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Class 10th

Chapter 3rd

Metals & Non- Metals

On the basis of their properties, all the elements can be divided into two main groups : metals and non-metals.

METALS

Metals are the elements that conduct heat and electricity, and are malleable and ductile.

Some of the examples of metals are : Iron, Aluminium, Copper, Silver, Gold, Platinum, Zinc, Tin, Lead, Mercury, Sodium, Potassium, Calcium and Magnesium

PHYSICAL PROPERTIES OF METALS

1. **Metals, in their pure state, have a shining surface. This property is called metallic lustre**

Example

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.

2. **Metals are generally hard. The hardness varies from metal to metal.**
3. **some metals can be beaten into thin sheets. This property is called malleability.**

Example

It we take a piece of aluminium metal, place it on a block of iron and beat it with a hammer four or five times, we will find that the piece of aluminium metal turns into a thin aluminium sheet, without breaking.

4. **The ability of metals to be drawn into thin wires is called ductility. Gold is the most ductile metal.**
5. **Metals are good conductors of heat and have high melting points. The best conductors of heat are silver and copper. Lead and mercury are comparatively poor conductors of heat**

Example

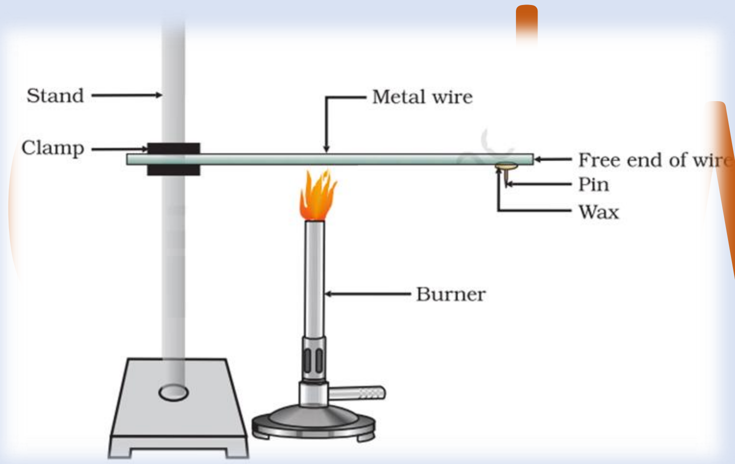
Take an aluminium or copper wire. Clamp this wire on a stand.

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?

the left end of aluminium rod is hot but the right end of rod is cold. So, heat now travels from the hotter left end of aluminium rod to its colder right end. As heat travels from the left side to the right side along the aluminium rod, it melts the wax which holds the nails.



6. Metals are good conductors of electricity

Example

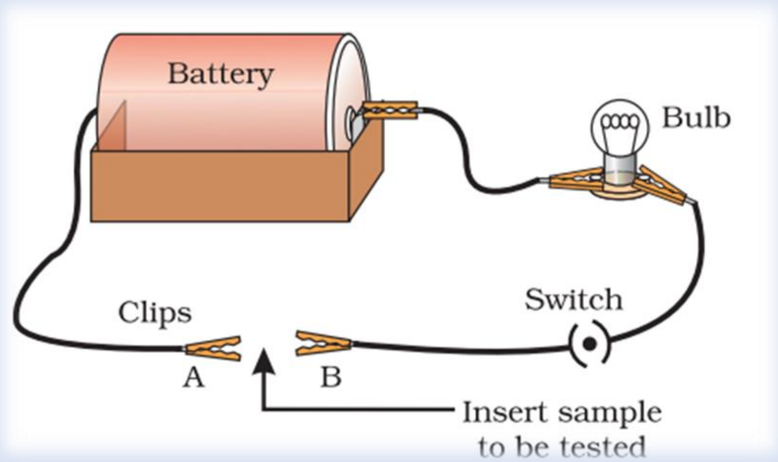
Set up an electric circuit as shown.

Place the metal to be tested in the circuit between terminals A and B as shown.

Let us now insert a piece of aluminium foil between the ends of crocodile clips A and B

We will see that the bulb lights up at once. This means that the aluminium foil allows electric current to pass through it. This shows that aluminium metal conducts electric current (or

electricity). In other words, aluminium metal is a good conductor of electricity



7. . Metals are solids at room temperature (except mercury which is a liquid metal.
8. Metals have high melting points and boiling points
9. Metals are sonorous. That is, metals make sound when hit with an object.

