue litmus paper becomes red when it is moistened with water. Because dry HCl reacts with water to form hydrogen ions (H^+) as

$$HCl\,(g) + H_2O \longrightarrow \underset{\text{Hydronium ion}}{H_3O^+} + Cl^-$$

Hydrogen ions (H^+) do not exist independently but it readily combine with H_2O to form hydronium ion, (H_3O^+).

$$\underset{\text{Hydrogen ion}}{H^+} + \underset{\text{water}}{H_2O} \longrightarrow \underset{\text{Hydronium ion}}{H_3O^+}$$

When dissolved in water HCl forms ions as

$$HCl\,(g) + H_2O\,(l) \longrightarrow H_3O^+\,(aq) + Cl^-(aq).$$

Similarly, H_2SO_4 in aqueous solution gives

$$H_2SO_4\,(l) + 2H_2O\,(l) \longrightarrow 2H_3O^+\,(aq) + SO_4^{2-}\,(aq)$$

Acetic acid also dissolves in water as

$$\underset{\text{acetic acid}}{CH_3COOH\,(l)} + H_2O\,(l) \longrightarrow H_3O^+\,(aq) + \underset{\text{acetate ion}}{CH_3COO^-(aq)}$$

As you know, the acidic reaction of acids are due to the presence of ionisable or replaceable hydrogens in them. Actually, it is the reactions of hydronium ions (H_3O^+) in solution.

3.2.3 What happens when a base is in aqueous solution ?

When solid sodium hydroxide is dissolved in water, it is ionised to form ions ;

$$\text{Na OH (s)} + \text{Water} \longrightarrow \text{Na}^{+}\text{(aq)} + \text{OH}^{-}\text{(aq)}$$

Similarly all soluble hydroxides of metals ionise to give hydroxide ion OH^- and the metal ion.

$$\text{Ca(OH)}_2\text{(s)} \xrightarrow{\text{water}} \text{Ca}^{2+}\text{(aq)} + 2\text{OH}^{-}\text{(aq)}$$

$$\text{KOH (s)} \xrightarrow{\text{water}} \text{K}^{+}\text{(aq)} + \text{OH}^{-}\text{(aq)}$$

Aqueous solution of alkalis are soapy to touch, bitter to taste and corrosive. You should not directly touch them as it causes burn on your skin. NaOH and KOH are highly corrosive and they are also known as **caustic soda** and **caustic potash** respectively.

You have seen above that soluble metal oxides also produce hydroxides in their aqueous solution .

Does the disolution of acids and alkalis in water involve heat change ?

Activity 3.11

- Take about 10 mL of water in a test tube
- Add about 10 drops of concentrated sulphuric acid to it slowly.
- Hold the test tube in your palm
- What do you feel – hot or cold ?
  Draw your conclusion from the observations you made.
- Now take about 3–4 pellets or flakes of NaOH in a test tube.
- Add about 5 ml of water into it and shake slowly to dissolve.
- Hold the test tube in your palm.
- What do you feel again– hot or cold ?

\

Now let us answer the following.

1. Is the dilution of acid with water exothermic or endothermic ?
2. 'NaOH dissolves in water with the absorption of heat'.
   Is it correct or not ? If not, give the corect statement.
3. Why does an aqueous solution of an acid conduct electricity ?
4. How is dilution affect on the concenctration of H_3O^+ ion of an acid ?
5. What will happen to the concentration of H_3O^+ in a given volume of an acid if it is mixed with a same amount of NaOH solution ?

3.3. Do all acids have the same strength ? Do all bases have the same strength ?

We have seen that all acids produce hydronium ion H_3O^+ in their aqueous solutions. The strength of an acid depends upon the amount of H_3O^+ ion it can produce in equimolar solutions. However, all acids do not produce the same amount of H_3O^+ ions. Therefore, all acids do not have the same strength.

The strength of a base (alkali) depends upon the number of OH^- ions it can produce in equimolar solution. All alkalis do not give the same amount of OH^- ion in solution. Thus, all alkalis are not of the same strength.

It is important to ascertain the strength of a given acid or base. This can be done with the help of a **universal indicator** which is a mixture of indicators. Strips of paper dipped in the universal indicator and dried are available in the market. With the help of this paper the strength of a given acid or alkali solution can be determined.

The strength of acid or base is expressed in a sacle called **pH scale**. The scale of pH values ranges from 0 to 14. Acidic solutions have pH value less than 7 and that of basic solutions above 7. You will learn more about pH in higher classes.

Acid solution that contains higher concentration of hydronium ions H_3O^+ is said to be stronger acid. Higher H_3O^+ ion concentration means the lesser pH value, i.e., solutions with lower pH values are stronger acids.

Solution with higher OH^- ion concentrations is more alkaline and therefore, such solution will have higher pH values. The scale of pH and the nature of the solution is given below:

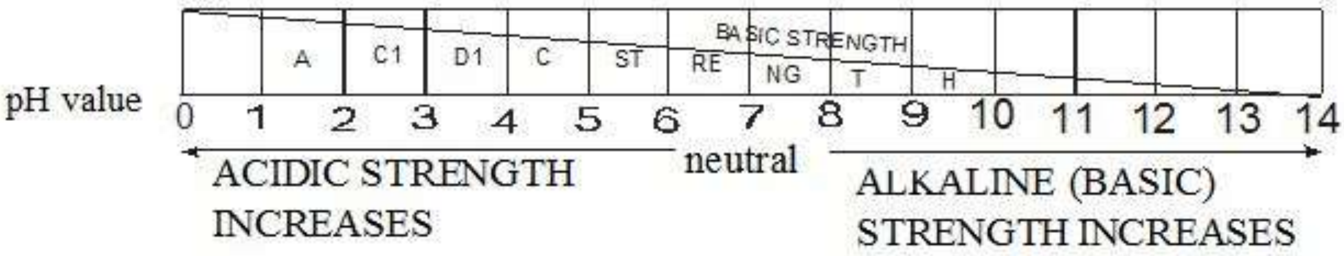


Fig. 3.5 pH scale.

Activity 3.12

- Collect the solutions listed in the Table 3.3. Fruit juices should be collected fresh and filter if necessary.
- Test the solutions with universal indicator paper.
- Match the colour of the test papers with the chart colour given along with the paper.
- Note the pH values of the solutions in the table

Table. 3.3.

Sl. No.	Sample solutions	Colour of P^H paper	P^H value	Nature of the substance
1	Tap water			
2	Mineral water			
3	Black tea			
4	Orange juice			
5	Sodium bicarbonate solution			
6	Sugarcane juice			

3.4. Different Kinds of Salts

We have come across that salts are the products obtained when an acid reacts with a base. We have also seen that all acids do not have the same strength. Similarly all bases do not have the same strength. Thus, the neutralisation of an acid by a base and vice versa may have different products. Let us discuss further.

Activity 3.12.

- Take about 10 mL or $\frac{1}{2}$ test tube of dilute HCl in a conical flask.
- Add two drops of phenophthalein indicator. (Note the colour, if any)
- Then add dil NaOH solution drop by drop, with constant swirling of the conical flask, till the colour of the mixture becomes light pink
- Why does the colour of the solution become pink ?

Here, HCl is neutralised by NaOH to form NaCl and water. The changes can be writen as,

$$HCl(aq) + H_2O \longrightarrow H_3O^+ (aq) + Cl^- (aq)$$

$$NaOH (aq) + H_2O \longrightarrow Na^+ (aq) + OH^- (aq)$$

H_3O^+ ions from the acid solution combine with OH^- ions from the base solution to form unionised water.

$$H_3O^+ + OH^- \longrightarrow 2H_2O$$

All the above changes can be written, in short, as

$$NaOH (aq) + HCl (aq) \longrightarrow NaCl (aq) + H_2O (l)$$

NaCl (aq) will ionise in the solution

$$NaCl (aq) \longrightarrow Na^+ (aq) + Cl^- (aq)$$

Here we get a salt sodium chloride by the complete neutralisation of HCl with NaOH. Such salts are called **normal salts**.

NaCl, Na_2SO_4, $MgCl_2$, KNO_3, $CuSO_4$ etc. are examples of normal salts.

3.4.1. Reaction between NaOH and H_2SO_4

The reaction between dilute NaOH and dilute H_2SO_4 may take place in the following steps:

$$H_2SO_4 (aq) + NaOH (aq) \longrightarrow \underset{\text{sodium hydrogen sulphate}}{NaHSO_4 (aq)} + H_2O (l)$$

$$NaHSO_4 (aq) + NaOH (aq) \longrightarrow \underset{\text{sodium sulphate}}{Na_2SO_4(aq)} + H_2O (l)$$

Even though sodium hydrogen sulphate, $NaHSO_4$, is a salt, it can further react with more NaOH to form the normal salt, sodium sulphate, Na_2SO_4. $NaHSO_4$ has got some acidic property and it can therefore, react with the base, NaOH. Salts like $NaHSO_4$ are called **acid salts**. They are obtained by the incomplete neutralisation of the acid by a base. Some other examples of acid salts are

Sodium bicarbonate (sodium hydrogen carbonate)	$NaHCO_3$
Sodium dihydrogen phosphate	NaH_2PO_4
Disodium hydrogen phosphate	Na_2HPO_4
Calcium bicarbonte	$Ca(HCO_3)_2$

3.4.2. Reaction between magnesium hydroxide Mg $(OH)_2$ and HCl

The reaction between $Mg(OH)_2$ and HCl can be written as

$$Mg\,(OH)_2 + HCl \longrightarrow Mg\,(OH)\,Cl + H_2O$$

Magnesium hydroxychloride

$$Mg\,(OH)\,Cl + HCl \longrightarrow Mg\,Cl_2 + H_2O$$

Magnesium chloride

Salts like Mg(OH) Cl are called **basic salts**. Such salts can further react with acids to form normal salts.

There are also some other types of salts about which you will learn in higher classes.

Hydrolysis of salts in water

When a salt is dissolved in water, the solution obtained may be neutral, acidic or alkaline which depends upon the nature of the salt taken.

Soluble normal salts derived from strong acids and strong bases give neutral solution. Thus NaCl solution or K_2SO_4 solution is neutral to litmus.

However, the normal salts derived from weak acid and a strong base such as Na_2CO_3 and normal salts derived from strong acid and a weak base such as $AlCl_3$ on hydrolysis give an alkaline (basic) or an acidic solution respectively.

$$Na_2CO_3 + 2H_2O \rightleftharpoons \underset{\text{(strong base)}}{2NaOH} + \underset{\text{(weak acid)}}{H_2CO_3}$$

$$AlCl_3 + 3H_2O \rightleftharpoons \underset{\text{(weak base)}}{Al(OH)_3} + \underset{\text{(strong acid)}}{3HCl}$$

Soluble acid salts usually give solutions acidic to litmus. Thus, NaHSO4 solution turns blue litmus red. But $NaHCO_3$ solution is slightly alkaline due to the formation of strong base NaOH. It turns red litmus blue.

$$NaHCO_3 + H_2O \rightleftharpoons NaOH + H_2CO_3$$

Basic salts are usually insoluble in water

SUMMARY

- Acids are sour to taste. Sour fruits and vegetables contain organic acids.
- Bases are bitter to taste (Do not taste bases directly).
- Acids neutralise bases to form salt and water.
- Indicator litmus is water soluble purple dye extracted from Lichen.
- Indicators change their colour in acidic and alkaline solutions.
- Most metals react with acids to liberate hydrogen gas.
- Hydrogen is lighter than air.
- Metals like Zn, Al and Sn react with sodium hydroxide to liberate hydrogen gas.
- Acids liberate carbon dioxide from metal carbonates.
- Metal oxides dissolve in water to form alkaline solutions.
- CO_2 turns lime water milky.
- Lime water reacts with carbon dioxide to form insoluble calcium carbonate.
- Excess CO_2 dissolves $CaCO_3$ to produce soluble $Ca(HCO_3)_2$.
- Acids neutralise bases to form salt and water.
- All acids have replaceble or ionisable hydrogen.
- Acids have different strengths.
- Organic acids are made of C, H and O.
- Organic acids are found plenty in vegetables & fruits .
- Bases are corrosive.
- In aqueous solutions, acids produce H_3O^+ ions while alkalis produce OH^- ions.

- Dry HCl cannot produce hydrogen ions H^+.
- pH scale is used to measure the strength of acids and alkalis. pH scale ranges from 0 to 14.
- Salts may be normal, acidic and basic salts.
- Normal salts are the products of complete neutralisation of an acid with a base.
- Acid salts can further react with more bases to form normal salts.
- Basic salts can react with more acids to form normal salts.
- Aqueous solutions of the normal salts of strong bases and strong acids are neutral.
- Aqueous solution of a normal salts of a strong bases and a weak acids is basic (alkaline).
- Aqueous solution of a normal salt of weak base and strong acid is acidic.

EXERCISES

1. Write the products obtained when Na_2SO_4 is hydrolysed. i.e., reacts with water ?

2. Write the formula of the acid and the base from which the salt Na_2CO_3 is obtained.

3. pH of samples of HCl and acetic acid are 2 and 3.5 respectively. Which one is a stronger acid?

4. Fe $(OH)_3$ is less soluble in water than Ca $(OH)_2$. Which one is a weaker base ?

5. Device an experimental set up to show that dilute hydrochloric acid is a good conductor of electricity.

★★★★★★

CHAPTER 4

TYPES OF CHEMICAL REACTION

In previous classes, we have learnt about physical and chemical changes. Changes which occur in physical properties such as states of matter, shape, size, colour, texture, electrical and magnetic condition etc, of matter without bringing any change in composition are called physical changes. Evaporation of water, melting of ice, dissolving sugar in water, breaking of wood into smaller pieces, magnetization of iron, glowing of electric bulb are examples of physical changes.

There are also many changes which results in the formation of new substances whose composition and properties are entirely different from those of the original substances. Let us consider the following changes occurring around us.

(i) Burning of paper producing ash, carbon dioxide and water vapour
(ii) Burning of candle wax and fuels like kerosene oil, petrol etc.
(iii) Rusting of iron rod, nail etc producing rust (hydrated iron oxide)
(iv) Souring of milk (formation of curd)
(v) Cooking of food
(vi) Digestion of food in our body
(vii) Taking of antacids when one gets stomach burn/ pain.

In all these changes, the original substances lose their nature and identity and form new substances. For example, when the paper is burnt, the products such as the ash, carbon dioxide, water vapour etc are entirely different from paper. Rust is a different substance from iron. These are examples of **chemical changes**. **Whenever such a chemical change occurs, we say that a chemical reaction has taken place**. In other words, a chemical change takes place as a result of chemical reaction. In a chemical reaction, the substances originally present and taking part in the reaction are called **reactants** and the new substances formed are called **products.** For example, in the burning of paper, the paper (cellulose) and air (oxygen) are the reactants and the ash, carbon dioxide and water vapour are the **products.**

Depending upon the nature of the reaction and the chemical change taking place, we can classify chemical reactions into various types.

4.1. TYPES OF CHEMICAL REACTIONS

We have learnt in class IX, that in a chemical reaction, atoms can neither be created nor destroyed as given by the Law of Conservation of Mass. During a chemical reaction atoms of one element do not change into those of another element. Nor do atoms disappear from the mixture or appear from elsewhere. We have also learnt in the previous chapter that atoms in molecules are held together by a force of attraction called **bond**. **In chemical reactions, the bonds present in the reactants are broken and new bonds are formed to produce new substances or products.** That is, there take place rearrangements or regroupings of atoms in various ways to give products.

4.1.1. Combination reactions

Activity 4.1.

[This activity is better demonstrated by the teacher. Students should not see directly to the white dazzling light produced on burning magnesium ribbon as it is harmful to the eye]

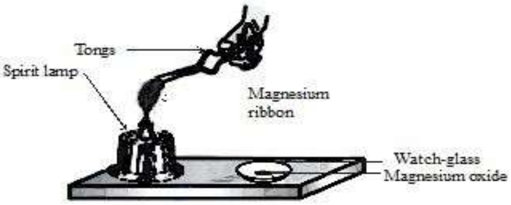


Fig.4.1. Burning of magnesium ribbon to produce magnesium oxide

- Take a strip of magnesium ribbon 2-3 cm long and clean it by rubbing with sand paper.
- Hold it with a pair of tongs in the flame of a spirit lamp or burner. (Do not observe directly the dazzling white light produced).
- Collect the ash formed in a watch glass. (Fig. 4.1.).
- What do you observe ?

It is observed that when heated strongly magnesium burns in air with dazzling white light. The white powder produced can be collected in a china dish. The white powder is magnesium oxide. It is formed due to the reaction between magnesium and oxygen present in the air.

The reaction can be represented by a chemical equation as follows :

$$2\,Mg \quad + \quad O_2 \longrightarrow MgO$$

$$\underbrace{\text{(Magnesium)} \quad \text{(Oxygen)}}_{\text{Reactants}} \qquad\qquad \underbrace{\text{(Magnesium oxide)}}_{\text{Product}}$$

Showing the physical states of the reactants and products, we may write the equation as follows :

$$2\,Mg\,(s) \;+\; O_2\,(g) \longrightarrow 2\,MgO\,(s)$$

The gases, liquid, aqueous and solid states of reactants and products are represented by the notations (g), (l), (aq) and (s) respectively.

A reaction in which a single product is formed from two or more reactants is known as a combination reaction.

The reaction of magnesium and oxygen to form magnesium oxide is an example of combination reaction between two elements.

Let us discuss some more examples of this type.

(i) Carbon or coke burns in air to produce carbon dioxide gas.

$$C(s) \;+\; O_2\,(g) \longrightarrow CO_2\,(g)$$

(ii) Hydrogen burns in oxygen to form water.

$$2\,H_2(g) \;+ O_2(g) \longrightarrow 2\,H_2O\,(l)$$

(iii) Hydrogen combines with chlorine in the presence of light to form hydrogen chloride.

$$H_2(g) \;+ Cl_2(g) \longrightarrow 2\,HCl\,(g)$$

(iv) Iron and sulphur combine on heating to form iron sulphide.

$$Fe\,(s) \;+\; S(s) \longrightarrow FeS(s)$$

(v) Nitrogen combines with hydrogen to form ammonia gas.

$$N_2(g) \;+ 3\,H_2(g) \longrightarrow 2\,NH_3\,(g)$$

Combination reactions also take place between two compounds to form a single product. Let us perform the following activity :

Activity 4.2.

- Take a small lump of calcium oxide (quick lime) in a beaker.
- Add water slowly to the lump of calcium oxide.
- Observe what happens.
- Touch the side of the beaker as shown in Fig 4.2.
- What do you feel ? Hot or cold ?

Fig. 4.2. Formation of slaked lime by the reaction of calcium oxide with water

When water is added to quick lime, they react vigorously to produce calcium hydroxide called slaked lime. A large amount of heat is evolved. The reaction can be represented by the following equation.

$$\underset{\text{quick lime}}{CaO(s) + H_2O\ (l)} \longrightarrow \underset{\text{slaked lime}}{Ca\ (OH)_2\ (aq)}$$

A few more examples are given below.

(i) Ammonia combines with hydrogen chloride to form ammonium chloride.

$$NH_3\ (g)\ + HCl\ (g) \longrightarrow NH_4Cl\ (s)$$

(ii) Calcium oxide reacts with sulphur trioxide to form calcium sulphate

$$CaO\ (s) + SO_3\ (g) \longrightarrow CaSO_4\ (s)$$

(iii) Sodium oxide reacts with water to form sodium hydroxide

$$Na_2O(s) + H_2O(l) \longrightarrow 2\ NaOH\ (aq)$$

Combination reaction also takes place between an element and a compound to form a new compound.

(i) Nitric oxide (a colourless gas) combines with oxygen of the air to form nitrogen dioxide which is a brown gas.

$$\underset{\text{(colourless)}}{2\ NO(g) + O_2(g)} \longrightarrow \underset{\text{(brown)}}{2\ NO_2(g)}$$

(ii) Sulphur dioxide reacts with oxygen (in the presence of catalyst) to form sulphur trioxide.

$$2\ SO_2\ (g) + O_2\ (g) \longrightarrow 2\ SO_3(g)$$

(iii) Carbon monoxide gas combines with oxygen of air to form carbon dioxide gas.

$$2\,CO\,(g) + O_2\,(g) \longrightarrow 2\,CO_2\,(g)$$

4.1.2. Decomposition Reactions

The reactions in which a compound is broken down into two or more simpler substances are known as **decomposition reactions**. Thus, these reactions are just opposite of combination reactions. Let us perform the following activity.

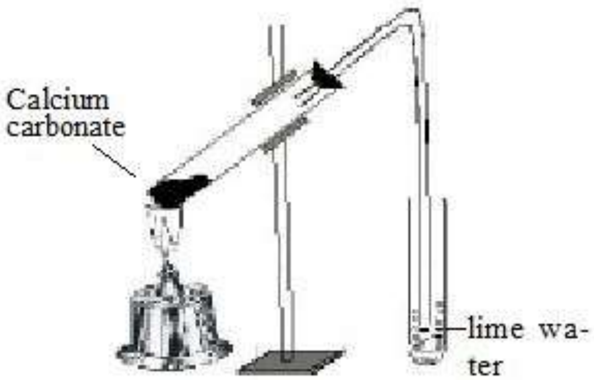


Fig.4.3. Heating of calcium carbonate and giving out carbon dioxide gas which turns lime water milky

Activity 4.3.

