By- Shailendra Shukla Sir

Chapter 3 Metals and Nonmetals

Metals

Those elements which have tendency to lose one or more electrons to form positive ions are called metals. For example, sodium (Na), potassium (K), magnesium (Mg), aluminium (Al) etc.

Physical Properties of Metals

(1) Metals have a shining surface. This property of metals is called metallic lustre.

Activity 3.1

- Take samples of iron, copper, aluminium and Magnesium.
- Note the appearance of each sample.
- Clean the surface of each sample by rubbing them with sand paper and note their appearance again.
- What do you observe ?

(2) Metals are usually hard and the hardness of metals varies from metal to metal.

Activity 3.2

- Take small pieces of iron, copper, aluminium, and magnesium.
- Try to cut these metals with a sharp knife and 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 and try to cut it with a knife.
- What do you observe ?

(3) Metals are malleable and ductile. Due to their malleability and ductility, metals can be given different shapes as per our needs.

Malleability-Some metals can be beaten into thin sheets. This property is called malleability. Gold and silver are the most malleable metals of nature ?

Activity 3.3

- Take pieces of iron, zinc, lead 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.

Ductility- Some of the metals can be drawn into thin wires. This ability of metals is called ductility. Gold is the most ductile metal. A wire of about 2 km length can be drawn from one gram of gold.

Activity 3.4

- List the metals whose wires you have seen in daily life.

(4) Metals are good conductors of heat and electricity.
(5) Metals have high melting points.
Can you name some metals that are used for making cooking vessels ? Do you know why these metals are used for making vessels ?

Activity 3.5

- Take an aluminium or copper wire. Clamp this wire on a stand as shown in Fig.
- Fix a pin to the free end of the wire using wax.
- 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 ?
- Note your observations. Does the metal wire melt ?

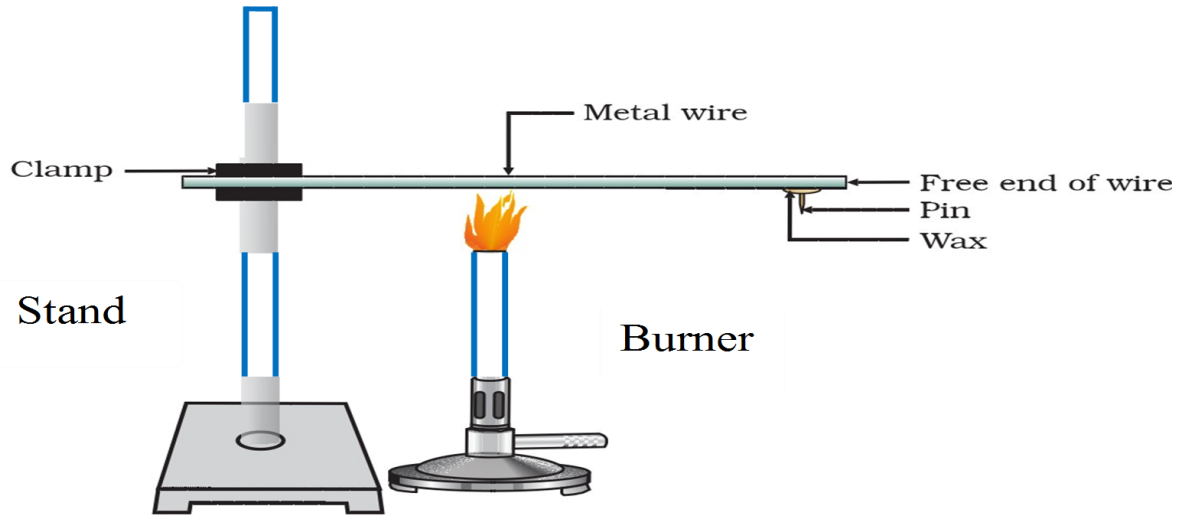


Fig.- Metals are good conductor of heat

This activity shows that, metals are good conductors of heat and they have high melting points. Silver and copper are the best conductors of heat while, lead and mercury are comparatively poor conductors of heat.

Activity 3.6

- Set up an electric circuit as shown in Fig.
- Place the metal to be tested in the circuit between terminals A and B as shown.
- Does the bulb glow ?
- What does this indicate ?

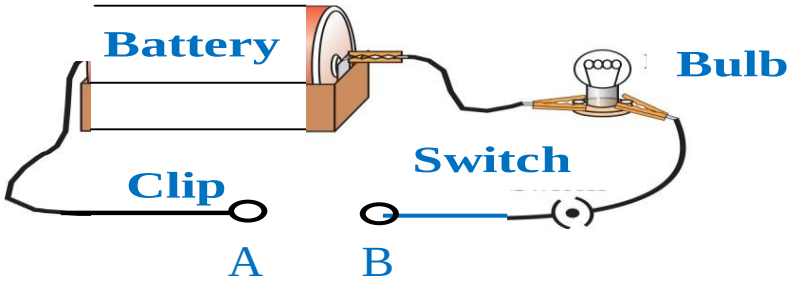


Fig.- Metals are good conductor of electricity

Electric wires remain coated with polyvinylchloride (PVC) or a rubber-like material. Why are the electric wires coated with such substances ?

