TOPIC

4

SEMESTER-I (Period-1)

C12CH4

Chemistry of Selected Metals and their Compounds

LEARNING OBJECTIVE

Upon completion of this topic, learners will:

- Describe the metallurgy, properties as well as uses of Al, Fe, Cu, Au and Sn
- Discuss alloys including the common uses of brass, bronze, steel, and duralumin and Analyze the production processes of cement and uses.

4.1 PROPERTIES OF SODIUM AND ITS COMPOUND

Sodium

Properties of Sodium

Sodium is a soft, silvery white metal, lighter than water. It can be easily cut with knife. It shines like silver when freshly cut. Its specific gravity or density is 0.97. It is a good conductor of heat and electricity. Its melting point is 97.5°C and boiling point is 892°C. it dissolves in liquid ammonia to give an intense blue solution. It is very reactive metal. It is rapidly tarnished on exposure to moist air first due to the formation of Na_2O which then rapidly changed into NaOH by the action of moisture and finally gets converted into Na_2CO_3 by the action of CO_2, if present. Thus, sodium is kept in kerosene.

$$4Na + O_2 \rightarrow 2N_2O$$

$$Na_2O + H_2O \rightarrow 2NaOH$$

$$2NaOH + CO_2 \rightarrow Na_2CO_3 + H_2O$$

Sodium burns violently in air or oxygen with a golden yellow flame forming a mixture of sodium monoxide and sodium peroxide.

$$4Na + O_2 \rightarrow 2Na_2O \text{ (Monoxide)}$$

$$2Na + O_2 \rightarrow Na_2O_2 \text{ (Peroxide)}$$

It decomposes water vigorously to liberate hydrogen and forms hydride when heated with dry hydrogen at 360°C.

$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$

$$2Na + H_2 \rightarrow 2NaH$$

It readily displaced hydrogen from acids, because it is strongly electropositive element.

$$2Na + 2HCl \rightarrow 2NaCl + H_2$$

$$2Na + 2HNO_3 \rightarrow 2NaNO_3 + H_2$$

It acts as a reducing agent, because of the ease with which it can lose its outermost electron. ($Na \rightarrow Na^+ + e$)

$$Al_2O_3 + 6Na \rightarrow 2Al + 3Na_2O$$

$$3CO_2 + 4Na \rightarrow 2Na_2CO_3 + C$$

$$SiO_2 + 4Na \rightarrow Si + 2Na_2O$$

$$AlCl_3 + 3Na \rightarrow Al + 3NaCl$$

$$BeCl_2 + 2Na \rightarrow Be + 2NaCl$$

$$B_2O_3 + 6Na \rightarrow 2B + 3Na_2O$$

When sodium is heated with ammonia at about 300–400°C, sodamide is formed.

$$2Na + 2NH_3 \rightarrow 2NaNH_2 + H_2$$

The blue coloured solution of sodium in ammonia is also used as a reducing agent.

When heated with mercury, it forms amalgam of varying composition, e.g., NaHg, Na_2Hg, Na_3Hg.

Uses

(1) it is used in the manufacture of a number of chemicals such as Na_2O_2, NaOH, NaCN, $NaNH_2$ etc.
(2) In industry it is used in the production of dyes, drugs, perfumes, and artificial rubber, etc.
(3) As reducing agent in the laboratory.
(4) Sodium amalgam is used as reducing agent in organic chemistry.
(5) In the detection of N, S and halogens in an organic compound.
(6) In petroleum refining and preparation of pharmaceuticals.
(7) Because of its high thermal conductivity, it is used for filling hollow valves of aircraft engines.

Sodium Nitrate ($NaNO_3$)

Preparation

1. $NaHCO_3 + HNO_3 \rightarrow NaNO_3 + H_2O + CO_2$
2. $Na_2CO_3 + 2HNO_3 \rightarrow 2NaNO_3 + CO_2 + H_2O$
3. $NaOH + HNO_3 \rightarrow NaNO_3 + H_2O$
4. $NH_4NO_3 + NaOH \rightarrow NH_4OH + NaNO_3$

Properties

(i) Sodium nitrate is a crystalline solid which is white.
(ii) It has two crystal structures rhombohedral and trigonal.
(iii) This compound has a sweet odour.
(iv) The solubility of $NaNO_3$ in water corresponds to 91.2 g/100 mL at a temperature of 25°C.
(v) This compound is also highly soluble in ammonia.
(vi) When dissolve in water, $NaNO_3$ dissociate into Na^+ and NO_3^-.
(vii) It is a very strong oxidizing agent, it reacts violently with reducing agents.
(viii) At high temperature, the compound is known to explosively decompose.