PHYSICAL PROPERTIES OF NON-METALS

1. Non-metals are neither malleable nor ductile. Non-metals are brittle (break easily)
2. Non-metals do not conduct heat and electricity
3. Non-metals are not lustrous (not shiny). They are dull
4. Non-metals are generally soft (except diamond which is an extremely hard non-metal)
5. Non-metals are not strong. They are easily broken
6. Non-metals have comparatively low melting points and boiling points (except diamond which is a non-metal having a high melting point and boiling point)
7. Non-metals have low densities, that is, non-metals are light substances
8. Non-metals are non-sonorous. They do not produce sound when hit with an object

CHEMICAL PROPERTIES OF METAL

What happen when Metal Burn in Air

Almost all metals combine with oxygen to form metal oxides

Metal + Oxygen → Metal oxide

Example

when copper is heated in air, it combines with oxygen to form copper(II) oxide, a black oxide.

$2Cu + O_2 \rightarrow 2CuO$

aluminium forms aluminium oxide.

$4Al + 3O_2 \rightarrow 2Al_2O_3$

metal oxides are basic in nature. But some metal oxides, such as aluminium oxide, zinc oxide 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.

Example

$Al_2O_3 + 6HCl \rightarrow 2AlCl_3 + 3H_2O$

$Al_2O_3 + 2NaOH \rightarrow 2NaAlO_2 + H_2O$

metal oxides are insoluble in water but some of these dissolves in water to form alkalis.

Example

Sodium oxide and potassium oxide dissolve in water to produce alkalis

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

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

Different metals show different reactivities towards oxygen. Metals such as potassium and sodium react so vigorously that they catch fire if kept in the open. Hence, to protect them and to prevent accidental fires, they are kept immersed in kerosene oil.

At ordinary temperature, the surfaces of metals such as magnesium, aluminium, zinc, lead, etc., are covered with a thin layer of oxide. The protective oxide layer prevents the metal from further oxidation.

Iron does not burn on heating but iron filings burn vigorously when sprinkled in the flame of the burner.

Copper does not burn, but the hot metal is coated with a black coloured layer of copper(II) oxide. Silver and gold do not react with oxygen even at high temperatures

<u>What Happen When Metal React with Water</u>

Metals react with water and produce a metal oxide and hydrogen gas. Metal oxides that are soluble in water dissolve in it to further form metal hydroxide.

Metal + Water → Metal oxide + Hydrogen

Metal oxide + Water → Metal hydroxide

Metals like potassium and sodium react violently with cold water. the reaction is so violent and exothermic that the evolved hydrogen immediately catches fire.

$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$ + heat energy

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$ + heat energy

The reaction of calcium with water is less violent. The heat evolved is not sufficient for the hydrogen to catch fire

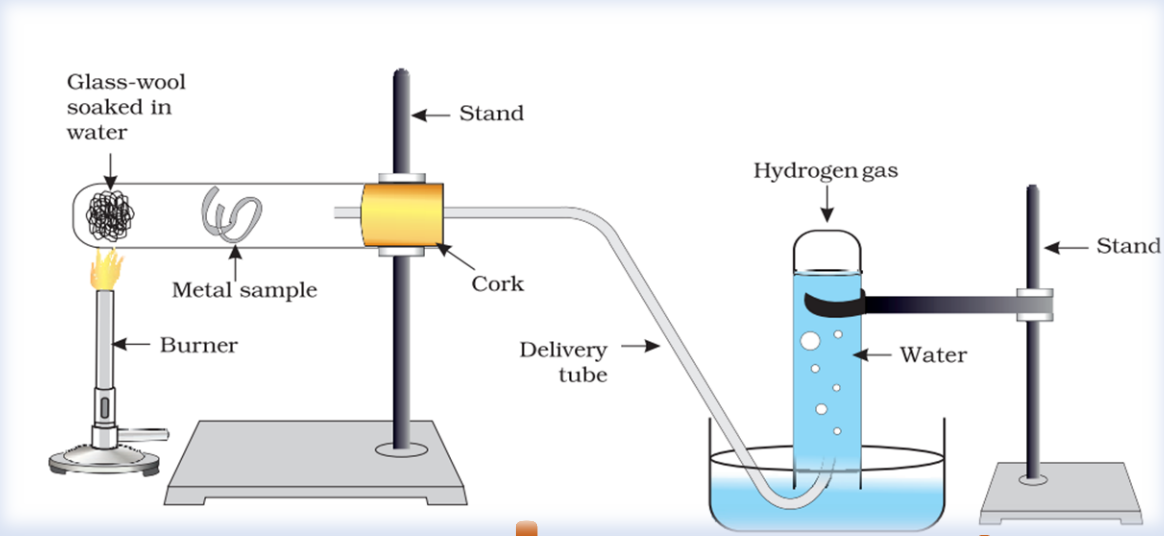
$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$