(5) Metals are sonorous i.e. they produce a sound on striking to the hard surfaces.

What happens when metals strike a hard surface ? Do they produce a sound ? The metals that produce a sound on striking to a hard surface are said to be sonorous. Can you now answer why school bells are made of metals ?

Chemical Properties of Metals

(1) Reaction of Metals with Oxygen (Air)

- **What happens when Metals are burnt in Air ?**
- **Do all metals react with air in the same manner ?**
- Almost all the metals combine with oxygen to form metal oxides.

$$\text{Metal} + \text{Oxygen} \rightarrow \text{Metal oxide}$$

- Na and K react vigorously with air and catch fire. Hence, to avoid contact with air, they are kept immersed in kerosene oil.

$$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$$
$$4K(s) + O_2(g) \rightarrow 2K_2O(s)$$

- Magnesium burns in air with a dazzling white flame.
$$2Mg(s) + O_{2(}g) \rightarrow 2MgO(s)$$
- Al, Zn, Pb etc. react slowly with air and form a protective layer over their surface, which protects them from corrosion.
$$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$$
$$2Zn(s) + O_2(g) \rightarrow 2ZnO(s)$$
- Iron does not burn on heating in air, but iron filings burn vigorously when sprinkled in the flame of the burner to form iron oxides.
$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$
$$3Fe(s) + 2O_2(g) \rightarrow Fe_3O_4(s)$$
- Copper does not burn, but the hot metal is coated with a black coloured layer of copper (II) oxide.
$$2Cu(s) + O_2(g) \rightarrow 2CuO\ (s)$$
(Copper) Copper (II) oxide (Black)
- Silver, platinum and gold being inert metals, do not react with air at all even at a very high temperature.
- Most of the metal oxides are basic in nature. They react with acids like HCl to form salt and water. For example,
$$Na_2O(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l)$$
$$CuO(s) + 2HCl(aq) \rightarrow CuCl_2(aq) + H_2O(l)$$
Copper (II) chloride (Blue green)
- However, some metal oxides, such as aluminium oxide (Al_2O_3), zinc oxide (ZnO), PbO, SnO etc. show both acidic as well as basic behaviour and react with both acids and bases to form salts and water. Such metal oxides are called amphoteric oxides. Aluminium oxide reacts with acids and bases in following manner -
$$Al_2O_3 + 6HC1 \rightarrow 2A1Cl_3 + 3H_2O$$
$$Al_2O_3 + 2NaOH \rightarrow 2NaAlO_2 + H_2O$$
(Sodium aluminate)
- Similarly, ZnO reacts with acids and bases as under-

$$ZnO(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2O(l)$$
$$ZnO(s) + 2NaOH(aq) \rightarrow Na_2ZnO_2(aq) + H_2O(l)$$
(Sodium zincate)

- Most of the metal oxides are insoluble in water, but some of them dissolve in water to form alkalis. For example, Sodium oxide and potassium oxide dissolve in water to produce alkalis as follows -

$$Na_2O(s) + H_2O(1) \rightarrow 2NaOH(aq)$$
$$K_2O(s) + H_2O(1) \rightarrow 2KOH(aq)$$

Activity 3. 9

- **Caution-** This activity needs the teacher's assistance. It would be better if students wear eye protection during the activity.
- Collect the samples of following metals - aluminium, copper, iron, lead, magnesium, zinc and sodium.
- Hold any of these samples with a pair of tongs and try burning over a flame.
- Repeat with the other metal samples.
- Collect the product if formed.
- Let the products and the metal surface cool down. Which metals burn easily ?
- What flame colour did you observe when the metal burnt ?
- How does the metal surface appear after burning ?
- Arrange the metals in the decreasing order of their reactivity towards oxygen.
- Are the products soluble in water ?

This activity reveal that-

- Sodium is the most reactive of the samples of metals taken here.
- The reaction of magnesium is less vigorous. It implies that, it is not as reactive as sodium.
- Aluminium is third most reactive metal amongst the sample taken.
- Here, burning in oxygen does not help us to decide about the relative reactivity of zinc, iron, copper and lead.

- **Anodising of Aluminium-** Aluminium develops a thin oxide layer over its surface, when exposed to air. This aluminium oxide coating makes it resistant to further corrosion. The resistance can be improved

further by making the oxide layer thicker.

- Anodising is a process of forming a thick aluminium oxide layer over the surface of aluminium.
- During anodising, a clean aluminium article is made the anode and electrolysis is carried out using dilute sulphuric acid as electrolyte.
- Here, the oxygen gas evolved at the anode, reacts with aluminium to make a thicker protective oxide layer over the surface of aluminium article.
- This oxide layer formed over the surface of aluminium article can be dyed easily to give it an attractive finish.

(2) Reaction with Water

- **What happens when Metals react with Water ?**
- Some metals react with water to form their oxides and release hydrogen gas. If the metal oxides formed is soluble in water, it dissolves further in water to form metal hydroxide.