Uses

1. Hybrid forms of aqua regia can be prepared with the help of $NaNO_3$. These bybrids also have the ability to dissolve gold.
2. This compound is widely used as a food additive since it acts as a preservative.
3. Sodium nitrate is used as an oxidizer in several types of fireworks.
4. It is also a component of some instant cold packs.
5. $NaNO_3$ is one of the components used for the storage and transfer of heat in some solar power plants
6. In order to promote the growth of Nitrosomonas bacteria, this compound is added to the wastewater in several wastewater treatment plants.
7. Sodium nitrate is also used in several rocket propellants and is known to be a substitute.

Sodium Sulphate (Na_2SO_4)

Preparation

1. $NaNO_3 + H_2SO_4 \rightarrow NaHSO_4 + HNO_3$
   Chile salt petre
   $NaCl + NaHSO_4 \rightarrow Na_2SO_4 + HCl$
2. $2NaCl + H_2SO_4 \rightarrow Na_2SO_4 + 2HCl$

Properties:

(i) It is a crystalline solid, soluble in water. The decahydrate when exposed to dry air shows efflorescence, i.e., loses water molecules of crystallization to air and forms anhydrous salt.

$$Na_2SO_4.10H_2O \xrightarrow[\text{Dry air}]{} \underset{\text{Anhyd. salt}}{Na_2SO_4} + 10H_2O$$

(ii) ***Reaction with H_2SO_4:*** When a solution having equivalent quantity of anhydrous sodium sulphate and conc. H_2SO_4 is cooled, sodium bisulphate in the form of prismatic crystals is separated.

$$\underset{\text{Conc.}}{Na_2SO_4(aq) + H_2SO_4} \xrightarrow{\text{cooled}} \underset{\substack{\text{Sod. bisulphate}\\ \text{(Prismatic crystals)}}}{2NaHSO_4}$$

(iii) It precipitates the insoluble sulphates of lead, barium, strontium, calcium, etc., from their soluble salt solutions:

$$Pb(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow 2NaNO_3 + \underset{\text{White ppt.}}{PbSO_4}\downarrow$$

Uses:

(i) In medicine as a purgative.
(ii) In the manufacture of craft paper, (brown paper for wrapping) paper board, soda glass, etc.
(iii) In the preparation of sodium sulphide, hypo, $NaHSO_4$, etc.

Sodium Chlorate ($NaClO_3$)

Preparation

$$NaCl + 3H_2O \rightarrow NaClO_3 + 3H_2$$

Properties

1. It is an odourless compound.
2. Its colour differs from light yellow to white crystalline solid.
3. It is very soluble in water and heavier than water. Hence, it can sink and break up at a fast rate.
4. While it is not an explosive by itself, yet it can cause powerful combustion on coming in contact with water.
5. It causes a highly exothermic reaction.
6. Powerful oxidizing reaction due to their inherent properties.
7. Sodium chlorate's boiling point is 300°C and the melting point in 248°C.
8. It is also soluble in some organic solvents like glycerol and methanol.
9. It is a powerful oxidizing agent. The hypochlorite ion helps to oxidize and bleach any chemical.
10. Its reaction with potassium bromide and hydrochloric acid gives products as potassium chloride, sodium chloride, bromine, and water.

$$NaClO_3 + 6KBr + 6HCl \rightarrow 6KCl + NaCl + 3H_2O + 3Br_2$$

11. Its chemical reaction with potassium iodide and hydrochloric acid produces products as sodium chloride, potassium chloride, iodine, and water.

$$NaClO_3 + 6KI + 6HCl \rightarrow NaCl + 3I_2 + 3H_2O + 6KCl$$

Uses

1. It is used for making herbicides, explosives, colours matches, inks, beautifiers, pharmaceuticals, defoliants, paper, and calfskin.
2. It is used as a dying mash in the creation of the paper.
3. Used in the Solvay process which uses up the salt along with H_2SO_4, and the presence of CH_3OH as the lessening specialist.
4. It is used as an oxidizing specialist in the large dye-making processes and an oxidizing and bleaching operator.
5. It is used in the medical field for preparing different drugs.
6. It is used in making fertilizers and explosives.

SOLVED QUESTIONS

Q1. Why alkali metals not found in nature?

Ans: Alkali metals are highly reactive hence they do not occur in the free state.

Q2. Why Na is less reaction than K?

Ans: Ionization enthalpy of K is less than Na.