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 bubbles of hydrogen gas sticking to its surface.

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.

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

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



<u>What Happen When Metal React with Acids</u>

Metals react with acids to give a salt and hydrogen gas.

Metal + Dilute acid → Salt + Hydrogen

Example

Magnesium reacts quite rapidly with dilute hydrochloric acid forming magnesium chloride and hydrogen gas :

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

Aluminium metal reacts rapidly with dilute hydrochloric acid to form aluminium chloride and hydrogen gas :

$2Al\ (s) + 6HCl\ (aq) \longrightarrow 2AlCl_3\ (aq) + 3H_2\ (g)$

Zinc reacts with dilute hydrochloric acid to give zinc chloride and hydrogen gas

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

Iron reacts slowly with cold dilute hydrochloric acid to give iron (II) chloride and hydrogen gas :

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

When a metal reacts with dilute nitric acid, then hydrogen gas is not evolved.

Nitric acid is a strong oxidising agent. So, as soon as hydrogen gas is formed in the reaction between a metal and dilute nitric acid, the nitric acid oxidises this hydrogen to water.

Very dilute nitric acid, however, reacts with magnesium and manganese metals to evolve hydrogen gas. This is because the very dilute nitric acid is a weak oxidising agent which is not able to oxidise hydrogen to water.

Aqua regia

Aqua-regia is a freshly prepared mixture of 1 part of concentrated nitric acid and 3 parts of concentrated hydrochloric acid. Thus, the ratio of conc. HNO3 and conc. HCl in aqua-regia is 1 : 3. Aqua-regia is a highly corrosive, fuming liquid Aqua-regia can dissolve all metals. For example, aqua-regia can dissolve even gold and platinum metals.

How Do Metal React with Solutions Of Other Metals salts?

When a more reactive metal is put in the salt solution of a less reactive metal, then the more reactive metal displaces (pushes out) the less reactive metal from its salt solution.

if metal A is more reactive than metal B, then metal A will displace metal B from its salt solution to form salt solution of metal A, and metal B will be set free.

Salt solution of metal B + Metal A $\longrightarrow$ Salt solution of metal A + Metal B

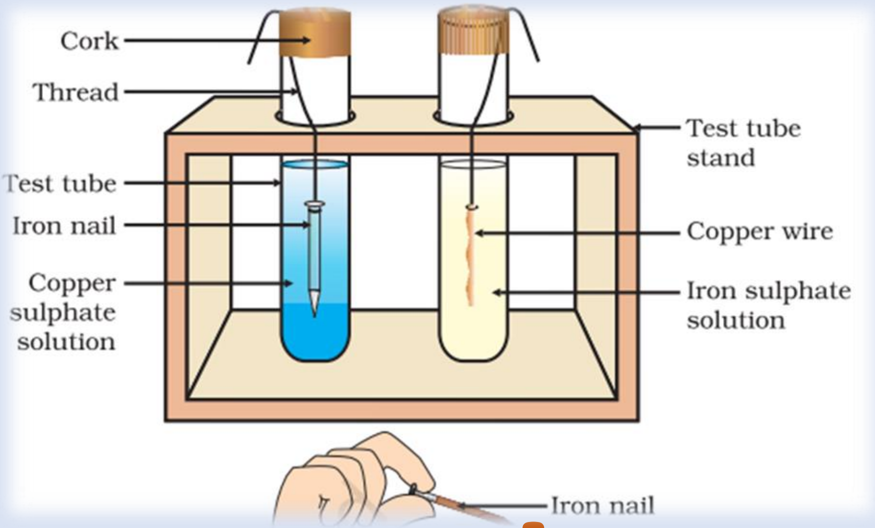
Example

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.