- Take about 4 g calcium carbonate in a hard glass test tube fitted with a delivery tube dipped in a test tube containing lime water as shown in Fig. 4.3.
- Heat the tube over the flame of a spirit lamp or burner.
- Observe the change in lime water

It is observed that lime water turns milky. When calcium carbonate is heated, it decomposes to carbon-dioxide gas.

The reaction that take place may be represented as follows.

$$CaCO_3(s) \xrightarrow{heat} CaO(s) + CO_2(g)$$

Carbon dioxide which comes out turns lime water milk

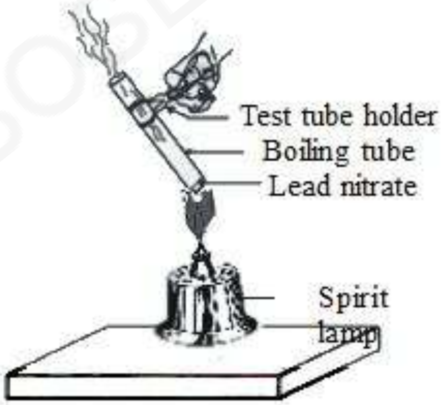


Fig.4.4. Decomposition of lead nitrate on heating

Activity 4.4.

- Take about 2 g of powdered lead nitrate in a test tube.
- Hold it in the test tube holder.
- Heat it over a flame, as shown in Fig. 4.4.
- Observe the brown fumes evolved.
- Note down the changes observed.

When lead nitrate is heated, it breaks down into lead monoxide, nitrogen dioxide and oxygen.

$$2\ \mathrm{Pb(NO_3)_2\,(s)} \xrightarrow{\text{heat}} 2\ \mathrm{PbO(s)} + 4\ \mathrm{NO_2\,(g)} + \mathrm{O_2\,(g)}$$

(brown fumes)

The colourless lead nitrate becomes yellow due to the formation of lead monoxide (PbO). Brown fumes are observed due to the evolution of nitrogen dioxide (NO_2) gas.

Similarly, when we heat green coloured crystals of ferrous sulphate in a dry test tube, we observe that the green colour of ferous sulphate becomes reddish brown due to the formation of iron (III) oxide. Sulphur dioxide and sulphur trioxide gases are also produced.

$$\mathrm{FeSO_4.\,7H_2O(s)} \xrightarrow{\text{heat}} \mathrm{FeSO_4(s)} + \mathrm{7H_2O\ (g)}$$

Ferrous sulphate crystals → Ferrous sulphate (anhydrous)

$$2\ \mathrm{Fe\,SO_4(s)} \xrightarrow{\text{heat}} \mathrm{Fe_2O_3\,(s)} + \mathrm{SO_2\,(g)} + \mathrm{SO_3\,(g)}$$

The reactions given above are examples of decomposition reactions caused by **heat**. They are called **thermal decomposition reactions**.

Decomposition reactions may also take place when electric current is passed through the compound in the molten state or in aqueous solution (decomposition by electricity). In certain cases, light also causes decomposition reactions. They are called *photochemical reactions.*

Decomposition by Electricity

Activity 4.5.

- Take a water voltameter with platinum electrodes.
- Pour sufficient water to the voltameter so that the platinum electrodes are well below the level of water.
- Fill two test tubes with water and close its mouth with your finger or thumb.
- Bring the test tubes with water upside down in the voltameter so that its mouth is under water. Then release your finger.
- Invert the test tubes with water over the two platinum electrodes so that the electrodes remain inside the tubes. (Fig 4.5).

- Now connect the electrodes to a 6 volt battery and see for bubbles of gas at both electrodes. If none appears, add a few mL of dilute sulphuric acid solution.
- Observe the formation of bubbles at both the electrodes. The gas in the bubbles displaces water in the test tubes.
- Compare the volume of the gases collected in both the test tubes.
- Test the gas over the anode with a glowing splinter.
- Try to light the gas collected over the cathode. Always bring out the gas collected over the cathode with mouth of the test tube downwards.
- What happens in each case ?
- Which gas is present in each test tube ?

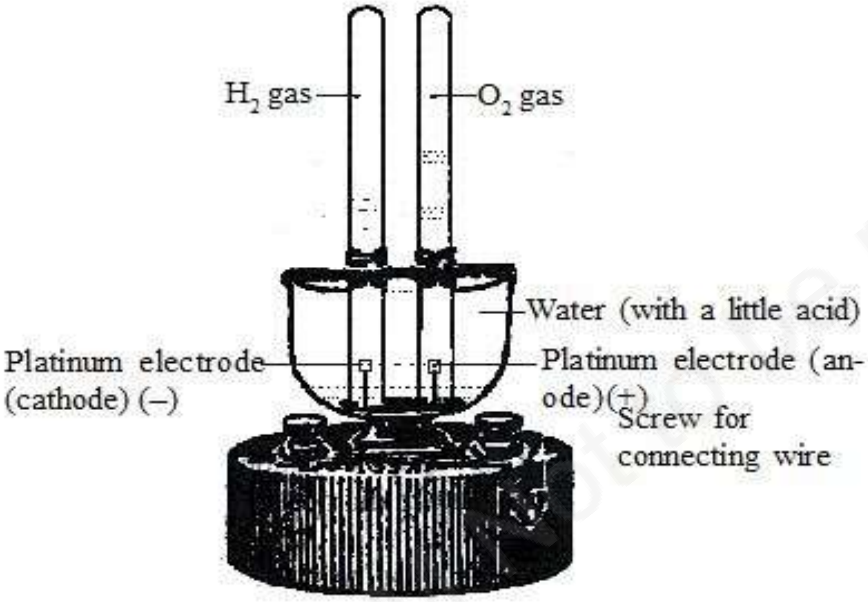


Fig. 4.5. Electrolysis of water

The following results are observed :

(i) Volume of gas collected over the cathode is double that of the gas collected over the anode.

(ii) The gas collected over the cathode burns with a ‘pop’ sound. It is hydrogen gas.

(iii) The glowing splinter introduced into the gas collected over the anode burst into flame. The gas is oxygen.

The above experiment shows that on passing electric current, water decomposes into hydrogen and oxygen according to the reaction :

$$2\,H_2O\,(l) \xrightarrow[\text{current}]{\text{electric}} 2\,H_2(g) + O_2\,(g)$$

Some more examples of decomposition reactions caused by passing electric current are given below.

(i) Electrolytic decomposition of molten sodium chloride

On passing electric current through molten sodium chloride, it decomposes to sodium and chlorine gas.

$$2\,NaCl\,(\text{fused}) \xrightarrow[\text{current}]{\text{electric}} 2\,Na(s) + Cl_2\,(g)$$

The reaction is used to obtain sodium metal on a large scale. Chlorine gas is obtained as a by-product.

(ii) Electrolytic decomposition of molten alumina (Al_2O_3)

When electric current is passed through molten alumina mixed with a little cryolite (it lowers melting point of alumina), alumina decomposes to give aluminium metal and oxygen gas.

$$2\,Al_2O_3\,(\text{fused}) \xrightarrow[\text{current}]{\text{electric}} 4\,Al + 3\,O_2$$

Decomposition by Light (Photochemical decomposition)

There are some reactions which take place in the presence of light. These are called **photochemical reactions**.

Activity 4.6

- Take a pinch of silver chloride or silver bromide on a watch glass.
- Observe its colour.
- Keep the watch glass and its contents in sunlight for some time.
- Observe the change in the colour of the silver chloride in the sunlight. (Fig. 4.6).

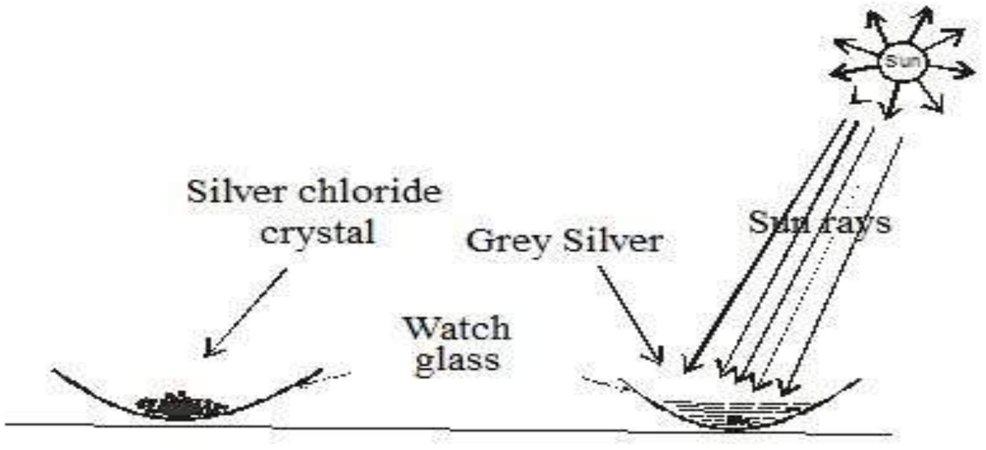


Fig. 4.6. Silver chloride turns grey in sun-light forming silver metal

It is observed that white silver chloride turns grey in sunlight. This is due to the decomposition of silver chloride into silver and chlorine by light.

$$2\ AgCl(s) \xrightarrow{\text{sunlight}} 2\ Ag(s) + Cl_2\ (g)$$

Silver bromide behaves similarly,

$$2\ AgBr \longrightarrow 2\ Ag + Br_2$$

Black and white photography is based upon these reactions. The photographic film is coated with these salts.

Exothermic and Endothermic Reactions

Every chemical reaction involves a change in energy. Either the reaction takes place and energy is liberated to the adjacent surroundings, or the reaction takes place with the absorption of energy i.e. energy is absorbed from the adjacent surroundings. The first type of reaction is said to be **exothermic** (heat comes out), and the second type **endothermic** (heat goes in). All chemical reactions are in one of these two categories.

As the root form **thermic** indicates, the original definition of these two terms was based only on heat energy. Today, however, *exothermic reactions refer to reactions which take place with the evolution of energy in any form —— heat, light, electrical, mechanical, etc. Endothermic reactions refer to reactions which are accompanied by the absorption of any form of energy.*

In activity 4.2, we have observed that a large amount of heat is evolved in the reaction. This makes the reaction mixture warm. Hence, the reaction of calcium oxide with water to form calcium hydroxide or slaked lime is an example of exothermic reaction.

Other examples of exothermic reactions are :

(i) Burning of natural gas (methane) and, other fuels.

$$\underset{\text{methane}}{CH_4\ (g)} + 2\ O_2\ (g) \xrightarrow{\qquad} CO_2(g) + 2\ H_2O(g) + \text{energy}$$

$$\underset{\text{(octane)}}{2\ C_8H_{18}} + 25\ O_2 \xrightarrow{\qquad} 16\ CO_2\ (g) + 18\ H_2O\ (g) + \text{energy}$$

Respiration is an Exothermic Process

We get energy from the food we eat. The food that we eat mainly contains starch (carbohydrate) from rice, potato, bread etc and proteins from meat, egg, milk, pulses etc. During digestion, food materials are broken down into simpler substances. Starch decomposes into simple sugars such as glucose. Similarly, protiens decompose to form simpler molecules called amino acids. The glucose combines with oxygen in the cells of our body and provides energy. The process is called respiration.

The reaction can be represented as follows :

$$\underset{\text{glucose}}{C_6H_{12}O_6\ (aq)} + 6\ O_2\ (g) \xrightarrow{\qquad} 6\ CO_2\ (g) + H_2O\ (g) + \text{energy.}$$

Decomposition Reactions are Endothermic Reactions

It is observed that decomposition reactions require energy in the form of **heat, light** or **electricity** for breaking down the reactants. As already learnt, those reactions in which energy is absorbed are called **endothermic reactions.**

However, there are also certain combination reactions in which heat is absorbed. For example, combination of nitrogen and oxygen to form nitric oxide (NO) takes place with the absorption of heat.

$$N_2(g) + O_2(g) + \text{heat} \xrightarrow{\qquad} 2NO\ (g)$$

A mixture of hydrogen and chlorine when exposed to sunlight forms hydrogen chloride.

$$H_2\,(g) + Cl_2\,(g) \xrightarrow{\text{sunlight}} 2\,HCl\,(g)$$

Photosynthesis is also an endothermic reaction. Chlorophyll present in green plants absorbs light and helps in the conversion of carbon dioxide and water into glucose

$$6\,CO_2(g) + 6\,H_2O\,(l) \longrightarrow \underset{\text{glucose}}{C_6H_{12}O_6} + 6\,O_2$$

Let us answer these.

1. A solution of an oxide 'X' is used for white washing.
   (i) What could be 'X'? Write its formula.
   (ii) Write the reaction of substance 'X' named in (i) above with water.
2. Why is the amount of gas collected in one of the test tubes in activity 4.5 double of the amount collected in the other ? Name this gas.

4.1.3. Displacement Reactions

Activity 4.7.

- Take two iron nails and clean them by rubbing with sand paper.
- Take two clean test tubes and mark A and B.
- Take about 10 mL of copper sulphate solution in each of them.
- Tie a thread on one iron nail and suspend it in test tube A. (Fig 4.7)

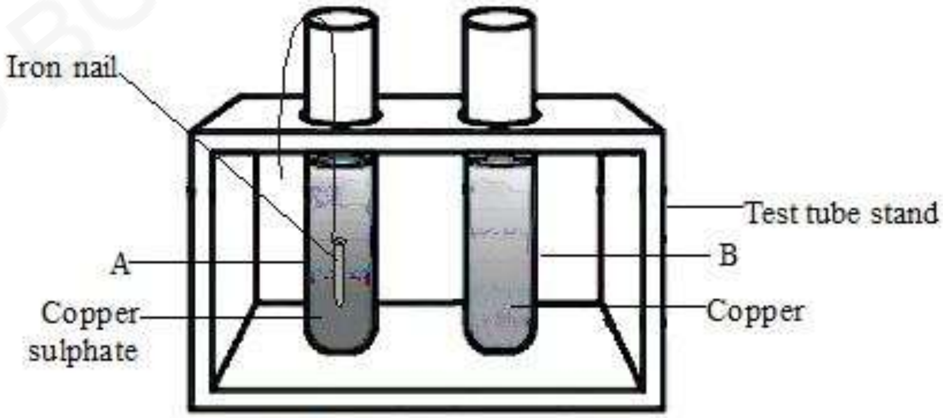


Fig. 4.7. Iron displaces copper from Copper sulphate solution

- Wait for about 30 minutes.
- Take out the nail from the solution.
- Observe the change in the colour of the solution in test tube A and compare it with solution in test tube B.
- Also, compare the colour of the iron nail that was suspended in the test tube with that of other nail.

It is observed that the blue colour of copper sulphate fades and iron nail becomes brownish in colour. These changes show that the following reaction takes place.

$CuSO_4$ (aq)	+	Fe(s) $\longrightarrow$	$FeSO_4$ (aq) +	Cu (s)
copper sulphate iron (blue)		iron	sulphate (light green)	brownish deposit on iron

Some more examples of displacement reactions are :

(i) When a piece of zinc is added to dilute sulphuric acid or dilute hydrochloric acid, hydrogen gas is evolved. The following reactions take place.

$$H_2SO_4 \text{ (dil)} + Zn \text{ (s)} \longrightarrow Zn\,SO_4 \text{ (aq)} + H_2 \text{ (g)}$$

$$2\,HCl \text{ (dil)} + Zn \text{ (s)} \longrightarrow ZnCl_2 \text{ (aq)} + H_2\text{(g)}$$

(ii) Zinc and magnesium displace copper from copper sulphate solution.

$$CuSO_4 \text{ (aq)} + Zn \text{ (s)} \longrightarrow Zn\,SO_4 \text{ (aq)} + Cu \text{ (s)}$$

$$CuSO_4 \text{ (aq)} + Mg \text{ (s)} \longrightarrow MgSO_4 \text{ (aq)} + Cu\text{(s)}$$

(iii) Copper displaces silver from silver nitrate solution

$$2\,Ag\,NO_3\text{(aq)} + Cu \text{ (s)} \longrightarrow Cu\,(NO_3)_2 \text{ (aq)} + 2\,Ag \text{ (s)}$$

(iv) When chlorine water is added to potassium iodide solution, iodine is liberated.

$$2\,KI\text{(aq)} + Cl_2 \text{ (aq)} \longrightarrow 2\,KCl \text{ (aq)} + I_2\text{(aq)}$$

Reactivity Series Of Metals Or Activity Series Of Metals

We have learnt that iron can displace copper from its solution as follows.

$$Fe(s) + Cu\,SO_4 \longrightarrow Cu(s) + FeSO_4(aq)$$

However, it may be noted that copper cannot displace iron from its solution. This means that the following reaction does not take place.

$$Cu\,(s) + FeSO_4(aq) \longrightarrow \text{No reaction}$$

This raises an important question, **"which metal can displace another metal from its salt solution"?**

We know that some metals are highly reactive while some others are less reactive. Based on the study of their reactions, metals can be arranged in order of their reactivities. **The arrangement of metals in the decreasing order of their reactivity is called reactivity series or activity series of metals. (Table 4.1).**

The metal which is higher in the reactivity series is more reactive than metals below it. Hence, in general, a more reactive metal (placed higher in the activity series) can displace the less reactive metal from its solution. Thus, iron displaces copper (less reactive) from its solution as

$$CuSO_4\,(aq) + Fe\,(s) \longrightarrow FeSO_4\,(aq) + Cu\,(s)$$

But copper cannot displace iron from $FeSO_4$ solution.

$$FeSO_4(aq) + Cu\,(s) \longrightarrow \text{No reaction}$$

However, copper can displace silver (less reactive) from its solution.

$$2\,Ag\,NO_3\,(aq) + Cu(s) \longrightarrow 2\,Ag\,(s) + Cu\,(NO_3)_2\,(aq)$$

Table 4.1 Activity series of metals

Element	Symbol	
Potassium	K	(Most reactive)
Barium	Ba	
Calcium	Ca	
Sodium	Na	
Magnesium	Mg	Metals more
Aluminium	Al	reactive than
		hydrogen
Zinc	Zn	
Iron	Fe	
Cobalt	Co	
Nickel	Ni	
Tin	Sn	
Lead	Pb	
Hydrogen	H	
Copper	Cu	
Mercury	Hg	Metals less
Silver	Ag	reactive than
Gold	Au	hydrogen
Platinum	Pt	(Least reactive)

4.1.4. Double Displacement Reactions

Activity 4.8.

- Take about 3 mL of barium chloride solution in a test tube.
- In another test tube, take about 3 mL of dilute sulphuric acid.
- Mix the two solutions.
- Observe what happens.

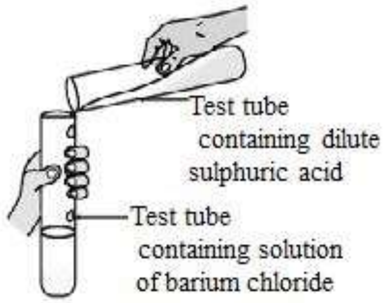


Fig 4.8. Formation of white

We observe that a white substance which is insoluble in water is formed. This insoluble substance is known as a **precipitate.** The white precipitate is barium sulphate which is formed due to the following reaction :

$$BaCl_2(aq) + H_2SO_4\,(aq) \longrightarrow BaSO_4\,(s) + 2\,HCl\,(aq)$$

The reactions which produce precipitates are also called **precipitation reactions.** What causes this ? The white precipitate of $BaSO_4$ is formed by the combination of SO_4^{2-} ion with Ba^{2+} ion. The other product formed is hydrogen chloride which is in solution.

Some more examples of double displacement reactions are given below.

(i) When silver nitrate solution is mixed with sodium chloride solution, a white precipitate of silver chloride is formed.

$$Ag\,NO_3(aq) + NaCl(aq) \longrightarrow AgCl\,(s) + NaNO_3(aq)$$

(ii) When dilute sulphuric acid is added to iron (ferrous) sulphide, a double displacement reaction takes place with the evolution of hydrogen sulphide gas ; ferrous sulphate formed remains in solution.

$$Fe\,S(s) + H_2SO_4\,(aq) \longrightarrow FeSO_4\,(aq) + H_2S\,(g)$$

(iii) When hydrogen sulphide gas is passed through a blue solution of copper sulphate, a black precipitate of copper sulphide is formed. Sulphuric acid formed remains in solution.

$$Cu\,SO_4\,(aq) + H_2S\,(g) \longrightarrow CuS\,(s) + H_2\,SO_4\,(aq)$$

Neutralisation Reactions

A double displacement reaction also takes place when an acid reacts with a base to form salt and water

$$\text{Acid} + \text{Base} \longrightarrow \text{Salt} + \text{Water}$$

These reactions are called neutralisation reactions.

