Metals & Non-Metals
Chemical Properties of Metals

❖ CHEMICAL PROPERTIES OF METALS

Metals are electropositive elements, so they ionise by loss of electrons and form positive ions (cations)

$$K \rightarrow K^{+} + e^{-}, \qquad Mg \rightarrow Mg^{+2} + 2e^{-}$$

The electropositive character of metals gives a certain characteristic chemical properties, these are discussed below.

(I) Reaction of Metals with Oxygen : Almost all metals combine with oxygen to form their respective oxides. Metals oxides are basic in nature.

Metal + Oxygen Metal oxide

When heated in air sodium burns with golden yellow flame. Potassium burn with pink-violet flame.

All the metals combine with oxygen, and form basic oxides

(a) $4Na + O_2 \;\; 2Na_2O$

$2Cu + O_2 \;\; 2CuO$

(b) Mg does not react with oxygen at room temperature. It combine with oxygen only on heating. Magnesium also combine with nitrogen and forms oxide & nitride.

$2Mg + O_2 \;\; 2MgO$

$6Mg + 2N_2 \;\; 2Mg_3N_2$ (Magnesium nitride)

(c) $2Zn + 2O_2 \;\; 2ZnO(s)$

(d) Iron does not burn in air even on strong heating but iron filings burn vigorously when sprinkled in the flame of burner. Iron reacts with oxygen or air to form Fe_3O_4.

$3Fe(s) + O_2(g) \;\; Fe_3O_4(s)$

It may be mentioned that oxides of some less electropositive metals are amphoteric in nature.

Aluminium & zinc are metals. These metals combine with oxygen and form amphoteric oxide. (Amphoteric oxides reacts with acids and bases)

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

$$2Zn + O_2 \ \ 2ZnO \ \ \text{(Amphoteric oxides)}$$

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

$$Al_2O_3 + 2NaOH \ \ 2NaAlO_2 + H_2O \ \text{(sodium aluminate)}$$

Alkali (I A group) & Alkaline earth metal (II A group) oxides are soluble in nature and forms metal hydroxides.

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

$$MgO\ (s) + H_2O\ (l) \rightarrow Mg(OH)_2$$

But most of the metal oxides are insoluble in nature.

- Different metals show different reactivities towards oxygen. Na & K – catches fire when they placed in moist air, So Na & K are kept in kerosene.
- Mg, Al, Zn & Pb reacts with oxygen and forms metal oxide. This oxide layer is called protective oxide layer, it prevent the further oxidation.
- Pb, Ag & Au – do not react with oxygen even at high temperature so they are called noble metals.

$$\boxed{Na > Mg > Zn > Fe > Cu}$$

Anodizing: is a process of forming a thick oxide layer of aluminium. During anodising, a clean aluminium article is made the anode and is electrolysed with dilute H_2SO_4. The oxygen gas evolved at the anode reacts with aluminium to make a thicker protective oxide layer. This oxide layer can be dyed easily to give Al – articles to an attractive finish.

(II) Action with water

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. But all metals do not react with water.

Metal + Water Metal oxide + Hydrogen

Metal oxide + Water Metal hydroxide

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

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

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

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

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

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 bubbles of hydrogen gas sticking to its surface.

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

Metals like aluminum, iron and zinc do not react either with cold or hot water. But they react with steam to form the metal oxide and hydrogen.

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

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

The reaction is reversible in nature.

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

Thus, the order of reactivity of some common metals with water.

$$\boxed{K > Na > Ca >> Mg > Zn > Fe > Cu}$$

(III) Reaction With Acids

All metals do not react with dilute hydrochloric acid and sulphuric acids. But when a metal reacts with any of these acids, a salt is formed and hydrogen gas is evolved. The metal replaces the hydrogen atoms in the acid to form a salt.