The copper metal produced in this reaction forms a red-brown layer on the iron strip, In this reaction, iron is displacing copper from copper sulphate solution. This displacement occurs because iron is more reactive than copper.



Reactivity Series

The arrangement of metals in a vertical column in the order of decreasing reactivities is called reactivity series of metals.

K	Potassium	Most reactive
Na	Sodium	
Ca	Calcium	
Mg	Magnesium	
Al	Aluminium	
Zn	Zinc	Reactivity decreases
Fe	Iron	
Pb	Lead	
[H]	[Hydrogen]	
Cu	Copper	
Hg	Mercury	
Ag	Silver	
Au	Gold	Least reactive

HOW DO METALS AND NON-METALS REACT

We learnt that noble gases, which have a completely filled valence shell, show little chemical activity. We, therefore,

explain the reactivity of elements as a tendency to attain a completely filled valence shell.

Example

sodium atom has one electron in its outermost shell. If it loses the electron from its M shell then its L shell now becomes the outermost shell and that has a stable octet.

The nucleus of this atom still has 11 protons but the number of electrons has become 10, so there is a net positive charge giving us a sodium cation Na^+.

$$\mathbf{Na} \longrightarrow \mathrm{Na}^+ + \mathrm{e}^-$$

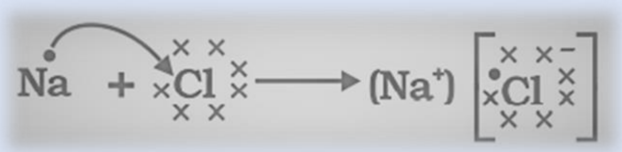
281 2 8

chlorine has seven electrons in its outermost shell and it requires one more electron to complete its octet.

$$\mathbf{Cl} + \mathrm{e}^- \longrightarrow \mathrm{Cl}^-$$

2 8 7 2 8 8

If sodium and chlorine were to react, the electron lost by sodium could be taken up by chlorine. After gaining an electron, the chlorine atom gets a unit negative charge, because its nucleus has 17 protons and there are 18 electrons in its K, L and M shells. This gives us a chloride anion Cl^-.



Sodium and chloride ions, being oppositely charged, attract each other and are held by strong electrostatic forces of attraction to exist as sodium chloride (NaCl).

<u>Ionic Compounds</u>

The compounds containing ionic bonds are known as ionic compounds. They are formed by the transfer of electrons from one

atom to another. The ionic compounds are made up of positively charged ions (cations) and negatively charged ions (anions).

PROPERTIES OF IONIC COMPOUNDS

1. **Physical nature:**

Ionic compounds are solids and are somewhat hard because of the strong force of attraction between the positive and negative ions. These compounds are generally brittle and break into pieces when pressure is applied.

2. **Melting and Boiling points:**

Ionic compounds have high melting and boiling points. This is because a considerable amount of energy is required to break the strong inter-ionic attraction.

3. **Solubility:**

Electrovalent compounds are generally soluble in water and insoluble in solvents such as kerosene, petrol, etc.

4. **Conduction of Electricity:**

The conduction of electricity through a solution involves the movement of charged particles. A solution of an ionic compound in water contains ions, which move to the opposite electrodes when electricity is passed through the solution. Ionic compounds in the solid state do not conduct electricity because movement of ions in the solid is not possible due to their rigid structure. But ionic compounds conduct electricity in the molten state. This is possible in the molten state since the electrostatic forces of attraction between the oppositely charged ions are overcome due to the heat. Thus, the ions move freely and conduct electricity.

OCCURRENCE OF METALS

The earth's crust is the major source of metals. Sea-water also contains salts of metals like sodium chloride, magnesium chloride, etc. Most of the metals are quite reactive and hence they do not occur as free elements in nature.