A few examples are given below:

(i) $NaOH + HCl \longrightarrow NaCl + H_2O$

(ii) $2\,NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + 2\,H_2O$

(iii) $KOH + HCl \longrightarrow KCl + H_2O$

Oxidation and Reduction Reactions (Redox reactions)

Activity 4.9.

- Heat about 1g of copper powder in a China dish (Fig 4.9)
- Observe what happens.

Fig 4.9. Oxidation of copper to copper oxide

It is observed that reddish brown copper powder becomes coated with a black substance. How is this black substance formed ?

This is because when copper is heated, it combines with oxygen and black copper oxide is formed.

$$2Cu + O_2 \xrightarrow{heat} 2CuO$$

We say that copper has been oxidized to copper (II) oxide.

If hydrogen gas is passed over this heated material of copper (II) oxide, the black coating on the surface turns brown. The reverse reaction takes place and copper is obtained.

$$CuO + H_2 \xrightarrow{heat} Cu + H_2O$$

We say that copper oxide has been reduced to copper.

If a substance gains oxygen during a reaction, it is said to be oxidised. If a substance loses oxygen during a reaction, it is said to be reduced.

During the above reaction, the copper (II) oxide is losing oxygen and is being reduced. The hydrogen is gaining oxygen and is being oxidised. In other words, one reactant gets oxidised while the other gets reduced during the reaction. Thus oxidation and reduction occur simultaneously. Such reactions in which oxidation and reduction occur simultaneously are called **oxidation - reduction reactions** or **redox** reactions.

Let us consider the following reaction.

$$MnO_2 + 4\ HCl \longrightarrow Mn\ Cl_2 + 2\ H_2O + Cl_2$$

In this reaction, HCl is oxidised to Cl_2 whereas MnO_2 is reduced to

$MnCl_2$. From this example, we can say that **if a substance gains oxygen or loses hydrogen during a reaction, it is said to be oxidised. If a substance loses oxygen or gains hydrogen during a reaction, it is said to be reduced.** The substance which oxidises other substances but itself undergoes reduction is called **oxidising agent.** The substance which reduces other substances but itself undergoes oxidation is called the **reducing agent**. In the above example, MnO_2 undergoes reduction and oxidises HCl. Hence MnO_2 is the oxidising agent, while HCl is the reducing agent.

Some more examples of redox reactions are given below.

(i) $ZnO + C \longrightarrow Zn\,(s) + CO$

(ii) $H_2S + Cl_2 \longrightarrow 2\,HCl + S$

According to electronic concept, **oxidation is a process which involves loss of one or more electrons. Reduction is a process which involve gain of electrons.** You will learn about these in higher classes.

Oxidation-Reduction Reactions in Everyday Life

There are a number of redox reactions taking place around us which affect our everyday life. We have observed that iron articles are bright and shining when new. But they get coated with a brown flaky substance when left exposed to moist air for a long time. This process is commonly known as rusting of iron. Rust is mainly hydrated ferric oxide ($Fe_2O_3. xH_2O$). It is formed due to the action of oxygen, water vapour or moisture and other gases present in the air on the surface of iron. The reaction may be represented by the following equation :

$$\underset{\text{iron}}{4\,Fe(s)} + \underset{\text{(from air)}}{3\,O_2\,(g)} + \underset{\text{moisture}}{2x\,H_2O\,(l)} \longrightarrow \underset{\text{hydrated ferric oxide (rust)}}{2Fe_2O_3 \cdot xH_2O\,(s)}$$

Rusting is an oxidation process. It causes damage to car bodies, bridges, iron railings, automobile parts and many other articles made of iron. The reason is that, the reddish brown crust of rust does not stick to the surface. It peels off exposing fresh surface for rusting. Thus, corrosion of iron is a continuous process which ultimately eats up the whole iron object.

Some other metals also lose their luster and shine after some time. The surface of articles made of copper and brass acquires a green coating of basic carbonate, $CuCO_3$, $Cu(OH)_2$ when exposed to moist air. This is due to the action of O_2, CO_2 and water vapour present in the air on the surface of copper.

$$\underset{\text{copper}}{2\,Cu\,(s)} + \underbrace{CO_2(g) + O_2(g) + H_2O\,(l)}_{\text{from moist air}} \longrightarrow \underset{\substack{\text{basic copper carbonate}\\ \text{(green)}}}{CuCO_3.\ Cu(OH)_2}$$

Likewise, the surface of silver metal gets tarnished on exposure to air. This is due to the formation of coating of black silver sulphide (Ag_2S) on the surface by the action of hydrogen sulphide gas present in the air.

$$4\,Ag(s) + 2H_2S(g) + O_2\,(air) \longrightarrow \underset{\substack{\text{silver sulphide}\\ \text{(black)}}}{2Ag_2S\,(s)} + 2H_2O$$

When a metal is reacted with substances around it such as oxygen, moisture, acids etc, it is said to corrode and the process is called corrosion. Rusting of iron, the formation of black coating on silver and the green coating on copper are examples of corrosion.

We have seen that iron objects such as window grills, iron gates, steel furnitures, bodies of cars, buses etc are painted with colours and the painted iron does not rust. Why is it so ?

Corrosion of iron (rusting) can be prevented if iron objects are not allowed to come in contact with damp air. Painting not only makes the iron objects beautiful but also prevents from corrosion by forming a barrier between the iron surface and the damp air.

Iron articles such as mechanical tools, machine parts etc can be prevented from corrosion by **greasing** and **oiling**.

Rusting of iron can be effectively prevented by coating iron with zinc. The process is called **galvanization**

Rancidity

Have you ever tasted or smelt the food materials containing fat or oil left for a longer time ?

When fats and oils in foods are oxidised, they become rancid and their smell and taste change. **The oxidation of fats or oils in foods resulting into bad smell and taste is called rancidity.** A common example is the rancidity of butter on keeping for a longer time.

To prevent rancidity, usually substances which prevent oxidation are added to foods containing fats and oils. These substances are called **anti-oxidants**. Keeping food in air tight containers also helps to slow down oxidation. We have noticed that bags of chips are flushed with gas such as nitrogen to prevent them from oxidation. In a number of cases, while packing the foods, the container is evacuated before sealing so that no air or oxygen is available inside for the oxidation of food. This is called **vacuum packing**.

4.3. SLOW AND FAST REACTIONS

We have observed that a piece of iron rusts in the course of time. We have also observed that the colour of many clothes fades slowly even if they are not washed. If we thrust our hand into a heap of rotting grass or straw, we find it very hot inside. These are all due to chemical reactions and are evidently very slow.

Food is digested in the stomach and intestines very slowly and takes hours for the digestion to complete. Setting of cement and fermentation of sugar to form alcohol are also examples of slow reactions.

When ethyl alcohol and acetic acid is heated in the presence of a little concentrated sulphuric acid, the reaction takes place after a few minutes and an ester called ethyl acetate is formed.

$$\underset{\text{Acetic acid}}{CH_3COOH\,(aq)} + \underset{\text{Ethyl alcohol}}{C_2H_5OH\,(l)} \longrightarrow \underset{\substack{\text{Ethyl acetate}\\ \text{(sweet smelling ester)}}}{CH_3COOC_2H_5\,(aq)} + H_2O\,(l)$$

It is a slow reaction. Generally reactions between covalent compounds are slow.

Fast reactions

Recall activity 4.6 where a white substance, barium sulphate is formed when a solution of barium chloride is mixed with dilute sulphuric acid solution. The precipitation is formed at once. So the reaction is very fast or instantaneous.

The reaction between silver nitrate solution and sodium chloride solution is also very fast. The white precipitate of silver chloride is formed at once. **Generally the reactions between ionic compounds in solution are fast.**

Burning of magnesium in air (oxygen) to form magnesium oxide and combination of nitric oxide (NO) with oxygen (O_2) to form nitrogen dioxide (NO_2) are also fast reactions.

Moderately Fast Reactions

In between very fast and very slow reactions, there are certain reactions which occur at moderate speed. These reactions may be called **moderately fast reactions**. The examples of such reactions are :

(i) Decomposition of hydrogen peroxide into water and oxygen

$$2\,H_2O_2\,(l) \longrightarrow 2\,H_2O\,(l) + O_2\,(g)$$

Hydrogen peroxide

(ii) Hydrolysis of sugar into glucose and fructose in presence of a little acid

$$C_{12}H_{22}O_{11} + H_2O \xrightarrow{\text{acid}} C_6H_{12}O_6(aq) + C_6H_{12}O_6(aq)$$

Sugar Glucose Fructose

Catalysts

The progress of a reaction can be altered (increases or decreases) by the use of specific substances called catalysts. *A catalyst is a substance which increases the rate of a chemical reaction without itself being consumed and the phenomenon is called catalysis. On the other hand, if a catalyst decreases (retards) the rate of a reaction, it is called a negative catalyst and the phenomenon is called negative catalysis.*

Activity 4.7.

- Take two beakers labelled A and B. In each of them take 10 mL of oxalic acid solution, 5 mL of dilute sulphuric acid solution and 10 mL of dilute potasium permanganate solution.
- Observe the colour of the solution in the two beakers.
- Now, in beaker A add one or two crystals of manganese (II) sulphate ($MnSO_4$) and observe carefully.
- Why is the colour of the solution in beaker A fades up faster and finally disappeared?
- Why is there slow change in the colour of the solution in the beaker B ?

What is the role of $MnSO_4$ in the above reaction ? It has speeded up the reaction between oxalic acid and potassium permanganate.

Catalysts are used in the preparation of a large number of industrially important compounds. For example, in the manufacture of ammonia from nitrogen and hydrogen, finely divided iron is used as a catalyst.

$$N_2\,(g) + 3\,H_2\,(g) \xrightarrow{\text{Fe (catalyst)}} 2\,NH_3\,(g)$$

Similarly, vanadium pentoxide (V_2O_5) is used a as a catalyst in the manufacture of sulphur trioxide.

$$2\,SO_2\,(g) + O_2(g) \xrightarrow[\text{catalyst}]{V_2O_5} 2\,SO_3\,(g)$$

A large number of reactions of industrial and biological importance takes place in the presence of catalyst, which would otherwise never occur in their absence. Most of the chemical reactions taking place in plants, animals and micro-organisms require their own specific catalyst called enzymes. Thus, enzymes are biocatalysts present in the living cells which catalyse reactions in living plants and animals. Enzymes are highly specific in their action. Each enzyme catalyses a particular reaction. For example, an enzyme called **invertase** can convert sucrose into glucose and fructose, but fails to in the case of very similar sugar maltose

$$\underset{\substack{\text{sucrose}\\\text{(cane sugar)}}}{C_{12}H_{22}O_{11}}+H_2O \xrightarrow{\text{invertase (catalyst)}} \underset{\text{glucose}}{C_6H_{12}O_6} + \underset{\text{fructose}}{C_6H_{12}O_6}$$

Maltose requires another enzyme called **maltase** for this purpose.

$$\underset{\text{maltose}}{C_{12}H_{22}O_{11}} + H_2O \xrightarrow{\text{maltase}} \underset{\text{glucose}}{2\,C_6H_{12}O_6}$$

In most of the cases, catalysts use to accelerate the reactions. In a few cases, however, catalysts are used to slow down the rate of reactions. Such catalysts are known as **negative catalysts.**

For example, glycerine is sometimes added to hydrogen peroxide as negative catalyst to slow down its rate of decomposition into hydrogen and oxygen.

Let us answer these.

1. Why does the colour of copper sulphate solution change when an iron nail is dipped in it ?
2. Give an example of a double displacement reaction other than the one given in activity 4.8.
3. Identify the substances that are oxidised and the substaces that are reduced in the following reaction.

   (i) $4\,Na(s) + O_2(g) \longrightarrow 2\,Na_2O(s)$

   (ii) $CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(l)$

SUMMARY

- A chemical reaction is a process in which one or more substances called reactants are changed to one or more new substances called products.
- Combination reactions are those reactions in which two or more substances combine to form a new single substance (compound).
- Decomposition reactions are those reactions in which a single substance splits into two or more substances.
- Reactions in which energy is given out along with the products are called exothermic reactions.
- Reactions in which energy is absorbed are called endothermic reactions
- A displacement reaction occurs when an element displaces another element from its compound.
- Two different atoms or group of atoms or ions are mutually exchanged in double displacement reactions.
- In a double displacement reaction, if one of the products is a precipitate, it is called precipitation reaction.
- When an acid reacts with a base to form salt and water, it is called neutralization reaction.
- Reactions may also involve the gain or loss of oxygen or hydrogen by substances. Oxidation is the process which involves gain of oxygen or loss of hydrogen. Reduction is loss of oxygen or gain of hydrogen.
- Depending upon the rate at which a chemical reaction proceeds, all chemical reactions may be classified into slow reactions, fast reactions and moderately fast reactions.
- The process of slowly eating away of the metal due to the action of the atmospheric gases on the surface of the metals resulting in the formation of compounds is called corrosion. Corrosion of iron is called rusting.
- Corrosion of metals can be prevented by providing protective layers with substances like grease, paint or metal coatings so that the metals do not come in contact with damp air.

- The oxidation of fatty foods forming products which have bad taste and bad smell is called rancidity.

EXERCISES

1. **Which of the statements about the reaction given below are NOT correct ?**

   $2\ PbO\ (s) + C(s) \longrightarrow 2\ Pb(s) + CO_2\ (g)$

   (a) Lead is reduced.
   (b) Carbon dioxide is oxidised.
   (c) Carbon is oxidised
   (d) Lead oxide is reduced.
      (i) (a) and (b)
      (ii) (a) and (c)
      (iii) (a), (b) and (c)
      (iv) all

2. $Fe_2O_3 + 2\ Al \longrightarrow Al_2\ O_3 + 2\ Fe$

   The above reaction is an example of a

   (a) combination reaction
   (b) double displacement reaction.
   (c) decomposition reaction.
   (d) displacement reaction.

3. What are combination reactions ? Give one example.

4. What are decomposition reactions ? Give one experiment to demonstrate a decomposition reaction.

5. Why are decomposition reactions called the opposites of combination reactions ?

6. Write one equation each for decomposition reactions where energy is supplied in the form of heat, light and electricity

7. What is the difference between displacement and double displacement reactions. Write equations for these reactions.

8. What is a displacement reaction ? Give one example.

9. In the refining of silver, silver can be recovered from silver nitrate solution by adding zinc metal. Write down the reaction invlolved.

10. What is a precipitation reaction ? Explain it by giving one example.

11. What are exothermic and endothermic reactions ? Give two examples each.

12. Why is respiration considered an exothermic reaction ?

13. Explain the terms **(a)** oxidation and **(b)** reduction with two examples each.

14. Why does painted iron not rust ?

15. Fat and oil containing food items flushed with nitrogen, why ?

★★★★★★

CHAPTER

5

METALS AND NON-METALS

In unit 1, we have learnt that the earliest classification of elements on the basis of their common features was the grouping of the elements into two broad categories– **metals** and **non-metals**. In the periodic table, the metals are placed on the left hand side and in the centre of the periodic table. On the other hand, the non-metals are placed on the right hand side of the table (Table 5.1)

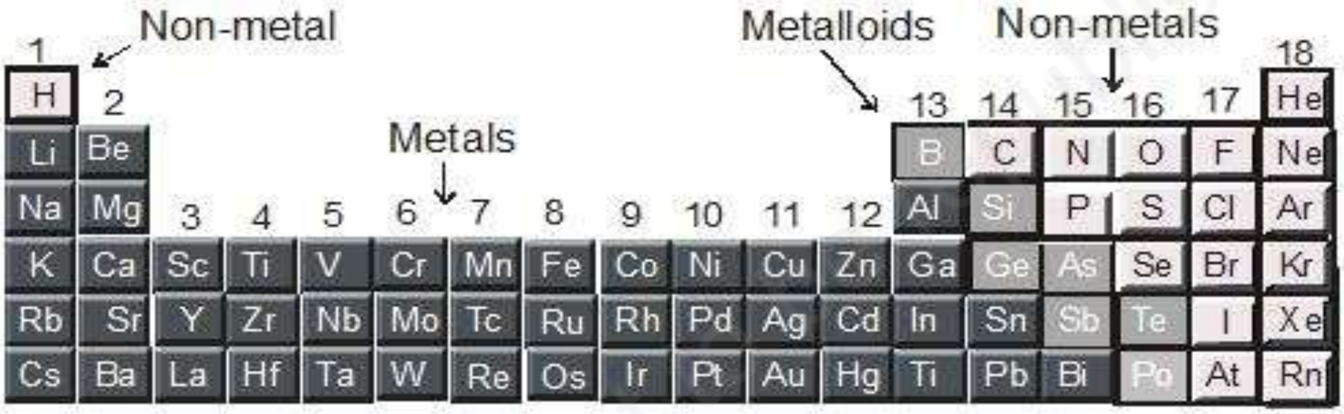


Table 5.1. Position of metals and non-metals in the Periodic Table

The metals and nonmetals are separated from each other in the periodic table by a zig-zag line. The elements close to zig-zag line show some properties of metals and some properties of non-metals. They are **metalloids**. The common examples of metalloids are germanium (Ge), arsenic (As), antimony (Sb), tellurium (Te) and polonium (Po). Although metals show a great diversity of properties, there are a few unifying characteristics, which set them apart from the non-metallic elements.

5.1. PHYSICAL PROPERTIES OF METALS AND NON-METALS

Some of the examples of metals are Iron, Aluminium, Copper, Silver, Gold, Tin, Lead, Mercury, Zinc, Sodium, Potassium, Calcium and Magnesium. All the metals are solids, except Mercury which is a liquid at room temperature.

Some of the examples of non-metals are Carbon, Sulphur, Phosphorus, Silicon, Hydrogen, Oxygen, Nitrogen, Chlorine, Bromine, Iodine, Helium, Neon and Argon.

Carbon occurs as diamond and graphite. Out of the non-metals, five are solids, one (bromine) is a liquid whereas remainings are gases.

Activity 5.1.

- Take samples of metals iron, copper, lead, silver, zinc and aluminium.
- Take samples of non-metals sulphur and graphite (carbon).
- Clean the surface of each sample of the metals by rubbing them with sand paper.
- Note their appearences.
- Carry out the same activities with the samples of non-metals.

Metals, in their pure state, have a shining surface. This propery is called **metallic lustre**. **Non-metals, in general, do not have such property.**

Activity 5.2.

- Take pieces of iron, aluminium, copper and zinc.
- Try to cut these metals with a sharp knife.
- Note your observations.
- Hold a piece of sodium metal with a pair of tongs. Dry it by pressing between the folds of a filter paper.
- Put it on a watch-glass or filter paper and try to cut it with a knife.
- Note your observations.
- Repeat the activity with non-metals, sulphur and graphite.
- Note the differences.

We find that **metals are generally hard.** Only a few metals like sodium are soft i.e.the hardness varies from metal to metal. **Non-metals are generally soft.**

Activity 5.3.

- Take pieces of iron, zinc, lead, aluminium and copper.
- Place any one metal on a block of iron and strike it four or five times with a hammer. What do you observe?
- Repeat with other metals.
- Record the change in the shape of these metals.
- Repeat the activities with the samples of non-metals-sulphur and graphite
- Note the differences.

We find that **metals can be beaten into thin sheets**. This property is called **malleability**. Gold and silver are the most highly malleable metals.

Non-metals are not malleable.

Activity 5.4.

- Take wires of copper, iron and aluminium. How can these wires be formed?
- Can we draw out wires of sulphur and graphite ?

Metals can be drawn into thin wires. This property is called **ductility**. Gold is the most ductile metal. One gram of gold can be drawn into very thin wires.

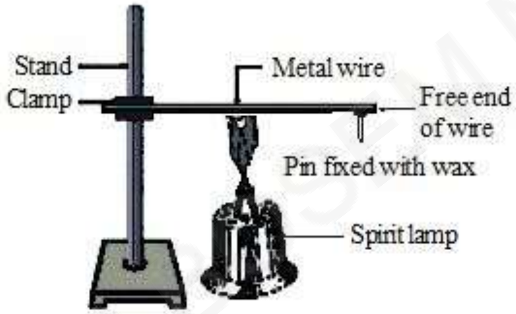


Fig. 5.1. Metals are good conductors of heat

Activity 5.5.

- Take a thick copper or aluminium wire. Clamp it horizontally as shown in Fig. 5.1
- Fix a pin with the help of wax (this can be done by keeping the pin over the wire and putting molten wax over it and allowing to cool).
- Heat the wire with a spirit lamp, candle or a burner near the place where it is clamped.
- What do you observe after some time?
- Does the metal wire melt ?

It will be observed that the wax melts and the pin falls. This shows that heat flows through the metal wire. Further, the wire does not melt even after long heating. The above activity shows that **metals are good conductors of heat and have high melting points**. Silver and copper are very good conductors of heat while lead and mercury are comparatively poor conductors. Non -metals do not have this property. *They are poor conductors of heat.*

You must have observed that if we connect a bulb to a battery through a metal wire, the bulb glows. We use copper or aluminium wires to carry electric current in our homes. This shows that **metals allow electric current to pass through them. Thus, metals are good conductors of electricity.** We have seen that wires which carry current in our homes have a Plastic coating of polyvinylchloride (PVC). This coating protects us from the current when we use the metal wires.

