

## 4. Non-Transition Metals

### 4.1 Introduction

The various stages of man's development are referred to by such terms as 'the stone age', 'the bronze age', 'the iron age', 'the age of steel', etc. Thus, the importance of metals cannot be overemphasised. At present, the level of iron and steel production and consumption in any country can be regarded as the barometer of economic development of the country. In fact, most of the goods and services which we enjoy today require the use of iron and steel in their manufacture.

A cursory look at the periodic table shows that many of the chemical elements are metals. All metals are solid at room temperature except mercury which occurs as a liquid.

### 4.2 Physical Properties of Metals

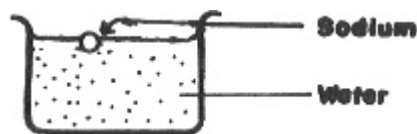
- (i) A typical metal is opaque and can be beaten into sheets (malleable).
- (ii) Metals can be drawn into wires (ductile) and can be polished,
- (iii) A metal is non-volatile, unlike the non-metals.
- (iv) They are generally good conductors of heat and electricity.
- (v) Most metals have relatively high densities with the exception of sodium, lithium, and magnesium.
- (vi) They also have metallic bonds which consist of positive nuclei embedded in seas of delocalised electrons.

### 4.3 Chemical Properties of Metals

#### ***Experiment 4.1:***

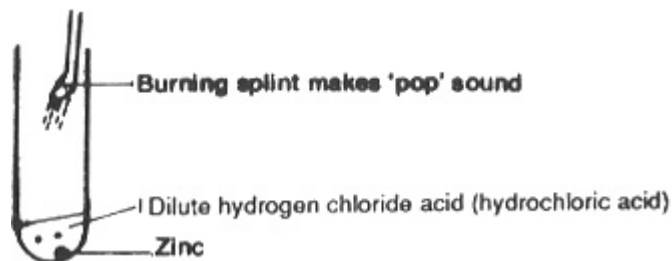
Investigating the Action of common Reactants on Metals:

- (a) Put some water in a beaker, carefully cut a piece of sodium metal and place in the water. Do not restrict the free movement of the piece of sodium. What do you observe? Repeat the above experiment using calcium, magnesium, iron and zinc metals.



*Figure 4.1: Reaction of sodium with water.*

- (b) Put some zinc granules in a test-tube. Add dilute hydrogen chloride acid (hydrochloric acid) to it. Identify the gas given off. Write an equation for the reaction. Repeat the experiment using magnesium ribbon, iron and copper wire. Do you observe any difference in the reactivities of the metals?



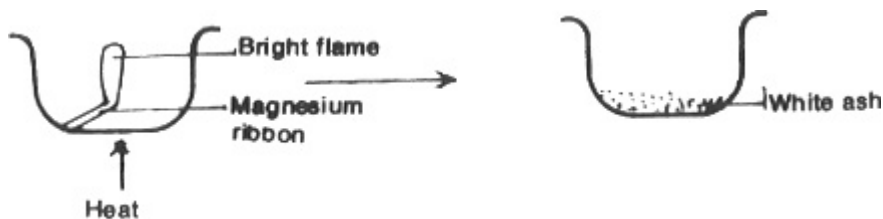
*Figure 4.2: Reaction of zinc with dilute hydrochloric acid*

- (c) Place a piece of sodium metal in an evaporating dish. What do you observe? Repeat the experiment using magnesium and copper.



*Figure 4.3: Exposing sodium to the air.*

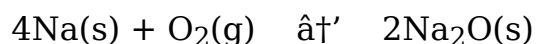
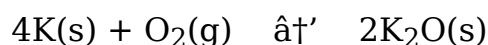
- (d) Put a piece of magnesium ribbon in an evaporating dish. Ignite it. What do you observe? Write an equation for the reaction.

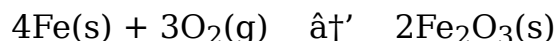
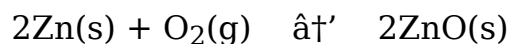
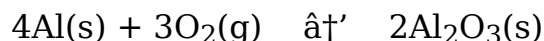
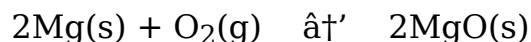
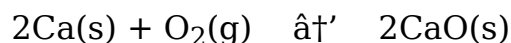


*Figure 4.4: Heating magnesium ribbon in air.*

### **Reaction with Oxygen:**

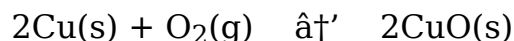
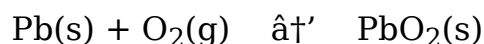
From the above experiments, we can arrange metals in order of their reactivities. Potassium, sodium and calcium burn in air or oxygen to form the corresponding oxides.





Magnesium, aluminium, zinc and iron do not react spontaneously with air or oxygen like the alkali metals, sodium and potassium.

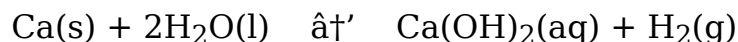
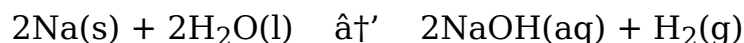
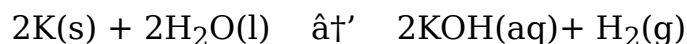
Lead, copper and mercury are oxidised when heated in air.



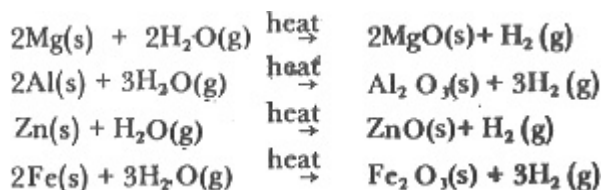
However, silver and gold are unaffected by air or oxygen. The reactions of metals with water follow almost the same pattern as their reactivities with oxygen.

## Reaction with Water

Potassium and sodium react violently with cold water liberating hydrogen. Calcium also decomposes cold water but not as violently as sodium and potassium.

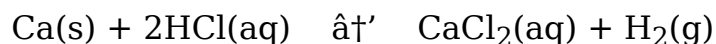


However, magnesium, aluminium, zinc and iron decompose only steam at red heat.



## Reaction with acids

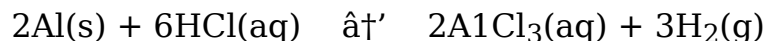
Potassium and sodium react very violently with both dilute and concentrated acids. **Do not react an alkali metal with an acid!** However, calcium reacts rapidly with dilute acids such as dilute hydrogen chloride acid (hydrochloric acid) and dilute trioxonitrate(V)



acid. Calcium will not dissolve in tetraoxosulphate(VI) acid due to the formation of insoluble calcium tetraoxosulphate(VI).

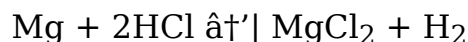
Aluminium is also attacked by dilute hydrogen chloride acid

(hydrochloric acid) but not by trioxonitrate(V) acid (an oxidizing acid). This is because a film of aluminium is formed when an oxidizing acid such as trioxonitrate(V) acid is used, which further reaction.

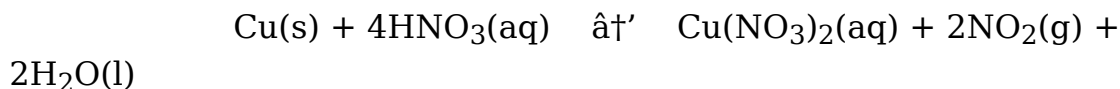
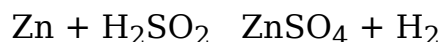


Aluminium is also attacked by hot concentrated tetraoxosulphate(VI)

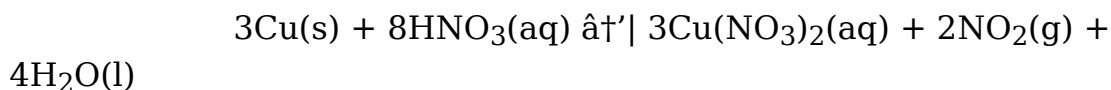
Magnesium, zinc and iron react with dilute acids, liberating hydrogen is a displacement reaction.



Lead, copper and mercury are only attacked by oxidizing acids.



(concentrated)



(dilute)

Silver and gold are not attacked by acids generally.

We can now draw up a table to show reactivity series of the metals:

K	Very reactive
Na	
Ca	
Mg	
Al	
Zn	
Fe	
Pb	
Cu	
Hg	
Ag	
Au.	Very unreactive

↑  
increasing reactivity

## 4.4 Occurrence of Metals

Metals occur naturally in rocks as minerals. Most are chemically combined with non-metals such as oxygen, sulphur, phosphorus and silicon. A few metals such as gold, silver, copper, mercury and members of the platinum family are often found free in nature. A mineral or mixture of minerals is described as an ore if it occurs in such a quantity that it is profitable to extract a metal from it. Naturally, ores are mixed with sandy impurities.

Most ores exist as oxides, trioxocarbonates(IV), trioxonitrates(V), sulphides, tetraoxosulphates(VI), halides, phosphates and silicates of

the metals. Some of these important ores are:

Oxides: Haematite ( $\text{Fe}_2\text{O}_3$ ), bauxite ( $\text{Al}_2\text{O}_3$ ), and cassiterite ( $\text{SnO}_2$ ).

Sulphides: zinc blende ( $\text{ZnS}$ ), galena ( $\text{PbS}$ ), iron pyrites ( $\text{FeS}$ ), copper pyrites ( $\text{CuS}$ ).

Trioxocarbonates(IV): limestone ( $\text{CaCO}_3$ ), calamine ( $\text{ZnCO}_3$ ).

Halides: rock salt ( $\text{NaCl}$ ), carnalite ( $\text{KCl} \cdot \text{MgCl}_2 \cdot 6\text{H}_2\text{O}$ ).

Free metals: silver and gold occur as the free metal.

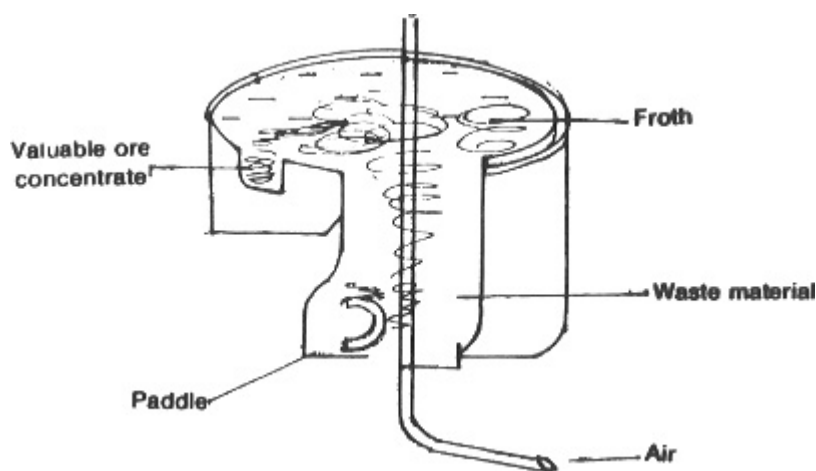
## 4.5 Principles of Extraction of Metals

Extraction of a metal from its ore is basically a reduction process. The method of reduction employed depends on the reactivity of the metal. Three steps are generally followed in the extraction process. These are, (a) concentration of the ore, (b) reduction; and (c) refining. We shall consider each of these processes a little further.