The compounds of metals found in nature are their oxides, carbonates, sulphides and chlorides, etc. In these compounds, the metals are present in the form of positive ions (or cations). Only a few less reactive metals (like copper, silver, gold and platinum) are found in the 'free state' as metals

EXTRACTION OF METAL

A metal is extracted from its ore. Some metals are found in the earth's crust in the free state. Some are found in the form of their compounds.

The metals at the bottom of the activity series are the least reactive. They are often found in a free state.

Example,

gold, silver, platinum and copper are found in the free state.

Copper and silver are also found in the combined state as their sulphide or oxide ores.

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.

The metals in the middle of the activity series (Zn, Fe, Pb, etc.) are moderately reactive. They are found in the earth's crust mainly as oxides, sulphides or carbonates.

Enrichment of Ore

Ores mined from the earth are usually contaminated with large amounts of impurities such as soil, sand, etc., called gangue. The impurities must be removed from the ore prior to the extraction of the metal. 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.

Extraction of Less Reactive Metals

Metals low in the activity series are very unreactive. The oxides of these metals can be reduced to metals by heating alone.

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.

$$2HgS\ (s) + 3O_2\ (g) \xrightarrow{\text{Heat}} 2HgO\ (s) + 2SO_2\ (g)$$

$$2HgO\ (s) \xrightarrow{\text{Heat}} 2Hg\ (l) + O_2\ (g)$$

Copper is a less reactive metal which is quite low in the reactivity series. Copper metal can be extracted just by heating its sulphide ore in air.

$$2Cu_2S + 3O_{2\,(g)} \xrightarrow{\text{Heat}} 2Cu_2O\ (s) + 2SO_{2\,(g)}$$

$$2Cu_2O + 2Cu_2S \xrightarrow{\text{Heat}} 6Cu\ (s) + So_{2\,(g)}$$

Extraction of Moderately Reactive Metals

The moderately reactive metals which are in the middle of reactivity series are extracted by the reduction of their oxides with carbon, aluminium, sodium or calcium.

it is easier to obtain metals from their oxides (by reduction) than from carbonates or sulphides.

The concentrated ores can be converted into metal oxide by the process of calcination or roasting.

The method to be used depends on the nature of the ore. A carbonate ore is converted into oxide by calcination whereas a sulphide ore is converted into oxide by roasting.

The chemical reaction that takes place during roasting and calcination of zinc ores can be shown as follows –

Roasting

$$2ZnS(s) + 3O_2\ (g) \xrightarrow{\text{Heat}} 2ZnO(s) + 2SO_2\ (g)$$

Calcination

$$ZnCO\ (s) \xrightarrow{\text{Heat}} ZnO(s) + CO\ (g)$$

The metal oxides are then reduced to the corresponding metals by using suitable reducing agents such as carbon. For example, when zinc oxide is heated with carbon, it is reduced to metallic zinc

$$ZnO(s) + C(s) \rightarrow Zn(s) + CO(g)$$

The highly reactive metals such as sodium, calcium, aluminium, etc., are used as reducing agents because they can displace metals of lower reactivity from their compounds.

Example

when manganese dioxide is heated with aluminium powder, the following reaction takes place –

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

Thermite Reaction.

The reduction of a metal oxide to form metal by using aluminium powder as a reducing agent is called a thermite reaction (or thermite process). The reactions of metal oxides with aluminium powder to produce metals are highly exothermic in which a large amount of heat is evolved.

This property of the reduction by aluminium is made use of in thermite welding for joining the broken pieces of heavy iron objects like girders, railway tracks or cracked machine parts.

$Fe_2O_3(s) + 2Al(s) \rightarrow 2Fe(l) + Al_2O_3(s)$ + Heat

Extraction of Highly Reactive Metals

The highly reactive metals such as potassium, sodium, calcium, magnesium and aluminium are placed high up in the reactivity series in its upper part.

The highly reactive metals are extracted by the electrolytic reduction of their molten chlorides or oxides.