Most of the metals produce sound when they strike a hard surface or hard object. The **metals that produce a sound on striking a hard surface is said to be sonorous**. Strike a hard surface with a graphite rod. Can you get a ringing sound ?

Can you say why school bells are made up of metals ?

Non-metals are not sonorous.

Some exceptions to general trends

A comparison of the general physical properties of metals and non-metals shows that there are many exceptions in the general properties. For example –

(i) **Physical state**: We have studied that metals are solids. But mercury is an exception as it is a liquid at room temperature.

(ii) **Melting and Boiling Points**: In activity 5.5, we have observed that metals have high melting points. But sodium, potasium, gallium and cesium have low melting points. Gallium has such a low melting point (30·15^{o}C) that it melts even when kept on our palm. We have also studied that non-metals have low melting and boiling points. But diamond which is a non-metal carbon has very high melting point.[Carbon, a non-metal, can exist in different forms. Each form is called an allotrope. Diamond and graphite are allotropes of carbon].

(iii) **Lustre**: We have learnt that non-metals do not have a shining appearance or lustre. However, iodine is a non-metal but it is lustrous.

(iv) **Hardness**: In activity 5.2, we have learnt that metals are generally hard, and hardness varies from metal to metal. We have found that sodium is a soft metal which can be easily cut with a knife. Similarly, lithium and potassium are soft. We have also studied that solid non-metals are soft. But carbon in the form of diamond is extremely hard.

(v) **Electrical conductivity :** We have studied that non-metals are poor conductors of electricity. But graphite is a conductor of electricity .

From the above discussions, it is obvious that we cannot classify the elements as metals or non-metals on the basis of their physical properties alone because there are many exceptions. We have to see if elements can be more clearly classified as metals and non-metals on the basis of chemical properties.

5.2. CHEMICAL PROPERTIES OF METALS & NON-METALS

5.2.1. Formation of Oxides

Activity 5.6.

- Take a magnesium ribbon and some sulphur powder.
- Recall activity 4.1. Burn the magnesium ribbon. Collect the ashes formed and dissolve them in water.
- Test the resultant solution with both blue and red litmus paper.
- Is the product formed on burning magnesium acidic or basic ?
- Take a little sulphur in a deflagrating spoon. Hold it over the flame of a spirit lamp till the sulphur melts and start burning.
- Now introduce the spoon into a jar containing some water at its bottom (Fig. 5.2)

Fig. 5.2. Burning of sulphur powder

- After the sulphur has burnt, take the deflagrating spoon out of the jar.
- Test the solution in the jar by dipping into it blue and red litmus paper.
- Is the product formed on burning sulphur acidic or basic ?
- Can we write the equations for these reactions ?

We can see that magnesium burns brilliantly with a dazzling white light. Magnesium combines with oxygen forming magnesium oxide.

$$Mg + O_2 \longrightarrow 2\ MgO$$

Magnesium oxide reacts with water to form an alkali called magnesium hydroxide. It turns red litmus blue.

$$MgO + H_2O \longrightarrow Mg\ (OH)_2$$

Almost all metals combine with oxygen to form metal oxides.

$$\text{Metal} + \text{Oxygen} \longrightarrow \text{Metal Oxide.}$$

For example,

$$2\ Ca + O_2 \longrightarrow 2\ CaO$$

$$4\ Al + 3O_2 \longrightarrow 2\ Al_2O_3$$

$$4\ K + O_2 \longrightarrow 2\ K_2O$$

Many metal oxides are insoluble in water but some of these dissolve in water to form alkalis. Sodium oxide and potassium oxide dissolve in water to produce alkalis as follows :

$$Na_2O(s) + H_2O\ (l) \longrightarrow 2\ NaOH\ (aq)$$

$$K_2O(s) + H_2O\ (l) \longrightarrow 2\ KOH\ (aq)$$

Metal oxides react with acids to form salt and water.Thus metal oxides are basic in nature . For example,

$$CuO + 2\ HCl \longrightarrow Cu\ Cl_2 + H_2O$$

$$MgO + H_2SO_4 \longrightarrow MgSO_4 + H_2O$$

But some metal oxides, such as aluminium oxide, zinc oxide, etc, show both acidic as well as basic behaviour. Such **metal oxides which react with both acids as well as bases to produce salts and water are known as amphoteric oxides**. Aluminium oxide reacts with acids and bases in the following manner.

$$Al_2O_3 + 6\,HCl \longrightarrow 2\,AlCl_3 + 3\,H_2O$$

$$Al_2O_3 + 2\,NaOH \longrightarrow \underset{\text{(Sodium meta aluminate)}}{2NaAlO_2} + H_2O$$

Non-metals react with oxygen to form acidic or neutral oxides:

When sulphur burns in air, it combines with the oxygen of air to form an acidic oxide called sulphur dioxide

$$S(s) + O_2(g) \longrightarrow SO_2(g)$$

Sulphur dioxide dissolves in water to form an acid called sulphurous acid.

$$SO_2(g) + H_2O(l) \longrightarrow \underset{\text{sulphurous acid}}{H_2SO_3(aq)}$$

Similar behaviour is shown by carbon.

$$C(s) + O_2(g) \longrightarrow CO_2(g)$$

$$CO_2(g) + H_2O\,(l) \longrightarrow \underset{\text{(carbonic acid)}}{H_2CO_3(aq)}$$

However, some oxides of non-metals are **neutral**. For example, carbon monoxide, nitric oxide (NO) and water are neutral oxides.

5.2.2 . Reaction with Water

Activity 5.7.

- Dry a piece of sodium metal by using a filter paper and cut a small piece of it.
- Place the piece of sodium metal in water taken in a trough.
- Observe what happens.
- Repeat the test with other samples of metals such as copper, iron, zinc etc, and also with non-metals, sulphur and graphite (carbon).
- Do they react with water in the same way as sodium ?

When sodium is placed in water, we find that the piece of sodium floats on water, starts moving on water making a hissing sound and then sodium catches fire with a little explosion. Sodium metal reacts vigorously with cold water forming sodium hydroxide and hydrogen gas. A lot of heat is also produced in the reaction. The heat burns hydrogen gas as well as the sodium metal. The reaction is highly exothermic.

$$2\,\mathrm{Na} + 2\,\mathrm{H_2O} \longrightarrow 2\,\mathrm{NaOH} + \mathrm{H_2} + \mathrm{Heat}$$

Potassium also reacts with cold water more vigorously.

$$2\,\mathrm{K} + 2\,\mathrm{H_2O} \longrightarrow 2\,\mathrm{KOH} + \mathrm{H_2} + \mathrm{Heat}$$

Calcium reacts with cold water to form calcium hydroxide and hydrogen gas. The reaction is less violent.

$$\mathrm{Ca} + 2\,\mathrm{H_2O} \longrightarrow \mathrm{Ca\,(OH)_2} + \mathrm{H_2} + \mathrm{Heat}$$

Magnesium does not react with cold water. It reacts slowly with hot water to form magnesium hydroxide and hydrogen but reacts very rapidly with steam to form magnesium oxide and hydrogen gas.

$$\mathrm{Mg(s)} + \underset{\text{(hot)}}{2\,\mathrm{H_2O\,(l)}} \longrightarrow \mathrm{Mg\,(OH)_2\,(aq)} + \mathrm{H_2}$$

$$\mathrm{Mg(s)} + \underset{\text{steam}}{\mathrm{H_2O\,(g)}} \longrightarrow \mathrm{MgO(s)} + \mathrm{H_2\,(g)}$$

Aluminium, zinc and red hot iron also react with steam to form their respective oxides and hydrogen.

$$\mathrm{Zn(s)} + \underset{\text{(steam)}}{\mathrm{H_2O\,(g)}} \longrightarrow \mathrm{Zn\,O(s)} + \mathrm{H_2\,(g)}$$

$$2\,\mathrm{Al} + \underset{\text{(steam)}}{3\,\mathrm{H_2O}} \longrightarrow \mathrm{Al_2\,O_3} + 3\,\mathrm{H_2}$$

$$\mathrm{Fe} + \underset{\text{(steam)}}{4\mathrm{H_2O}} \longrightarrow \underset{\text{Iron (II, III) oxide}}{\mathrm{FeO.\,Fe_2O_3}} + 4\,\mathrm{H_2}$$

Non- metals do not react with water or steam to evolve hydrogen gas.

5.2.3. Reaction with Dilute Acids

Activity 5.8.

- Take small pieces of metals such as magnesium, zinc, iron, aluminium, copper etc except sodium and potasium. [sodium and potasium need not be taken as they react vigorously even with cold water]
- Put the samples separately in test tubes containing dilute hydrochloric acid.
- Observe the rate of formation of hydrogen gas bubbles carefully.
- Which metals react vigorously with dilute hydrochloric acid ?

Metals react with dilute acids to give salt and hydrogen gas

$$\text{Metal} + \text{Dilute acid} \longrightarrow \text{Salt} + \text{Hydrogen}$$

$$Mg\,(s) + 2\,HCl\,(aq) \longrightarrow MgCl_2\,(aq) + H_2\,(g)$$

$$Zn + 2\,HCl \longrightarrow ZnCl_2 + H_2$$

$$Fe + 2\,HCl \longrightarrow Fe\,Cl_2 + H_2$$

Similarly, $Mg(s) + H_2SO_4(aq) \longrightarrow MgSO_4\,(aq) + H_2(g)$

The rate of formation of bubbles was very fast in the case of magnesium. The reactivity decreases in the order $Mg > Al > Zn > Fe$.

In the case of copper, no bubble is seen. This shows that copper does not react with dilute HCl. Silver and gold metals also do not react with dilute acids.

We also find that non-metals do not react with dilute acids.

Hydrogen gas is not evolved when a metal reacts with nitric acid. It is because HNO_3 is a strong oxidising agent. It oxidises the H_2 produced to water and itself gets reduced to one of the nitrogen oxides (nitrous oxide N_2O, nitric oxide, NO, nitrogen dioxide NO_2). But magnesium reacts with very dilute HNO_3 to evolve H_2 gas.

Let us answer these.

1. Name five metals and five non-metals

2. Name one metal and one non-metal which exist in liquid state at room temperature.

3. Give an example of a metal which

- **(i)** can be easily cut with a knife,
- **(ii)** is the best conductor of heat,
- **(iii)** is a poor conductor of heat.

4. What is meant by saying that metals are malleable and ductile ?

5. Which property of copper and aluminium make them suitable for making electric wires.

6. Name a non-metal which conducts electric current.

7. Write equations for the reactions of
   - **(a)** potasium with water
   - **(b)** red-hot iron with steam
   - **(c)** zinc with dilute sulphuric acid.
8. How do metals differ from non-metals in their reaction with oxygen and the product is dissolved in water ?

5.3. OCCURRENCE OF METALS

The major source of metals is the earth's crust. Sea water also contains some soluble salts of metals like sodium chloride, magnesium chloride, etc. The most abundant metal in the earth's crust is aluminium. The second most abundant metal is iron and third one is calcium.

Metals occur in nature in free (native state) or in combined state, i.e. in the form of their compounds. We have already studied the reactivity series of metals in Unit 4. The metals at the bottom of the activity series are the least reactive. They are often found in a free state. For example, gold, silver, platinum and copper are found in the free state. Copper and silver are also found in nature as their sulphides or oxides. The metals at the top of the activity series (K, Na, Ca, Mg and Al) are so reactive that they are never found in nature as free elements. They are found in the combined state as oxides, suphides, halides or carbonates. The metals in the middle of the activity series i.e., Zn, Fe, Pb etc are moderately reactive. They are also found in the earth's crust mainly as oxides, or carbonates. It is found that many metals occur as oxides. This is because oxygen is a very reactive element and is very abundant on the earth. Thus we find that the occurrence of metals are closely related to their reactivity (Fig. 5.3).

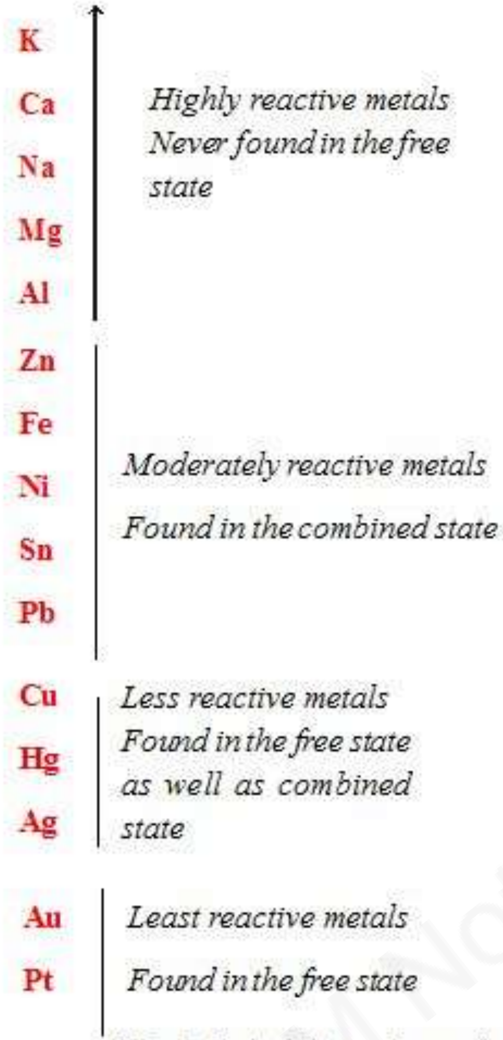


Fig. 5.3. Activity series and occurences of metals.

The natural materials in which the metals or their compounds are found in earth are called minerals. Some minerals may contain a large percentage of metal whereas others may contain a small percentage of the metal. Thus, all the minerals cannot be used to extract metals. Those minerals from which metals can be extracted profitably are called **ores**. For example, aluminium occurs in the earth's crust in the form of two minerals, bauxite ($Al_2O_3 . 2H_2O$) and clay ($Al_2O_3 . 2SiO_2 . 2H_2O$). Out of these two, aluminium can be profitably extracted from bauxite but not from clay. Therefore, the ore of aluminium is bauxite. Thus, all ores are minerals but all minerals are not ores.

The ores of some common metals are given below :

Metal	Name of the ore	Name of compound present in the ore	Formula of the ore
1. Sodium (Na)	Rock salt	Sodium chloride	Na Cl
2. Aluminium (Al)	Bauxite	Aluminium oxide	$Al_2O_3 . 2H_2O$
3. Copper (Cu)	(i) Cuprite	Copper (I) Oxide	Cu_2O
	(ii) Copper glance	Copper (I) sulphide	Cu_2S
	(iii) Copper pyrite	Copper iron sulphide	$CuFeS_2$
4. Iron (Fe)	Haematite	Iron (III) oxide	Fe_2O_3
5. Zinc (Zn)	(i) Zinc blende	Zinc sulphide	ZnS
	(ii) Calamine	Zinc carbonate	$ZnCO_3$
6. Mercury (Hg)	Cinnabar	Mercurry (II) sulphide	HgS
7. Lead (Pb)	Galena	Lead sulphide	PbS

5.4. EXTRACTION OF METALS (METALLURGY)

As we have learnt above, metals are obtained from their ores. An ore may contain the metal in the elementary state or in the form of its compounds along with earthy and sandy impurities. **The unwanted impurities such as the earthy, sandy and rocky materials associated with the ores are called gangue or matrix.** So, after mining of the ore from the ground, it is converted into pure metal. Several steps are involved in this process. **The various steps involved in the extraction of the metal from its ores followed by refining of the metal is called metallurgy.** The steps involved in the metallurgy varies from metal to metal. The actual process of metallurgy employed for a particular metal depend upon the nature of the ore, nature of the metal and the types of impurities present. However, metallurgy of most of the metals involves the following steps

(i) Concentration or enrichment of the ore.

(ii) Extraction of the metal from the concentrated ore.

(iii) Refining or purification of the impure metal.

Step 1. Concentration or Enrichment of the ore

The first step in metallurgy is to remove the unwanted impurities or gangue from the ore. As the mined ore is in the form of big lumps, it is first broken into small pieces in mechanical crushers and then pulverised to a fine powder in special mills. The processes used for removing the gangue from the ore are based on the differences between the physical or chemical properties of the gangue and the ore. Different separation techniques are employed.

(a) Gravity Separation or Hydraulic Washing

The crushed and powdered ore is taken in large wooden tanks or spread on special types of tables having grooves on the top. It is then washed with a stream of water. The lighter gangue particles are washed away while the heavier ore particles are left behind. The method is used for the concentration of oxide ores of heavy metals like lead, tin and iron.

(b) Froth Floatation Process

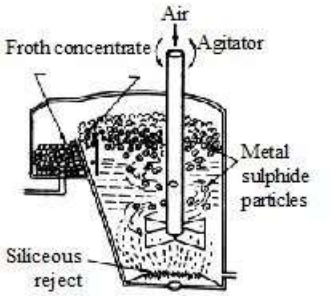


Fig. 5.4. Apparatus for froth floatation concentration

A convenient method widely used in the concentration of sulphide ores is the **froth floatation process**. Sulphides are more readily wetted by oil than by water, while those of gangue are preferentially wetted by water. The crushed and powdered ore is wetted with oil and then agitated in a tank of soapy water. Air is bubbled through the mixture (Fig. 5.4), bubbles stick to oil coated particles, and they float on the surface as **froth**. The gangue being wetted by water settles down to the bottom of the tank. The froth on the top of the tank is allowed to flow off, and the concentrated ore is recovered. Sulphide ores such as copper pyrites ($CuFeS_2$), galena (PbS), zinc blende (ZnS) etc are concentrated by this method.

(c) Electromagnetic Separation

This method is used for the concentration of the ore when either the ore or the impurities (gangue) is magnetic in nature.

The arrangement consists of two rollers over which a leather belt moves. One of the rollers is an electromagnet. The powdered ore is dropped over the belt on the side of non-magnetic roller. As the ore particles move attracted by the magnetic roller and fall into a heap on the inner side. The non magnetic particles simply falls down from the edge of the magnetic roller (Fig

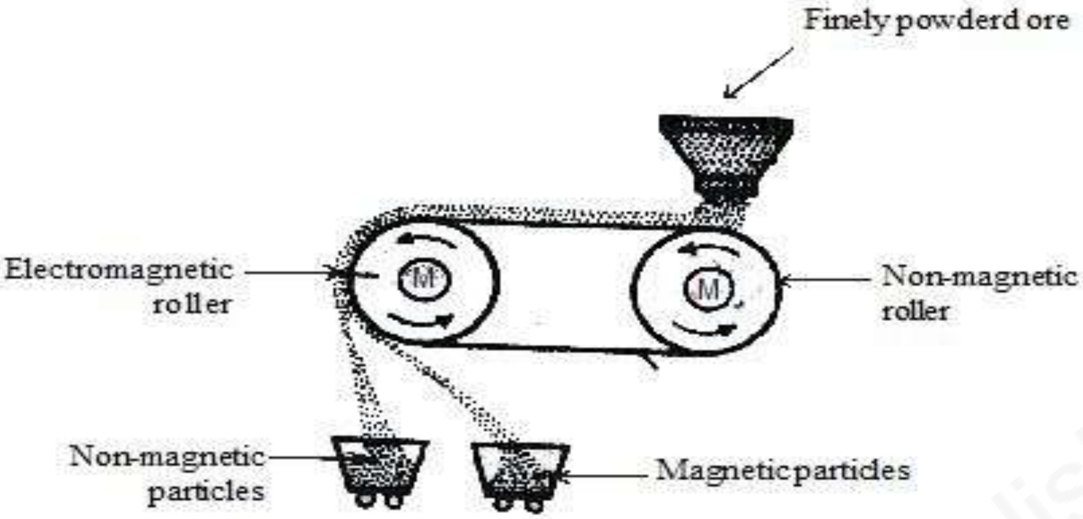


Fig. 5.5. Electromagnetic methods of concentration

5.5). Some ores like tinstone, magnetite etc. are separated by this method.

(d) Leaching

This is a chemical method for the concentration of the ore. In this process, a particular mineral in the ore is dissolved selectively by employing certain acids, bases or other reagents , separating it from the gangue.

For example, bauxite, (Al_2O_3) an aluminium ore is concentrated by leaching with hot sodium hydroxide solution. The aluminium oxide dissolves in the excess base forming soluble sodium meta-aluminate

$$Al_2O_3 + 2\,Na\,OH \longrightarrow \underset{\text{sodium meta aluminate}}{2\,Na\,AlO_2} + H_2O$$

The solution is filtered to remove insoluble gangue particles. The filtrate when diluted with water and agitated gives precipitate of aluminium hydroxide.

$$NaAlO_2 + 2\,H_2O \longrightarrow Al(OH)_3\downarrow + NaOH$$

The precipitate is dried and then heated strongly to get pure aluminium oxide (alumina).

$$2\,Al\,(OH)_3 \xrightarrow{\text{heat}} Al_2O_3 + 3\,H_2O$$

(pure
alumina)

Step II Extraction of the Metal from the Concentrated Ore (Reduction)

On the basis of reactivity, we have classified the metals into three categories. They are (i) highly reactive metals (ii) moderately reactive metals and (iii) metals of low reactivity (Fig.5.3) Different techniques are used for obtaining the metals in each category. You are familliar with the process of oxidation and reduction explained in previous chapter. Obtaining metals from their compounds is also reduction process.