### (a) Concentration of the Ore:

This is a pre-treatment of the metal ore to remove earthy impurities. It may involve (i) oil flotation, (ii) magnetic separation, or (iii) hydraulic washing.

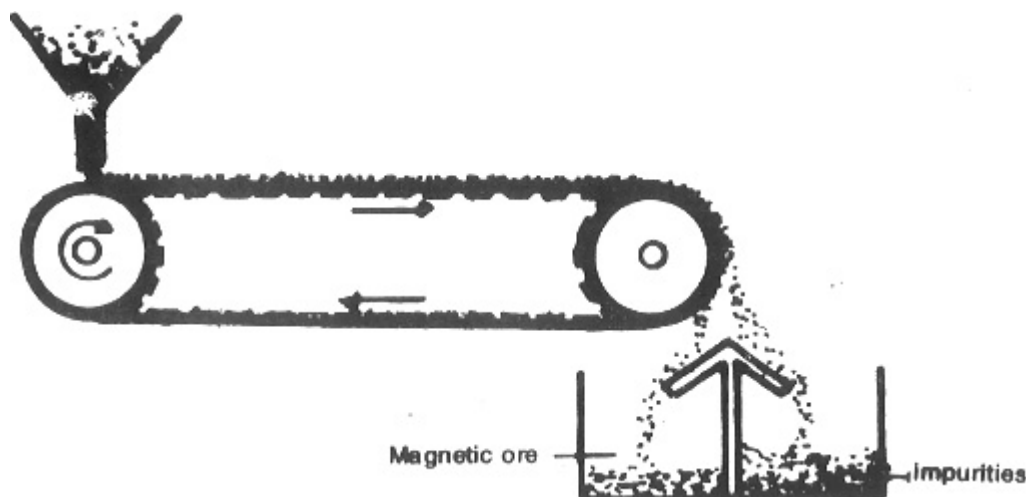
- (i) **Oil flotation:** The finely powdered ore is suspended in water. Oil collectors (e.g. fatty acids or pine oil) and froth stabilizers (e.g. cresols or aniline) are added. The mixture is stirred with rotating paddles. The earthy impurities, called gangue, which are wetted by water sink down, while the mineral particles stick to the oil droplets floating on water. They are then skimmed off from the surface of the water.



*Figure 4.5: Oil flotation method.*

- (ii) **Magnetic separation:** If the mineral, but not the earthy impurities (gangue), is attractable by a magnetic field, then this method can be used. The powdered ore is spread thinly on a

rubber belt which is drawn over pulleys through a magnetic field. The earthy impurities fall off as the belt becomes vertical. The mineral particles, however, remain on the belt until the belt passes out of the influence of the magnetic field. Thus, they are collected in a different container from that in which the earthy impurities are collected.



*Figure 4.6: Magnetic separation method*

(iii) **Hydraulic washing:** This is based on the difference in densities between mineral ore and earthy impurities. The ground ore is placed on top of a sloping table which is kept vibrating by motors. A stream of water is made to flow across the table at right angles to the slope. At each jerk of the inclined table, less dense particles are thrown up. They get moved across the table by the stream of water. Thus, only the heavier mineral ore particles reach the bottom of the table.

(b) **Reduction**

(i) **Reduction by electrolysis:** Metals which are higher up in the electrochemical series are usually extracted by electrolysis of the fused or molten salts. Examples include sodium, potassium, calcium and magnesium, which are extracted by the electrolysis of their fused chlorides using carbon as anode. Carbon anode is usually used because chlorine will attack any other electrode. Aluminium is obtained from aluminium oxide by electrolysis.

(ii) **Reduction of oxides:** Metals which are less reactive are extracted by using reducing agents such as carbon (coke) or carbon(II) oxide, CO. An example is the reduction of iron(III) oxide in the blast furnace using coke and carbon(II) oxide. Zinc oxide is also reduced in the same manner. This process is often referred to as chemical reduction. Lead, tin, iron and zinc are obtained by chemical reduction



(iii) **Reduction of sulphides:** The sulphides are "roasted" in air to obtain the oxide followed by reduction process which may be by self reduction or with a reducing agent. Self reduction is often referred to as thermal reduction. Mercury(II) sulphide is heated to produce the metal, mercury. Zinc sulphide is reduced in two steps:

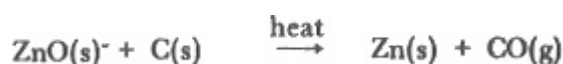


TABLE 4.1: **Summary of extraction process**

Metal	Reactivity	Common ores	Method of extraction
Potassium	Very reactive	Chloride	Electrolysis of fused chloride.
Sodium	Very reactive	Chloride	Electrolysis of fused chloride.
Calcium	Very reactive	Trioxocarbonate (IV)	Electrolysis of fused chloride.
Aluminium	Very reactive	Oxide	Electrolysis of oxide.
Zinc	Moderately reactive	Sulphide	Roasting of ore in oxygen/air to convert to oxide, and reduction with coke.
Iron	Moderately reactive	Oxide	Reduction with coke.
Tin	Reactive	Sulphide	Reduction with coke.
Lead	Reactive	Sulphide	Roasting of ore in air to convert to oxide, and reduction with coke.
Copper	Least reactive	Sulphide	Heating in limited air.
Mercury	Least reactive	Sulphide	Roasting of ore in air.
Silver	Least reactive	Occur on the free element	
Gold	Least reactive		

From the foregoing, it is evident that the nature of the metal determines the reduction method to be used. This is related to a number of factors such as the electro-positivity of the metal, the reduction potential,  $E^0$  and the position of the metal in the electrochemical series. It also depends on the metal-oxygen bond. If the bond is very strong, it is hard to break it. The more negative  $\Delta G$  becomes, the stronger the bond.

In reduction involving coke or carbon(II) oxide, there is competition between the carbon and the metal for oxygen. The element whose  $\Delta G$  value for oxide formation is the most negative wins the competition and combines with the oxygen.  $\Delta G$  values vary with

temperature. For most metals, the negative value of  $\Delta G$  becomes lower as the temperature increases. That is, the formation of the metal oxide is less favoured at high temperatures. But the formation of the oxide of carbon is favoured by high temperature. At 2000°C carbon is the most powerful reducing agent.

The purity of the product is one of the factors considered in choosing an extraction method. At times, a small amount of impurity is desirable. Hence, a small amount of carbon in iron is advantageous, and steel usually contains carbon which enhances its physical properties.

## 4.6 General Properties of Alkali Metals

Alkali metals belong to group 1 of the periodic table. They occur in nature in the combined state with other elements, mostly as chlorides such as LiCl, KCl, NaCl. They exhibit oxidation number of +1, and are good conductors of heat and electricity.

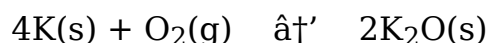
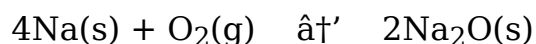
### Physical Properties

- (i) Melting and boiling points decrease down the group.
- (ii) The metals are soft and the softness decreases down the group.
- (iii) They are good conductors of heat and electricity.
- (iv) They are malleable and ductile.
- (v) The freshly cut surfaces have metallic lustre but easily tarnish on exposure to air.

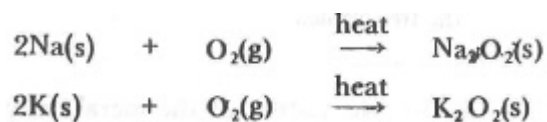
### Chemical Properties

All the alkali metals are very reactive. The reactivities increase down the group.

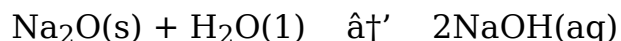
**Reaction with air:** When the alkali metals are exposed to air, they combine with oxygen to form the normal oxides.



However, if they are heated in air or oxygen they form the peroxides.

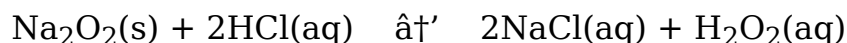


The normal oxides are basic and dissolve in water to form strong alkaline solutions.

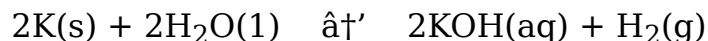
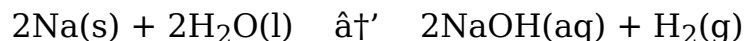


The peroxides dissolve in water or dilute acidic solutions to produce hydrogen peroxide.

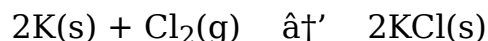
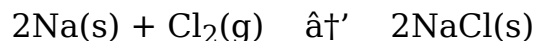




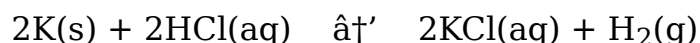
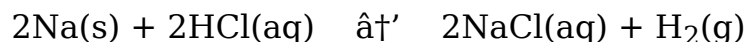
**Reaction with water:** They react violently with water to form alkaline solutions.



**Reaction with chlorine:** When a piece of sodium is dropped in a jar of chlorine gas it reacts with the chlorine to form sodium chloride.



**Reaction with dilute acids:** Alkali metals react violently with dilute acids, displacing hydrogen from the acids.



Alkali metal chlorides are electrolytes. They are ionic in solution and the extraction of the metals involves the electrolysis of the fused chlorides and not their aqueous solutions because hydrogen is liberated instead of the alkali metals when the aqueous solutions are electrolysed.

## 4.7 Sodium and Its Compounds

Sodium is the most commonly found alkali metal in the laboratory. It exists in the +1 oxidation state in all its compounds. It is not found free in nature. The first scientist to isolate sodium was Sir Humphrey Davy in 1802, by electrolysing molten sodium hydroxide. Sodium is now extracted by the electrolysis of fused sodium chloride.

Sodium occurs as sodium chloride in sea water or on land as rock salt. It also occurs as sodium trioxocarbonate(IV) and sodium trioxonitrate(V).

### Extraction of Sodium

Sodium is extracted from fused sodium chloride by electrolysis in the Downs's cell. Since the melting point of sodium chloride is high, about 800°C, calcium chloride is added to lower the melting point to about 600°C. Carbon anode and iron cathode are used as the electrodes.

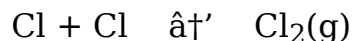
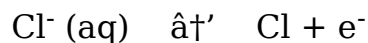
During the electrolysis, sodium ions migrate to the cathode where each sodium ion gains an electron to become sodium atom which is deposited as sodium metal.



Molten sodium is collected at the cathode

Molten sodium is less dense than the fused sodium chloride and hence floats on top.

At the carbon anode, chloride ions lose one electron per ion to form chlorine atoms which combine in pairs to form chlorine gas molecules.



Carbon is used as the anode because it is not attacked by chlorine. The same method is used for the extraction of potassium.

