Metals and Non-metals

Corrosion

Alloys

Alloys are homogeneous mixtures of metal with other metals or nonmetals. Alloy formation enhances the desirable properties of the material, such as hardness, tensile strength and resistance to corrosion.
Examples of few alloys -
Brass: copper and zinc
Bronze: copper and tin
Solder: lead and tin
Amalgam: mercury and other metal

Corrosion

Gradual deterioration of a material usually a metal by the action of moisture, air or chemicals in the surrounding environment.
Rusting:
$4Fe(s) + 3O_2(from\ air) + xH_2O(moisture) \rightarrow 2Fe_2O_3.xH_2O(rust)$

Corrosion of copper:
$Cu(s) + H_2O(moisture) + CO_2(from\ air) \rightarrow CuCO_3.Cu(OH)_2(green)$

Corrosion of silver:
$Ag(s) + H_2S(from\ air) \rightarrow Ag_2S(black) + H_2(g)$

Prevention of Corrosion

Prevention:
1. Coating with paints or oil or grease: Application of paint or oil or grease on metal surfaces keep out air and moisture.

2. Alloying: Alloyed metal is more resistant to corrosion. Example: stainless steel.

3. Galvanization: This is a process of coating molten zinc on iron articles. Zinc forms a protective layer and prevents corrosion.

4. Electroplating: It is a method of coating one metal with another by use of electric current. This method not only lends protection but also enhances the metallic appearance. Example: silver plating, nickel plating.

5. Sacrificial protection: Magnesium is more reactive than iron. When it is coated on the articles made of iron or steel, it acts as the cathode, undergoes reaction (sacrifice) instead of iron and protects the articles.

Physical Properties

Physical Properties of Metals

- Hard and have a high tensile strength
- Solids at room temperature
- Sonorous
- Good conductors of heat and electricity
- Malleable, i.e., can be beaten into thin sheets
- Ductile, i.e., can be drawn into thin wires
- High melting and boiling points (except Caesium (Cs) and Gallium (Ga))
- Dense, (except alkali metals). Osmium - highest density and lithium - least density
- Lustrous
- Silver-grey in colour, (except gold and copper)

Non-Metals

Nonmetals are those elements which do not exhibit the properties of metals.

Physical Properties of Nonmetals

- Occur as solids, liquids and gases at room temperature
- Brittle
- Non-malleable
- Non-ductile
- Non-sonorous
- Bad conductors of heat and electricity

Exceptions in Physical Properties

- Alkali metals (Na, K, Li) can be cut using a knife.
- Mercury is a liquid metal.
- Lead and mercury are poor conductors of heat.
- Mercury expands significantly for the slightest change in temperature.
- Gallium and caesium have a very low melting point
- Iodine is non-metal but it has lustre.
- Graphite conducts electricity.
- Diamond conducts heat and has a very high melting point.

Chemical Properties

Chemical Properties of Metals

- Alkali metals (Li, Na, K, etc) react vigorously with water and oxygen or air.
- Mg reacts with hot water.
- Al, Fe and Zn react with steam.
- Cu, Ag, Pt, Au do not react with water or dilute acids.

Reaction of Metals with Oxygen (Burnt in Air)

Metal + Oxygen $\rightarrow$ Metal oxide (basic)

- Na and K are kept immersed in kerosene oil as they react vigorously with air and catch fire.

$4K(s) + O_2(g) \rightarrow 2K_2O(s)$ (vigorous reaction)

- Mg, Al, Zn, Pb react slowly with air and form a protective layer that prevents corrosion.

$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$ (Mg burns with a white dazzling light)

$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$

- Silver, platinum and gold don't burn or react with air.

Basic Oxides of Metals

Some metallic oxides get dissolved in water and form alkalis. Their aqueous solution turns red litmus blue.

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

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

Amphoteric Oxides of Metals

Amphoteric oxides are metal oxides which react with both acids as well as bases to form salt and water.

For example - Al_2O_3, ZnO, PbO, SnO

$Al_2O_3(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2O(l)$

$Al_2O_3(s) + 2NaOH(aq) \rightarrow 2NaAlO_2(aq) + H_2O(l)$

$ZnO(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2O(l)$

$ZnO(s) + 2NaOH(aq) \rightarrow Na_2ZnO_2(aq) + H_2O(l)$

Reactivity Series

The below table illustrates the reactivity of metals from high order to low order.