(a) Extraction of Metals which are low in the Activity series

Metals in this group are unreactive. The sulphides and oxides of these metals can be reduced to metals by heating alone. For example, cinnabar (HgS) is an ore of mercury. When it is heated in air, it is first converted into mercuric oxide (HgO). Mercuric oxide is then reduced to mercury on further heating.

$$2\,HgS\,(s) + 3\,O_2\,(g) \xrightarrow{\text{heat}} 2\,Hg\,O(s) + 2\,SO_2(g)$$

$$2\,HgO \xrightarrow{\text{heat}} 2\,Hg\,(l) + O_2\,(g)$$

Similarly, when copper glance (Cu_2S), an ore of copper, is heated strongly in air, it gives copper according to the following reactions :

$$2\,Cu_2S + 3\,O_2 \longrightarrow 2\,Cu_2\,O + 2\,SO_2\uparrow$$

$$2\,Cu_2\,O + Cu_2S \longrightarrow 6\,Cu + SO_2\uparrow$$

(b) Extraction of Metals in the Middle of the Activity Series. (Fe, Zn, Pb, etc.)

These metals are found in nature in the form of their oxide, sulphide or carbonate ores. Further, it is easier to obtain the metal from its oxide than sulphides or carbonates. Therefore, prior to reduction, the sulphide and carbonate ores are first converted into the corresponding metal oxides. This is done by either of the following two methods.

(i) Calcination (for carbonate ores)

The carbonate ores are changed into oxides by heating strongly in limited air. The process is known as calcination. For example, calamine, a carbonate ore of zinc can be converted to zinc oxide by calcination.

$$ZnCO_3\,(s) \xrightarrow{\text{heat}} ZnO(s) + CO_2\,(g)$$

(ii) Roasting (for sulphide ores)

It is the process of heating the ores strongly in the presence of excess air. As a result, the sulphide ores are converted into oxides

For example,

$$\underset{\text{(zinc blende)}}{2\,ZnS\,(s)} + 3\,O_2\,(g) \xrightarrow{\text{heat}} 2\,ZnO(s) + 2\,SO_2\,(g)$$

$$\underset{\text{(galena)}}{2\,PbS\,(s)} + 3\,O_2\,(g) \xrightarrow{\text{heat}} 2\,PbO(s) + 2\,SO_2(g)$$

Reduction to the Metal

To reduce the metal oxide to the metal, the roasted or calcined ore is mixed with suitable quantity of coke (carbon), or another element which has greater affinity for oxygen. They act as reducing agents. Oxides of zinc, iron and lead are reduced to their respective metals by heating with coke.

$$ZnO\,(s) + C(s) \xrightarrow{\text{heat}} Zn\,(s) + CO(g)$$

$$Fe_2O_3(s) + 3\,C(s) \xrightarrow{\text{heat}} 2\,Fe(s) + 3\,CO(g)$$

Carbon monoxide formed also acts as reducing agent. It further reduces the metal oxide to metal.

$$Fe_2O_3\,(s) + 3\,CO(g) \longrightarrow 2\,Fe\,(s) + 3CO_2(g)$$

$$ZnO(s) + CO\,(g) \longrightarrow Zn\,(s) + CO_2\,(g)$$

The process by which the metal oxides in the fused state are reduced by carbon to the free metal involving melting is also called **smelting**.

Certain metallic oxides cannot be reduced by carbon because the affinity of oxygen for the metal is greater than its affinity for carbon. Such metallic oxides can be reduced with the help of highly reactive metals such as sodium, magnesium, aluminium, etc. Such metals can act as reducing agents. For example, oxides of manganese and chromium are reduced to their corresponding metals by heating with aluminium powder.

$$3\,MnO_2\,(s) + 4\,Al\,(s) \xrightarrow{\text{heat}} 3\,Mn\,(l) + 2\,Al_2O_3\,(s) + \text{Heat}$$

$$Cr_2O_3\,(s) + 2\,Al\,(s) \longrightarrow 2\,Cr\,(l) + Al_2O_3\,(s) + \text{Heat}$$

The reaction is **highly exothermic**. The heat evolved is so high that the metal is obtained in the molten state. As the reaction using aluminium libertate a large amount of heat, the process is known as **aluminothermic process**.

Similarly, when iron (iii) oxide (Fe_2O_3) is heated with aluminium powder, the heat evolved is so high that iron obtained melts.

$$Fe_2O_3\,(s) + 2\,Al\,(s) \xrightarrow{\text{heat}} 2\,Fe\,(l) + Al_2O_3\,(s)$$

The reaction is , therefore, used for welding the broken parts of iron machinery, join railway tracks or cracked machine parts.

(c) Extraction of Metals of High Reactivity (K, Na, Ca, Mg, Al)

These highly reactive metals cannot be obtained by reduction of their oxides by heating with carbon or aluminium. This is because these highly reactive metals have greater affinity for oxygen than for carbon or aluminium. These metals are obtained by the electrolysis of their fused salts. The method is called **electrolytic reduction.** As an example, we may consider the manufacture of sodium metal by the electrolysis of fused sodium chloride (NaCl). Various reactions taking place during electrolysis may be represented as given below.

When sodium chloride melts, it splits into sodium ion (Na^+) and chloride ion (Cl^-).

$$NaCl\,(s) \xrightarrow{\text{fusion}} Na^+\,(l) + Cl^-(l)$$

When electricity is passed through the melt, Na^+ ions go to the negatively charged electrode (cathode) and reduced to sodium atoms. Chloride (Cl^-) ions move towards the positively charged electrode (anode) and are oxidised to chlorine gas.

$$\text{At cathode,}\quad Na^+ + e^- \longrightarrow Na \quad [\text{ reduction }]$$

$$\text{At anode,}\quad 2\,Cl^- \longrightarrow Cl_2 + 2\,e^- \quad [\text{ oxidation }]$$

Similarly, calcium and magnesium are also obtained by the electrolysis of their fused chlorides. Aluminium is obtained by the electrolytic reduction of the molten alumina (Al_2O_3).

Step III . Refining of Impure Metals

The metals obtained by various reduction processes are not very pure. They are called crude metals. They contain impurities, which must be removed to obtain pure metals. The method used for refining depends upon the nature of the metals and the nature of the impurities present.

The most widely used method for refining impure metals is **electrolytic refining**. The procedure generally followed in this method is as follows .

(i) The impure metal, taken in the form of a thick block is made the anode in the electrolytic cell.
(ii) A thin sheet of pure metal is made the cathode.
(iii) A solution of the salt of the metal is taken as the electrolyte (electrolytic bath).

When an electric current is passed through the solution, the pure metal from the anode passes into the solution in the form of metal ions. An equivalent amount of metal ions from the solution is deposited as pure metal on the cathode. Thus, as the electrolysis proceeds the size of anode keeps on decreasing while that of the cathode keeps on growing. The soluble impurities go into the solution, whereas, the insoluble impurities settle down below the anode as **anode mud**.

As an example, we may consider the electrolytic **refining of copper** (Fig. 5.6)

The block of impure copper is made the anode. A thin sheet of pure copper is made the cathode . The electrolyte is a solution of copper sulphate acidified with dilute sulphuric acid. On passing electric current, pure copper from the anode passes into solution as copper (II), Cu^{2+} ions. Equivalent amounts of Cu^{2+} ions from the solution are deposited on the cathode as pure copper.

$$\text{At Anode, } Cu\,(s) \longrightarrow Cu^{2+}\,(aq) + 2\,e^-$$
$$\text{At Cathode, } Cu^{2+}(aq) + 2\,e^- \longrightarrow Cu\,(s)$$

Thus, as the electrolysis proceeds, anode becomes thinner whereas cathode becomes thicker. The impurities present in the crude copper either go into the solution or fall down below the anode. The less reactive metals like

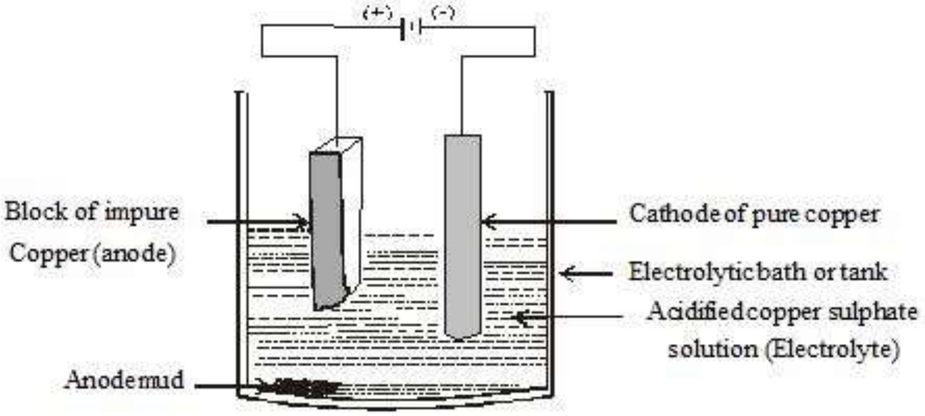
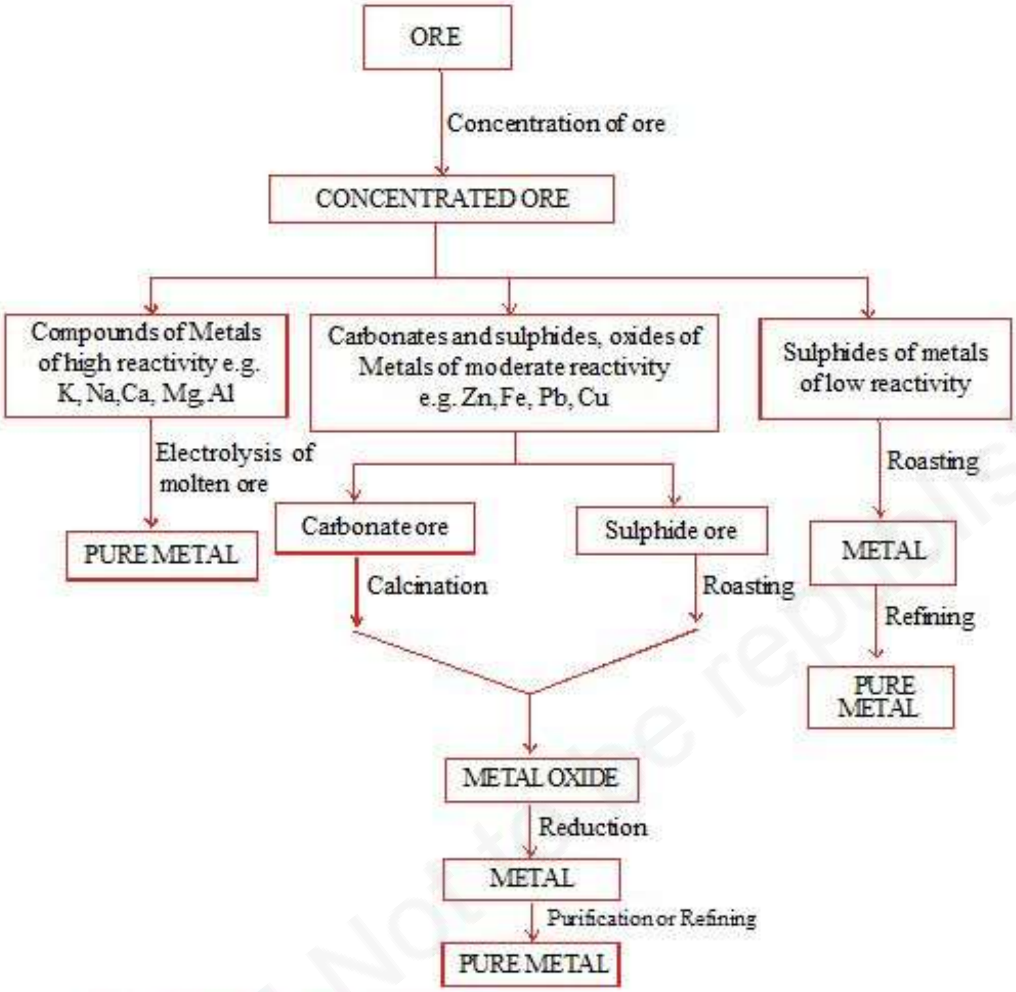


Fig.5.6. Electrolytic refining of Copper

gold, silver etc present in the impure coppers falls below the anode as such in the form of **"anode mud"**.

Many metals, such as zinc, tin, nickel, silver, gold, etc. are refined electrolytically in this manner.

The different steps involved in the extraction of pure metal from ores may be summarised as follows.



5.5. ALLOYS AND AMALGAM

We are familiar with many household implements like knife, dagger, spades, springs, saws etc. made of steel. We are also familiar with various kitchen utensils like dishes, bowls, spoons, nuts, screws etc. made of stainless steel or brass. Statues are made of bronze. They are made of materials called **alloys**.

A homogenous mixture of two or more metals or metal and non-metal is called an alloy. An alloy containing mercury as one of the constituent metals is known as **amalgam**. For example, sodium amalgam, zinc amalgam etc. Although there are only about 75 naturally occuring metal elements, there are thousands of different combinations of them, each of which has its own special properties. This shows that metals have the ability to mix with and combine with one another to form practically a large number of alloys. In general, an alloy is the solid which result when two or more metals, or metal and non-metal are melted together to form a homogenous mixture and then allowed to cool.

Alloys have properties of their own that often differ markedly from those of the elements of which they are made. For example, solder is a low melting alloy of tin and lead. Its melting point is lower than that of any of its constituents. There are also alloys which have higher melting points than any of their constituents.

Alloys are usually harder than the parent metals. Pure iron is quite soft and ductile in comparison with steel alloys which are composed mainly of iron and small percentage of carbon. The tensile strength of duralumin (an alloy of Al with Cu, Mg & Mn) is five times that of pure aluminium. Gold is too soft even for jewelery. It must be hardened by alloying with copper or some other metal.

One of the chief differences between pure metals and alloys is corrosion resistance. Alloys are usually more resistant to corrosion, hence they are more useful for materials that are exposed to the action of corrosive agents. However, alloys are generally poorer conductors of heat and electricity than pure metals. Very pure copper is needed for electrical wires, because even traces of certain impurities decrease the conductivity a great deal.

In spite of the important differences between the elemental metals and alloys, there are some important similarities. Alloys usually, have to some degree the metallic properties of heat and electrical conductivity, ductility and malleability, and they show metallic lustre.

Table 5.4. gives some of the commonly used alloys, their composition and uses.

Table 5.4. Some common alloys

Name	Composition	Properties	Uses
1. Steel	Iron (98.4 –99.2%) Carbon (0.2– 1.6%)	Hard, tough and strong	Fabrication, cutting blades, saws, springs, ships,bridges, vehicles etc
2. Stainless Steel	Iron (74%) Chromium (18%) Nickel (8%)	Corrosion-resistant, hard	Kitchenware, stainless cutlery, ornamental pieces, surgical instruments
3. Brass	Copper (60–80%) Zinc (20– 40%)	Malleable, strong. corrosion resistant, can be easily cast	Utensils, hardware, radiator cores, screws, nuts, bolts, musical instruments
4. Bronze	Copper (75–90%) Tin (10-25%)	Light, strong, highly resistant to corrosion	Statues, coins, medals, cooking utensils, valves, rods
5. Duralumin	Aluminium (95%) Copper (4%) Mangnese (0.5%) Magnesium (0.5%)	Light, strong, corrosion resistant, great tensile strength	Aeroplanes, automobile parts, pressure cookers
6. Solder	Lead (40 – 60 %) Tin (40– 60 %)	Lower melting point than either Pb or Sn	Soldering, joining electrical wires.

A great deal of scientific research is going on at present to the study of alloys for use in high-temperature jet and rocket engines and nuclear energy power plants. When the methods for the production of rarer metals are perfected, these researches will undoubtedly lead to the formation of more valuable alloys than we have today.

Let us Answer these.

1. Name two metals which are found in nature in the free state.
2. **Define the following terms.**
   (i) Mineral (ii) Ore (iii) Gangue
3. Differentiate between roasting and calcination used in metallurgy.
4. Name two sulphide and two oxide ores.
5. Name the systematic steps involved in metallurgy.

6. Write the chemical process used for obtaining a metal from its oxide.

7. Give an example of a sulphide ore which is reduced to metal by heating alone. Give the relevant equations involved.

8. Why aluminium cannot be obtained by reduction of its oxide with coke ?

9. Explain how the following metals can be obtained from their compounds by reduction process.

   (a) Metal "X" which is low in activity series

   (b) Metal "Y" which is moderately reactive

   (c) Metal "Z" which is high up in the reactivity series

   Give one example of each type.

10. An alloy "solder" is used in electronic industry to join different electronic components. What are the constituents of this alloy ?

11. Which alloy is used for making bodies of aircrafts ? What is its approximate composition?

SUMMARY

- Elements can be classified as metals and non-metals.
- Metals have metallic lustre, malleability ductililty and are good conductors of heat and electricity. They are solids at room temperature except mercury which is a liquid.
- Non-metals have properties different from those of metals. They are neither malleable nor ductile. They are bad conductors of heat and electricity, except graphite which conducts electric current.
- Metals combine with oxygen to form basic oxides. Aluminium oxide and zinc oxide show the properties of both basic and acidic oxides. These oxides are known as amphoteric oxides.
- Non-metals combine with oxygen to form acidic or neutral oxides.
- Different metals have different reactivities with water and dilute acids.
- Metals occur in nature as free elements or in the form of their compounds.
- The extraction of metals from their ores and then refining them for use is known as metallargy.
- Several steps are involved in the extraction of pure metal from ores. Different techniques are to be used for obtaining the metals from their ores.

- Metallurgy of most of the metals involve three steps–
  (i) Concentration or enrichment of the ore (ii) Extraction of the metal from the concentrated ore and (iii) Refining or purification of the impure metal.
- A homogenous mixture of two or more metals, or metal and non-metal is called an alloy.

EXERCISES

1. Give five points to distinguish metals from non-metals.

2. You are given a hammer, a battery, a bulb, wires and a switch.
   (a) How would you use them to distinguish between samples of metals and non-metals.
   (b) Assess the usefulness of these tests.

3. Frying pans and boilers are generally made from aluminium but not from steel. Why? Give two reasons.

4. Distinguish between metals and non-metals on the basis of their chemical properties. (Give three points).

5. What are amphoteric oxides ? Give examples of two amphoteric oxides.

6. Name two metals which will displace hydrogen from dilute acids, and two metals which will not.

7. Differentiate between :
   (i) Mineral and ore
   (ii) Alloy and amalgam
   (iii) Calcination and roasting.

8. What are the common methods of ore concentration ?

9. Name the process that is used to concentrate sulphide ores. Describe the stages that are involved in the conversion of concentrated sulphide ore into the corresponding metal.

10. How is impure metal refined electrolytically ?

11. What is an alloy ? Write the compostion and two uses of brass and bronze.

★★★★★★

CHAPTER 6

CARBON AND ITS COMPOUNDS

Coal, charcoal and graphite are black substances. Graphite is used for making pencils and dry cells. These substances are made up of carbon. Diamond is a bright, beautiful, hard and colourless substance which is used in jewellery. It is very rare and costly. Diamond is pure carbon. Carbon can exist in distinctly different forms which have quite different properties.

Many substances around us are made up of carbon. Clothes, foods, medicines, paper, plastic, cooking gas, electric cables and even transport vehicles and aeroplanes have components made up of carbon compounds. All living things have structures mainly based on carbon compounds.

The amount of carbon present in the earth's crust and in the atmosphere is quite low. Only about 0.02 % of carbon is present in the earth's crust as minerals in the form of carbonates, bicarbonates, coal and petroleum and about 0.03 % of the atmosphere is carbon dioxide. The importance of carbon is quite large inspite of its meagre abundance in the earth. In this chapter, we will discuss the uniqueness of carbon and its properties. This will illustrate the usefulness of carbon and its compounds.

6.1. Allotropes of Carbon

As seen above carbon occurs in nature, in physically distinct forms. However chemically they are the same. When heated strongly in air diamond and graphite leaves no residue but all are converted into carbon dioxide. Graphite is soft, slippery and greyish black in colour. It is a good conductor of electricity. However, diamond is transparent, very hard and is a non-conductor of electricity.

The difference between them lies in the manner in which the carbon atoms are bonded to one another. In diamond each carbon atom lies at the centre of a regular tetrahedron and is covalently bonded with four carbon atoms located at its four corners (Fig 6.1a). All the bonding electrons of the carbon atoms are thus used leaving no free electrons. This makes diamond an extremely poor conductor of electricity.