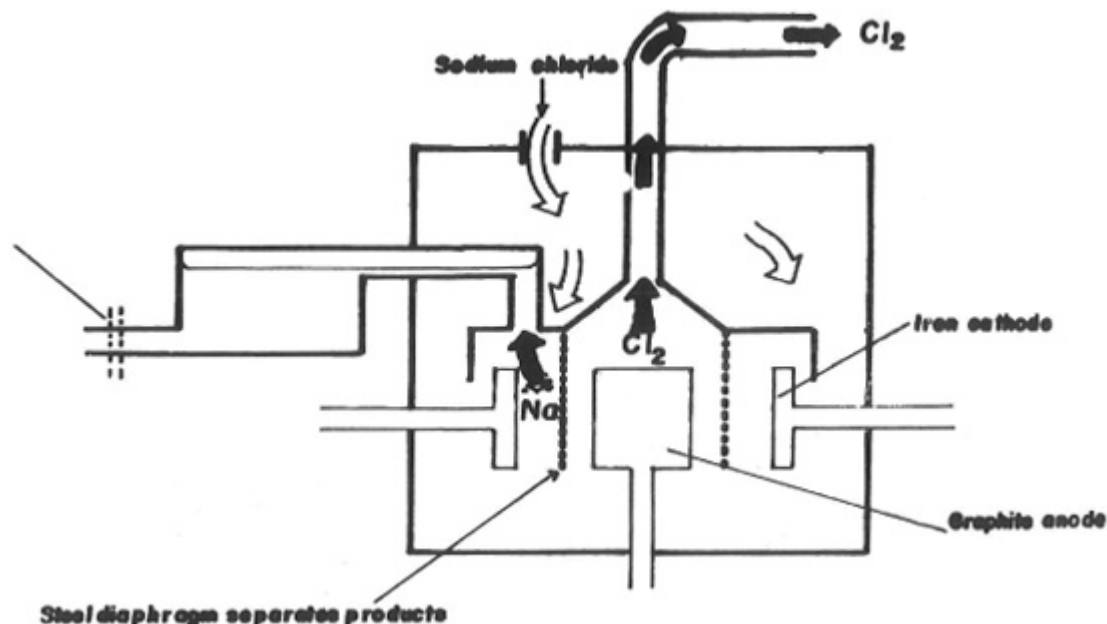


Figure 4.7: The Downs's cell

## Physical Properties of Sodium

- (i) It is a soft metal.
- (ii) It is silvery white and shiny and readily tarnishes on exposure to air.
- (iii) It is a good conductor of heat and electricity.
- (iv) It is very malleable.
- (v) It has a low density and floats on top of water.
- (vi) It has a low melting point of  $98^\circ\text{C}$ .

## Chemical Properties of Sodium

### Experiment 4.2:

Investigating the action of common Reagents on Sodium Metal.

- (a) Place a piece of sodium in a deflagrating spoon and keep it in the open air for some time. What do you observe? Does it burn? Write an equation for the reaction.

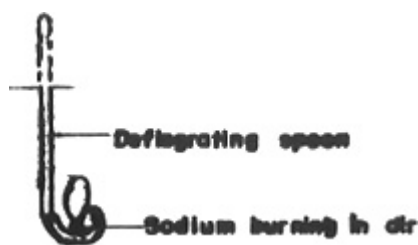


Figure 4.8: Exposing sodium metal in air.

- (b) Fill a gas jar with oxygen. Drop a piece of sodium metal into it. Does it burn more brightly in it? Write an equation for the reaction.

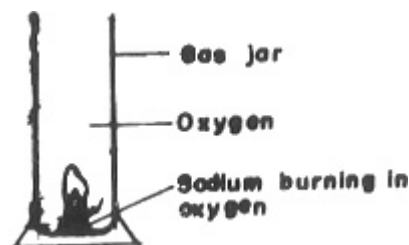


Figure 4.9: Burning sodium in oxygen.

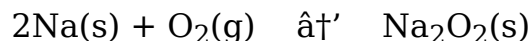
- (c) Put some water in a beaker. Drop a small piece of sodium metal into it. What do you observe as the metal darts about on the water? What gas is given off? Write an equation for the reaction. Test the resulting solution with red litmus paper.
- (d) Collect some chlorine gas in a gas jar and place a piece of sodium metal in it. What do you observe? Write an equation for the reaction.

The results of these experiments illustrate some of the properties of sodium. These and other properties of sodium will now be discussed.

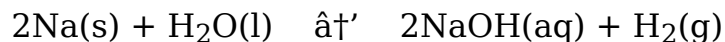
- (a) Sodium is a very reactive metal.
- (b) Sodium reacts with oxygen of the air to form sodium oxide.



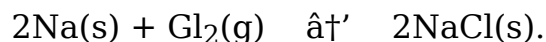
- (c) It burns in a gas jar of excess oxygen to form sodium peroxide.



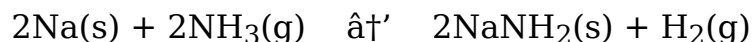
- (d) It reacts violently with water. When placed in water, it darts about, displacing hydrogen from water. The resultant solution is alkaline. Some heat is also given off.



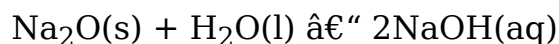
- (e) It reacts with chlorine to form sodium chloride.



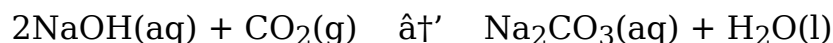
- (f) It reacts with ammonia gas to form solid sodamide, with evolution of hydrogen.



(g) When exposed to air, sodium is attacked by oxygen to form sodium monoxide. The sodium monoxide absorbs moisture from the air to form sodium hydroxide.



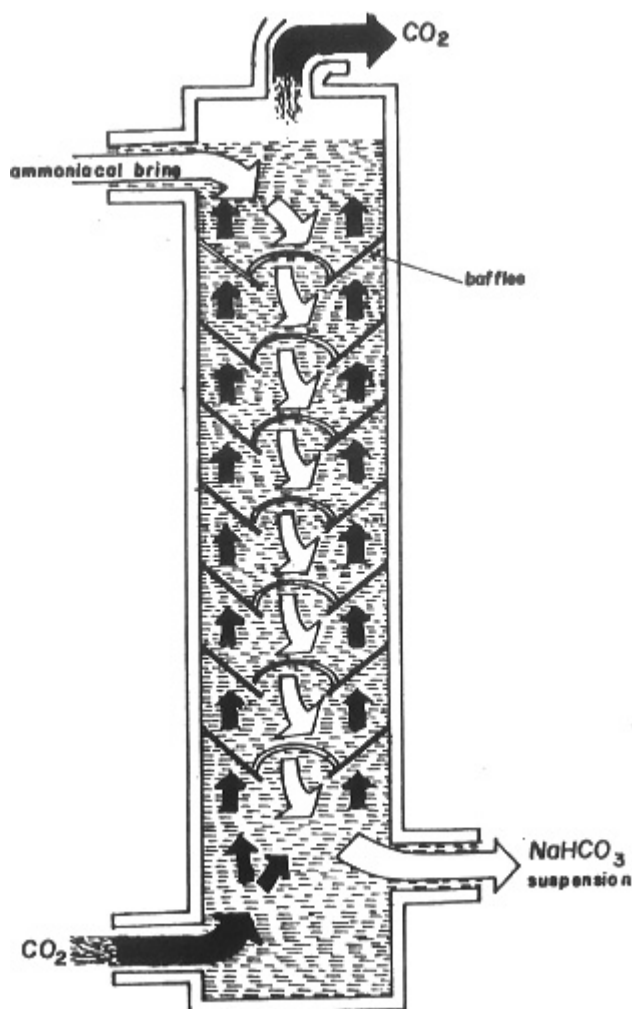
The sodium hydroxide later combines with the carbon(IV) oxide of the air to form sodium trioxocarbonate(IV).



This is why sodium is stored under paraffin.

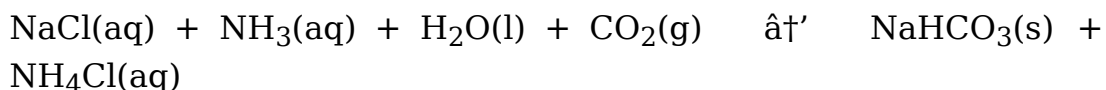
### Sodium Trioxocarbonate(IV), $\text{Na}_2\text{CO}_3$

Sodium trioxocarbonate(IV) is manufactured by the Solvay process. In this process, ammoniacal brine (concentrated sodium chloride solution saturated with ammonia) is pumped into a Solvay tower (Figure 4.10) fitted with perforated mushroom-shaped plates. The perforated plates provide a fine spray of the solution so that a large surface area is exposed for reaction to take place. Carbon(IV) oxide is blown up the tower.

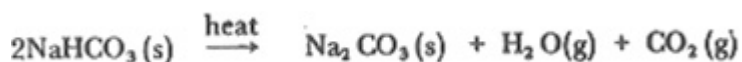


*Figure 4.10: Solvay tower for manufacture of  $\text{Na}_2\text{CO}_3$*

Sodium hydrogen trioxocarbonate(IV) is formed. The sodium hydrogen trioxocarbonate(IV) is not very soluble, and so precipitates.



The precipitate is filtered off and washed to remove ammonium compounds. It is then converted to sodium trioxocarbonate(IV) by heating. The carbon(IV) oxide given off is recycled.



On crystallization, washing soda is obtained.



Ammonia is recovered from the ammonium chloride left in the solution after precipitation, as well as from the solution washed out from the precipitate, by heating it with calcium hydroxide.



The liberated ammonia is thus re-introduced into the system. The only waste product from the Solvay process is calcium chloride.

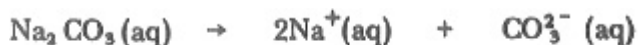
### **Properties of sodium trioxocarbonate(IV)**

*Experiment 4.3:* Investigating the properties of Sodium Trioxocarbonate (IV).

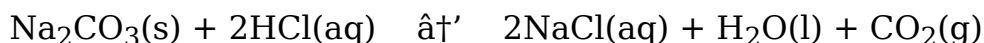
- (i) Put some sodium trioxocarbonate(IV) in a test-tube. Heat it strongly. Test for the liberation of any gas.
- (ii) Add water to another sample and shake. Does it dissolve? Test the solution with red litmus paper.
- (iii) Put some sodium trioxocarbonate(IV) in a test-tube. Add dilute hydrochloric acid. Identify any gas given off.

These experiments illustrate some of the properties of sodium trioxocarbonate(IV). In general,

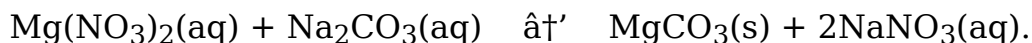
- (a) sodium trioxocarbonate(IV) is not decomposed by heating.
- (b) it is soluble in water. It hydrolyses in water to give an alkaline solution which turns red litmus paper blue. It is therefore used in volumetric analysis.



- (c) Sodium trioxocarbonate(IV) reacts with dilute acids to give off carbon(IV) oxide.

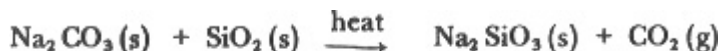


- (d) Sodium trioxocarbonate(IV) is used to precipitate insoluble trioxocarbonate(IV) salts.