  Metal + Water → Metal oxide + Hydrogen

  Metal oxide + Water → Metal hydroxide
- Metals like sodium and potassium react violently with cold water. Here, the reaction is so violent and exothermic that, the hydrogen evolved catches fire immediately.

  $2Na(s) + 2H_2O(1) \rightarrow 2NaOH(aq) + H_2(g) + Heat$

  $2K(s) + 2H_2O(1) \rightarrow 2KOH(aq) + H_2(g) + Heat$
- The reaction of calcium with cold water is less violent. The heat evolved is not sufficient for the hydrogen to catch fire.

  $Ca(s) + 2H_2O(1) \rightarrow Ca(OH)_2(aq) + H_2(g) + Heat$

  Here, calcium starts floating because the bubbles of hydrogen gas formed stick to the surface of the metal.
- Magnesium does not react with cold water. It reacts with hot water to form magnesium hydroxide and hydrogen. It also starts floating due to the sticking of bubbles of hydrogen gas to its surface.

  $Mg(s) + 2H_2O(1) \rightarrow Mg(OH)_2(aq) + H_2(g) + Heat$
- Metals like aluminium, iron and zinc do not react either with cold or hot water. But they react with steam to form the metal oxide and hydrogen.

  $2Al(s) + 3H_2O(g) \rightarrow Al_2O_3(s) + 3H_2(g)$

  $3Fe(s) + 4H_2O(g) \rightarrow Fe_3O_4(s) + 4H_2(g)$

$$Zn(s) + H_2O(g) \rightarrow ZnO(s) + H_2(g)$$

- Metals such as lead, copper, silver and gold do not react with water at all.

- **Activity 3.10**
- **Caution-** This Activity needs the teacher's assistance.
- Collect the samples of following metals - aluminium, copper, iron, lead, magnesium, zinc and sodium.
- Put small pieces of the samples separately in beakers half-filled with cold water.
- Which metals reacted with cold water ? Arrange these metals in the increasing order of their reactivity with cold water.
- Did any metal produce fire on reacting with water ?
- Does any metal start floating after some time ?
- Put the metals that did not reacted with cold water in beakers half-filled with hot water.
- For the metals that did not reacted with hot water, arrange the apparatus as shown in Figure and observe their reaction with steam.

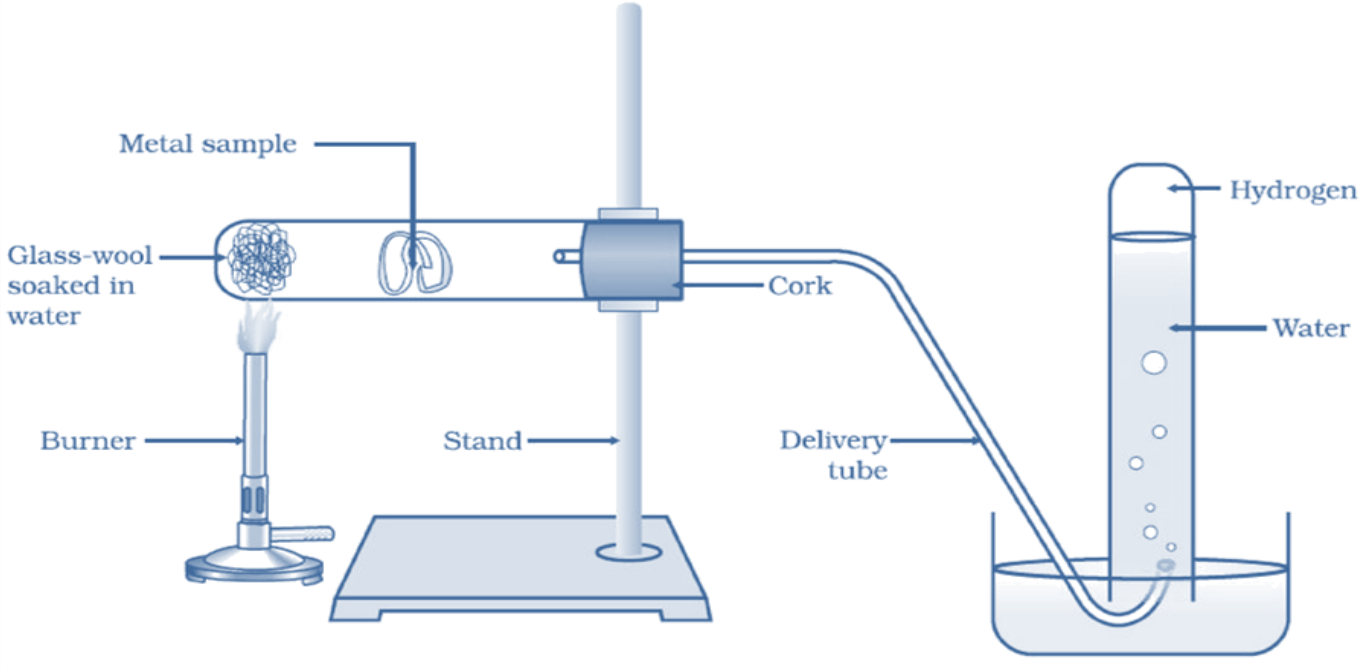


Fig.- Reaction of Metals with Steam

- Which metals did not react even with steam ?
- Arrange the metals in the decreasing order of reactivity with water.