Metal + dilute acid Metal salt + Hydrogen

(a) Sodium reacts with dilute acid with explosive violence

$$\underset{\text{Sodium metal}}{2Na(s)} + \underset{\text{Hydrochloric acid}}{2HCl(aq)} \longrightarrow \underset{\text{Sodium chloride}}{2NaCl\,(aq)} + H_2(g)$$

This reaction shows that sodium is a very reactive metal.

(b) Magnesium metals reacts rapidly with dilute hydroxhloric acid to form magnesium chloride and hydrogen gas.

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

(c) Zn reacts with dil. HCl, but less rapidly than Mg.

$$Zn(s) + 2HCl(aq) \longrightarrow \underset{\text{Zinc chloride}}{ZnCl_2(aq)} + H_2(g)$$

This shows that zinc is less reactive than magnesium

(d) Iron react very slowly with dil HCl to form ferrous chloride and hydrogen.

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

(e) $Cu(s) + HCl(aq)$ No reaction. Thus, order of reactivity is

$$\boxed{Na > Mg > Zn > Fe > Cu}$$

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 is itself reduced to any of oxides of nitrogen (N_2O, NO, NO_2). But Mg and Mn react with very dilute HNO_3 to evolve H_2 gas.

$$Mn(s) + 2HNO_3(aq)\ \ Mn\ (NO_3)_2\ (aq) + H_2(g)$$
(Very dilute)

The rate of formation of bubbles was the fastest in the case of magnesium. The reaction was also the most exothermic in this case. The reactivity decreases in the order $Mg > Al > Zn > Fe$. In the case of copper, no bubbles were seen and the temperature also remained unchanged. This show that copper does not react with dilute HCl.

(IV) Reaction with Chlorine

Most of the metals react with chlorine to form chlorides. These chlorides are ionic (or electrovalent) in character. During the formation of these chlorides, metal loses electrons and becomes positively charged whereas chlorine atoms accept electrons and become negatively charged ions (chloride ions). During this reaction metal undergoes oxidation whereas chlorine undergoes reduction. Some examples are given below:

$$\underset{\text{Sodium}}{2Na(s)} + Cl_2(g) \longrightarrow \underset{\text{Sodium chloride}}{2NaCl(s)};$$

$$\underset{\text{Calcium}}{Ca(s)} + Cl_2(g) \longrightarrow \underset{\text{Calcium chloride}}{CaCl_2(s)}$$

$$\underset{\text{Magnesium}}{Mg(s)} + Cl_2(g) \longrightarrow \underset{\text{Magnesium chloride}}{MgCl_2(s)}$$

$$\underset{\text{Zinc}}{Zn(s)} + Cl_2(g) \longrightarrow \underset{\text{Znic chloride}}{ZnCl_2(s)}$$

(V) Reaction with Hydrogen

Most of the metals do not form compounds with hydrogen because metals form compounds by loss of electrons which are accepted by the other element. But hydrogen usually forms compounds with other elements by loss of electrons or by sharing of electrons. It does not accept electrons. However, a few reactive metals such as sodium, potassium and calcium react with hydrogen to form ionic hydrides.

$$\underset{\text{Sodium}}{2Na(s)} + H_2(g) \longrightarrow \underset{\text{Sodium hydride}}{2NaH(s)};$$

$$\underset{\text{Potassium}}{2K(s)} + H_2(g) \longrightarrow \underset{\text{Potassium hydride}}{2KH(s)}$$

$$\underset{\text{Calcium}}{Ca(s)} + H_2(g) \longrightarrow \underset{\text{Calcium hydride}}{CaH_2(s)}$$

These hydrides are highly unstable compounds. They are decomposed with water.

$$NaH(s) + H_2O(l) \quad NaOH(aq) + H_2(g)$$

(VI) Reaction of Metals with Solutions of Other Metal Salts

Reactive metals can displace less reactive metals from their compounds in solution or molten form.

We have seen in the previous sections that all 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.

It is simple and easy if metal A displaces metal B from its solution,

it is more reactive than B.