Q3. When an Sodium dissolve in liquid ammonia the solution acquires different colours. Explain the reason for this type of colour change

Ans: When an alkali metal is dissolved in liquid ammonia it produces a blue coloured conducting solution due to formation of ammoniated cation and ammoniated electron as given below:

$$\underset{\text{Almali metal}}{Na} + (x + y)NH_3(l) \rightarrow \underset{\text{Ammoniated cation}}{[Na(NH_3)_x]^+} + \underset{\text{Ammoniated electron}}{[e(NH_3)_y]^-}$$

When the concentration is above 3M, the colour of solution is copper-bronze. This colour change is because the concentrated solution contains clusters of metal ions and hence possess metallic lustre.

Calcium

Properties:

(i) Calcium is a moderately hard, silvery white metal. Its melting point is 851°C and boiling point is 1480°C. It is malleable and ductile and is a good conductor of electricity.

(ii) *With air*: On heating it burns in air forming the oxide and a little calcium nitride.

$$2Ca + O_2 \xrightarrow{\text{Burning}} 2CaO$$

$$3Ca + N_2 \xrightarrow{\text{Burning}} Ca_3N_2$$

(iii) With hydrogen:

$$\underset{\text{Hot}}{Ca} + H_2 \xrightarrow{\Delta} \underset{\substack{\text{Calcium hydride}\\ \text{(Hydrolith)}}}{CaH_2}$$

(iv) *With carbon*:

$$Ca + 2C \xrightarrow{\Delta} \underset{\text{Calcium carbide}}{CaC_2}$$

(v) *With CO_2*:

$$5Ca + 2CO_2(g) \xrightarrow[\text{heated}]{} 4CaO + CaC_2$$

(vi) *With chlorine*:

$$Ca + Cl_2 \rightarrow CaCl_2$$

(vii) *With water*: Ca decomposes cold water.

$$Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2\uparrow$$

(viii) *With ammonia*: Calcium metal absorbs ammonia gas forming $Ca(NH_3)_6$ which on heating forms calcium amide.

$$Ca + 6NH_3 \rightarrow Ca(NH_3)_6$$

$$Ca(NH_3)_6 \xrightarrow{\Delta} \underset{\substack{\text{Calcium}\\ \text{amide}}}{Ca(NH_2)_2} + 4NH_3 + H_2$$

The amide on further strong heating forms calcium nitride.

$$3Ca(NH_2)_2 \rightarrow Ca_3N_2 + 4NH_3$$

(ix) *Reducing action*:

$$Cr_2O_3 + 3Ca \xrightarrow{\Delta} 3CaO + 2Cr$$

Uses:

(i) As a drying agent for the preparation of absolute alcohol.
(ii) In the preparation of calcium hydride and metallic uranium.
(iii) For removing last traces of air from noble gases and to obtain higher vaccum.
(iv) As a deoxidant for copper, cast iron and steel and for removing sulphur from petroleum.
(v) For the preparation of beryllium, chromium, vanadium, thorium by the reduction of their compounds.

4.2 CALCIUM CARBONATE (LIMESTONE), $CaCO_3$

Calcium carbonate occurs in nature in different forms such as *limestone, marble* and *chalk*. While the first two are hard rocks, the

third one is a soft white amorphous mass. It also occurs as *dolomite* in combination with magnesium carbonate.

In laboratory, $CaCO_3$ can be obtained by passing CO_2 gas through lime water or by adding sodium carbonate to a solution of some soluble salt of calcium.

$$Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 \downarrow + H_2O$$

$$CaCl_2 + Na_2CO_3 \longrightarrow CaCO_3 \downarrow + 2NaCl$$

The product thus obtained is known as *precipitated chalk.*

Calcium carbonate is a white fluffy powder which is almost insoluble in water. Limestone on calcination gives CaO and CO_2.

$$CaCO_3 \longrightarrow CaO + CO_2$$

Uses

1. Limestone is used for the *preparation of lime and cement.*
2. Limestone is also used *as flux during smelting of iron ores.*
3. Marble is used *as building material.*
4. Specially precipitated calcium carbonate is used in the manufacture of high quality paper.
5. Precipitated chalk is used *in medicines and in tooth-pastes.*

4.3 CALCIUM OXIDE (QUICK LIME OR LIME), CaO

Quick lime is prepared by heating limestone ($CaCO_3$) in a special type of kiln at 1070–1270 K.

$$CaCO_3 \rightleftharpoons CaO + CO_2$$

Since the reaction is reversible, carbon dioxide gas is flushed out by a forced draught in order to shift the equilibrium in favour of products.

Calcium oxide is a white amorphous solid which melts at 2870 K. It is extremely stable and does not decompose even when fused. On exposure to atmosphere, it absorbs moisture and carbon dioxide.

$$CaO + H_2O \longrightarrow Ca(OH)_2$$

$$Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 + H_2O$$

Quick lime when heated in oxyhydrogen flame emits a brilliant light (*lime-light*).