(3) Reaction of Metals with Acids

- **What happens when Metals react with Acids ?**
- Most of the metals react with dilute acids to form metal salts and release hydrogen gas.

Metal + Dilute acid → Metal Salt + Hydrogen

$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

$Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$

- Do all metals react with acids in the same manner ?
- **Activity 3. 11**
- Collect the samples of following metals - aluminium, copper, iron, lead, magnesium and zinc.
- If the samples are tarnished, rub them to clean with sand paper.
- Put the samples separately in test tubes containing dilute hydrochloric acid.
- Suspend thermometers in the test tubes, so that their bulbs are dipped in the acid.
- Observe the rate of formation of bubbles carefully.
- Which metals reacted vigorously with dilute hydrochloric acid ?
- With which metal did you record the highest temperature ?
- Arrange the metals in the decreasing order of reactivity with dilute acids.
- Write equations for the reactions of magnesium, aluminium, zinc and iron with dilute hydrochloric acid.

Here, the rate of formation of bubbles is the fastest in the case of magnesium. The reaction is also the most exothermic in this case. The reactivity of these metals with HCl decreases in the following order-

Mg > Al > Zn > Fe

In the case of copper, no bubbles are seen and the temperature also remaines unchanged. This shows that copper does not react with dilute HCl.

Hydrogen gas does not evolve 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 oxides of nitrogen such as N_2O, NO, NO_2 etc.

However, magnesium (Mg) and manganese (Mn) react with very dilute HNO_3 to evolve H_2 gas.

- **Aqua regia** (Latin for 'royal water') is a freshly prepared mixture of concentrated hydrochloric acid and concentrated nitric

acid in the ratio of 3 : 1. It can dissolve gold, even though neither of these acids can do so alone. Aqua regia is a highly corrosive, fuming liquid. It is one of the few reagents that is able to dissolve gold and platinum.

(4) Reaction of Metals with Salts of Other Metals-

- **How do Metals react with Solutions of other Metal Salts ?**

Activity 3.12

- Take a clean wire of copper and an iron nail.
- Put the copper wire in a solution of iron sulphate and the iron nail in a solution of copper sulphate taken in test tubes.
- Record your observations after 20 minutes.
- In which test tube did you find that a reaction has occurred ?
- On what basis can you say that a reaction has actually taken place ?
- Can you correlate your observations with the activities 3.9, 3.10 and 3.11 ?
- Write a balanced chemical equation for the reaction that has taken place.
- Name the type of reaction.
- Which metal, copper or iron, is more reactive according to your observations ?

A more reactive metal can displace less reactive metals from their compounds in solution or molten form.

In fact, all the metals are not equally reactive. We checked the reactivity of various metals with oxygen, water and acids. But all metals do not react with these reagents. So, we were not able to put all the metal samples we had collected in decreasing order of their reactivity. Displacement reactions give a better evidence about the reactivity of metals. It is simple and easy that, if metal A displaces metal B from its solution, it implies that metal A is more reactive than metal B.

Reactivity Series

If different metals are arranged in decreasing the order of their activities a series is obtained which is called the reactivity series. A reactivity series consisting of some most common metals is as under-

Symbol	Element
K	Potassium (Highly active metal)
Ba	Barium
Ca	Calcium
Na	Sodium
Mg	Magnesium
Al	Aluminium
Zn	Zinc
Fe	Iron
Ni	Nickel
Sn	Tin
Pb	Lead
H	Hydrogen
Cu	Copper
Hg	Mercury
Ag	Silver
Au	Gold
Pt	Platinum

Non-Metals

Those elements which have tendency to gain electrons to form negative ions are called non-metals. For example, Carbon (C), Sulphur (S), Phosphorous (P), Silicon (Si), Hydrogen (H), Oxygen (O), Nitrogen (N), Chlorine (Cl), Bromine (Br), Neon (Ne) and Argon (Ar) etc.

Physical properties of non-metals

- **Hardness- Non-metals are generally soft, but the diamond is an exception which is the hardest substance of nature.**
- **State- Non-metals may be solid, liquid or gas.**
- **Lustre- Non-metals have a dull appearance. Diamond and iodine are the exceptions which are lustrous and therefore, they have a shining surface.**
- **Sonority- Non-metals are not sonorous, i.e., they do not produce sound on striking with other objects.**
- **Conduction- Non-metals are a bad conductor of heat and electricity. Graphite which is an allotrope of carbon is an exception. It is a good conductor of electricity.**
- **Malleability and ductility- Non-metals are not malleable and ductile but are brittle in nature. They get broken on biting with hammer.**
- **Melting and boiling point- Non-metals generally have low melting and boiling points. Diamond is an exception which has a very high melting point and boiling point.**
- **Density- Most of the non-metals have low density.**

Chemical properties of Non-metals

1. **Reaction with Oxygen- Non-metals react with oxygen to form their oxides.**

Non-metal + Oxygen → Non-metallic oxide

For example,

When carbon is burnt in excess of air, carbon dioxide is formed with release of a large amount of heat.

$$C + O_2 \rightarrow CO_2 + Heat$$

Carbon Oxygen Carbon dioxide

When carbon is burnt in an insufficient supply of air, it forms carbon monoxide.

$$2C + O_2 \rightarrow 2CO + Heat$$

Carbon Oxygen Carbon monoxide

Carbon monoxide is a toxic substance. If it is inhaled in excess, it may be fatal for the living beings.