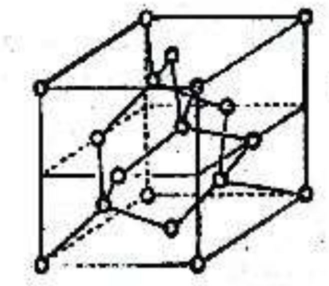


Fig. 6.1 (a) Structure of Diamond

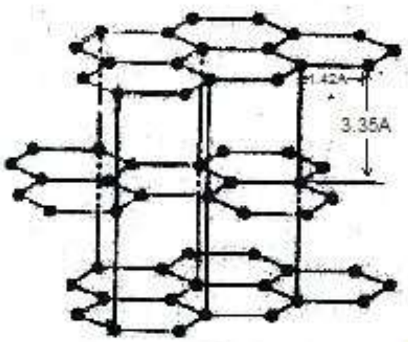


Fig. 6.1 (b) Structure of Graphite

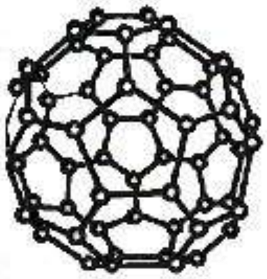


Fig. 6.1 (c) Structure of Fullerene

The tightly bonded three dimensional tetrahedral arrangement of carbon atoms produce a rigid network and makes diamond the hardest substance known. The melting point of diamond is 3500°C which is very high.

Each carbon in graphite is covalently bonded to only three neighbouring carbon atoms forming layers of hexagonal networks (Fig 6.1b), separated by a comparatively larger distance. Because of this distance between two successive layers, the possibility of a covalent bond being formed between carbon atoms lying in opposite layers is ruled out. The fourth valence electron thus remains almost free. This enable the easy flow of electrons through graphite and makes it a good conductor of electricity. These layers slide over one another. This makes graphite soft to feel, greasy to touch and thus graphite has lubricating properties. Thus powdered graphite is used as lubricants in machines where high temperature makes other lubricants unsuitable.

Graphite and diamond are the allotropes of carbon. Allotropism is the phenomenon of existence of an element in two or more physically different forms but with similar chemical properties and those forms are called **allotropes**.

Carbon also has another allotropic form known as fullerene (Fig 6.1c). It is a form in which sixty carbon atoms are arranged in the shape of a football. Since this looks like the geodisic dome designed by the US architect Buckminster Fuller, the molecule is named as **fullerene**.

Since carbon has four valence electrons, it is capable of bonding with four other carbon atoms or atoms of some other monovalent element. Such bonds are single bonds. Carbon forms compounds with hydrogen, oxygen, nitrogen, sulphur, phosphorus, chlorine, bromine and many other elements giving rise to a large number of compounds. These compounds have specific properties which are characteristics of the elements rather than carbon present in the molecule.

6.2. Catenation – Unique propety of Carbon

The number of carbon compounds whose formulas are known to us is about three million. The number of compounds formed by all the other elements put together is less than one lakh.

What makes the number of carbon compounds so large ?

Carbon has the unique property to form bonds with other carbon atoms to form long chains. This property of carbon is called **catenation**. The carbon chains may be straight, branched or even rings. The carbon atoms may be linked by double or triple bond in addition to single bonds. The carbon-carbon bond is very strong and hence stable.

The bonds that carbon forms with most other elements are very strong and therefore, these compounds are exceptionally stable. One reason for the formation of strong bonds by carbon is its small size. This enables the nucleus to hold on to the shared electron pairs strongly. Generally, the covalent bonds formed by elements having larger sizes are much weaker.

6.2.1. Hydrocarbons Saturated hydrocarbons

Compounds of carbon and hydrogen only are called hydrocarbons. Methane, CH_4 is the simplest possible hydrocarbon. Carbon is tetravalent while hydrogen is monovalent. Each of the four valence electrons of carbon shares with one electron of hydrogen to acquire the noble gas configuration. Thus, methane is formed as shown in Fig. 6.2

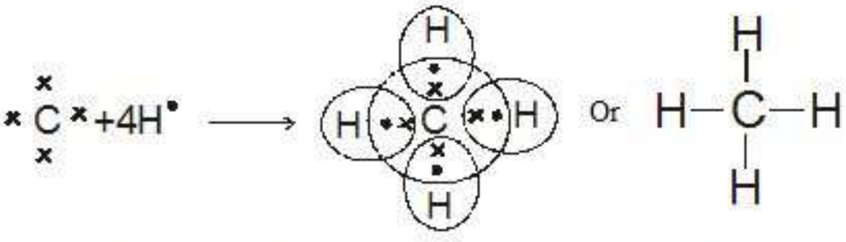


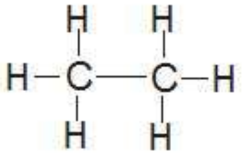
Fig. 6.2. Electron dot and cross structure of methane

Methane occurs in the nature in the form of gas. It is formed by the decay of plants and animals i.e., organic matters and is also known as marsh gas. Natural gas found in earth's crust is always associated with petroleum. Methane is widely used as fuel and is a major component of bio-gas and Compressed Natural Gas (CNG).

The next hydrocarbon after methane is ethane with the molecular formula C_2H_6. To arrive at the structure first write the two carbon atoms linked by a single bond,

C — C

Then, in the second step, use the hydrogen atoms to satisfy the remaining three valencies of each carbon,



The electron dot structure of ethane is shown in the Fig 6.3

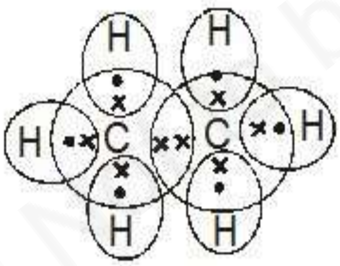


Fig 6.3. Electron dot structure of ethane molecule.

Now give answers for the following

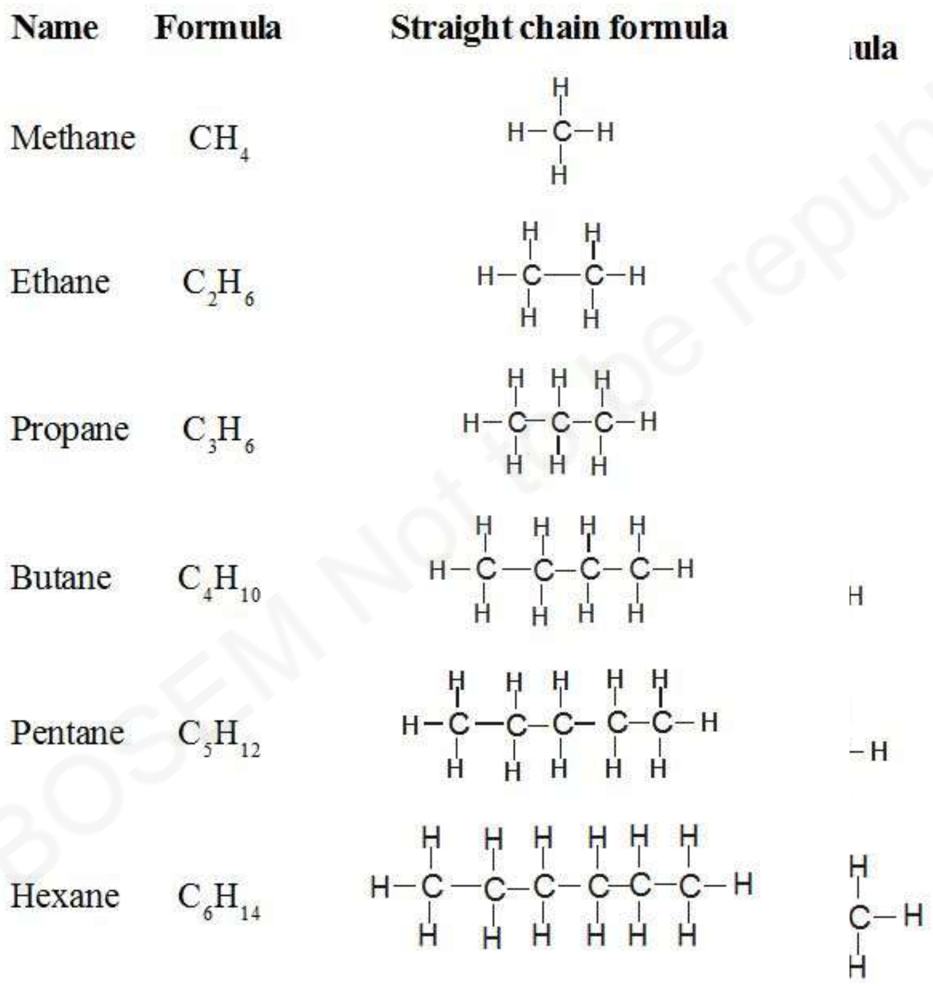
1. Draw the structure of propane with the molecular formula C_3H_8.

2. Draw the electron dot structure of chloromethane with the molecular formula CH_3Cl.

The hydrocarbons ethane C_2H_6, propane C_3H_8 etc. contains carbon -carbon single bonds only. Such hydrocarbons along with methane CH_4 are known as **saturated hydrocarbons** or simply **alkanes**. They are **open chain** compounds.

Alkanes are represented by the general formula $C_n\ H_{2n+2}$ when n = 1,2,3 .....etc. Alkanes when arranged in order of increasing molecular mass constitute a series in which any two consecutive alkanes differ by $–CH_2–$. Such a series is called **homologous series**. The names and structures of a few alkanes are given in Table 6.1.

Table 6.1. Names and Structures of some Alkanes

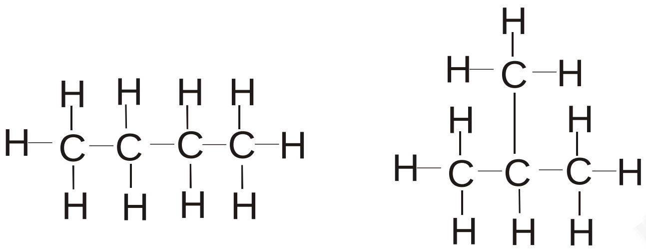


Name	Formula	Straight chain formula
Methane	CH_4	
Ethane	C_2H_6	
Propane	C_3H_6	
Butane	C_4H_{10}	
Pentane	C_5H_{12}	
Hexane	C_6H_{14}	

Will it be possible to write structures of butane C_4H_{10} other than the straight chain structure given above ? We can write two different carbon skeletons of C_4H_{10} using the principle of carbon-carbon single bonds only. They are (Fig 6.4)

***Fig 6.4 Two possible carbon skeletons of* some C_4H_{10}**

Satisfying the remaining valencies of carbon with hydrogen atoms we have the complete structures of the molecules of C_4H_{10} (Fig. 6.5).



(a) (b)

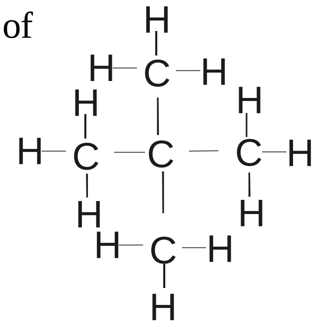
***Fig. 6.5 Complete structure of* C_4H_{10}**

The structure (a) has all the four carbon atoms in a single line or chain. Its name is n-butane (read as normal butane). The structure (b) has three carbon atoms in a single chain and one carbon atom is branched. It is called iso-butane. If the carbon atoms in the longest chain are numbered as

then the structure (b) is also called as 2-Methyl propane. Why is it so?

Let us answer these.

1. Give the name of



2. **(a)** Write the possible carbon skeletons of C_5H_{12}.

   (b) Then write the possible structures of C_5H_{12}.

   The two structures of C_4H_{10}, although they have the same molecular formula, are different from each other. Such compounds with the same molecular formula but different structure are called **structural isomers**. They can be writen in short as

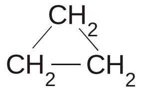
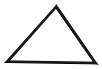
$CH_3CH_2CH_2CH_3$ and

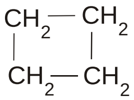
What is the relationship between the two structures ?

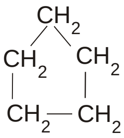
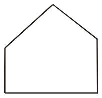
In addition to straight and branched chains, the same hydrocarbon may have carbon atoms in the form of a ring. Such hydrocarbons are called **cycloalkanes**. They have the general formula C_nH_{2n} i.e., they have two hydrogen atoms less than the corresponding saturated hydrocarbons.

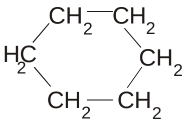
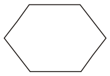
Why is it so? Examples of some cycloalkanes are given in Table 6.2

Table 6.2.

NAME	**FORMULA**	**STRUCTURE**		
Cyclopropane	C_3H_6		or	

Cyclobutane C_4H_8  or 

Cyclopentane  or 

Cyclohexane  or 

Unsaturated Hydrocarbons

In section 6.2 we have seen that carbon-carbon bonds may also be double or triple. Such open chain hydrocarbons containing double or triple bonds are called **unsaturated hydrocarbons**. They are so called as they have fewer number of hydrogen atoms when compared to the saturated hydrocarbons.

Alkenes are those unsaturated hydrocarbons which contain double bond. Such compounds have the general formula C_nH_{2n} (same as cycloalkanes). The first member of this group is ethene (or ethylene) which have the formula C_2H_4. Let us write the electron dot structure of ethene, C_2H_4. It is shown in Fig 6.6.

Fig. 6.6 (a) Electron dot structure of ethene (b) Structural formula

Table 6.3 Shows the homologues of ethene, name and the structural formulas.

Table 6.3

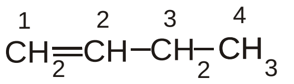
Molecular Formula	Structure	Name
C_2H_4	$CH_2 = CH_2$	Ethene
C_3H_6	$CH_3—CH = CH_2$	Propene
C_4H_8	(i) $CH_2 = CH—CH_2 — CH_3$	But–1– ene
	(ii) $CH_3 — CH = CH — CH_3$	But – 2 – ene

Let us answer these.

1. Why is the first member of alkene C_2H_4?
2. Why does butene have two names ?

Isomerism in alkene

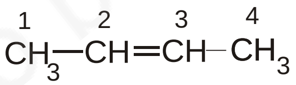
The two structures of butene C_4H_8 have the same carbon chain but differ in the position of carbon-carbon double bond. The carbon atoms in the chain are numbered starting from the end which is nearer to the double bond. Thus



$\overset{1}{C}H_2{=}\overset{2}{C}H{-}\overset{3}{C}H_2{-}\overset{4}{C}H_3$

But - 1-ene (1– Butene)

Fig 6.7 (a)



$\overset{1}{C}H_3{-}\overset{2}{C}H{=}\overset{3}{C}H{-}\overset{4}{C}H_3$

But - 2- ene (2-Butene)

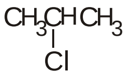
Fig 6.7 (b)

In Fig 6.7 (a) the double bond lies between the C_1 and C_2 and therefore it is called **But–1–ene.** This indicates that the position of double bond is between C_1 and C_2. In Fig 6.7 (b) the double bond lies between C_2 and C_3 and therefore it is called **But-2-ene.** This indicates that double bond lies between C_2 and C_3

Such compounds having the same carbon chain but differ in the position of double bond (which is a characteristic property) are called **postion isomers**. Let us answer the following.

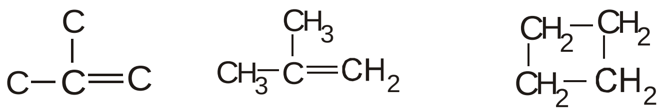
1. Write the position isomers of pentene C_5H_{10}.

2. Identify the relationship between

(i) $CH_3CH_2CH_2Cl$ and  $CH_3\underset{\displaystyle Cl}{\underset{|}{C}}HCH_3$

(ii) CH_3CHCl_2 and CH_2ClCH_2Cl

Examine the structures given in Fig 6.8 (a) and 6.8 (b)


Skeleton

(a) Isobutene

(2-Methyl propene)

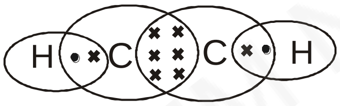
(b) Cyclobutane

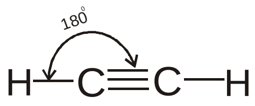
Fig 6.8 Different structures of C_4H_8

They have the same molecular formula C_4H_8 as butenes shown in Figs 6.7. But the structure of carbon chains are all different. Thus the relationship between But-1-ene or But-2-ene with isobutene (2-Methyl propene) or cyclobutane is structural isomerism. Thus butene can show both **structural** and **position isomerisms.**

Alkynes:

Hydrocarbons with carbon-carbon triple bond are called alkynes. They have the general formula $C_nH_{2n-2.}$ The first member of the series is ethyne (or acetylene) with the formula C_2H_2. Let us draw the structure of this molecule.





Some of the compounds of the family are

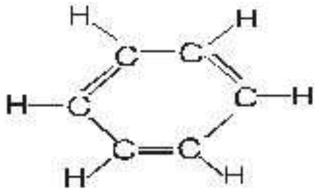
Formula	Structure	Name
C_2H_2	$H-C\equiv C-H$	Ethyne
C_3H_4	$CH_3-C\equiv C-H$	Propyne
C_4H_6	(i) $H-C\equiv C-CH_2-CH_3$	But-1-yne
	(ii) $CH_3-C\equiv C-CH_3$	But-2-yne

Question : **(i)** Why does Butyne has two isomers ?

(ii) Is there any other structural isomer of Butyne ?

Alkenes and alkynes are highly reactive organic compounds. They are used for making many useful chemicals and other substances like plastic, rubbers etc.

Yet there is another type of hydrocarbons which are known as **arenes** or **benzene** and its **derivatives**. They are cyclic and contain at least one benzene ring. Benzene itself is C_6H_6. The simplest representation of benzene is



Arenes are known as aromatic hydrocarbons. They are very stable hydrocarbons inspite of their general formula C_nH_{2n-6}.

We are familiar with cooking gas, motor spirit, petrol, diesel, kerosene, lubricating oil. They are all hydrocarbons and their source is petroleum which is a black liquid found in the crust of the earth.

6.3. Petroleum (Mineral Oil)

Petroleum is a dark coloured viscous liquid found deep in the earth's crust. Chemically it is a mixture of various types of hydrocarbons along with some oxygen, sulphur and nitrogen containing compounds.

Petroleum is believed to be formed by the decay and decomposition of marine animals as well as that of plant materials of the prehistoric forests. In other words, petroleum is of animal and plant origin. It is thought that due to some upheavals like earthquake, volcanic eruptions etc, these prehistoric forests, and sea animals get buried under the earth's crust. Due to prolonged action of high pressure and high temperature in the interior of the earth for ages, the organic matter decomposed into petroleum.

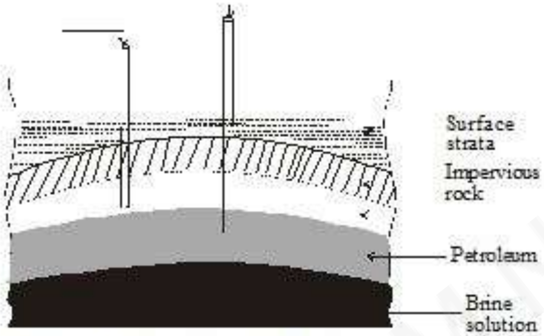


Fig. 6.9. Mining of Petroleum

A diagrametic representation of the occurrence and mining of petroleum is given in Fig 6.9. Petroleum is pumped out through the deep wells bored into the crust.

The process of separating crude petroleum into useful fractions by fractional distillation is known as refining. Fractions condensing at different boiling points are collected separately and are called products of petroleum. They are– gas, gasoline or petrol, kerosene, gas oil, diesel oil, lubricating oil, vaseline, paraffin, wax and asphalt.

Uses

Gas is used as a fuel, gasoline as a motor fuel, kerosene as household fuel and lighting purpose, gas oil and diesel oil as furnace fuel and for diesel engines, lubricating oil and grease for lubrication, vaseline for softening skin, paraffin wax for making candles, and asphalt for road surfacing. LPG (liquefied petroleum gas) is widely used for cooking.

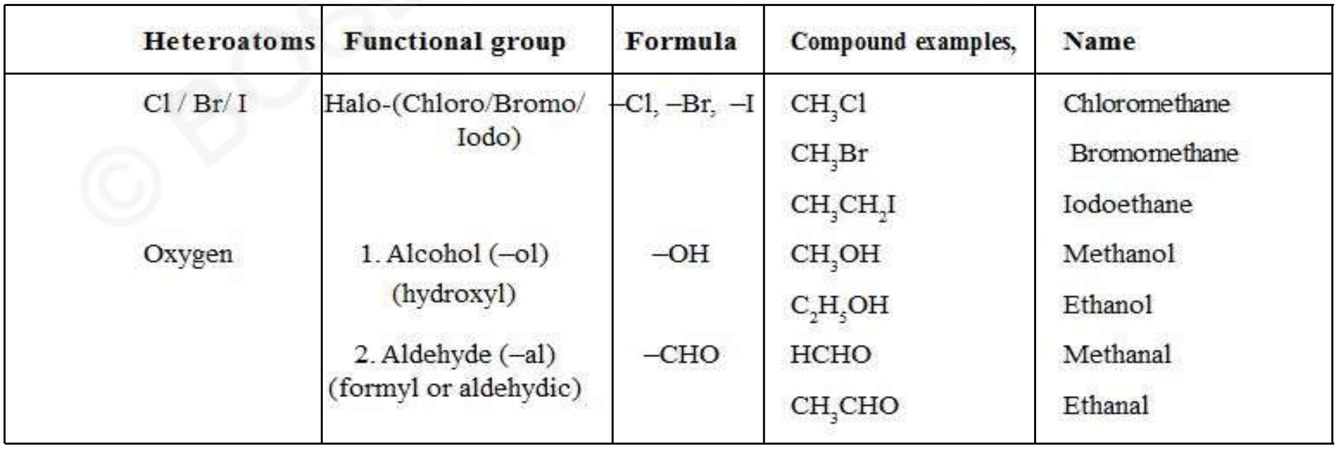
6.4. Some compounds of carbon other than hydrocarbons

Carbon can form covalent bonds with many other elements like oxygen, sulphur, halogens– fluorine, chlorine, bromine, iodine etc. In a hydrocarbon one or more hydrogens can be replaced by these elements to form stable compounds.These hetero atoms may also be present as a group combined with other atoms.