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

Reaction of Metals with Water or Steam

$Metal + Water \rightarrow Metal\ hydroxide\ or\ Metal\ oxide + Hydrogen$

$2Na + 2H_2O(cold) \rightarrow 2NaOH + H_2 + heat$
$Ca + 2H_2O(cold) \rightarrow Ca(OH)_2 + H_2$
$Mg + 2H_2O(hot) \rightarrow Mg(OH)_2 + H_2$
$2Al + 3H_2O(steam) \rightarrow Al_2O_3 + 3H_2$
$Zn + H_2O(steam) \rightarrow ZnO + H_2$
$3Fe + 4H_2O(steam) \rightarrow Fe_3O_4 + 4H_2$

Reaction of Metals with Acid

$Metal + dilute\ acid \rightarrow Salt + Hydrogen\ gas$
$2Na(s) + 2HCl(dilute) \rightarrow 2NaCl(aq) + H_2(g)$
$2K(s) + H_2SO_4(dilute) \rightarrow K_2SO_4(aq) + H_2(g)$

Only Mg and Mn, react with very dilute nitric acid to liberate hydrogen gas.
$Mg(s) + 2HNO_3(dilute) \rightarrow Mg(NO_3)_2(aq) + H_2(g)$
$Mn(s) + 2HNO_3(dilute) \rightarrow Mn(NO_3)_2(aq) + H_2(g)$

Displacement Reaction

A more reactive element displaces a less reactive element from its compound or solution.

How Do Metal React with Solution of Other Metal Salts

$Metal\ A + Salt\ of\ metal\ B \rightarrow Salt\ of\ metal\ A + Metal\ B$

$Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$
$Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)(aq) + 2Ag(s)$

Reaction of Metals with Bases

$Base + metal \rightarrow salt + hydrogen$
$2NaOH(aq) + Zn(s) \rightarrow Na_2ZnO_2(aq) + H_2(g)$
$2NaOH(aq) + 2Al(s) + 2H_2O(l) \rightarrow 2NaAlO_2(aq) + 2H_2(g)$

Extraction of Metals and Non-Metals

Applications of Displacement Reaction

Uses of displacement reaction

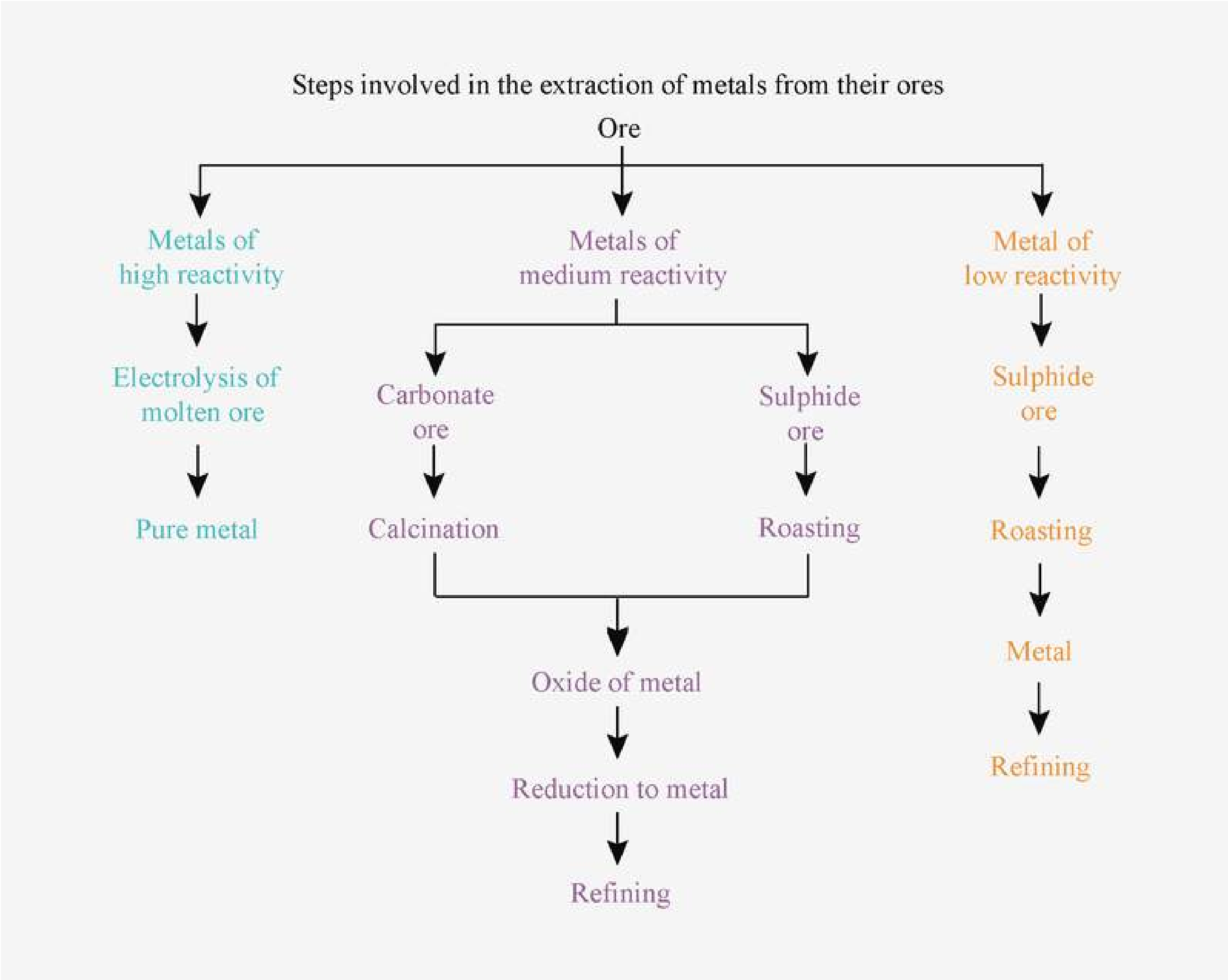
1. Extraction of metals
2. Manufacturing of steel
3. Thermite reaction: $Al(s) + Fe_2O_3(s) \rightarrow Al_2O_3 + Fe(molten)$