Metal A + Salt solution of B Salt solution of A + Metal B

Certain metals have the capacity to displace some metals from the aqueous solution of their salts. These reaction are known as metals displacement reactions. It may be noted that

a metal placed higher in the activity series can displace the metal which occupies a lower position form the aqueous solution of its salt or a less reavtive metal can displace more reactive metal from its salt solution.

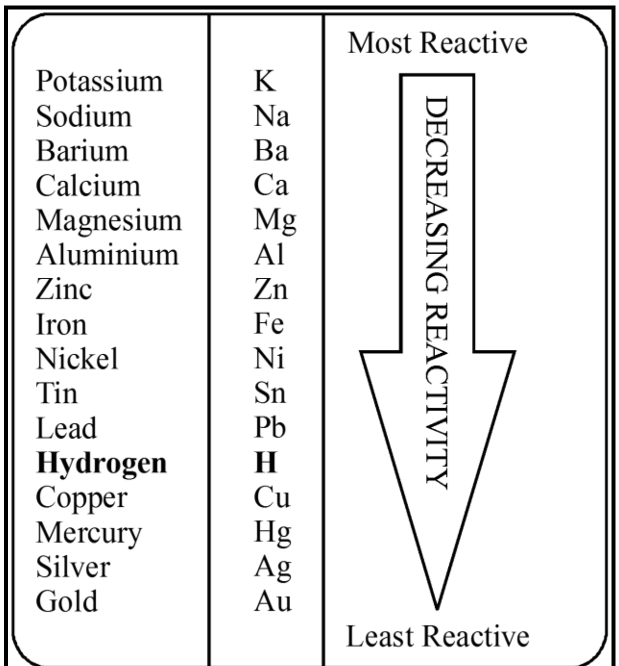
For example, if we look at the activity series, we find that zinc occupies a much higher position in the activity series as compared to copper. It is expected to displace copper present in the aqueous solution of its salt (e.g. $CuSO_4$)

$$Zn(s) + CuSO_4(aq) \quad ZnSO_4(aq) + Cu(s)$$

Activity/Reactivity Series of Metals

The reactivity of metals differs from metal to metal. Some of the metals are reactive, while others are less reactive towards chemical reagents. The elements that can lose electrons easily and form positively charged ions are more reactive. The elements that cannot lose electrons easily are less reactive.

Metals can be arranged in the decreasing order of their reactivity in a series. This series is called the reactivity or activity series of metals. The series has been derived from the reactions discussed above and many other similar reactions.



Activity series of some metals

1. The activity series of metals provides a list of metals arranged in order of their decreasing chemical activity. The most active metal, potassium is at the top of the list and the least active metal, gold, is at the bottom.
2. The ease with which a metal loses electrons and forms positive ions in solution, decreases as we go do down the activity series from potassium to gold.
3. Hydrogen is included in the activity series even though it is not a metal. The hydrogen ion (H^+), like other metal ions, has a positive charge in most chemical reactions.

Significance of Activity Series

1. The metals above hydrogen in the activity series have greater tendency than hydrogen to give up electrons in their solutions. Such metals are called electropositive metals.

   For example, potassium (K), the first metal in the series is the most electropositive, while gold (Au), the last metal of the series is the least electropositive.

2. The metals above hydrogen in the series can liberate hydrogen when treated with an acid solution. Thus, magnesium and zinc react with dilute solutions of sulphuric acid to produce hydrogen gas.

$$Mg + H_2SO_4 \qquad MgSO_4 + H_2;$$

$$Zn + H_2SO_4 \qquad ZnSO_4 + H_2$$

3. A more electropositive metal can replace a less electropositive metal from the solution of a salt of the less electropositive metal. For example, when an iron rod is dipped into a solution of copper sulphate, reddish coloured copper is deposited on the iron rod.