Sulphur reacts with oxygen to form sulphur dioxide It catches fire when exposed to air.

$$S + O_2 \rightarrow SO_2 + \text{Heat}$$

Sulphur Oxygen Sulphur dioxide

When hydrogen reacts with oxygen water is formed.

$$2H_2 + O_2 \rightarrow 2H_2O$$

Hydrogen Oxygen Water

Non-metallic oxides are acidic in nature i,e, they dissolve in water to form acids. For example,

Carbon dioxide dissolve in water to form carbonic acid.

$$CO_2 + H_2O \rightarrow H_2CO_3$$

Carbonic acid

Sulphur dioxide dissolve in water to form sulphurous acid.

$$SO_2 + H_2O \rightarrow H_2SO_3$$

Sulphurous acid

Sulphur dioxide reacts with oxygen to form sulphur trioxide which dissolve in water to form sulphuric acid.

$$2SO_2 + O_2 \rightarrow 2SO_3$$

Sulphur trioxide

$$SO_3 + H_2O \rightarrow H_2SO_4$$

Sulphuric acid

Thus, the aqueous solution of non-metal oxides turns blue litmus red.

2. Reaction with Chlorine-

Non-metal react with chlorine to form their respective chlorides.

Non-metal + Chlorine → Non-metal chloride

For example,

Hydrogen reacts with chlorine to form hydrogen chloride.

$$H_2 + Cl_2 \rightarrow 2HCl$$

Hydrogen chloride

Similarly, phosphorous reacts with chlorine to form phosphorous trichloride.

$$P_4 + 6Cl_2 \rightarrow 4PCl_3$$

Phosphorus trichloride

3. Reaction with Hydrogen- **Non-metals reactive with hydrogen to form covalent hydrides.**

Non-metal + Hydrogen → Covalent Hydride

For example,

Sulphur combines with hydrogen to form Hydrogen sulphide.

$$S + H_2 \rightarrow H_2S$$

Hydrogen sulphide

Similarly, nitrogen combines with hydrogen in presence of iron to form ammonia.

$$N_2 + 3H_2 \rightarrow 2NH_3$$

Ammonia

- **Non-metals do not react with water (or steam) to evolve Hydrogen gas.**
- **Non-metals in general, do not react with dilute acids.**

4. Reaction of Metals with Non-metals

Metals combine with non-metals to form ionic compounds. The formation of the ionic compounds involve the formation of electrovalent bonds.

Chemical Bonding

Valency

The term 'valency' has been derived from the Latin word 'valentia' meaning 'combining power' or 'combining capacity'. Thus,

"The combining capacity of an atom is called its valency."

Valency of an element depends upon its electronic configuration and especially upon the number of electrons in its outermost shell (valence shell) and is determined using following rules-

(1) If an atom contains 8 electrons in its outermost shell i.e. it has complete octet, then its valency is zero. For example, inert gases like Ne, Ar, Kr etc.

	K L	
Ne (10) –	2, 8	Valency = 0
	K L M	
Ar (18) –	2, 8, 8	Valency = 0

Helium is its exception, where first shell is outermost shell, which can accommodate only a maximum of two electrons. Hence, for He duplet configuration is stable configuration and its valency is zero.

(2) If only outermost shell of an atom is incompletely filled, then its valency depends upon the number of electrons in its outermost shell. Here,

(a) If the outermost shell contains 1 - 4 electrons, then its valency is equal to the number of electrons in the outermost shell. For example,

	K	L	M	
Li (3) -	2,	1		Valency = 1
Be (4) -	2,	2		Valency = 2
B (5) -	2,	3		Valency = 3
C (6) -	2,	4		Valency = 4
Na (11) -	2,	8,	1	Valency = 1
Mg (12) -	2,	8,	2	Valency = 2
Al (13) -	2,	8,	3	Valency = 3

(b) If it contains 4 - 7 electrons, then its valency is equal to (8 - n), where n is number of electrons in the outermost shell of the atom. For example,

	K	L	
N (7) -	2,	5	Valency = (8-5) = 3
O (8) -	2,	6	Valency = (8-6) = 2
F (9) –	2,	7	Valency = (8-7) = 1

(3) If outer two shells of an atom are incompletely filled, then valency of the atom depends upon the number of electrons in both of these two shells. These atoms show variable valencies. For example, transition metals.

