Metals & Non-Metals
Chemical Properties of Metal, Ionic Compounds

❖ 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)

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

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

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

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

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.

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

(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. $CuSO_4$)

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

❖ Ionic Compounds

Ionic compound. Ionic compounds are **formed between metals and non-metals.** The metals have a tendency to lose their valence electrons. So, they combine with the non-metals which have a tendency to gain electron(s) and form ionic bonds.

Ionic Compound Definition: An ionic compound is **a compound formed by ions bonding together through electrostatic forces**. Examples: Table salt, NaCl, is an ionic compound.

- Binary Ionic Compounds Containing a Metal and a Nonmetal. A binary compound is a compound formed from two different elements. ...
- Ionic Compounds Containing a Metal and a Polyatomic Ion. Metals combine with polyatomic ions to give ionic compounds. ...
- Acids and Acid Salts. ...
- Binary Compounds Between Two Nonmetals.