Calcium oxide combines with acidic oxides on heating

$$CaO + SiO_2 \longrightarrow \underset{\text{Calcium silicate}}{CaSiO_3}$$

$$6CaO + P_4O_{10} \longrightarrow \underset{\text{Calcium phosphate}}{2Ca_3(PO_4)_2}$$

Uses

1. CaO is used in the manufacture of calcium carbide.

$$CaO + 3C \xrightarrow[\text{Electric furnace}]{3300\ K} CaC_2 + CO$$

2. It is used as building material.
3. It is used in the purification of sugar.
4. It is used in tanning industry.
5. It is used in softening of hard water.
6. It is used in drying of gases and alcohol.
7. CaO is used in preparation of $Ca(OH)_2$, slaked lime, which finds many industrial uses.

4.4 CALCIUM HYDROXIDE (SLAKED LIME), $Ca(OH)_2$

Calcium hydroxide is prepared by adding limited amount of water to quick lime, CaO. The process is known as *slaking of lime*.

$$CaO + H_2O \longrightarrow Ca(OH)_2$$

Calcium hydroxide is a white amorphous powder. It is sparingly soluble in water. A suspension of slaked lime in water is called **milk of lime.** A clear aqueous solution of calcium hydroxide is known as **lime water.**

With chlorine, slaked lime gives bleaching powder.

$$2Ca(OH)_2 + 2Cl_2 \longrightarrow \underbrace{CaCl_2 + Ca(OCl)_2}_{\text{Bleaching powder}} + 2H_2O$$

Uses

1. Milk of lime is used for white washing due to its disinfectant properties.
2. Slaked lime is used in glass making, in tanning industry, in sugar industry and for the manufacture of bleaching powder.

3. Lime water is well known laboratory reagent for the detection of CO_2. When carbon dioxide is passed through lime water, calcium carbonate is formed due to which lime water turns milky.

$$Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 + H_2O$$

On passing more carbon dioxide milkiness disappears due to the formation of calcium bicarbonate.

$$CaCO_3 + H_2O + CO_2 \longrightarrow Ca(HCO_3)_2$$

The clear solution on heating again gives milkiness due to decomposition of calcium bicarbonate into calcium carbonate.

$$Ca(HCO_3)_2 \longrightarrow CaCO_3 + CO_2 + H_2O$$

4. Slaked lime is used in the manufacture of calcium hydrogen sulphate, $Ca(HSO_4)_2$ which is used in paper industry.
5. Slaked lime is used as building material. **Mortar** is formed by adding water to a mixture of sand and slaked lime to form a thick paste. As this mixture dries, it hardens due to reaction between $Ca(OH)_2$ and atmospheric CO_2 to form $CaCO_3$.

$$Ca(OH)_2(s) + CO_2(g) \longrightarrow CaCO_3(s) + H_2O(g)$$

4.5 CALCIUM SULPHATE, $CaSO_4$

Calcium sulphate occurs in nature as *anhydrite,* $CaSO_4$ and as *gypsum*, $CaSO_4.2H_2O$.

Calcium sulphate is sparingly soluble in water.

Uses

1. Gypsum is used for making *plaster of paris*.
2. Gypsum is also used for the treatment of soil.
3. Anhydride is used for the manufacture of sulphuric acid.

4.6 PLASTER OF PARIS, HEMIHYDRATE OF CALCIUM SULPHATE, $(CaSO_4)_2H_2O$

It is prepared by heating gypsum ($CaSO_4.2H_2O$) to 393 K.

$$2CaSO_4.2H_2O \xrightarrow{390\ K} (CaSO_4)_2H_2O + 3H_2O$$

The temperature should not be allowed to rise above 393 K because above this temperature the whole of water of crystallisation is lost. The resulting anhydrous $CaSO_4$ is called *dead burnt plaster* because it loses the property of setting with water.

Uses

1. On mixing with water it changes into plastic mass and solidifies due to rehydration. This is called *setting of plaster of paris*. During the process of setting, it undergoes slight expansion (about 1%). Consequently, it produces a very sharp impressions of the mould into which it is put. Therefore, it is used for producing moulds for industries such as pottery, ceramics.
2. For setting broken or fractured bones in the body.
3. For making statues, models and other decorative material.

Calcium Chloride ($CaCl_2$)

Calcium chloride, $CaCl_2.6H_2O$: Calcium chloride is present in small amounts in sea water

Preparation:

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O + CO2$$

$$\downarrow \text{ Crystallization}$$

$$CaCl_2.6H_2O$$

Properties:

(i) $CaCl_2.6H_2O$ is a colourless, crystalline, highly deliquescent solid melting at 772°C. It is extremely soluble in water.