Electronic theory of valency
or
Cause of chemical combination
(Octet Theory)

Electronic theory of valency was for the first time proposed by Kossel and Lewis. The main postulates of this theory are as under-

(1) Those elements which contain 8 electrons in their outermost shell i.e. which have complete octet, are highly stable and do not take part in chemical combination e.g. inert gases like Ne, Ar, Kr etc.

	K	L
Ne (10) –	2,	8

	K	L	M
Ar (18) –	2,	8,	8

Thus, octet configuration is the stable electronic configuration of the elements. Helium is its exception, where first shell is outermost shell, which cannot accommodate more than two electrons. Hence, for Helium, duplet configuration (having 2 electron in first shell) is the stable configuration.

(2) Those elements which do not contain 8 electrons in their outermost shell i.e. which have incomplete octet, are unstable and therefore, they give, take or share electrons to complete their octet. Hence, they have tendency to take part in chemical combination.

Thus, the tendency of elements to complete their octet i.e. to gain stability by acquiring stable electronic configuration of their nearest inert gas, is the principal cause of chemical combination.

Types of chemical bond

Chemical bonds are mainly of following three types-

(1) Electrovalent bond-

"The bonds formed by transfer of one or more electrons between the atoms are called electrovalent bonds."

The number of electrons given or taken by one atom of an element during the formation of these bonds is called electrovalency of the element."

Here, the atom which gives electron, becomes positively changed while, that which takes electron, becomes negatively changed. The oppositely changed ions thus formed, get bonded together by strong electrostatic force of attraction.

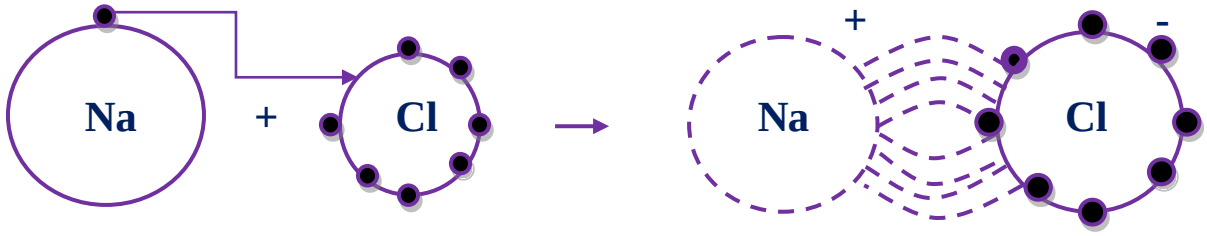
Examples-

(1) Formation of Sodium chloride (NaCl)

The electronic configuration of sodium and chlorine atoms are as under-

	K	L	M
Na (11) -	2,	8,	1
Cl (17) -	2,	8,	7

During formation of NaCl, sodium atom transfers its valence electron to the outermost shell of chlorine atom. It results in the formation of Na^+ and Cl^- ions, which get bonded together by electrostatic force of attraction. Here, the electrovalency of both sodium and chlorine atoms is 1.



Na - atom	Cl - atom	Na^+ ion	Cl^- ion
(2, 8, 1)	(2, 8, 7)	(2, 8)	(2, 8, 8)

Fig.- Formation of NaCl (diagrammatic)

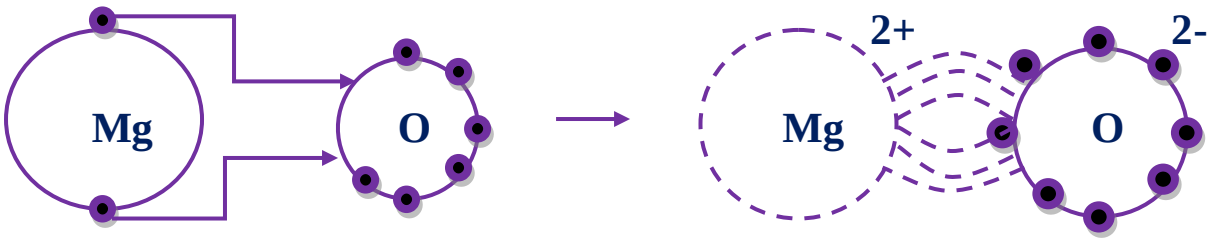
Thus, both sodium and chlorine atoms acquire the stable electronic configuration of their nearest inert gas i.e. Ne and Ar respectively.

(2) Formation of Magnesium oxide (MgO)

The electronic configuration of magnesium and oxygen atoms are as under-

	K	L	M
Mg (12) -	2,	8,	2
O (8) -	2,	6	

During formation of MgO, magnesium atom transfers its both the valence electrons to the outermost shell of oxygen atom. It results in the formation of Mg^{2+} and O^{2-} ions, which get bonded together by electrostatic force of attraction. Here, the electrovalency of both magnesium and oxygen atoms is 2.



Mg - atom	O - atom	Mg^{2+} ion	O^{2-} ion
(2, 8, 2)	(2, 6)	(2, 8)	(2, 8)

Fig.- Formation of MgO (diagrammatic)

Thus, both magnesium and oxygen atoms acquire the stable electronic configuration of their nearest inert gas i.e. Ne.