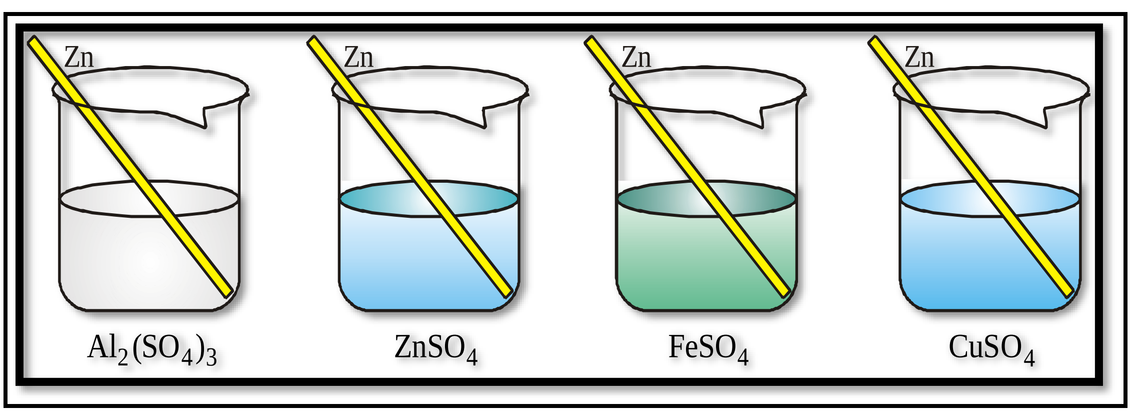
$$Fe + CuSO_4 \qquad FeSO_4 + Cu$$

   This is because iron is more electropositive than copper.

Example :

Four students A, B, C and D noted that initial colour of the solutions in beakers I, II, III and IV.

After inserting zinc rods in each solution and leaving it undisturbed for two hours, noted the colour of each solution again.



They recorded their observations in the form of table given below:

Student	Colour of the solution	I	II	III	IV
A	Initial Final	Colourless Colourless	Colourless Colourless	Light green Colourless	Blue Colourless
B	Initial Final	Colourless Colourless	Light Yellow Colourless	Light green Light green	Blue Colourless
C	Initial Final	Colourless Light blue	Colourless Colourless	Light green Colourless	Blue Light Blue
D	Initial Final	Light green Colourless	Colourless Colourless	Light green Dark green	Blue Colourless

Which students made the correct observation in all the four beakers?

Solution :

The student A had recorded the correct observations in all the four beakers.

(i) In beaker I, zinc is not in a position to displace aluminium metal, since it is placed below it. Therefore, the solution will remain colourless.

(ii) In beaker II, zinc is in contact with $ZnSO_4$. Therefore, no chemical reaction is possible. The solution will remain colourless.

(iii) In beaker III, zinc will displace iron present in ferrous sulphate solution (light green) to form zinc sulphate which is colourless.

(iv) In beaker IV, zinc will displace copper from copper sulphate solution (blue) to form zinc sulphate which is colourless.

From the above discussion, it is quite clear that the student A has made correct observations initially and also after the experiment.

(VII) Reaction of Metals with Non-Metals

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.

We know that a 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^+.

On the other hand chlorine has seven electrons in its outermost shell and it requires one more electron to complete its octet. If sodium and chlorine were to react, the electrons 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 $C1^-$.

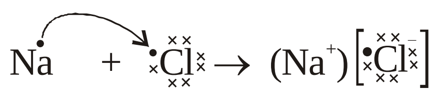
So both these elements can have a give and take relation between them as follows.

$$\underset{2,\,8,1}{Na} \longrightarrow \underset{2,\,8}{Na^+} + e^-$$

(Sodium cation)

$$\underset{2,\,8,7}{Cl} + e^- \longrightarrow \underset{2,8,8}{Cl^-}$$

(chloride anion)



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). It should be noted that sodium chloride does not exist as molecules but aggregates of oppositely charged ions.

Let us see the formation of one more ionic compound, magnesium chloride

$$\underset{2,8,2}{Mg} \qquad \underset{2,8}{Mg^{2+}} + 2e^-$$

(Magnesium cation)

$$\underset{2,8,7}{Cl} + e^- \qquad \underset{2,8,8}{Cl^-}$$

(Chloride anion)