Example

sodium, magnesium and calcium are obtained by the electrolysis of their molten chlorides. The metals are deposited at the cathode (the negatively charged electrode), whereas, chlorine is liberated at the anode (the positively charged electrode). The reactions are –

At cathode $Na^+ + e^- \longrightarrow$ **Na**

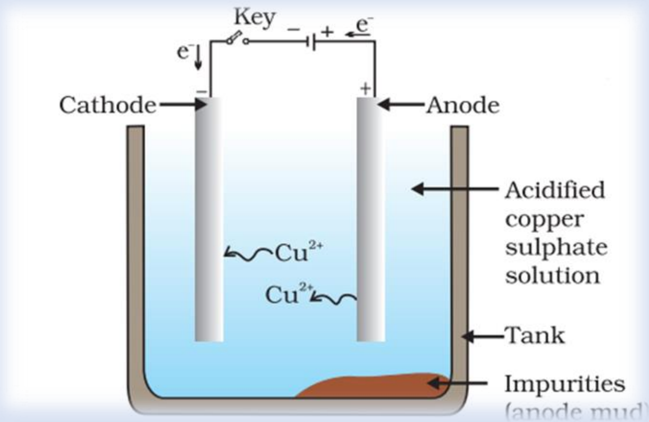
At anode $2Cl^- \longrightarrow Cl_2 + 2e^-$

Refining of Metals

Electrolytic Refining

Many metals, such as copper, zinc, tin, nickel, silver, gold, etc., are refined electrolytically. In this process, the impure metal is made the anode and a thin strip of pure metal is made the cathode. A solution of the metal salt is used as an electrolyte.

On passing the current through the electrolyte, the pure metal from the anode dissolves into the electrolyte. An equivalent amount of pure metal from the electrolyte is deposited on the cathode. The soluble impurities go into the solution, whereas, the insoluble impurities settle down at the bottom of the anode and are known as anode mud.



CORROSION

The eating up of metals by the action of air, moisture or a chemical (such as an acid) on their surface is called corrosion.

Example

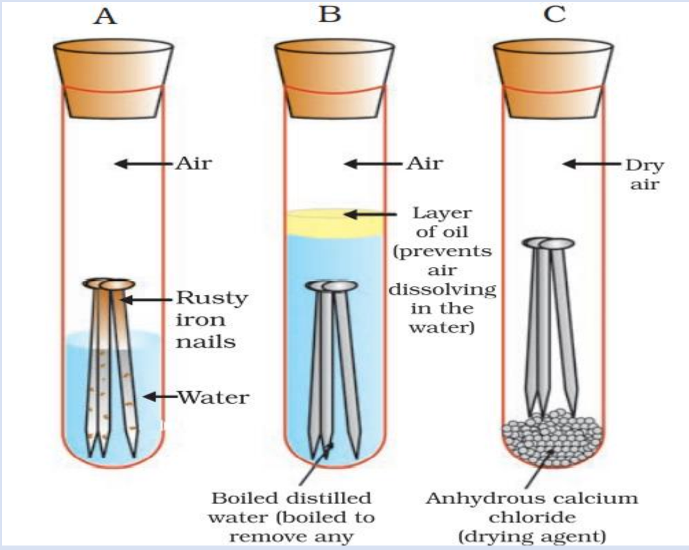
Take three test tubes and place clean iron nails in each of them.

Label these test tubes A, B and C. Pour some water in test tube A and cork it.

Pour boiled distilled water in test tube B, add about 1 mL of oil and cork it. The oil will float on water and prevent the air from dissolving in the water.

Put some anhydrous calcium chloride in test tube C and cork it. Anhydrous calcium chloride will absorb the moisture, if any, from the air. Leave these test tubes for a few days and then observe.

t iron nails rust in test tube A, but they do not rust in test tubes B and C. In the test tube A, the nails are exposed to both air and water. In the test tube B, the nails are exposed to only water, and the nails in test tube C are exposed to dry air.



Prevention Of Corrosion

The rusting of iron can be prevented by painting, oiling, greasing, galvanising, chrome plating, anodising or making alloys.

Galvanisation is a method of protecting steel and iron from rusting by coating them with a thin layer of zinc. The galvanised article is protected against rusting even if the zinc coating is broken.

Alloying is a very good method of improving the properties of a metal.

For example, iron is the most widely used metal. But it is never used in its pure state. This is because pure iron is very soft and stretches easily when hot. But, if it is mixed with a small amount of carbon (about 0.05 %), it becomes hard and strong. When iron is mixed with nickel and chromium, we get stainless steel, which is hard and does not rust. Thus, if iron is mixed with some other substance, its properties change