(3) Formation of Magnesium chloride ($MgCl_2$)

The electronic configuration of magnesium and chlorine atoms is as under-

	K	L	M
Mg (12) -	2,	8,	2
Cl (17) -	2,	8,	7

During formation of $MgCl_2$, magnesium atom transfers its both the valence electrons to the outermost shell of two chlorine atoms (one each). It results in the formation of Mg^{2+} and Cl^- ions, which get bonded together by electrostatic force of attraction. Here, the electrovalency of magnesium atom is 2 while, that of chlorine atoms is 1.

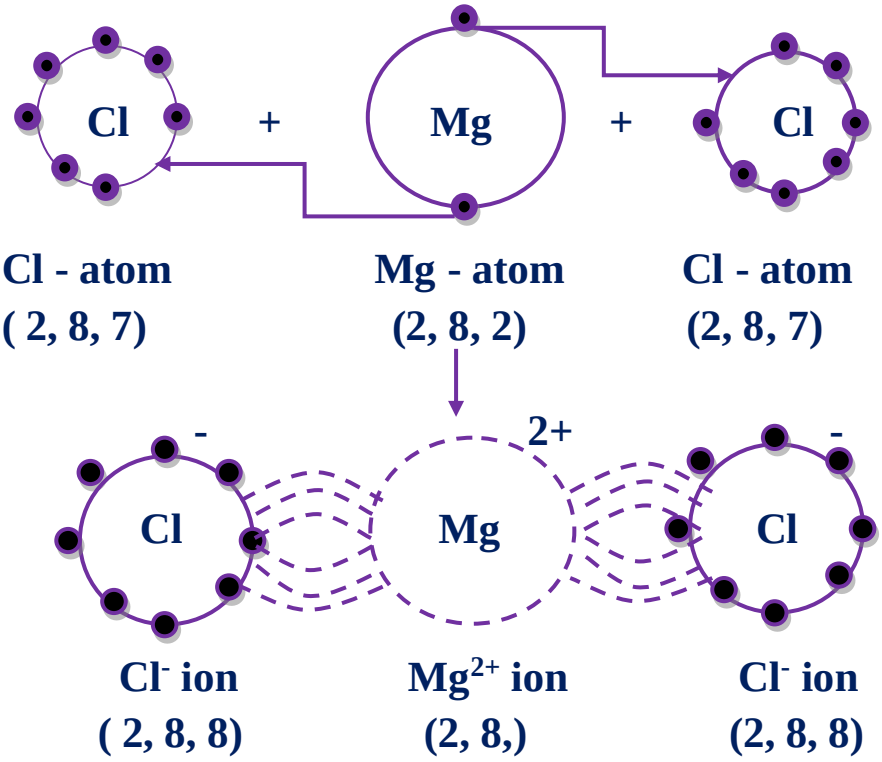


Fig.- Formation of $MgCl_2$ (diagrammatic)

Thus, both magnesium and chlorine atoms acquire the stable electronic configuration of their nearest inert gas.

Conditions of formation of electrovalent bonds

(1) When a highly electropositive element (metal) combines with highly electronegative element (nonmetal), then electrovalent bond is formed between them.

(2) If an element can complete its octet by giving one or more electrons and other can do so by gaining one or more electrons, then electrovalent bond is formed between them.

Properties of Electrovalent compounds

(1) Electrovalent compounds are ionic in nature i.e. they are consisting of oppositely charged ions bonded together by strong electrostatic force of attraction.

(2) Since, strong electrostatic force of attraction operates between oppositely charged constituent ions, these compounds have very high melting and boiling points.

(3) These compounds have high thermal stability and therefore, do not decompose on heating.

(4) Being ionic in nature, these compounds are highly soluble in polar solvents like water.

(5) When these compounds are dissolved in water or melted, they get dissociated into their constituent ions. For example,

$$NaCl(s) \xrightarrow{Water} Na^{+}(aq.) + Cl^{-}(aq)$$

This phenomenon is called dissociation.

(6) Due to presence of ions, the aqueous solutions and molten forms of these compounds are good conductor of electricity.

(7) Electrovalent compounds are, in general, solid at room temperature.

(8) Electrovalent bonds are non-directional in nature. Hence, electrovalent compounds do not have a definite geometry.

(2) Covalent bond

"The bonds formed by equal sharing of electrons between the atoms are called covalent bonds."

The number of electrons shared by one atom of an element for the formation of a covalent bond is called covalency of the atom.

Here, shared electron pair lies simultaneously in the outermost shell of both the bonded atoms and provides them stability.

Depending upon the number of shared electron pairs between the bonded atoms, being 1, 2 and 3, covalent bonds are called single, double and triple bonds respectively and are represented by placing single, double and triple lines between the bonded atoms respectively.