The thermite reaction is used in welding of railway tracks, cracked machine parts, etc.

Occurrence of Metals

Most of the elements especially metals occur in nature in the combined state with other elements. All these compounds of metals are known as **minerals**. But out of them, only a few are viable sources of that metal. Such sources are called **ores**.
Au, Pt - exist in the native or free state.

Extraction of Metals



Metals of high reactivity - Na, K, Mg, Al.
Metals of medium reactivity - Fe, Zn, Pb, Sn.
Metals of low reactivity - Cu, Ag, Hg

Roasting

Converts sulphide ores into oxides on heating strongly in the presence of excess air.
It also removes volatile impurities.
$2ZnS(s) + 3O_2(g) + Heat \rightarrow 2ZnO(s) + 2SO_2(g)$

Calcination

Converts carbonate and hydrated ores into oxides on heating strongly in the presence of limited air. It also removes volatile impurities.

$ZnCO_3(s) + heat \rightarrow ZnO(s) + CO_2(g)$
$CaCO_3(s) + heat \rightarrow CaO(s) + CO_2(g)$
$Al_2O_3.2H_2O(s) + heat \rightarrow 2Al_2O_3(s) + 2H_2O(l)$
$2Fe_2O_3.3H_2O(s) + heat \rightarrow 2Fe_2O_3(s) + 3H_2O(l)$

Extracting Metals Low in Reactivity Series

By self-reduction- when the sulphide ores of less electropositive metals like Hg, Pb, Cu etc., are heated in air, a part of the ore gets converted to oxide which then reacts with the remaining sulphide ore to give the crude metal and sulphur dioxide. In this process, no external reducing agent is used.

1. $2HgS(Cinnabar) + 3O_2(g) + heat \rightarrow 2HgO(crude\ metal) + 2SO_2(g)$
   $2HgO(s) + heat \rightarrow 2Hg(l) + O_2(g)$

2. $Cu_2S(Copper\ pyrite) + 3O_2(g) + heat \rightarrow 2Cu_2O(s) + 2SO_2(g)$
   $2Cu_2O(s) + Cu_2S(s) + heat \rightarrow 6Cu(crude\ metal) + SO_2(g)$

3. $2PbS(Galena) + 3O_2(g) + heat \rightarrow 2PbO(s) + 2SO_2(g)$
   $PbS(s) + 2PbO(s) \rightarrow 2Pb(crude\ metal) + SO_2(g)$