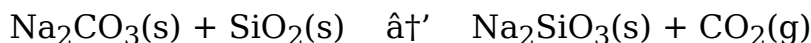


**Uses of sodium trioxocarbonate(IV):** It has many important uses like:-

- (a) It is used for water softening.
- (b) Sodium hydrogen trioxocarbonate(IV) is used in making baking powder. Baking powder is used in the production of bread. It causes bread to swell (rise) Baking powder contains sodium hydrogen trioxocarbonate(IV), powder as a diluent and tartaric acid. The rise in the bread is due to the production of carbon(IV) oxide which causes the bread to rise.
- (c) It is used in the manufacture of glass. In this process, it is fused with calcium trioxocarbonate(IV), silicon(IV) oxide, and a little carbon as reducing agent.



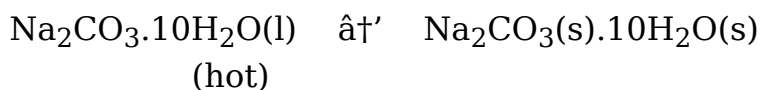
- (d) It is used in the manufacture of water glass by fusing it with silicon(IV) oxide to produce sodium trioxosilicate(IV).



Water glass is used in fire-proofing and for preserving eggs.

### **Sodium Trioxocarbonate(IV) Decahydrate, $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$**

This is also called washing soda. It is prepared from anhydrous sodium trioxocarbonate(IV) by crystallizing it from a hot concentrated solution.



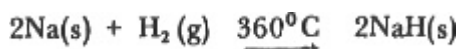
The crystals separate as colourless transparent crystals.

Sodium trioxocarbonate(IV) decahydrate is efflorescent. It loses nine molecules of water of crystallization on exposure to the air.



### **Sodium Hydride, NaH**

Sodium reacts with hydrogen at  $360^\circ\text{C}$  to form sodium hydride, NaH.



Sodium hydride is a crystalline ionic compound.

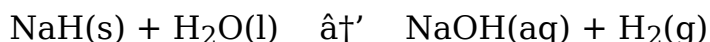
**Properties of sodium hydride:** In sodium hydride, which is an ionic

hydride, hydrogen exists as the hydride ion,  $\text{H}^-$ .

- (i) When sodium hydride is electrolysed in the fused state, hydrogen gas is liberated at the anode.
- (ii) It ignites when exposed to moist air.
- (iii) It burns when heated in air.
- (iv) It reacts with moist carbon(IV) oxide to form sodium methanoate.

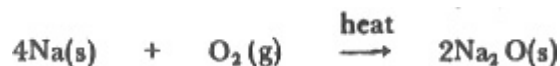


- (v) It dissolves in water to form sodium hydroxide and hydrogen.

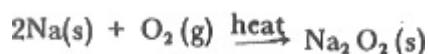


## Oxides of Sodium

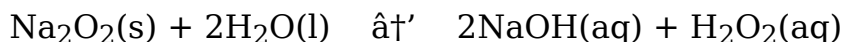
When sodium is heated in a limited supply of oxygen, sodium oxide is formed.



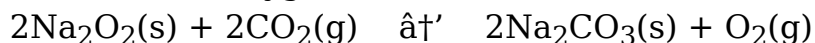
In excess oxygen, sodium peroxide is formed.



Sodium peroxide reacts with water to form hydrogen peroxide.



If it is warmed, oxygen is produced. Sodium peroxide absorbs carbon(IV) oxide from the air to form sodium trioxocarbonate(IV) with the liberation of oxygen.

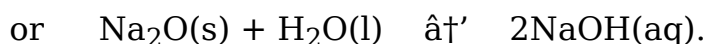
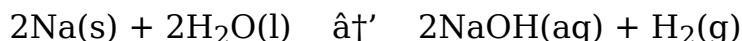


Sodium peroxide is used for the preparation of hydrogen peroxide and oxygen.

## Sodium Hydroxide, NaOH

Sodium hydroxide is a whitish, deliquescent solid. It is very corrosive, hence it is called caustic soda. The solution in water has a soapy or slippery feel.

**Laboratory preparation of sodium hydroxide:** Sodium hydroxide can be prepared in the laboratory by dropping a piece of sodium metal in a beaker of water. The resultant solution turns red litmus paper blue.



**Industrial production of sodium hydroxide:** Sodium hydroxide is produced industrially by electrolysis of concentrated sodium chloride solution (brine) using the diaphragm cell. The brine is contained in a

porous diaphragm closely covered on the outside by a steel gauze which is the cathode. The graphite anode dips into the concentrated sodium chloride solution in the porous diaphragm, Figure 4.11.

During the process of electrolysis, sodium ions and hydrogen ions migrate to the cathode. Hydrogen ions are discharged in preference to sodium ions.

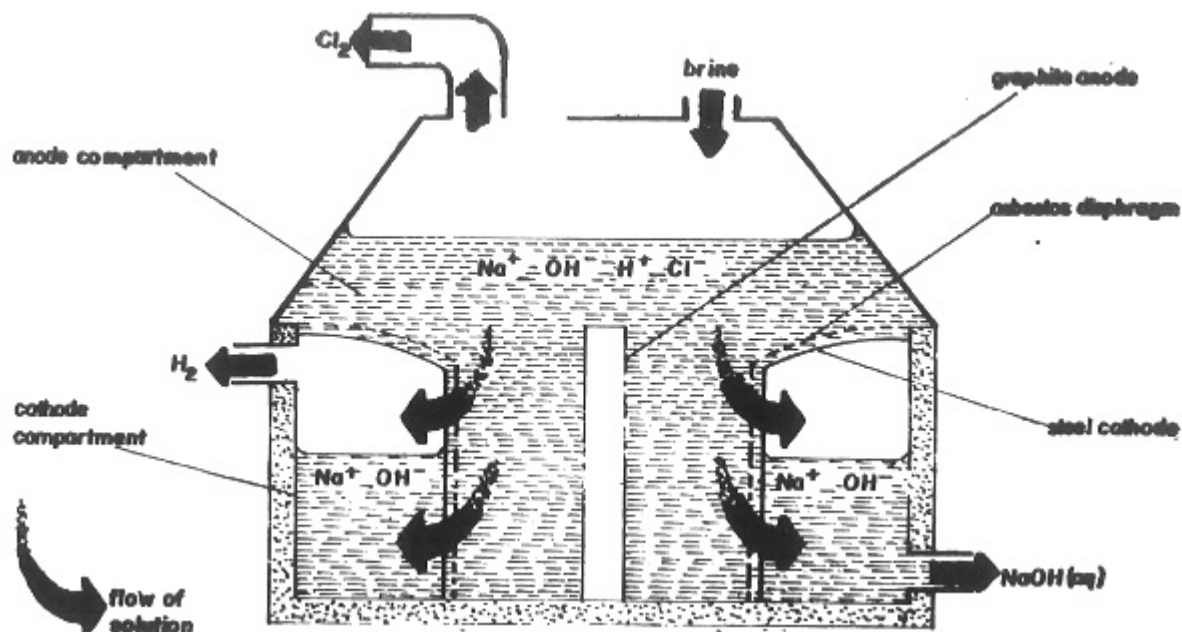
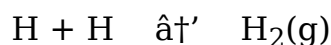
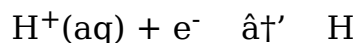
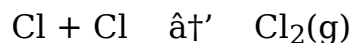
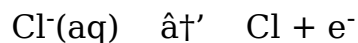


Figure 4.11: The Diaphragm cell.

Chloride and hydroxyl ions migrate to the anode where the chloride ions are preferentially discharged.



The discharge of chloride and hydrogen ions at the anode and cathode respectively, leaves behind sodium and hydroxide ions. The hydroxide ions are obtained from the dissociation of water molecules. The sodium ions collect at the cathode. The sodium and hydroxide ions form sodium hydroxide solution with water at the cathode.

The sodium hydroxide solution drips from the diaphragm and collects at the bottom of the outer compartment of the diaphragm, from where the solution is removed.

The solution is then concentrated by heating and later cooled to obtain solid sodium hydroxide.

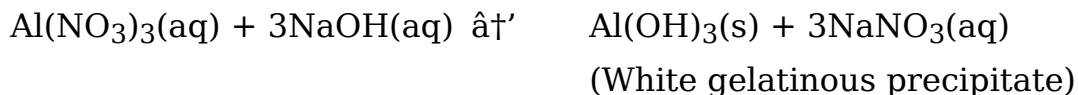
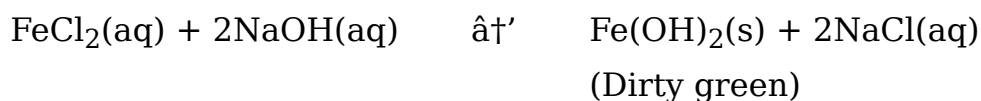
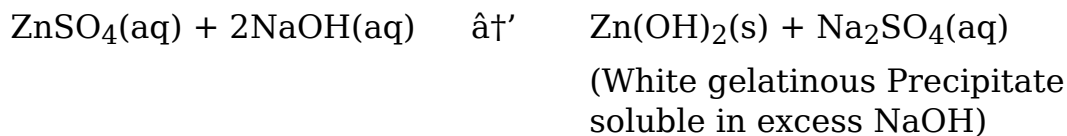
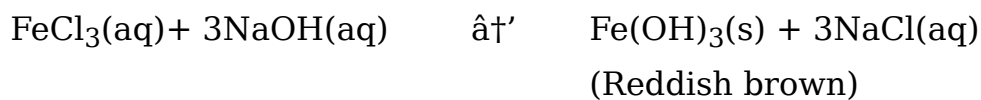
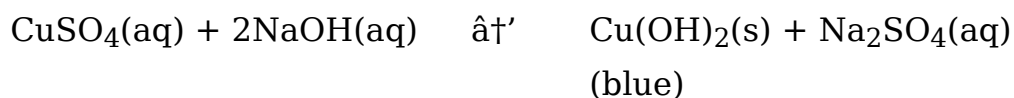
*Experiment 4.4: Investigating the Reactions of Sodium Hydroxide with common Reagents.*

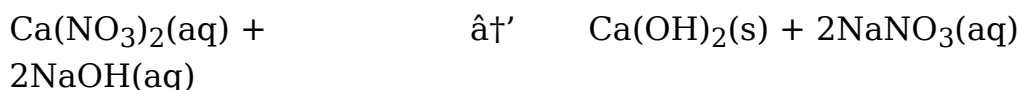


- Place a few pellets of solid sodium hydroxide in a watch glass. What is the colour of the pellets? Leave the pellets in the open air for some time. Do the pellets become wet? Do they dissolve after some time?
- Add a few pellets of sodium hydroxide to water in a test-tube and shake. Do they dissolve? Test the solution with red litmus paper. Feel the test-tube. Does it become hot?
- Secure six clean test-tubes labelled 1 to 6. Pour solutions of copper(II) tetraoxosulphate(VI), iron(III) chloride, zinc tetraoxosulphate(VI), iron(II) tetraoxosulphate(VI), aluminium trioxonitrate(V) and calcium trioxocarbonate(V) into each of the test-tubes in turn. Add few drops of sodium hydroxide to each solution in turn and record your observations. Then add excess sodium hydroxide solution. Again record your observations. Draw a table of tests, observations and inferences.
- Put some solid ammonium chloride in a test-tube. Add sodium hydroxide solution to it and heat. Test the gas given off with wet red litmus paper.

**Properties and reactions of sodium hydroxide:** The results of Experiment 4 illustrate some of the properties of sodium hydroxide.