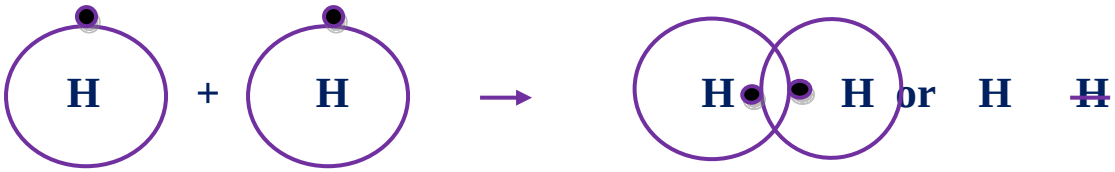
Examples-

(1) Formation of H_2 molecule

The electronic configuration of hydrogen atom is as under-

	K
H (1)	- 1

Thus, it needs one electron to complete its duplet i.e. stable electronic configuration of its nearest inert gas helium. Hence, two hydrogen atoms share one electron each to form a shared electron pair, which lies simultaneously in outermost shell of both the hydrogen atoms and provides them stability.



H- atom **H- atom** **H_2 molecule**

Fig.- Formation of H_2 molecule

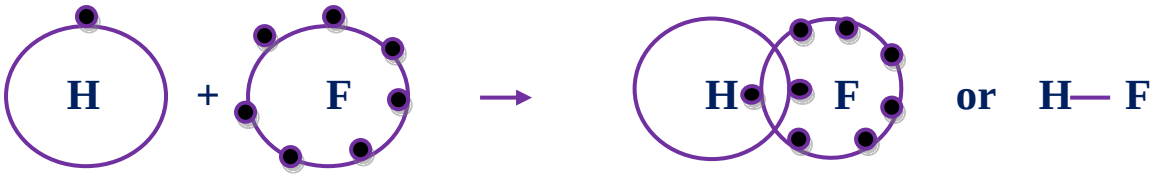
Here, both the hydrogen atoms give one electron in shared pair, therefore, their covalency is 1.

(2) Formation of HF molecule

The electronic configuration of hydrogen and fluorine atoms are as under-

	K	L
H (1) -	1	
F (9) -	2,	7

Here, both hydrogen and fluorine atoms need one electron to acquire the stable electronic configuration of their nearest inert gas. Hence, they share one electron each to form a shared electron pair, which lies simultaneously in outermost shell of both the atoms and provides them stability.



H- atom **F- atom** **HF molecule**

Fig.- Formation of HF molecule

Here, both hydrogen and fluorine atoms give one electron each in shared pair, therefore, their covalency is 1.

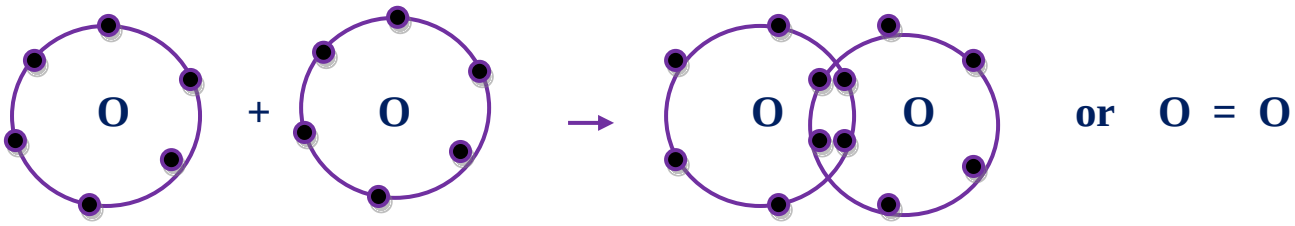
(3) Formation of O_2 molecule

The electronic configuration of oxygen atom is as under-

	K	L
O (8) -	2,	6

Here, both oxygen atoms need two electrons to acquire the stable electronic configuration of their nearest inert gas i.e. neon. Hence, they share two

electrons each to form two shared electron pairs, which lie simultaneously in outermost shell of both the atoms and provide them stability.



O- atom **O- atom** **O_2 molecule**

Fig.- Formation of O_2 molecule

Here, both the oxygen atoms give two electrons each in shared pairs, therefore, their covalency is 2.

Conditions of formation of covalent bonds

(1) When two highly electronegative elements (nonmetals) combine together, then covalent bond is formed between them.

(2) If both the atoms require one or more electrons to complete their octet, then covalent bond is formed between them.

Properties of Covalent compounds

(1) Covalent compounds are molecular in nature i.e. they are consisting of neutral molecules, bonded together by weak vander Waal forces.

(2) Since, weak vander Waal forces operate amongst constituent molecules, these compounds have relatively low melting and boiling points.

(3) Covalent compounds are usually insoluble in water. However, they are fairly soluble in non-polar solvents like benzene, carbon tetrachloride, chloroform etc.

(4) Due to absence of ions, the aqueous solutions and molten forms of these compounds are in general, bad conductor of electricity.

(5) These compounds are thermally unstable and therefore, get decomposed on heating.

(6) Covalent bonds are directional in nature i.e. they are formed in some definite directions and provide a definite shape to the molecules. Thus, covalent compounds have a definite geometry.

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