Extracting Metals in the Middle of Reactivity Series

Smelting - it involves heating the roasted or calcined ore(metal oxide) to a high temperature with a suitable reducing agent. The crude metal is obtained in its molten state.
$Fe_2O_3 + 3C(coke) \rightarrow 2Fe + 3CO_2$

Aluminothermic reaction - also known as the Goldschmidt reaction is a highly exothermic reaction in which metal oxides usually of Fe and Cr are heated to a high temperature with aluminium.
$Fe_2O_3 + 2Al \rightarrow Al_2O_3 + 2Fe + heat$
$Cr_2O_3 + 2Al \rightarrow Al_2O_3 + 2Cr + heat$

Extraction of Metals Towards the Top of the Reactivity Series

Electrolytic reduction:

1. Down’s process: Molten NaCl is electrolysed in a special apparatus.

At the **cathode** (reduction) -
$Na^+(molten) + e^- \rightarrow Na(s)$
Metal is deposited.

At the **anode** (oxidation) -
$2Cl^-(molten) \rightarrow Cl_2(g) + 2e^-$
Chlorine gas is liberated.

2. Hall’s process: Mixture of molten alumina and a fluoride solvent usually cryolite, (Na_3AlF_6) is electrolysed.

At the **cathode** (reduction) -
$2Al^{3+} + 6e^- \rightarrow 2Al(s)$

Metal is deposited.

At the **anode** (oxidation) -
$6O^{2-} \rightarrow 3O_2(g) + 12e^-$
Oxygen gas is liberated.

Enrichment of Ores

It means removal of impurities or gangue from ore, through various physical and chemical processes. The technique used for a particular ore depends on the difference in the properties of the ore and the gangue.

Refining of Metals

Refining of metals - removing impurities or gangue from crude metal. It is the last step in metallurgy and is based on the difference between the properties of metal and the gangue.

Electrolytic Refining

Metals like copper, zinc, nickel, silver, tin, gold etc., are refined electrolytically.
Anode – impure or crude metal
Cathode – thin strip of pure metal
Electrolyte – aqueous solution of metal salt

From anode (oxidation) - metal ions are released into the solution
At cathode (reduction) - equivalent amount of metal from solution is deposited
Impurities deposit at the bottom of the anode.

The Why Questions

Electronic configuration

Group 1 elements - Alkali metals

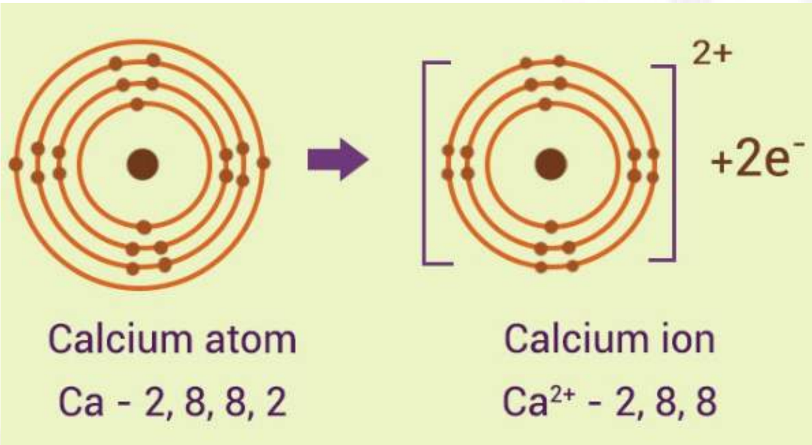
$Element$	$Electronic\ configuration$
$Lithium(Li)$	$2, 1$
$Sodium(Na)$	$2, 8, 1$
$Potassium(K)$	$2, 8, 8, 1$
$Rubidium(Rb)$	$2, 8, 18, 8, 1$