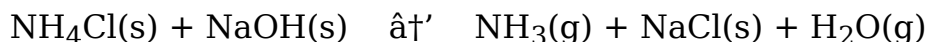
- Sodium hydroxide pellets on exposure to air absorb moisture and dissolve to form a saturated solution. That is, sodium hydroxide is deliquescent.
- It dissolves in water to form an alkaline solution which turns red litmus paper blue.
- Its solution precipitates insoluble metallic hydroxides from solutions of the metallic salts.



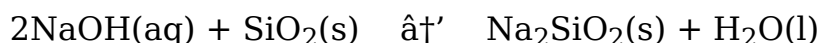


(White powdery precipitate)

- (d) When heated with ammonium salts, a colourless, pungent and alkaline gas (ammonia) is given off. The gas turns moist red litmus paper blue.



- (e) Sodium hydroxide attacks glass. This is why bottles containing sodium hydroxide are usually covered with rubber or plastic stoppers.



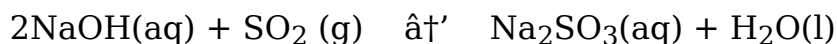
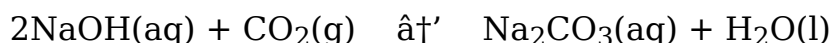
- (f) Sodium hydroxide absorbs chlorine. With cold, dilute sodium hydroxide solution, sodium monochlorate(I) is formed.



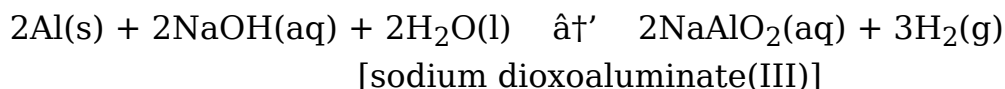
With hot, concentrated sodium hydroxide solution, sodium trioxochlorate(V) is formed.



- (g) Sodium hydroxide also absorbs carbon(IV) oxide and sulphur(IV) oxide.



- (h) Sodium hydroxide solution reacts with amphoteric metals such as aluminium and zinc.



Thus, sodium hydroxide is never stored in aluminium or zinc containers because sodium hydroxide attacks such containers.

**Uses of sodium hydroxide:** Sodium hydroxide is used:

- (i) in the paper industry.
- (ii) in the soap industry.
- (iii) in the manufacture of rayon.
- (iv) in volumetric analysis.
- (v) in the laboratory as a strong alkali.
- (vi) in the purification of bauxite during the extraction of aluminium.

**TABLE 4.2: Action of sodium hydroxide solution on some common cations.**

---

Cation (in solution)	Observations
1. Copper(II) ions.	Blue precipitate, insoluble in excess sodium hydroxide solution.
2. Iron(II) ions.	Greenish gelatinous precipitate, insoluble in excess sodium hydroxide solution.
3. Iron(III) ions	Reddish-brown precipitate, insoluble in excess sodium hydroxide solution
4. Zinc ions	White gelatinous precipitate, soluble in excess Sodium hydroxide to form sodium dioxoaluminate(111).
5. Aluminium ions	White gelatinous precipitate, soluble in excess sodium hydroxide to form sodium dioxoaluminate(III).
6. When heated with ammonium salt.	Pungent smell, gas given off turns red litmus blue, and forms white fumes with concentrated hydrochloric acid vapour

## Sodium Chloride, NaCl

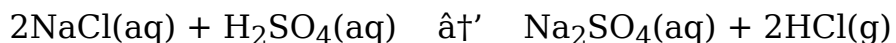
Sodium chloride occurs as rock-salt. It is mined as a solid or pumped out of the earth as brine. It can also be extracted from salt wells or sea water by evaporation or by common ion effect.

*Experiment 4.5: Investigating the Reactions of Sodium Chloride with some common Reagents*

- Place a few crystals of sodium chloride in a watch glass and leave it till the next day in the laboratory. Note the colour. Do the crystals absorb moisture from the atmosphere? -
- Put some sodium chloride crystals in a test-tube. Add water and shake. Do the crystals dissolve? Test the solution with red and blue litmus papers.
- Add dilute tetraoxosulphate(IV) acid to a few crystals of sodium chloride in a test-tube and heat. Identify any gas given off.

## Properties of sodium chloride

- It is a white crystalline solid.
- Pure sodium chloride is not deliquescent. It is the presence of magnesium chloride impurities in table salt that makes it deliquescent.
- Sodium chloride is soluble in water, forming a neutral solution.
- Sodium chloride reacts with concentrated tetraoxosulphate(VI) acid in the cold to liberate hydrogen chloride gas.



## Uses of sodium chloride

- (i) It is used to preserve food.
- (ii) Sodium chloride as table salt is an important part of our diet as it supplies essential minerals to our body.
- (iii) It is used in the manufacture of sodium hydroxide and sodium trioxocarbonate(IV).
- (iv) It is used in the manufacture of chlorine and sodium.
- (v) It is used in soap making for the [“salting out” process].

### **Sodium Trioxonitrate(V), $\text{NaNO}_3$**

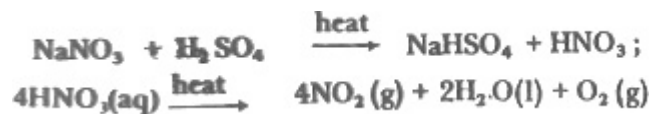
Sodium trioxonitrate(V) occurs naturally in Chile. It can be prepared in the laboratory by titration using 2 M sodium hydroxide and 2 M trioxonitrate(V) acid.

*Experiment 4.6: Investigating the Reactions of sodium trioxonitrate(V) with common Reagents*

- (i) Place some sodium trioxonitrate(V) crystals in an evaporating dish. Note the colour.
- (ii) Put some sodium trioxonitrate(V) crystals in a test-tube. Add water and shake. Do the crystals dissolve? Test the solution with litmus paper.
- (iii) Heat some sodium trioxonitrate(V) crystals in a test-tube. Identify any gas given off.
- (iv) Add concentrated tetraoxosulphate(VI) acid to the crystals in a test-tube and warm. identify any gas given off.

### **Properties of sodium trioxonitrate(V)**

- (i) Sodium trioxonitrate(V) is a colourless crystal which is not deliquescent.
- (ii) The crystals are soluble in water and the solution is neutral to litmus papers.
- (iii) When it is strongly heated, oxygen is given off.
- (iv) Sodium trioxonitrate(V) liberates trioxonitrate(V) acid and brown nitrogen(IV) oxide when it is heated with concentrated tetraoxosulphate(VI) acid.



Uses of sodium tetraoxosulphate(V)

- (i) It is used extensively in the manufacture of dyes.
- (ii) It is used in making sodium dioxonitrate(III).
- (iii) It is used in making nitrogenous fertilizers.

## **4.8 General Properties of Alkaline Earth**

# Metals

The group 2 elements “beryllium, calcium, magnesium, strontium and barium are commonly called the alkaline earth metals. The most common among them are calcium and magnesium. Calcium occurs naturally as calcium trioxocarbonate(IV),  $\text{CaCO}_3$ , and gypsum,  $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ . Magnesium occurs as magnesium chloride in sea water and as the mineral, dolomite,  $\text{MgCO}_3 \cdot \text{CaCO}_3$ .

**TABLE 4.3. General properties of alkaline earth metals**

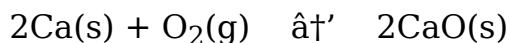
	Appearance	Melting point ( $^{\circ}\text{C}$ )	Boiling point ( $^{\circ}\text{C}$ )	Conductivity	Ionization energy ( $\text{kJ mol}^{-1}$ )	Relative atomic mass.
Beryllium	Grey solid	1278	2970	Good	900	9
Magnesium	White lustrous solid	649	1100	Good	740	24
Calcium	Grey solid	350	1490	Good	590	40
Strontium	Grey	770	1380	Good	550	87.6
Barium	Grey	710	1637	Good	500	137.4

## Physical Properties

- (i) They are all solid metals whose melting points increase down the group.
- (ii) They are good conductors of heat and electricity.
- (iii) The sizes increase down the group.
- (iv) The changes in physical properties are not regular.

## Chemical Properties

- (a) All the metals except magnesium, tarnish in air to form the corresponding oxides.

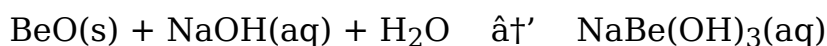
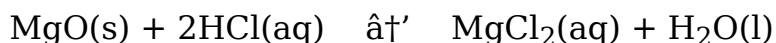
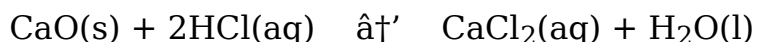
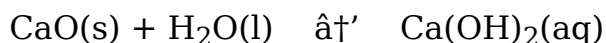


When heated in air, the elements form oxides of the general formula  $\text{MO}$ .

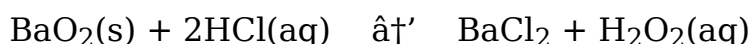
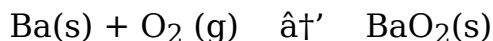
- (b) Magnesium reacts with steam to form magnesium hydroxide. Calcium, strontium and barium form hydroxides with cold water. The reaction is not as violent as we have with the alkali metals.



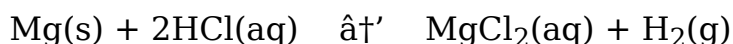
- (c) Their oxides are generally basic and their solubilities increase down the group. Beryllium oxide is amphoteric. It dissolves in sodium hydroxide to form sodium beryllate. Magnesium oxide is sparingly soluble in water. Calcium, strontium and barium oxides are more soluble in water and form weak alkaline solutions.



- (d) Strontium and barium form peroxides of the general formula  $\text{MO}_2$ . These peroxides react with dilute acids to form hydrogen peroxide. They are true peroxides.

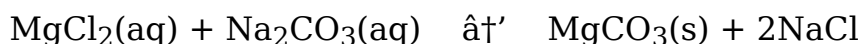
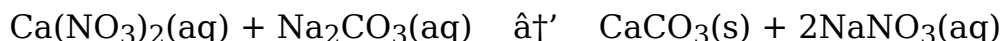


- (e) They combine directly with chlorine or hydrogen chloride acid (hydrochloric acid) to form the chloride. Their chlorides are of the general formula  $\text{MCl}_2$ . Beryllium chloride is partly covalent in character. Magnesium chloride is ionic but forms basic chloride,  $\text{Mg(OH)Cl}$ , on heating. Magnesium chloride thus exhibits some covalent character.



Calcium chloride, strontium chloride, and barium chloride are all ionic. They conduct electric current in the molten state. They are all soluble in ethanol except barium chloride which is insoluble. The points of the chlorides increase down the group.

- (f) The trioxocarbonate(IV) salts are formed by double decomposition.



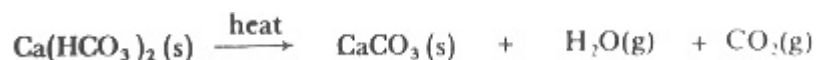
All the trioxocarbonate(IV) salts are insoluble in water. Beryllium, Calcium and magnesium trioxocarbonate(IV) salts decompose on heating, to give the corresponding oxides.