(ii) *Effect of heat*: When heated at 200°C, $CaCl_2.6H_2O$ gives a white porous mass of $CaCl_2.2H_2O$. On fusion anhydrous $CaCl_2$ is left as residue.

$$CaCl_2.6H_2O(s) \xrightarrow{200°C} CaCl_2.2H_2O + 4H_2O$$

$$\downarrow \text{ Strong heating}$$

$$CaCl_2 \text{ (Anhydrous)}$$

(iii) The anhydrous salt is extremely hygroscopic but absorbs ammonia also giving the compound $CaCl_2 8NH_3$.

$$\underset{\text{Anhydrous}}{CaCl_2} + 8NH_3 \rightarrow CaCl_2.8NH_3$$

(iv) $CaCl_2$ readily dissolves in alcohol by forming $[Ca(C_2H_5OH)_4]Cl_2$. Thus, is cannot be used for drying alcohol.

(v) A mixture of $CaCl_2$ and ice has a very low temperature (– 54°C).

Uses:

(1) Fused, anhydrous calcium chloride is used as a desiccating agent.

(2) In medicine, in freezing mixtures.
(3) Anhydrous $CaCl_2$ is used for preparing metallic calcium.
(4) It is used to keep cotton thread moist during spinning.
(5) For sprinkling on the roads for prevention of dust.

SOLVED QUESTIONS

Q1. *Chlorination of calcium hydroxide produces bleaching powder. Write its chemical equation.*

Ans. Bleaching powder is obtained by passing Cl_2 into $Ca(OH)_2$. Though bleaching powder is often written as $Ca(OCl)_2$ it is actually a mixture.

$$3Ca(OH)_2 + 2Cl_2 \longrightarrow \underset{\text{Bleaching powder}}{Ca(OCl)_2 \,.\, Ca(OH)_2 \,.\, CaCl_2 \,.\, 2H_2O}$$

Q2. *What is quick lime, slaked lime and lime water? What happens when carbon dioxide gas is passed through lime water?*

Ans. Quick lime is calcium oxide, CaO.

Slaked lime is calcium hydroxide $Ca(OH)_2$.

Lime water is a clear solution of calcium hydroxide in water.

When carbon dioxide gas is passed through lime water, it becomes milky due to formation of calcium carbonate.

$$Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 \downarrow + H_2O$$

If carbon dioxide gas is passed in excess, the milkiness disappears due to formation of $Ca(HCO_3)_2$, which is soluble in water.

$$CaCO_3 + CO_2 + H_2O \longrightarrow Ca(HCO_3)_2.$$

Q3. *How would you distinguish between magnesium and calcium?*

Ans. Mg, when heated in a flame does not impart any characteristics colour to the flame whereas calcium inparts brick red colour to the flame.

Reactivity of Aluminium

(i) Action of air: It is not affected by dry air. In contact with moist air, a film of oxide is formed at the surface which acts as a protective coating against further corrosion.

On heating in air, it burns readily giving a brilliant light and large amount of heat is liberated.

$$4Al + 3O_2 \xrightarrow{\text{heated}} 2Al_2O_3, \Delta H = -3230 \text{ kJ}$$

(ii) *Action of water*: Pure aluminium remains unaffected by pure water. Impure aluminium is readily corroded by water containing dissolved salts.

It decomposes boiling water giving out hydrogen gas.

$$2Al + 6H_2O \xrightarrow[\text{Boiling}]{} 2Al(OH)_3 + 3H_2$$

(iii) *Reaction with acids*: Aluminium dissolves in dilute HCl evolving hydrogen gas.

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

It reacts with hot sulphuric acid to give out SO_2 gas.

$$2Al + 6H_2SO_4 \longrightarrow Al_2(SO_4)_3 + 6H_2O + 3SO_2$$

Nitric acid of any concentration has no action with aluminium. The metal becomes passive in contact with HNO_3 due to surface oxidation.

Reactivity of Iron

(i) *Reaction with air*: On strong heating iron begins to burn in oxygen forming ferrosoferric oxide.

$$3Fe + 2O_2 \xrightarrow{\text{heat}} Fe_3O_4$$

(ii) *Reaction with water*: Red hot iron decomposes steam forming Fe,O_4 and hydrogen.

$$3Fe + 4H_2O \xrightarrow[\text{Red hot-}\ \text{Steam}]{} Fe_3O_4 + 3H_2$$

(iii) Reaction with acids: Standard oxidation potential of iron is + 0.44 volt. Hence, it dissolves in dilute acids evolving hydrogen gas.