Group 2 elements - Alkaline earth metals

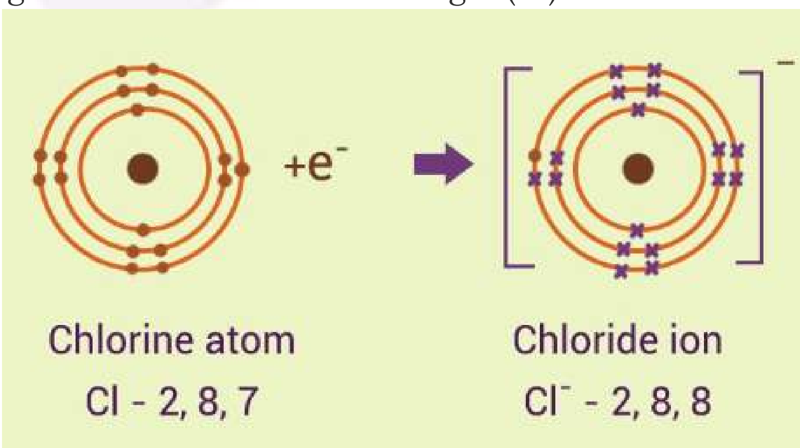
Element	*Electronic configuration*
Beryllium(Be)	2, 2
Magnesium(Mg)	2, 8, 2
Calcium(Ca)	2, 8, 8, 2
Stronium(Sr)	2, 8, 18, 8, 2

How Do Metals and Nonmetals React

Metals lose valence electron(s) and form cations.
Non-metals gain those electrons in their valence shell and form anions.
The cation and the anion are attracted to each other by strong electrostatic force, thus forming an ionic bond.
For example: In Calcium chloride, the ionic bond is formed by oppositely charged calcium and chloride ions.
Calcium atom loses 2 electrons and attains the electronic configuration of the nearest noble gas (Ar). By doing so, it gains a net charge of +2.



The two Chlorine atoms take one electron each, thus gaining a charge of -1 (each) and attain the electronic configuration of the nearest noble gas (Ar).



Ionic Compounds

The electrostatic attractions between the oppositely charged ions hold the compound together.
Example: $MgCl_2, CaO, MgO, NaCl, etc.$

Properties of Ionic Compound

Ionic compounds

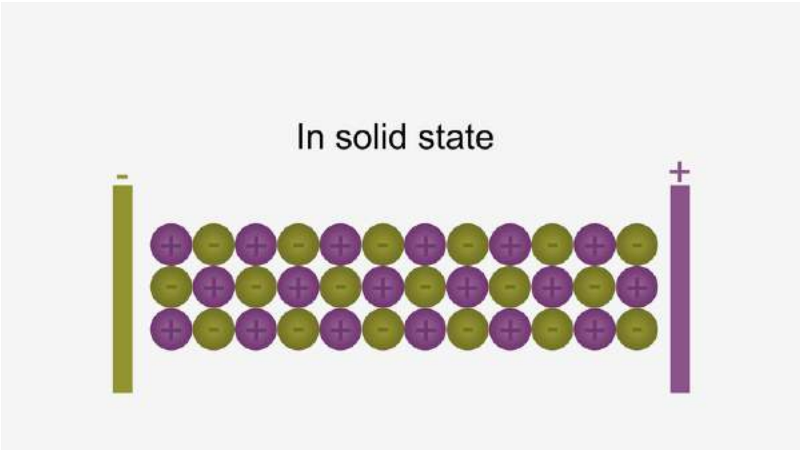
1. Are usually crystalline solids (made of ions).
2. Have high melting and boiling points.
3. Conduct electricity when in aqueous solution and when melted.
4. Are mostly soluble in water and polar solvents.

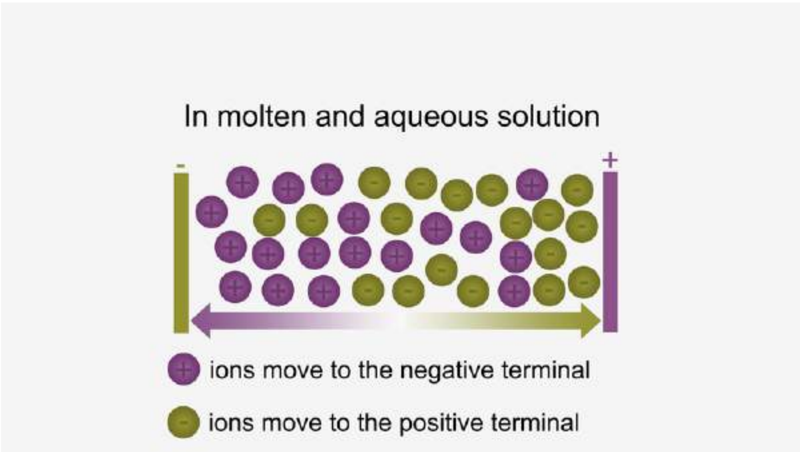
Physical Nature

Ionic solids usually exist in a regular, well-defined crystal structures.