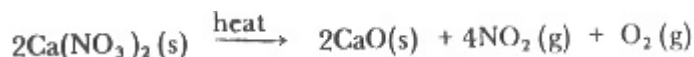
Strontium and barium trioxocarbonate(IV) salts are not decomposed by heat.

- (g) They all form hydrogen trioxocarbonate(IV) salts of the general formula  $\text{M(HCO}_3)_2$ , and are soluble in water. The hydrogen trioxocarbonate(IV) salts of calcium and magnesium cause hardness of water.

The hydrogen trioxocarbonate(IV) salts are decomposed by heat.



- (h) All form trioxonitrate(V) salts of the general formula  $M(\text{NO}_3)_2$ , which are soluble in water. The trioxonitrate(V) salts are hydrated except for barium trioxonitrate(V). All trioxonitrate(V) salts of the alkaline earth metals decompose on heating to form oxides of the metals, nitrogen(IV) oxide, and oxygen.



- (i) Tetraoxosulphate(VI) salts of the alkaline earth metals are sparingly soluble in water. Solubility decreases from magnesium down the group. This is due to the decrease of energy of hydration with increasing ionic size.

**TABLE 4.4: Properties of alkaline earth metal compounds**

Property	Chlorides	Trioxocarbonate(IV) salts	Hydrogen-trioxocarbonate(IV) salts.	Tetraoxo-sulphate(IV) salts
Solubility in water	All soluble (except beryllium)	All insoluble.	All soluble	All insoluble
Action of heat	No effect	Decomposed	Decomposed	No effect
Ionic character	Ionic, except beryllium which shows some covalent character	Ionic	Ionic	Ionic
pH	Neutral	Basic	Basic	Neutral

## 4.9 Calcium

### Occurrence

Calcium occurs in nature as limestone, chalk, marble, calcite and aragonite which are forms of calcium trioxocarbonate(IV),  $\text{CaCO}_3$ . Limestone deposits occur in many states of Nigeria. Large deposits occur at Nkalagu in Enugu state, Ukpilla (near Auchi) in Edo state, Ewekoro (near Abeokuta) in Ogun state. It also occurs in Sokoto and Cross River states. Calcium also occurs as dolomite, a double salt with formula  $\text{CaCO}_3 \cdot \text{MgCO}_3$ , and as the following tetraoxosulphate(VI) salts: anhydrite ( $\text{CaSO}_4$ ) and gypsum ( $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ ). Calcium is the most abundant of the alkaline earth metals.

### Extraction

Calcium metal is extracted from calcium chloride by the electrolysis of the fused chloride salt to which a little calcium fluoride is added. The

addition of calcium fluoride is made to reduce the melting point to  $700^{\circ}\text{C}$ . A graphite anode and an iron cathode are used. As calcium is deposited at the iron cathode the latter is raised so that eventually the effective cathode is calcium. Figure 4.12 illustrates the electrolytic cell for the extraction.

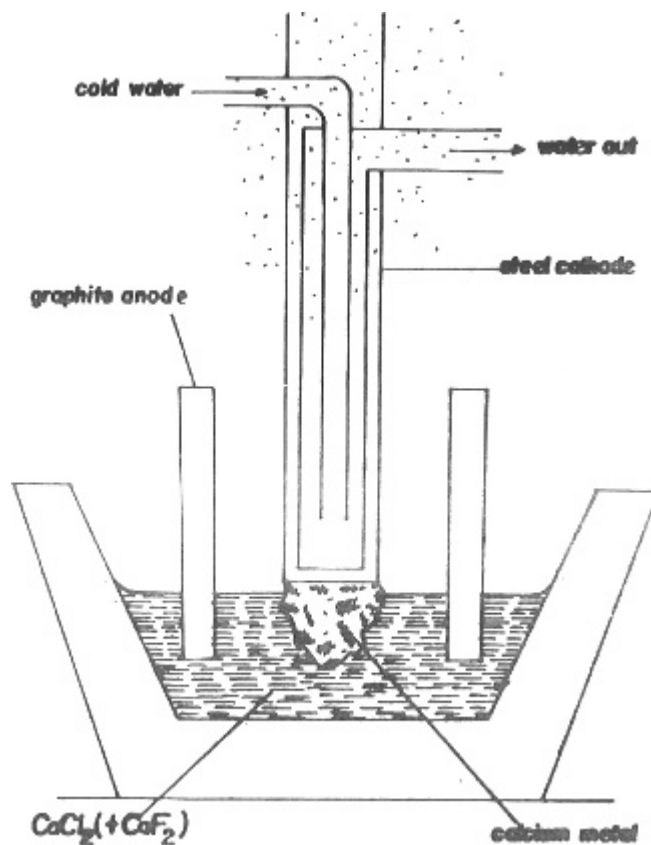


Figure 4.12: Extraction of calcium

## Compounds of Calcium

### Calcium oxide, CaO

This is produced by heating calcium trioxocarbonate(IV); limestone, to about  $900^{\circ}\text{C}$  in a lime kiln.





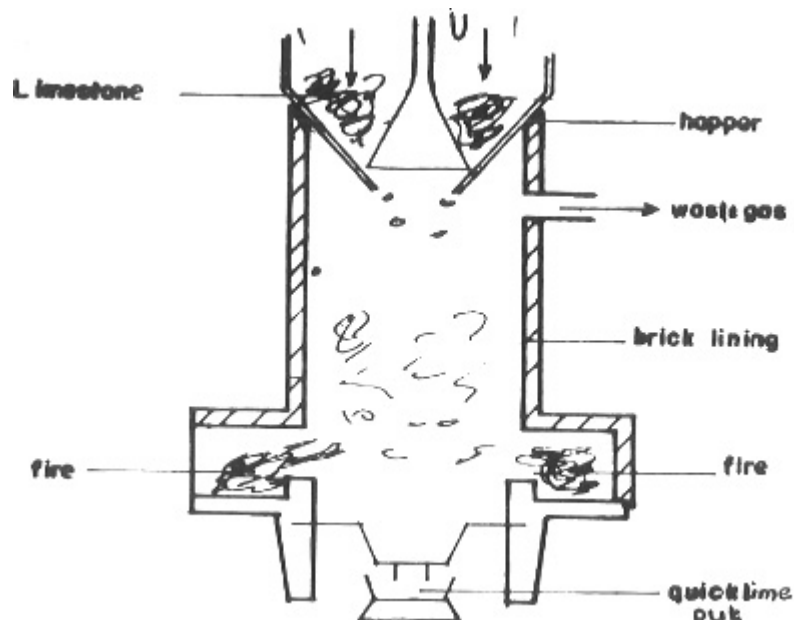


Figure 4.13: Lime Kiln

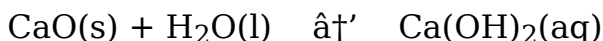
Producer gas is used to produce a high temperature flame to decompose the limestone.

*Experiment 4.7: Investigating the Reactions of Calcium Oxide with common Reagents*

- Put a large piece of calcium oxide in a hard glass test-tube. Add water to it. Describe what happens.
- Test the solution of calcium oxide with litmus papers.
- Filter the solution obtained in (i), add more water and again filter. Is calcium oxide soluble in water?

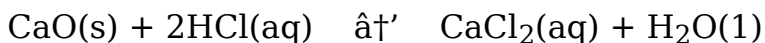
**Properties of calcium oxide:**

- Calcium oxide is sparingly soluble in water. The dissolution of calcium oxide in water is exothermic. As water is added to a large piece of calcium oxide, it cracks and later swells, giving off steam. A lot of heat is generated.

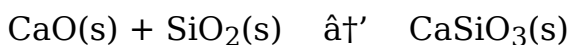


On filtering the resultant solution, a fine white powder is obtained. This powder is called slaked lime.

- Calcium oxide is a strong base. Its solution turns red litmus blue.



- In the blast furnace, calcium oxide reacts with silicon(IV) oxide,  $\text{SiO}_2$ , to form calcium trioxosilicate(IV), which is slag.

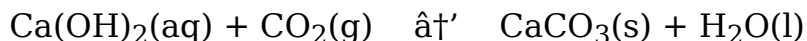


## Uses of calcium oxide

- (i) It is used for making calcium carbide, one of the reactants in the preparation of ethyne.  
$$\text{CaO(s)} + 3\text{C(s)} \rightarrow \text{CaC}_2\text{(s)} + \text{CO(g)}$$
- (ii) Because of its basic character, it is used for lining furnaces especially in the conversion of iron to steel.
- (iii) Calcium oxide is used to dry ammonia in the laboratory because of its basic and hygroscopic character.
- (iv) It is used to neutralize soil acidity in agriculture.
- (v) It is used in the manufacture of mortar, cement and plasters.

**Cement:** Cement is made by heating an equimolar mixture of limestone and clay (aluminium trioxosilicate(IV)). A kind of clinker is thus formed. The clinker is ground to a fine powder. When this powder is mixed with sand and water, it sets to a hard mass.

**Mortar:** Mortar is made by mixing together, one part of quicklime, CaO, with four parts of sand. Enough water is added to give a thick paste of calcium hydroxide and sand. The setting process occurs when the calcium hydroxide formed combines with carbon(IV) oxide to form calcium trioxocarbonate(IV). The mixture is kept in the open air to facilitate loss of water.



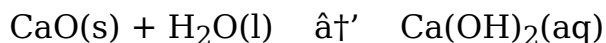
The sand prevents cracking as the mortar sets. The mortar finally sets into a hard mass.

**Plaster:** Plaster consists of mortar and some binding material such as fibre. It is used for walls and ceilings.

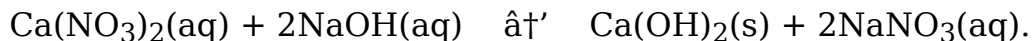
**Concrete:** Concrete is made up of cement mixed with gravel or shells and water. Concrete is used for building bridges, pillars, floors, etc.

## Calcium Hydroxide, Ca(OH)<sub>2</sub>

Calcium hydroxide is often referred to as slaked lime. It is manufactured industrially by adding water to calcium oxide.



In the laboratory, it may be prepared by adding dilute sodium hydroxide to calcium trioxonitrate(V) solution in a test-tube.



*Experiment 4.8: Investigating the Reactions of Calcium Hydroxide with common Reagents*

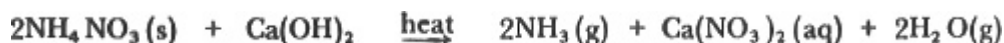
- (i) Add calcium hydroxide solution to ammonium chloride in a test-tube and heat. Identify any gas given off.

- (ii) Dip red litmus paper into a solution of calcium hydroxide in a test-tube. What do you observe?
- (iii) Pass carbon(IV) oxide into a solution of calcium hydroxide in a test-tube for a short time, and later in excess. What do you observe?

### Properties of calcium hydroxide

- (a) Calcium hydroxide is weakly alkaline. It turns red litmus paper blue. However, it is a strong base.