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

$$Fe + H_2SO_4 \xrightarrow[\text{Dil.}]{} FeSO_4 + H_2$$

Dilute HNO_3 reacts with iron as follows:

$$3Fe + 8HNO_3 \xrightarrow{\text{Hot}} 3Fe(NO_3)_2 + 2NO + 4H_2O$$

$$4Fe + 10HNO_3 \xrightarrow[\text{Dil.}]{\text{cold}} 4Fe(NO_3)_2 + NH_4NO_3 + 3H_2O$$

Properties of Copper

1. It is a reddish, brown coloured shining metal.
2. It melts at 1083C and boils at 2325°C.
3. Its specific gravity is 8.94.
4. It is the best electrical conductor next to silver.
5. It is malleable, ductile and tenacius metal.
6. Molten copper absorbs sulphur dioxide which is given out on cooling producing blisters on the surface of the metal.
7. *Action of Air*: Copper is not acted upon by dry air. When exposed to moist air for a long time, a green layer of basic copper carbonate is deposited on the surface.

$$2Cu + H_2O + CO_2 + O_2 \rightarrow \underset{\text{Basic copper carbonate (green)}}{CuCO_3.Cu(OH)_2}$$

(Moisture) (air)

$$2Cu + O_2 \xrightarrow[\text{below 1100°C}]{\Delta} \underset{\text{cupric oxide}}{2CuO}$$

$$4Cu + O_2 \xrightarrow[\text{above 1100°C}]{\Delta} \underset{\text{Cupric oxide}}{2Cu_2O}$$

8. *Reaction with acids*: Copper does not react with non-oxidising acids (for example, dil. HCl and dil. H_2SO_4). However, it is dissolved in these acids in the presence of air (oxygen).

$$Cu + H_2SO_4 + O_2 \rightarrow CuSO_4(aq) + H_2O$$

dil. (air)

$$Cu + 2HCl + O_2 \rightarrow CuCl_2(aq) + H_2O$$

conc. air

$$Cu + 2H_2SO_4 \xrightarrow[\text{conc.}]{\Delta} CuSO_4 + 2H_2O + SO_2 \uparrow$$

Copper reacts with HNO_3 (Oxidising acid) under different conditions to give different products.

$$Cu + \underset{\text{(conc. and cold)}}{4HNO_3} \rightarrow Cu(NO_3)_2 + \underset{\text{Brown}}{2NO_2}\uparrow + 2H_2O$$

$$3Cu + \underset{\text{(50\% nad cold)}}{8HNO_3} \rightarrow 3Cu(NO_3)_2 + \underset{\text{(Nitric oxide)}}{2NO} + 4H_2O$$

9. *Reaction with chlorine and sulphur*:

$$\underset{\text{hot}}{Cu} + Cl_2 \xrightarrow{\Delta} \underset{\text{(copper (II) chloride)}}{CuCl_2}$$

$$Cu + S \xrightarrow{\Delta} \underset{\text{copper sulphide}}{CuS}$$

10. *Reaction with ferric salt*:

$$Cu(s) + 2\underset{\text{(ferric ion)}}{Fe^{3+}}(aq) \longrightarrow Cu^{++}(aq) + 2\underset{\text{(ferrous ion)}}{Fe^{2+}}$$

11. *Aciton of aqueous ammonia*: Metallic copper is dissolved in NH_4OH in the presence of air forming the complex salt tetrammine copper (II) hydroxide.

$$Cu + H_2O \ {}^*O_2 \rightarrow Cu(OH)_2$$

(air)

$$Cu(OH)_2 + 4NH_3 \rightarrow \underset{\text{Deep blue complex}}{[Cu(NH_3)_4]\ (OH)_2}(aq)$$

12. *Action of water*: Water has no action on copper. When steam is passed over white hot copper. Cu2O is formed.

$$\underset{\text{White hot}}{2Cu} + \underset{\text{Steam}}{H_2O} \rightarrow Cu_2O + H_2$$

Uses of Copper

1. Copper forms many useful alloys with metals such as Zn, Sn, Ni, etc.
2. It is extensively used for making electric cables and other electric appliances.
3. For electroplating and electrotyping.
4. For making utensils.
5. Copper is alloyed with gold and silver for making ornaments.

Compound of Copper

Copper (II) sulphate or cupric sulphate $CuSO_4.5H_2O$

Preparations

1. $Cu + H_2SO_4 + ½O_2 \rightarrow CuSO_4(aq) + H_2O$

$$\downarrow \text{ crystallization}$$

$$CuSO_4.5H_2O$$

2. $CuO + H_2SO_4 \rightarrow CuSO_4(aq) + H_2O$

Properties:

(i) It is a blue colourd crystalline solid. It crystallizes in triclinic lattice. It is soluble in water. It crystallizes with five molecules of water of crystallization.