Functional Groups

The heteroatoms and the group attached to the hydrocarbon part forming a stable molecule characterise the specific properties of the compound irrespective of the nature and length of the carbon chain. Hence these heteroatoms and groups are called **functional groups**. Example of some of these functional groups are given in Tables 6.4.

Table 6.4 Some functional groups of Carbon compounds



Heteroatoms	Functional group	Formula	Compound examples,	Name
Cl / Br/ I	Halo-(Chloro/Bromo/ Iodo)	–Cl, –Br, –I	CH_3Cl	Chloromethane
			CH_3Br	Bromomethane
			CH_3CH_2I	Iodoethane
Oxygen	1. Alcohol (–ol) (hydroxyl)	–OH	CH_3OH	Methanol
			C_2H_5OH	Ethanol
	2. Aldehyde (–al) (formyl or aldehydic)	–CHO	HCHO	Methanal
			CH_3CHO	Ethanal

3. Ketone (–one)	–CO–	$CH_3–CO–CH_3$	Propanone
4. Carboxylic acid (–oic acid)	–COOH	HCOOH CH_3COOH	Methanoic acid Ethanoic acid (Acetic acid)

Alcohols are a family of carbon compounds which contain the functional group –OH attached to a saturated carbon as C OH. These compounds have similar chemical properties inspite of the differences in the number of carbon chain. They are represented by the general formula $C_nH_{2n+1}OH$.

Let us answer the following.

1. Write the homologous series of alcohols upto four carbon atoms.
2. Write any two isomers of C_4H_9OH.

6.5. Nomenclature of Carbon Compounds

The name of an organic compound essentially consists of the following three parts.

1. Word root 2. Suffix and 3. Prefix.

Word root : It is the basic unit of the name. It denotes the number of carbon atoms present in the **main chain** (the longest possible continuous chain of carbon atoms including the functional group and carbon-corbon multiple bonds) of the molecules. For molecules upto four carbon atoms special word root are used but for chains of five or more carbon atoms Greek roots are used as given below :

Chain length	Word Root
C_1	Meth –
C_2	Eth –
C_3	Prop –
C_4	But –
C_5	Pent –
C_6	Hex –
C_7	Hept–
C_8	Oct–
C_9	Non–
C_{10}	Dec–

The general name of the word root for any carbon chain is **alk–**.

Suffix : There are two types of suffixes.

(a) Primary suffix

A primary suffix is added to the the word root to indicate whether the carbon chain is of single bonds only (saturated) or contains carbon - carbon multiple bonds (unsaturated). The three basic primary suffixes are:

Nature of carbon chain	Primary suffix	General Name
i) Saturated	– ane	Alkane
ii) Unsaturated with a double bond	– ene	Alkene
iii) Unsaturated with a a triple bond	– yne	Alkyne

It is illustrated as in the following example.

Oraganic compound	Word Root	Primary Suffix	Name
$CH_3CH_2CH_3$	Prop	– ane	Propane
$CH_3CH=CH_2$	Prop	– ene	Propene
$CH_3-C\equiv CH$	Prop	– yne	Propyne

b) Secondary Suffix

A secondary suffix is added to the primary suffix to indicate the nature of the functional group present in the molecule. Some secondary suffixes are given below:

Family of compound	Functional group	Secondary suffix
Alcohols	– OH	– ol
Aldehydes	– CHO	– al
Ketones	– CO –	–one
Carboxylic acids	– COOH	– oic acid

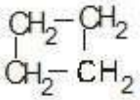
The examples given below illustrate the use of word root, primary suffix and secondary suffix while naming an organic compound.

Compound	Word Root	Primary Suffix	Secondary Suffix	Name
CH_3OH	Meth	an (e)*	ol	Methanol
CH_3CH_2OH	Eth	an (e)*	ol	Ethanol
$CH_3CH_2CH_2OH$	Prop	an (e)*	ol	Propanol (Propan-1-ol)
CH_3COOH	Eth	an (e)*	oic acid	Ethanoic acid

* The terminal 'e' from the primary suffix has been dropped because the secondary suffix begins with a vowel.

Prefix: There are two types of prefixes. They are

a) Primary prifix: It is used to distinguish open chain compounds from ring compounds (closed chain compounds). For example, in the case of cyclic (ring) compounds, a primary prefix, **Cyclo** is used immediately before the word root. Thus the name of the compound is

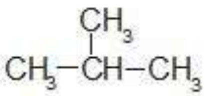


is Cyclo + but + ane = Cyclobutane

(Primary prefix) (Word root) Primary suffix (Name)

If the prefix 'Cyclo' is not used, it simply indicates open chain (acyclic) compound.

b) Secondary prefix: Certain groups are not considered as functional groups but are treated as substituents. These are called secondary prefixes. These prefixes are added immediately before the word root or primary prefix in **alphabetical order**. Some important secondary prefixes are given below

Group	Secondary prefix	Example	Name
$-F$	Fluoro	CH_3F	Fluoromethane
$-Cl$	Chloro	CH_3CH_2Cl	Chloromethane
$-CH_3$	Methyl	 $CH_3-\overset{\overset{CH_3}{\vert}}{C}H-CH_3$	2–Methyl propane

Note : *Fluromethane, chloroethane are single word name.*

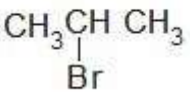
Thus, the complete name of an organic compound in the IUPAC system (International Union of Pure and Applied Chemistry) consists of the following parts :

Secondary prefix + primary prefix + word root + primary suffix + secondary suffix .

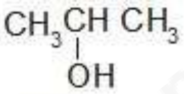
Questions

1. **Let us name the following compounds :**

(a) $CH_3CH_2CH_2Br$

b) 
$CH_3\underset{\substack{|\\ Br}}{CH}CH_3$

(c) $CH_3CH_2CH_2OH$

d) 
$CH_3\underset{\substack{|\\ OH}}{CH}CH_3$

(e) CH_3CHO

f) CH_3CH_2COOH

2. **Draw the structures of**

a) 1-Bromobutane b) 2-bromobutane c) Ethanoic acid

d) Propanone f) Butanal

6.6. Chemical Properties Of Carbon Compounds

Let us try to understand some of the properties of carbon compounds. Carbon and its compounds are combustible substances.

6.6.1. Combustion

All the allotropic forms of carbon burn in oxygen to form carbon dioxide producing large amounts of heat and light (energy).

$$C + O_2 \longrightarrow CO_2 + \text{Energy}$$

Some fuels like coal, charcoal and wood burn leaving some residue. But fuels like LPG, kerosene, alcohol etc. burn without leaving any residue. Some of the carbon compounds burn with sooty flame while some others burn with clean flames. All compounds of carbon burn to form carbon dioxide and water. Thus

$$\underset{\text{methane}}{CH_4} + 2\,O_2 \longrightarrow CO_2 + 2\,H_2O + \text{Energy}$$

$$\underset{\text{ethanol}}{CH_3CH_2OH} + 3O_2 \longrightarrow 2CO_2 + 3\,H_2O + \text{Energy}$$

Activity 6.1

- Hold a metal scale or plate over the flame of a burning candle and observe. Is there any deposit of soot ?
- Repeat the above step using a spirit lamp.
- Repeat again using a diesel oil lamp
- Record your observation and draw a conclusion.

Saturated hydrocarbons usually burn with a clear flame whereas unsaturated hydrocarbons burn with coloured (yellow to orange) flame with a lot of soots. Can you explain why is it so ?

When the supply of air (oxygen) is limited even the saturated hydrocarbons burn with a sooty flame. Why ? The gas and kerosene stoves used at home have air inlets to mix with the fuel to burn them completely. If the bottoms of the cooking vessels become black , it means that air holes are blocked and the fuel is wasted (incomplete burn). Fuel such as coal, petrol and diesel contain some compounds of sulphur and nitrogen. Their combustion produces oxides of sulphur and nitrogen. These oxides are the major pollutants of air. Besides fine particles of carbon often missed combustion and are released in the air as particulates. These particulates cause pollution.

6.6.2. Oxidation

Combustion is also an oxidation process. The final product of oxidation of an organic compound is carbon dioxide and water. Many organic compounds may form intermediate products when reacted with oxygen or any other oxidising agent. Oxidation, in its simplest sense, may be defined as the additon of oxygen or removal of hydrogen from a carbon compound.

Activity 6.2.

- Take about 3 mL of ethanol in a test tube and warm it gently on a water bath as in Fig 6.10.
- Add drop by drop a solution of 1% alkaline $KMnO_4$ (potassium permagnanate).
- Note the change in the colour of $KMnO_4$ solution.
- Why does not the colour of potassium permaganate disappear when added in excess?

Note

The above activity may be performed by using acidic potassium dichromate solution ($K_2Cr_2O_7$ + Conc H_2SO_4). However, precaution must be taken not to touch this solution by hand.

Alkaline $KMnO_4$ or acidified $K_2Cr_2O_7$ are oxidising agents i.e. they can supply oxygen atoms which can be added to organic compounds.

The reaction between ethanol and alkaline $KMnO_4$ produces an aldehyde which in turn is oxidised further to acid. The reaction occurs in the two steps

$$\text{i) } \underset{\text{ethanol}}{CH_3CH_2OH} + \underset{\text{(from } KMnO_4)}{[O]} \xrightarrow{\text{alk } KMnO_4} \underset{\text{(ethanal)}}{CH_3CHO + H_2O}$$

$$\text{ii) } CH_3\ CHO + [O] \xrightarrow{\text{alk } KMnO_4} \underset{\text{(ethanoic acid)}}{CH_3COOH}$$

Nascent oxygens (oxygen atoms) are released within the reaction medium (*in situ*). The first step is oxidation of ethanol by the removal of two hydrogens, and the second step is the oxidation by adding oxygen atom to the intermediate product ethanal. Intermediate product is readily converted to ethanoic acid.

6.6.3. Addition Reactions

(i) **Hydrogenation (addition of hydrogen):** Unsaturated hydrocarbons add hydrogen in presence of catalyst like palladium, platinum, nickel etc. to form saturated hydrocarbons. Thus, ethene adds a molecule of hydrogen to form ethane.

$$\underset{\text{ethene}}{CH_2=CH_2 + H_2} \xrightarrow{\text{Ni cat.}} \underset{\text{ethane}}{CH_3-CH_3}$$

This reaction is commonly used in the hydrogenation of vegetable oils which contain unsaturated carbon chains. On hydrogenation vegetable oils give vegetable ghee.

(ii) **Bromination (Halogenation) :** On treatment with bromine water, ethene adds a molecule of bromine to form dibromoethane. (Orange colour of bromine water is discharged during the reaction).

$$\underset{\text{ethene}}{CH_2=CH_2 + Br_2\ (\text{water})} \longrightarrow \underset{\text{1, 2 - dibromoethane}}{CH_2Br-CH_2Br}$$

This reaction is used as a test for the unsaturated organic compounds (pressence of double or triple bonds)

(iii) **Addition of Hydrobromic acid :** Ethene readily reacts with hydrobromic acid HBr to produce bromoethane

$$\underset{\text{ethene}}{CH_2=CH_2 + HBr\ (aq)} \longrightarrow \underset{\text{bromoethane}}{CH_3\ CH_2Br}$$

6.6.4. Substitution Reaction: Saturated hydrocarbons are very stable and unreactive towards most reagents. But, in the presence of diffused sunlight (indirect light), chlorine can react with methane to produce various products. One or all the hydrogens of CH_4 can be replaced by chlorine atoms. Such a reaction is known as s**ubstitution reaction.**

$$\underset{\text{Methane}}{CH_4 + Cl_2} \longrightarrow \underset{\text{Chloromethane}}{CH_3Cl + HCl}$$

$$CH_3Cl + Cl_2 \longrightarrow \underset{\text{Dichloromethane}}{CH_2\ Cl_2 + HCl}$$

$$CH_2Cl_2 + Cl_2 \longrightarrow \underset{\text{Trichloromethane (Chloroform)}}{CHCl_3 + HCl}$$

$$CHCl_3 + Cl_2 \longrightarrow \underset{\text{Tetrachloromethane (Carbon tetrachloride)}}{CCl_4 + HCl}$$

However, the reaction is very complex and very fast.

6.7. Properties of ethanol CH_3CH_2OH

Ethanol (ethyl alcohol) is one of the important compounds of carbon. Let us see some of its properties.

Ethyl alcohol is a colourless volatile liquid with a boiling point of 78°C. It has a typical pleasant smell. It is inflammable and burns with a blue flame. It is miscible with water in all proportions. It is also miscible with many organic solvents. Ethanol mixed with petrol, is used as a fuel for internal combustion engines. It is also used as the solvent for lacquers.

Ethyl acohol is the active ingredient of all alcoholic drinks. Depending upon the type of drinks the alcohol content ranges from 3 – 50%. When consumed, the alcohol is oxidised and supplies energy which in turn stimulates the human system. However alcohol affects Central Nervous System (CNS) as it is neurodepressant. When taken in small amounts, apparent excitation and euphoria are experienced. Hesitation, caution, self-respect and restraint are lost first. Mood and feelings are altered. When taken in larger amounts it causes disorganization of thought, impairment of memory and other faculties and drowsiness.

Alcohol has been believed to produce CNS depression by a generalized membrane action by altering the state of membrane lipids. A person , therefore, loses control over his body after drinking alcohol.

Some simple reaction of ethanol

(i) Reaction with sodium metal

Ethanol reacts with sodium metal to produce sodium ethoxide and hydrogen gas.

$$\underset{\text{Ethanol}}{2CH_3CH_2OH} + 2Na \longrightarrow \underset{\text{Sodium ethoxide}}{2CH_3CH_2ONa} + H_2$$

(ii) Dehydration of ethanol

Ethanol on heating with concentrated sulphuric acid, H_2SO_4 at 170°C, loses a water molecule giving ethene.

$$\underset{\text{ethanol}}{CH_3CH_2OH} \xrightarrow[170^\circ C]{\text{Conc. } H_2SO_4} \underset{\text{ethene}}{CH_2 = CH_2} + H_2O$$

(iii) Oxidation of ethanol : Oxidation of ethanol with alkaline $KMnO_4$ gives ethanoic acid which is called acetic acid (Refer to 6.10)

Oxidation of fermented liquors which contain about 12–15% ethyl alcohol, with air under the influence of bacteria *Mycoderma aceti* gives dilute acetic acid solution (containing about 4–7% acetic acid) called vinegar. Vinegar is widely used as a preservative of pickles and as table food.

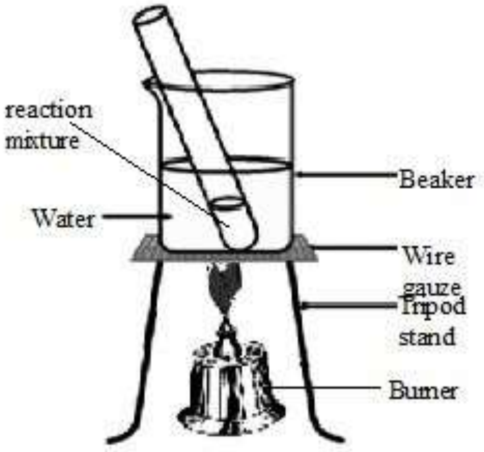


Fig. 6.11 Formation of ester

6.8. Properties of Ethanoic acid

Ethanoic acd is commonly called acetic acid. Acetic acid is a colourless, corrosive liquid with a sharp vinegar odour and sour taste. Its freezing point is 16.6 °C and boiling point is 118°C. Its vapours are suffocating and cause damage to lungs. Below its melting point acetic acid is converted into ice like solid hence it is called glacial acetic acid. It is miscible with water, ether and ethanol.

Activity 6.3

- Take about 2ml of ethanol and equal volume of glacial acetic acid in a test tube,.Add 2–3 drops of concetrated sulphuric acid.
- Heat the test tube on a water bath as shown in the Fig. 6.11. for about 5 minutes
- Pour the contents of the tube in about 20 ml cold water taken in a beaker and smell the mixture.
  What is your observation ?

Some of the reactions of acetic acid are

i) Reaction with metals, alkalis etc. Acetic acid reacts with metal, metal oxides, alkalis, carbonates and bicarbonates to form corresponding salts.

$2CH_3COOH + 2Na \longrightarrow 2CH_3\ COONa + H_2$

Acetic acid → Sodium acetate

$2CH_3COOH + Zn \longrightarrow (CH_3COO)_2\ Zn + H_2$

Zinc acetate

$$CH_3COOH + NaOH \longrightarrow CH_3COONa + H_2O$$

$$2CH_3COOH + Na_2CO_3 \longrightarrow 2CH_3COONa + H_2O + CO_2$$

$$CH_3COOH + Na\,HCO_3 \longrightarrow CH_3COONa + H_2O + CO_2$$

(ii) Reaction with alcohol (Esterification reaction)

Acetic acid reacts with ethyl alcohol in presence of an acid catalyst to form ester, ethyl acetate,

$$\underset{\text{(Ethanoic acid)}}{\underset{\text{Acetic acid}}{CH_3COOH}} + \underset{\text{(Ethanol)}}{\underset{\text{Ethyl alcohol}}{CH_3CH_2OH}} \xrightarrow{H^+} \underset{\text{(Ethyl ethanoate)}}{\underset{\text{Ethyl acetate}}{CH_3\,COOC_2H_5}} + H_2O$$

Ethyl acetate (ester) is a sweet-smelling liquid. It is used in making perfumes and as flavouring agents.

SUMMARY

- Graphite, Diamond and Fullerene are the allotropes of carbon.
- Carbon forms largest number of compounds.
- Graphite is a good conductor of electricity whereas diamond is a non-conductor.
- Carbon can bond with other carbon atoms. This property carbon is called catenation.
- Compounds of carbon and hydrogen are called hydrocarbons.
- Hydrocarbons with C–C single bonds only are called saturated hydrocarbons.
- Unsaturated hydrocarbons contain C–C multiple bonds.
- Isomers are compounds with the same molecular formula but having different structures.
- Any two successive members of a homologous series differ by a $-CH_2-$ unit.
- Hydrocarbons may be of straight chain or ring structures.
- Cycloalkanes have the same general formula C_nH_{2n} as alkenes.
- Butene shows position isomerism.
- Petroleum is the major source of hydrocarbons.
- Asphalt is used for road surfacing.

- Complete burning of organic compounds produce carbon dioxide as the end product.

- Saturated hydrocarbons burn with a clean flame.
- Unsaturated hydrocarbons usually burn with a sooty flame.
- Oxidation of ethyl alcohol with alkaline $KMnO_4$ or acidic $K_2Cr_2O_7$ produces acetic acid.
- Hydrogenation of vegetable oils give, vegetable ghee.
- Saturated hydrocarbons undergo substitution reaction with chlorine in presence of indirect sunlight.
- Ethanol produces depression of Central Nervous System.
- Ethanol on dehydration with conc H_2SO_4 produces ethene.
- Vinegar is dilute aqueous solution of acetic acid.
- Vapours of acetic acid cause damage to lungs.
- Acetic acid reacts with ethanol to give sweet smelling ester ethyl acetate.

EXERCISES

1. Why does graphite conduct electricity ?
2. Why is diamond so hard ?
3. What is CNG ? Give an use of it.
4. Draw the electron dot structure of ethane and count the number of covalent bond in a molecule of it.
5. Draw the structure of propane and find out the number of hydrogens which have similar properties.
6. Draw the structure of 2–methyl propane. How many types of carbon and hydrogens are there ?
7. What is the general formula of cycloalkanes ?
8. Can you write But–3–ene? If not why ?
9. Draw the structure of benzene molecule.
10. What are important fractions obtained from petroleum ? Mention their important uses.
11. Define functional groups. Write the functional groups of alcohol , aldehyde and carboxylic acids.
12. What is a word root ?

13. What are primary and secondary prefixes?

14. What are primary and secondary suffixes ?

15. How is oxidation different from combustion ?

16. Write equation for esterification reaction.

17. What is the formula of chloroform ?
Mention one use of chloroform.

★★★★★★

CHAPTER 7

MATERIALS OF COMMON USE

Can we live without common salt ? Common salt (or table salt) is chemically sodium chloride, NaCl. We have seen that it is the product obtained by the neutralisation of HCl with sodium hydroxide NaOH. We use common salt in our food. Can you say some more uses of common salt ?