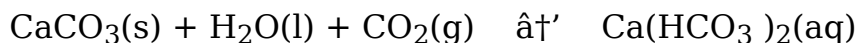
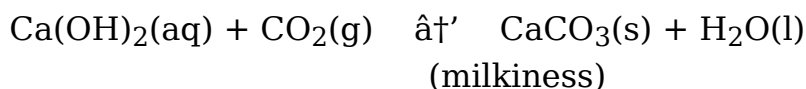
As a strong base, it liberates ammonia when heated with any ammonium salt.



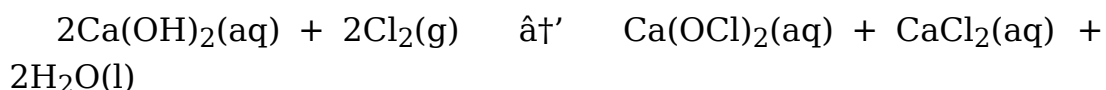
- (b) It loses water when strongly heated to a temperature of above  $400^\circ\text{C}$ .



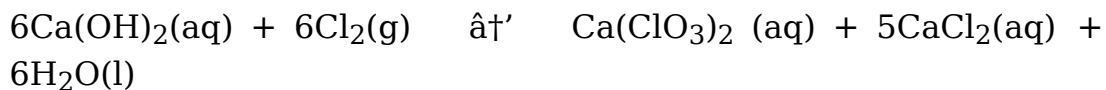
- (c) When carbon(IV) oxide is passed into lime-water (calcium hydroxide solution), the limewater turns milky. The milkiness disappears when the carbon(IV) oxide is passed for a long time.



- (d) When chlorine gas is passed into a cold solution of calcium hydroxide, calcium monoxochlorate(I), calcium chloride and water are formed.

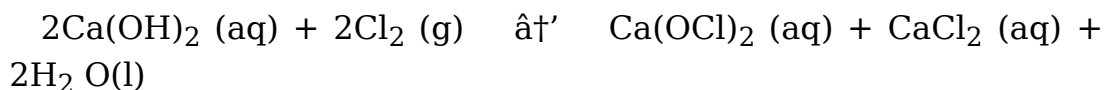


However, if hot calcium hydroxide solution is saturated with chlorine gas, the products are calcium trioxochlorate(V), calcium chloride and water.



### Uses of calcium hydroxide

- (i) It is used in agriculture to reduce soil acidity.
- (ii) It is used for making bleaching powder.

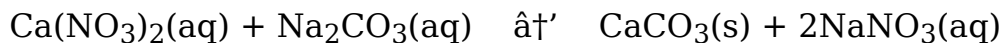


- (iii) It is used to soften hard water.
- (iv) It is used for making plaster and mortar.

- (v) It is used in the laboratory as limewater for detecting carbon(IV) oxide gas.

### **Calcium Trioxocarbonate(IV), $\text{CaCO}_3$**

It can be prepared in the laboratory by mixing calcium trioxonitrate(V) with sodium trioxocarbonate(IV).



Calcium trioxocarbonate(IV) is white in colour.

*Experiment 4.9: Investigating Reaction of Calcium Trioxocarbonate(IV) with common Reagents.*

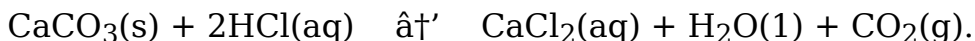
- Put some powdered calcium trioxocarbonate(IV) in a test-tube and heat strongly. Identify any gas given off.
- Put some powdered calcium trioxocarbonate(IV) in a test-tube. Add water to it and shake. Does it dissolve?
- Add dilute hydrogen chloride acid (hydrochloric acid) to calcium trioxocarbonate(IV) in a test-tube. Identify any gas given off.

### **Properties of calcium trioxocarbonate(IV)**

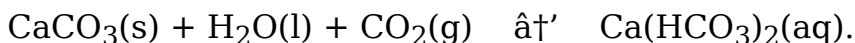
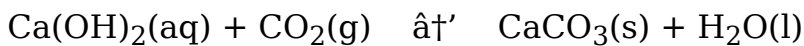
- (a) Calcium trioxocarbonate(IV) decomposes to calcium oxide and carbon(IV) oxide when heated above  $500^\circ\text{C}$ .



- (b) Calcium trioxocarbonate(IV) is insoluble in water.
- (c) When dilute hydrogen chloride acid(hydrochloric acid) is added to it, carbon(IV) oxide is liberated.



- (d) When excess carbon(IV) oxide is passed into limewater, calcium hydrogen trioxocarbonate(IV) is formed. This is because the calcium trioxocarbonate(IV) which was first formed dissolved on the passing of excess carbon(IV) oxide to form calcium hydrogentrioxocarbonate(IV).



### **Uses of calcium trioxocarbonate(IV)**

- In making toothpaste as precipitated chalk;
- In the manufacture of sodium trioxocarbonate(IV) in the Solvay process to generate carbon(IV) oxide;
- In the smelting of iron in the blast furnace;
- Occurs as marble, which is used in building, and limestone, used in the manufacture of cement;

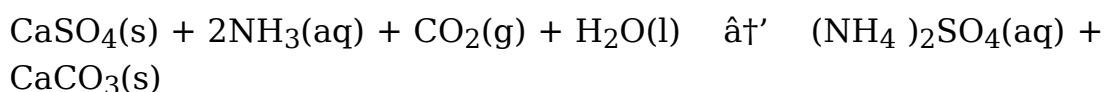
- (v) In the preparation of putty.
- (vi) In the manufacture of quicklime.

### **Calcium Tetraoxosulphate(VI), $\text{CaSO}_4$**

It occurs as the minerals: "gypsum,  $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ , and anhydrite,  $\text{CaSO}_4$ . Calcium tetraoxosulphate(VI) is stable to heat and is insoluble in water.

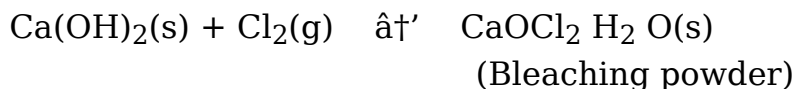
Gypsum is used to manufacture "plaster of Paris". When gypsum is heated, it is converted to "plaster of Paris" ( $2\text{CaSO}_4 \cdot \text{H}_2\text{O}$ ). "Plaster of Paris" is used for making plaster casts in surgery to keep broken bones or joints in fixed positions, and as plasters.

Anhydrite is used to produce sulphur(IV) oxide for the manufacture of tetraoxosulphate(VI) acid. It is also used to manufacture ammonium tetraoxosulphate(VI) fertilizers:

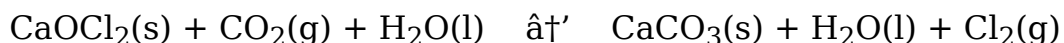


### **Bleaching Powder:**

This is prepared by passing chlorine over solid calcium hydroxide (slaked lime). The slaked lime absorbs the chlorine to form bleaching powder.



Chlorine is displaced from bleaching powder by dilute acids. Even carbon dioxide of the air decomposes bleaching powder. As a result, the powder deteriorates on exposure to air.

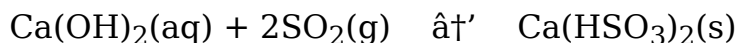


Bleaching powder usually smells of chlorine because of its reaction with air.

Freshly prepared bleaching powder contains about 36% of recoverable chlorine. Bleaching powder is used in dye works and in laundries. Chlorine gas is also transported as bleaching powder.

### **Calcium Hydrogen Trioxosulphate(IV), $\text{Ca}(\text{HSO}_3)_2$**

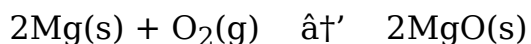
This is manufactured by passing sulphur(IV) oxide into a paste of calcium hydroxide and water (milk of lime).



Calcium hydrogen trioxosulphate(IV),  $\text{Ca}(\text{HSO}_3)_2$  is used in bleaching wood pulp which is used in paper making.

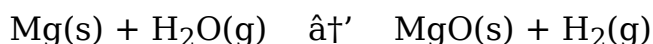
## **4.10 Magnesium**

Magnesium metal is obtained by the electrolysis of fused magnesium chloride. It has high melting and boiling points, and is very reactive. The metal burns in air with a very bright flame to form magnesium oxide.



It reacts with dilute acids to liberate hydrogen. It also reacts with chlorine to form the chloride.

Magnesium is not attacked by cold water, but it reacts with steam to form the oxide and hydrogen.



Magnesium is used in making alloys, flares and fireworks.

### **Magnesium Tetraoxosulphate(VI), $\text{MgSO}_4$**

This occurs as the mineral kieserite,  $\text{MgSO}_4 \cdot \text{H}_2\text{O}$ . Magnesium tetraoxosulphate(IV) heptahydrate,  $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$  is commonly known as Epsom salt. It is a colourless crystal which is used as a mild purgative in medicine.

## **4.11 Aluminium**

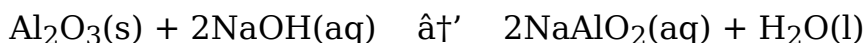
Aluminium belongs to group 3 of the periodic table. It is the third most abundant element in the earth's crust, being widely distributed in rocks and clays as aluminosilicates.

The most important ore of aluminium is bauxite,  $\text{Al}_2\text{O}_3 \cdot 2\text{H}_2\text{O}$ . Bauxite occurs in Ghana as one of that country's most important minerals. Aluminium also occurs as cryolite,  $\text{Na}_3\text{AlF}_6$ , feldspar,  $\text{K}_2\text{Al}_2\text{Si}_6\text{O}_{16}$ , mica and kaolin.

### **Extraction**

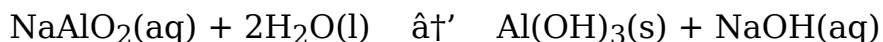
Aluminium is extracted mainly from its bauxite ore. Its extraction from the other ores would be very costly. Bauxite is always associated with oxides of iron and silicon. These impurities are first removed before reducing the oxide ore because their presence would make the electrolytic reduction difficult.

**Purification of bauxite:** Bauxite is roasted in air to convert  $\text{Fe}^{2+}$  to  $\text{Fe}^{3+}$ . The product is ground to powder, then boiled with aqueous sodium hydroxide solution under pressure. Aluminium oxide goes into solution as sodium dioxoaluminate(III).



The impurities (iron(III) oxide, silicon oxide, etc) which are insoluble, are filtered off. The filtrate is diluted, then seeded with crystals of aluminium hydroxide. This process precipitates aluminium hydroxide from solution. The precipitate is filtered off, washed and heated

strongly to obtain pure aluminium oxide.



**Electrolysis:** The pure aluminium oxide is dissolved in molten cryolite in an iron tank with carbon lining (Figure 4.14). The carbon lining acts as the cathode. A graphite anode is dipped into the electrolyte.