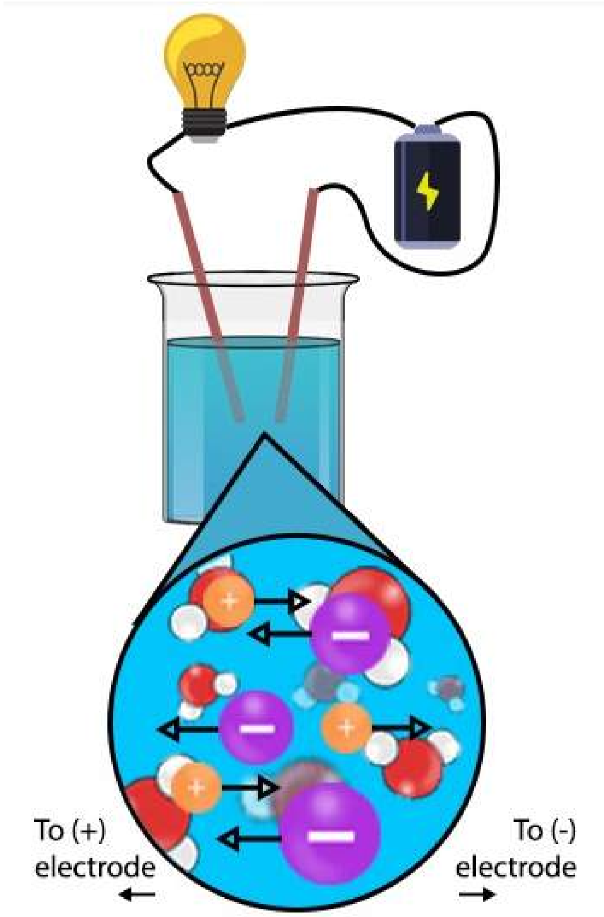
Electric Conduction of Ionic Compounds

Ionic compounds conduct electricity in the molten or aqueous state when ions become free and act as charge carriers.
In solid form, ions are strongly held by electrostatic forces of attractions and not free to move; hence do not conduct electricity.





For example, ionic compounds such as NaCl does not conduct electricity when solid conduct electricity but when dissolved in water or in molten state, it will conduct electricity.

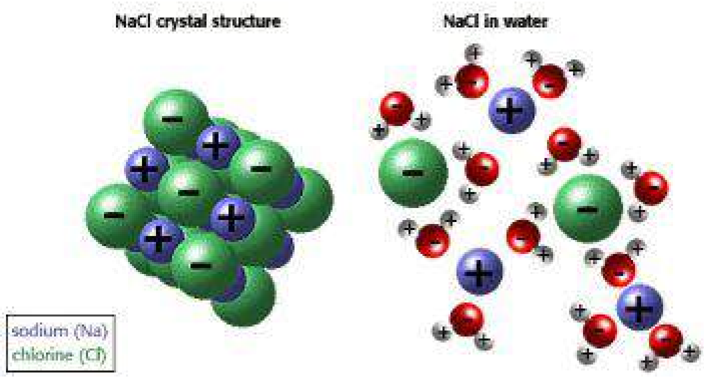


Salt solution conducts electricity

Melting and Boiling Points of Ionic Compounds

In ionic compounds, the strong electrostatic forces between ions require a high amount of energy to break. Thus, the melting point and boiling point of an ionic compound are usually very high.

Solubility of Ionic Compounds



Most ionic compounds are soluble in water due to the separation of ions by water. This occurs due to the polar nature of water.

For example, $NaCl$ is a 3-D salt crystal composed of Na^+ and Cl^- ions bound together through electrostatic forces of attractions. When a crystal of NaCl comes into contact with water, the partial positively charged ends of water molecules interact with the Cl^- ions, while the negatively charged end of the water molecules interacts with the Na^+ ions.
This ion-dipole interaction between ions and water molecules assist in the breaking of the strong electrostatic forces of attractions within the crystal and ultimately in the solubility of the crystal.