(ii) *Effect of heat*: On heating at 100°C, it loses four, out of five molecules of water to form monohydrate. It on further heating at 230°C gives anhyd. $CuSO_4$ Anhydrous $CuSO_4$ is a non-crystalline white powder.

$$\underset{\text{Pentahydrate (Blue)}}{CuSO_4.5H_2O} \xrightarrow[-4H_2O]{100°C} \underset{\text{Monohydrate (Bluish white)}}{CuSO_4.H_2O} \xrightarrow{230°C} \underset{\text{Anhydrous salt white power}}{CuSO_4}$$

Anhydrous copper sulphate (white powder) is used for testing water. Water turns the white anhydrous salt into blue coloured crystalline salt.

$$\underset{\text{White powder}}{CuSO_4} \xrightarrow{\text{Water}} \underset{\text{Blue, pentahydrate crystalline salt}}{CuSO_4.5H_2O}$$

Anhydrous $CuSO_4$ decomposes at 750°C:

$$2CuSO4 \xrightarrow[750°C]{\Delta} \underset{\text{Cupric oxide (Black)}}{2CuO(s)} + 2SO_2 \uparrow + O2 \uparrow$$

(iii) *Hydrolysis*: Aqueous solution of copper sulphate is acidic due to cationic hydrolysis.

$$CuSO4 \longrightarrow Cu^{++} + SO_2^{-4}$$

$$2H_2O \longrightarrow 2OH^- + 2H^+$$

$$\uparrow\downarrow \qquad \downarrow\uparrow$$

$$\underset{\text{weak base}}{Cu(OH)_2} \quad \underset{\text{strong acid}}{H_2SO_4}$$

Uses:

(i) Used as fungicide
(ii) In electroplating as electrolyte
(iii) In dyeing and calico printing
(iv) In making green pigment
(v) In the preparation of felling solution

Cupric oxide, CuO

It is called black oxide of copper and is found in nature as tenorite.

Preparation:

(i) $Cu(OH)_2 \xrightarrow{\Delta} CuO + H_2O$

(ii) On a commercial scale, it is obtained by heating/calcination of malachite which is found in nature.

$$CuCO_3.Cu(OH)_2 \rightarrow 2CuO + CO_2 + H_2O$$

Properties:

(a) *Physical*:

It is black powder insoluble in water and stable to moderate heating.

(b) *Chemical*:

(i) The oxide dissolves in acids HCl, H2SO4 of HNO3 forming corresponding salts.

$$CuO + 2H^+ \longrightarrow Cu^{2+} + H_2O$$

(ii) When heating to 1100 – 1200C, it is converted into cuprous oxide with evolution of oxygen.

$$4CuO \longrightarrow 2Cu_2O + O_2$$

(iii) It is reduced to metallic copper by reducing agents like hydrogen, carbon and carbon monoxide

$$CuO + H_2 \longrightarrow Cu + H_2O$$

Uses

(i) Used as pigmenting agent in ceramic compounds.
(ii) It is widely used in laboratories for the preparation of various copper salts.
(iii) It is used in the manufacture of wood preservatives
(iv) It is used in the welding process.
(v) Cuo used in the manufacture of lithium batteries.

$CuCl_2.2H_2O$ (Cupric chloride)

Preparation:

1. $2Cu + 4HCl + O_2 \longrightarrow 2CuCl_2 + 2H_2O$
2. $CuO + 2HCl \longrightarrow CuCl_2 + H_2O$
3. $Cu(OH)_2\ CuCO_3 + 4HCl \longrightarrow 2CuCl_2 + 3H_2O + CO_2$

Properties:

Physical

1. It is deliquescent compound and is readily soluble in water.
2. The dilute solution is blue but concentrated solution, is however, green
3. It changes to yellow when concentrated HCl is added. The blue colour is due to complex cation $[Cu(H_2O)_4]^{2+}$ and yellow colour due to complex anion
4. The aqueous solution is acidic due to its hydrolysis
   $$CuCl_2 + 2H_2O \longrightarrow Cu(OH)_2 + 2HCl$$
5. The anhydrous salt on heating forms Cu_2Cl_2 and Cl_2
   $$2CuCl_2 \longrightarrow CuCl_2 + Cl_2$$
   $$3CuCl_2.2H_2O \xrightarrow{\Delta/\text{Strong}} CuO + Cu_2Cl_2 + 2HCl + Cl_2 + 5H_2O$$
6. It is readily reduced to Cu2Cl2 by copper turnings or sulphur dioxide gas or nascent hydrogen.
   $$CuCl_2 + Cu \longrightarrow Cu_2Cl_2$$
   $$2CuCl_2 + SO_2 + 2H_2O \longrightarrow Cu_2Cl_2 + 2HCl + H_2SO_4$$
   $$2CuCl + 2H \longrightarrow Cu_2Cl_2 + 2HCl$$
7. A pale blue precipitate of basic cupric chlroide $CuCl_2.3Cu(OH)_2$ is obtained when NaOH is added.
   $$CuCl_2 + 2NaOH \longrightarrow Cu(OH)_2 + 2NaCl$$
   $$CuCl_2 + 3Cu(OH)_2 \longrightarrow CuCl.3Cu(OH)_2 \downarrow$$

It dissolves in ammonium hydroxoide forming a deep blue solution. On evaporation of this solution deep blue crystals of tetraammine cupric chloride are obtained.

$$CuCl_2 + 4NH_4OH \longrightarrow [Cu(NH_3)_4]\ Cl_2.H_2O + 3H_2O$$

Uses

1. It is used a catalyst for organic and inorganic reaction
2. Textile dlycing and printing mordant

3. Glass and ceramic pigment
4. Wood preservative
5. Disinfectant insecticide and fungicide.

EXERCISE

A. Multiple Choice Questions

In each of the following questions only one option is correct. Select the correct option.

1. The similarity in the properties of alkali metals is due to:
(a) their same atomicity
(b) similar outer shell configuration
(c) same energy of outershell
(d) same principal quantum number of outer shell.

2. The formula of soda ash is:
(a) Na_2CO_3 (b) $Na_2CO_3.5H_2O$
(c) $Na_2CO_3.10H_2O$ (d) $Na_2CO_3.2H_2O$.

3. On dissolving moderate amount of sodium metal in liquid NH_3 at low temperature, which one of the following does not occur?
(a) Blue coloured solution is obtained.
(b) Na^+ ions are formed in the solution.
(c) Liquid ammonia becomes good conductor of electricity.
(d) Liquid ammonia remains diamagnetic.

4. A metal M readily forms water soluble sulphate MSO_4, water insoluble hydroxide $M(OH)_2$ and oxide MO which becomes inert on heating. The hydroxide is soluble in NaOH. Then M is:
(a) Be (b) Mg
(c) Ca (d) Sr.

5. The by-product obtained in Solvay process for the manufacture of sodium carbonate is ..... .
(a) NaCl (b) $CaCl_2$
(c) NaOH (d) $NaHCO_3$.

B. Fill in the Blanks

Fill in the following blanks with appropriate items:

1. Sodium is ..... electropositive than potassium.

2. Most abundant ore of sodium is ..... .

3. Alkaline earth metals are ..... reducing agents than alkali metals.

4. The most abundant alkaline earth metal is ..... .

5. The chemical formula of chile salt petre is ..... .
6. Sodium reacts with excess of oxygen to form ..... .

C. True/False Statements

State whether the following statements are true or false:

1. Alkali metals are generally extracted by electrolysis of their ores.
2. Sodium is more electropositive than magnesium.
3. Melting points of alkaline earth metals are lower than those of corresponding alkali metals.
4. Caesium is the lightest alkali metal.
5. In Castner-Kellner cell, sodium hydroxide is formed in the central compartment.
6. The chemical formula of plaster of paris is $CaSO_4 . 2H_2O$.

D. Answer the following Questions

1. What happens when
   (i) Calcium is burnt in air
   (ii) Quick lime is heated with silica
   (iii) Chlorine reacts with slaked lime
   (iv) Calcium nitrate is heated?
2. What happens when
   (i) Sodium metal is dropped in water?
   (ii) Sodium metal is heated in free supply of air?
   (iii) Sodium peroxide dissolve in water?
3. State as to why sodium is found to be more useful than potassium.
4. *Explain the following:*
   (i) Sodium metal is stored under kerosene
   (ii) *Sodium wire is used to dry benzene but cannot be used to dry ethanol.*
5. *Describe the nature and properties of solutions of alkali metals in liquid ammonia.*
6. *Explain why alkaline earth metals have higher first ionization enthalpies than the corresponding alkali metals.*
7. What happen when copper react with
   (i) 50% cold HNO_3
   (ii) HCl
8. Complete the following reaction
   (i) $Al + 3H_2O \longrightarrow ..... + H_2$
   (ii) $Fe + H_2SO_4 \longrightarrow ..... + H_2$