7.1. Sources of common salt, NaCl

Sodium chloride or common salt, NaCl occurs abundantly in nature as rock salt. Rock salt forms cubic crystals which are colourless when pure, but often have yellow tinge due to the presence of impurities.

Sodium chloride also occurs in solution as brine containing upto about 25% NaCl. Brine also contains chlorides and sulphates of calcium and magnesium. There are brine wells at Ningel, Chandrakhong and Waikhong villages of Thoubal district in Manipur. However, the amount of sodium chloride in these brine solutions is quite low. A small quantity only of *Meitei Thum* is produced now. Sea water also contains about 2.5 % of common salt.

In India, common salt is generally obtained from the sea by evaporation of sea water in open fields or shallow iron pans utilising the heat of sun. Such salts are not of pure quality. Pure sodium chloride is prepared by passing hydrochloric acid gas into a saturated solution of impure salt. The pure salt is precipitated and obtained on filtration. It is then dried. Table salt sold in the market is now iodised to prevent iodine deficiency diseases.

Activity 7.1

☞ Take two cups of crushed ice.
☞ Insert a thermometer and note the temperature.
☞ Mix the ice thoroughly with a cup of common salt.
☞ Note the temperature again with the thermometer.
☞ What change do you observe ?
☞ What can be inferred from the observation ?

Ice and common salt mixture is often used as a freezing mixture and the temperature can even go down upto about – 20°C. Why do we use freezing mixture ?

Common salt gives a characteristic flavour to food. It is also used for preserving fish, meat and other foods like pickles etc. Sodium chloride is also used for salting out soap, regenerating water softeners and in the "Salt-glazing" of earthenwares.

7.1.1. Common salt – a raw material for various chemicals

Common salt, a basic component of food for us is also a key raw material for various materials (chemicals) of our use such as sodium hydroxide, sodium carbonate (washing soda), sodium, chlorine, bleaching powder, sodium bicarbonate, sodium sulphate, sodium hypochlorite and sodium chlorate. Let us see how some of these useful materials are obtained from sodium chloride.

Sodium hydroxide, NaOH

On passing electricity through brine (aqueous solution of sodium chloride) in specially designed electrolytic cell, sodium chloride is decomposed to produce sodium hydroxide, hydrogen and chlorine. All these products have industrial applications

$$2\ NaCl\ (aq) + 2\ H_2O\ (l) \xrightarrow{Electricity} 2\ NaOH\ (aq) + Cl_2\ (g) + H_2(g)$$

Hydrogen gas (H_2) is given off at the cathode and sodium hydroxide (NaOH) is formed arround the cathode while chlorine gas (Cl_2) is given off at the anode.

Hydrogen gas so obtained can be used as fuel and for the manufacture of ammonia for further use in the manufacture of fertilisers. Chlorine gas is used for sterilizing water (water treatment), in swimming pools; manufacture of PVC (poly vinylchloride) plastics, pesticides and CFC (Chlorofluro Carbon). CFC is used in refrigerators. However, its use as refrigerent is not encouraged as it causes depletion of ozone layer.

Can you think of any other use of hydrogen and chlorine ?

Hydrogen and chlorine can be made to combine to form hydrogen chloride (HCl).

$$H_2(g) + Cl_2(g) \longrightarrow 2HCl\ (g)$$

Sodium hydroxide is used for making soaps and detergents, paper and artificial fibre etc.

Sodium hydroxide can also combine with chlorine to produce sodium hypochlorite which is a bleaching agent for paper and fabrics.

$$2\,NaOH(aq) + Cl_2\,(g) \longrightarrow NaCl\,(aq) + NaOCl\,(aq) + H_2l\,(O)$$

sodium hypochlorite solution.

Sodium hypochlorite is the basic ingredient of all commercial bleaching solutions sold in the market. It is used to remove ink stains and for bleaching clothes etc. Such bleaching agents have a strong smell of chlorine.

The bleaching action of such a solution is due to the chlorine liberated in solution. Chlorine can also kill germs present in water. These properties are due to the fact that chlorine acts as an oxidising agent.

What is the white powdery substance used to disinfect drains in hospitals ? It is also used to kill germs in drinking water.

7.1.2. Bleaching powder (chloride of lime), Ca (OCl) Cl

The white substance which have a peculiar smell, used as a disinfectant in public places like hospitals, drains, market places etc. is bleaching powder . The chlorine gas produced by the electrolysis of sodium chloride can be used for the manufacture of bleaching powder. Chemically, it is calcium chloro-hypochlorite,and is obtained commercially by passing chlorine gas over slaked lime at about 35–45°C.

$$\underset{\text{slaked lime}}{Ca(OH)_2 + Cl_2} \longrightarrow \underset{\text{Bleaching powder}}{Ca\,(OCl)\,Cl + H_2O}$$

Formula of bleaching powder is usually written as above or simply $CaOCl_2$ as it explains all its reactions and agrees with available chlorine content. However, it is a mixture of calcium hypochlorite and basic calcium chlorides, $Ca(OCl)_2 \cdot 4H_2O$ and $CaCl_2 \cdot Ca(OH)_2 \cdot H_2O$

Bleaching powder has strong smell of chlorine. It absorbs moisture from the air. It does not dissolve completely in water but, always leaves a white residue of lime. The aqueous suspension, however, contains calcium hypochlorite $Ca\,(OCl)_2$ in the solution.

Carbon dioxide and moisture of the atmosphere decompose bleaching powder liberating chlorine.

$$Ca(OCl)Cl + CO_2 \longrightarrow Ca\,CO_3 + Cl_2$$
$$Ca(OCl)\,Cl + H_2O \longrightarrow Ca(OH)_2 + Cl_2$$

Let us answer the following.

1. Why does bleaching powder act as a disinfectant ?
2. Why does bleaching powder act as a bleaching agent ?

Uses: Bleaching powder is used

- **(i)** as a disinfectant,
- **(ii)** for sterilising water to make water free from disease causing germs and bacteria,
- **(iii)** for bleaching pulp in paper industry, linen and cotton in textile industry,
- **(iv)** as an oxidising agent in chemical industries.

Commercial samples of bleaching powder contains 35–37 % available chlorine. Available chlorine is the amount of chlorine set free by the sample of bleaching powder when decomposed. This free chlorine is responsible for the oxidising and bleaching properties of bleaching powder.

7.1.3. Baking Soda (sodium hydrogen carbonate), $NaHCO_3$

Baking soda is usually kept in all kitchens. It is used in making *Pakoras* and special Manipuri dish *Uti*. Chemically it is sodium hydrogen carbonate $NaHCO_3$. It is also known as *Soda bicarb*

When an excess of carbon dioxide is bubbled through a nearly saturated solution of common salt containing ammonia, ammonium hydrogen carbonate is formed. This will react with more of sodium chloride to form sparingly soluble sodium hydrogen carbonate and ammonium chloride.

$$NaCl + H_2O + CO_2 + NH_3 \longrightarrow \underset{\text{Sodium hydrogen carbonate}}{NaHCO_3} + NH_4Cl$$

Sodium hydrogen carbonate is separated by filtration. It is then washed with cold water and dried in air.

Sodium hydrogen carbonate is a white solid. It is sparingly soluble in water. The aqueous solution of sodium hydrogen carbonate is slightly hydrolysed and the solution is alkaline in reaction.

$NaHCO_3 + H_2O$ □ $NaOH + H_2CO_3$

Uses

(i) Sodium hydrogen carbonate is used in medicine to reduce acidity in the stomach.

(ii) A mixture of $NaHCO_3$ and mild edible acid like tartaric acid is known as baking powder. When heated or mixed with water the following reaction takes place

$2NaHCO_3 \xrightarrow{\text{heat}} Na_2CO_3 + H_2O + CO_2$

$NaHCO_3 + H^+$ (from tartaric acid) $\xrightarrow{\text{water}}$ Na-tartarate (salt) $+ H_2O + CO_2$

Carbon dioxide so produced during the reaction causes bread or cake to swell making them soft and fluffy.

(iii) sodium hydrogen carbonate is used in soda -acid fire extinguishers.

7.1.4. Washing Soda (Sodium carbonate) $Na_2CO_3 . 10H_2O$)

Washing soda is another very important chemical obtained from sodium chloride. Sodium carbonate is also found naturally in lakes of Magadi in East Africa, and in lakes of lower Egypt. The Magadi deposits, containing about 200 million tons of native soda, is an important source of soda.

Formerly sodiun carbonate was obtained from the ash of plants growing on the sea-shore. This method is no longer used nowadays.

A French chemist Nicholas Leblanc prepared washing soda directly from sodium chloride in the following steps:

(i) Sodium chloride is treated with concentrated sulphuric acid to form sodium sulphate which is called **salt-cake**.

$$2NaCl + H_2SO_4 == \underset{\text{salt-cake}}{Na_2SO_4} + 2HCl$$

The HCl gas liberated is absorbed in water giving commercial hydrochloric acid.

(ii) The salt-cake is grounded, and mixed with its own weight of limestone and half its weight of coal and strongly heated to 800–1000°C . The salt cake is reduced by coke to sodium sulphide, which then reacts with calcium carbonate to form sodium carbonate :

$Na_2SO_4 + 4C == Na_2S + 4CO$
$Na_2S + CaCO_3 == Na_2CO_3 + CaS$

The product called black-ash, contains about 45% sodium carbonate. The Leblanc process is no longer used nowadays but salt-cake is prepared for using in glass industry.

The ammonia soda or Solvay Process

Sodium carbonate is prepared by this process. In the section 7.1.3, you have seen that $NaHCO_3$ is prepared from brine solution by bubbling CO_2 into it after being saturated with ammonia (NH_3). From the solution sparingly soluble $NaHCO_3$ is separated by filtration and dried. $NaHCO_3$ is then calcined (heated at high temperature) to produce Na_2CO_3:

$$2\ NaHCO_3 == Na_2CO_3 + H_2O + CO_2$$

Carbon dioxide so liberated is used for further reaction with brine.

Washing soda has the formula $Na_2CO_3\ .10\ H_2O$. It is prepared by crystallising the solution below 32°C. It is commercially known as **washing soda.**

Anhydrous Na_2CO_3 is called soda ash and it is an amorphous powder. It absorbs moisture form air, forming sodium carbonate monohydrate, $Na_2CO_3\ .H_2O$ which is known **as crystal carbonate.**

Uses of washing soda

- **(i)** Sodium carbonate is used as cleansing agent for domestic purposes and that is why it is called washing soda.
- **(ii)** It is used in the manufacture of glass, soap, paper and borax.
- **(iii)** It is used for making hard water soft.

You have seen amorphous sodium carbonate and crystalline sodium carbonate. Why is Na_2CO_3 amorphous while $Na_2CO_3\ .H_2O$ and $Na_2CO_3.10\ H_2O$ are crystalline ?
Let us find out the reason.

7.1.5. Water of Crystallisation

Na_2CO_3 is an amorphous white solid while $Na_2CO_3 \cdot H_2O$ and $Na_2CO_3 \cdot 10\ H_2O$ are crystalline solids. It is evident that H_2O molecules play an important role in the crystal structures of solids. Let us see this fact.

Activity 7.2

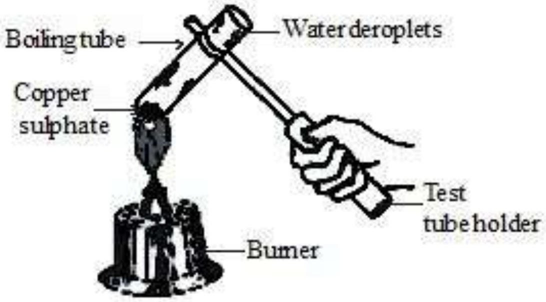


Fig. 7.1. To show that water of crystallisation can be lost on heating.

- Take a dry boiling tube and place in it some crystals of copper sulphate.
- Heat the copper sulphate crystals in the tube over a spirit lamp.
- What is the colour of copper sulphate now ?
- What do you observe near the mouth of the boiling tube ? (Droplets of water are collected near the mouth of the tube)
- Can you explain the collection of water drops?
- Cool the tube, then add 1–2 drops of water. What happened to the colour of the crystal?

Blue colour is restored again.

Crystalline copper sulphate is blue and is called blue vitriol. It has the formula $CuSO_4.5\ H_2O$. These five molecules of water present in copper sulphate is called its **water of crystallisation.** When the blue crystals are heated, these water molecules are lost and the blue clour becomes colourless. When it is treated with water the lost water of crystallisation is regained and it is again blue in colour. Many a crystal contain water of crystallisation.

Examples are : Green vitriol $FeSO_4 \cdot 7H_2O$, (Ferrous sulphate), Zinc sulphate $ZnSO_4 \cdot 7\ H_2O$ (white vitriol), Gypsum $CaSO_4 \cdot 2\ H_2O$.

All crystals need not have water of crystallisation. For example NaCl, NH_4Cl, KCl, KI, sugar etc.

Gypsum is naturally occuring calcium sulphate, $CaSO_4 \cdot 2\ H_2O \cdot$ Do you know the white powder used for making casts and used in surgical bandages?

7.1.6. Plaster of Paris, $2CaSO_4 \cdot H_2O$.

When gypsum ($CaSO_4 \cdot 2H_2O$) is heated to about 120°–130°C it loses water molecule to form calcium sulphate hemihydrate $CaSO_4 \cdot \frac{1}{2} H_2O$. It can also be written as $(CaSO)_2 \cdot H_2O$ i.e. two calcium sulphate molecules share one water molecule. This substance is called Plaster of Paris.

$$\underset{\text{gypsum}}{2CaSO_4 . 2H_2O} \xrightarrow{\text{heat}} \underset{\text{plaster of paris}}{(CaSO_4)_2 \cdot H_2O} + 3H_2O$$

Uses : Plaster of paris is a white powder. It can be converted back to gypsum when treated with water. The newly formed product, on rehydration, sets into a hard solid. Plaster of paris is, therefore, used for making casts in moulds, and as a plaster for supporting fractured bones in the right position. It is also used for wall plaster.

7.2. Soaps and Detergents

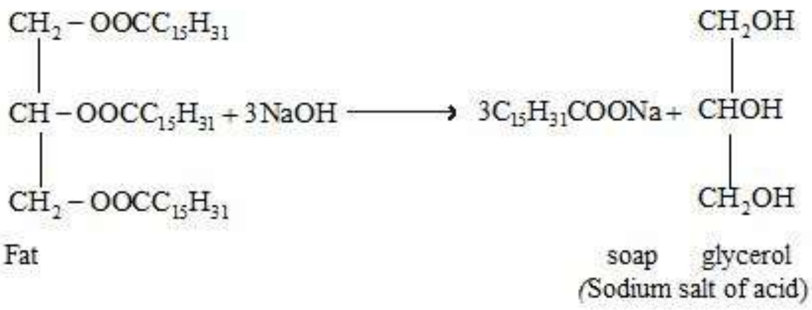
Soap has been used as a cleansing agent for more than two thousand years. It is used for washing clothes. We use toilet soaps in the form of solid cakes or in the solution form. Can you think of any other material which can be used for washing clothes ?

Activity 7.3

- Collect some ash obtained by burning straw or dried plantain leaves or dried stems of pea, cowpea etc, in a beaker.
- Add sufficient water to leach it and stir. Allow it to settle for some time.
- Filter the mixture.
- Soak your dirty handkerchief in the filtrate for some time and wash it.
- Does the hadkerchief become clean ?

The ash is rich in carbonates of potassium and sodium. The solution is actually alkaline and therefore, it can remove oil and grease from dirty clothes and linens.

Soap is made from oil or fat which are esters of higher fatty acids and glycerol. When fats and oils are heated with a solution of sodium hydroxide, they are decomposed (hydrolysed) to form sodium salt of the fatty acid and glycerol. An example is given below:



$$\begin{array}{l} CH_2-OOCC_{15}H_{31} \\ | \\ CH-OOCC_{15}H_{31} + 3NaOH \longrightarrow 3C_{15}H_{31}COONa + \begin{array}{c} CH_2OH \\ | \\ CHOH \\ | \\ CH_2OH \end{array} \\ | \\ CH_2-OOCC_{15}H_{31} \end{array}$$

Or $(C_5H_3\ COO)_3\ C_3H_5 + 3Na\ OH \longrightarrow 3C_{15}H_{31}COONa + C_3H_5(OH)_3$

The soap is precipitation by the addition common salt. It is called salting out of soap. The soap is separated from the mixture by filtration and the filtrate is used for the recovery of glycerol. The soap is then mixed with colours, perfumes etc. and cast into various shapes for use.

Activity 7.4.

- Take two test tubes and label them A and B.
- In each of them, put 10mL of water and two drops of cooking oil
- To the tube A add a few drops of soap or detergent solution.
- Shake both the test tubes vigourouslx y closing the month of the tube with your thumb. After some time leave the tubes undisturbed.
- In which test tube the oil layer appears again? Why ?

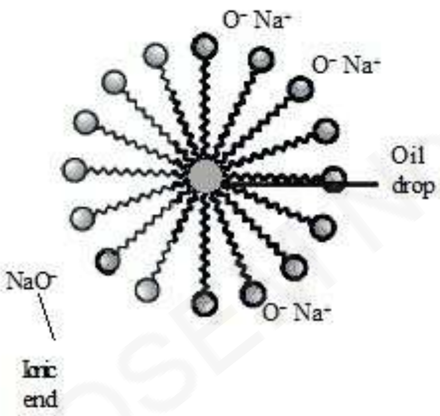


Fig. 7.2. Formation of Micelles

This activity demonstrates the cleansing action of soap or detergent. Dirts are oily in nature and oil is not soluble in water. The molecules of soap which are sodium salt of long chain fatty acid have two ends. One end which is water soluble is ionic and the other end is non-ionic hydrocarbon part which is fat or oil soluble. The soap molecules, thus form structures called micelles (Fig. 7.2) where the hydrocarbon end lies inside the oil drop and the ionic-end lies in the aqueous region. When the dirty cloth is soaked in soap solution and then beaten, the oily drop is converted into innumerable pieces on being pulled out by the soap melecules. Thus the dirt (oil) is passed into water and the cloth is washed clean.

Now a days detergents have replaced soap for washing cloths. Detergents are of many types. They are generally ammonium or sodium sulphonate salts of long chain hydrocarbons which are generally obtained from petroleum. A typical example is $C_{17}H_{35}–OSO_3Na$. The cleansing action of detergents is similar to that of soaps. One advantage of detergent is that it can form copious lather even with hard water and therefore, washes clothes more effectively. Soap forms insoluble scum with hard water which contains calcium (Ca^{2+}) and magnesium (Mg^{2+}) ions. This is due to the formation of calcium and magnesium salts of fatty acids. However, one disadvantage of synthetic detergents is that , they are not biodegrable. Therefore, they may pollute soily river, take and other water bodies. Soaps are biodegradable.

Let us answer these

1. Why are beating, scrubbing or agitation required while washing clothes ?
2. Device a simple method for detection of hard or soft water.
3. Can detergent be used to detect hard water ?

SUMMARY

- We use common salt for the preparation of food daily.
- Rock salt is natural common salt.
- Brine wells are available in Manipur.
- A mixture of crushed ice and common salt is called freezing mixture.
- NaCl is a starting material for the manufacture of NaOH, Na_2CO_3, $NaHCO_3$, and Na_2SO_4.
- Cl_2, H_2 and NaOH are the products obtained from the electrolysis of NaCl.
- Bleaching powder is prepared by passing chlorine over slaked lime.
- Bleaching powder is a strong antiseptic and bleaching agent.
- The bleaching and antiseptic properties of bleaching powder are due to Cl_2 liberated on hydrolysis

- Baking powder is $NaHCO_3$ mixed with some edible organic acid.
- Washing soda is prepared from NaCl solution.
- Washing powder contains 10 molecules of water of crystallisation.
- Water of crystallisation may be removed by heat
- Anhydrous salts may regain colour by absorbing water.
- Plaster of Paris is prepared from gypsum.
- Soaps and detergents are used for cleansing.
- Soap is sodium or potasium salt of long chain fatty acids.
- Detergents are ammonium or sodium sulphonates of long chain hydrocarbons.

EXERCISES

1. How is pure sodium chloride obtained from the impure salt ?
2. What is the product obtained by reacting NaOH with HCl ?
3. How is freezing mixture made ?
4. Give important uses of sodium chloride.
5. Name the products obtained when aqueous sodium chloride is electrolysed.
6. What are the main uses of chlorine ?
7. What is the use of sodium hypochlorite ?
8. How is bleaching powder prepared ?
9. How does bleaching powder decompose in moist air ?
10. Give the important use of bleaching powder
11. Give a method for obtaining sodium hydrogen carbonate.
12. What happens when $NaHCO_3$ reacts with dil HCl ?
13. What is salt cake ?
14. Describe Solvay process.
15. What is the colour of anhydrous copper sulphate ?
16. What are green vitriol and blue vitriol ?
17. How is plaster of Paris made ?
18. What is the difference between soap and detergents ?

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