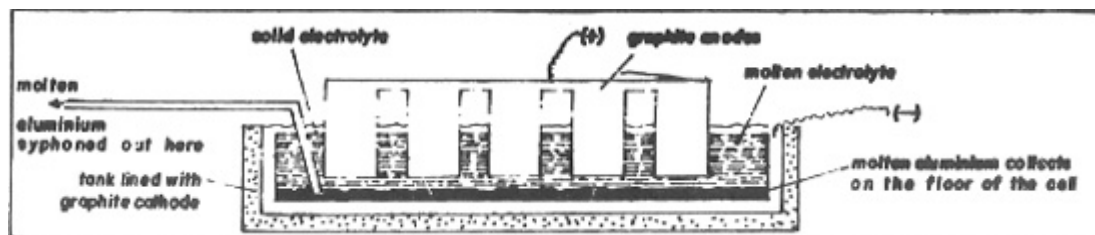
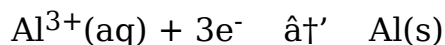
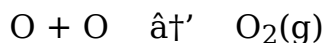
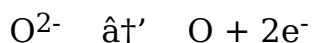


Figure 4.14: Electrolysis of bauxite

The current which passes through the electrolyte maintains it in the molten state at about  $900^\circ\text{C}$ . The ions present in the electrolyte are  $\text{Al}^{3+}$ ,  $\text{O}^{2-}$  from the pure aluminium oxide, and  $\text{Na}^+$ ,  $\text{Al}^{3+}$ ,  $\text{F}^-$  from the molten cryolite.  $\text{Al}^{3+}$  ions being less electropositive than  $\text{Na}^+$ , gain electrons at the cathode to be discharged.



Since molten aluminium is denser than the electrolyte, the free aluminium formed sinks to the bottom of the tank. From there, it is tapped and allowed to solidify. At the anode,  $\text{O}^{2-}$  ions lose electrons to get discharged. Oxygen gas is therefore formed at the anode. It oxidises the carbon anode to carbon(IV) oxide. Thus,



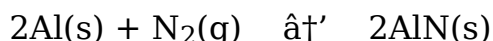
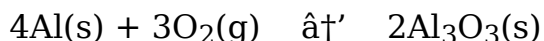
the anode has to be changed from time to time.

*Experiment 4.10: Investigating the Reactions of Aluminium with common Reagents.*

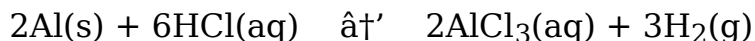
- (i) Expose a piece of aluminium metal or aluminium foil. What is the colour of the metal?
- (ii) Leave aluminium foil in the atmosphere for some hours. What do you observe?
- (iii) Add dilute hydrochloric acid to aluminium metal in a test-tube. Identify any gas given off.
- (iv) Place some aluminium powder in a test-tube, add some sodium hydroxide and warm. What do you observe? Identify any gas given off.

## Properties of Aluminium:

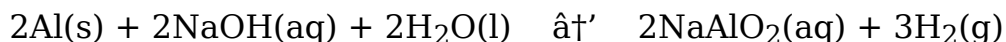
- (a) Aluminium is a silvery-white metal with a high melting point ( $660^{\circ}\text{C}$ ) and a density of  $2.68\text{ g/cm}^3$
- (b) It can be beaten into sheets, and is ductile.
- (c) On exposure of the metal to air, a thin film of aluminium oxide forms on the surface and coats the metal. This protects the metal from further reaction.
- (d) Aluminium burns in air at about  $800^{\circ}\text{C}$  to form aluminium oxide and aluminium nitride.



- (e) Aluminium reacts with both dilute and concentrated hydrogen chloride acid (hydrochloric acid). The reaction is more rapid when concentrated hydrogen chloride acid (hydrochloric acid) is used.

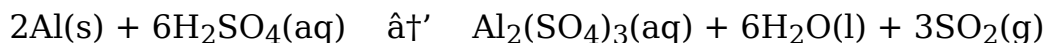


- (f) Aluminium reacts with sodium hydroxide solution to form sodium dioxoaluminate(III), liberating hydrogen gas. It is an amphoteric metal.



- (g) Aluminium does not react with either dilute or concentrated trioxonitrate(V) acid. This is because a coating of oxide is formed over the aluminium and this prevents further action. The aluminium is thus rendered passive.

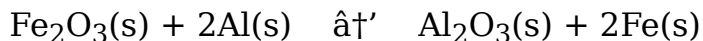
- (h) Aluminium reacts only with hot concentrated tetraoxosulphate(VI) acid. It has no reaction with the dilute acid.



- (i) If aluminium metal is placed in a solution of mercury(II) trioxonitrate(V) solution, it displaces the mercury.

- (j) Aluminium has a small size and a high charge of +3. These two factors are responsible for its high hydration energy. It is heavily hydrated, with six molecules of water:  $\text{Al}(\text{H}_2\text{O})_6^{3+}$

- (k) Aluminium has a high affinity for oxygen, and generates a lot of heat. This is why aluminium is used in the "thermit" process for the extraction of iron. It reduces iron(III) oxide to the metal.



The "thermit" process is used to weld cracks in steel.

- (l) Aluminium is a good reducing agent.

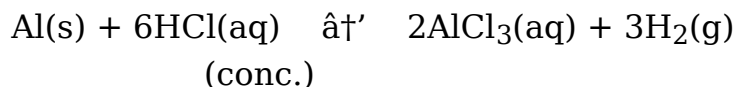


## Uses of Aluminium

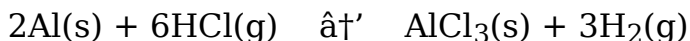
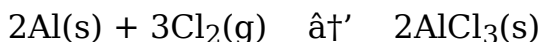
- (i) It is used for making cooking utensils. This is due to its cheapness, low density, high conductivity for heat, and resistance to attack by cooking solutions.
- (ii) It is used in overhead electric cables because of its low density and high electrical conductivity.
- (iii) It is used in making light alloys with high tensile strength. These alloys are used in aircraft construction. They are Duralumin (Al, Mg, Cu, Mn), Magnalium (Mg, Al), and aluminium bronze (Al, Cu).
- (iv) In making aluminium paint. The powdered metal is mixed with oils to make paints.
- (v) In the "thermit" process for the extraction of iron. A mixture of iron(III) oxide and aluminium powder is fired by burning a piece of magnesium ribbon stuck in the mixture. The burning piece of magnesium provides the heat required.
- (vi) It is used in preparing alums used in the treatment of town water supply.

## Aluminium Chloride, $\text{AlCl}_3$

Aluminium chloride can be prepared by dissolving aluminium in concentrated hydrogen chloride acid (hydrochloric acid). By concentrating the solution, hexahydrate crystals of aluminium chloride are obtained.



The anhydrous salt cannot be prepared from the hexahydrate by heating it in air. The crystals react with water. Anhydrous aluminium chloride is prepared by heating aluminium in a stream of dry chlorine or hydrogen chloride.



When the aluminium chloride is produced, it sublimes and collects in the bottle carrying the drying apparatus. The calcium chloride prevents moisture from coming in contact with the aluminium chloride.

By vapour density measurements, aluminium chloride is known to have the formula  $\text{AlCl}_3$  at  $400^\circ\text{C}$  and below. Above  $400^\circ\text{C}$ , it exists as a dimer with the formula  $\text{Al}_2\text{Cl}_6$ .

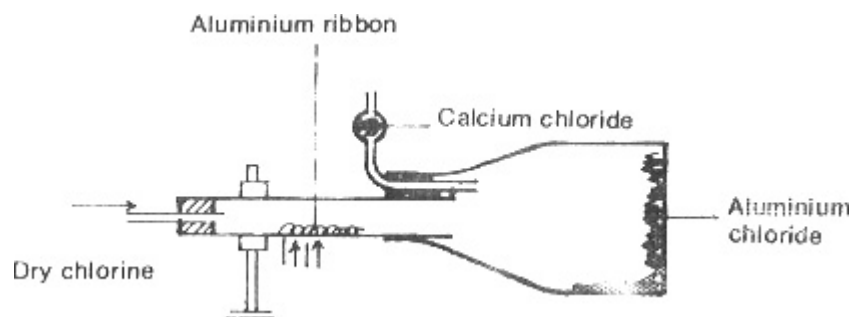
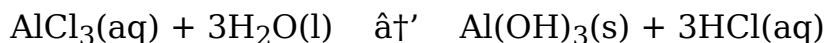


Figure 4.15: Preparation of anhydrous aluminium chloride

## Properties of aluminium chloride

- It is a pale yellow solid. The yellow colour is due to the presence iron(III) chloride impurities
- Anhydrous aluminium chloride sublimes on heating.
- It fumes in moist air, liberating hydrogen chloride gas.
- Anhydrous aluminium chloride is very deliquescent.
- It is readily hydrolysed in water to produce an acidic solution.



If the solution is evaporated to dryness aluminium hydroxide residue will be left.

- Aluminium chloride has a high degree of covalent character.

## Uses of aluminium chloride

- As a catalyst in Friedel-Crafts reaction in organic chemistry.
- It is used in cracking processes in the petroleum industry.

## Alums

Alums have the general formula  $\text{Y}_2\text{SO}_4 \cdot \text{M}_2(\text{SO}_4)_3 \cdot 24\text{H}_2\text{O}$ , where Y is  $\text{Na}^+$ ,  $\text{K}^+$  or  $\text{NH}_4^+$ , and M is  $\text{Al}^{3+}$ ,  $\text{Fe}^{3+}$  or  $\text{Cr}^{3+}$ . Alums are hydrated double tetraoxosulphate(VI) salts. Alums have sharp tastes. The most common is potash alum. It has the formula  $\text{K}_2\text{SO}_4 \cdot \text{Al}_2(\text{SO}_4)_3 \cdot 24\text{H}_2\text{O}$ .

### Experiment 4.11: Preparation of Potash Alum, $\text{K}_2\text{SO}_4 \cdot \text{Al}_2(\text{SO}_4)_3 \cdot 24\text{H}_2\text{O}$

Weigh out equimolar quantities of crystals of potassium tetraoxosulphate(VI) and aluminium tetraoxosulphate(VI) in two separate beakers.

$\text{K}_2\text{SO}_4$	$\text{Al}_2(\text{SO}_4)_3$	Formulae
174 g	342 g	Molar mass

About 0.05 mole of each salt may be used.

Dissolve each separately with as little hot water as possible. To prevent the hydrolysis of the aluminium tetraoxosulphate(VI), acidify the solution with dilute tetraoxosulphate(VI) acid. Mix the two hot

solutions, and heat. Leave to cool. Crystals of potassium aluminium alum will separate out. Note the shape of the crystals. Filter off the crystals. Wash crystals with distilled water and dry between two filter papers.

## Uses of alum

- (i) Alums are used in the purification of town water supply.
- (ii) Alums are used in medicine.

## Aluminium Hydroxide, $\text{Al}(\text{OH})_3$

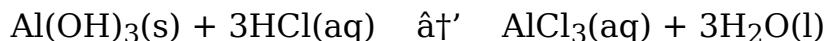
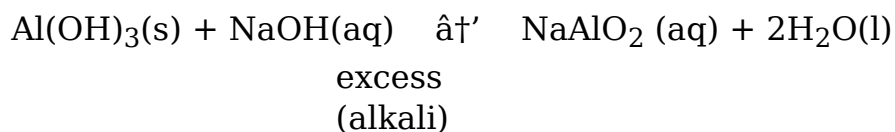
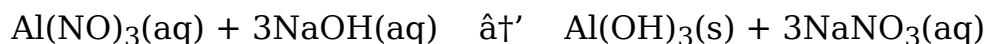
This is prepared by the action of sodium hydroxide solution on a soluble aluminium salt.

*Experiment 4.12: Preparation of Aluminium Hydroxide and study of its properties.*

Put some aluminium trioxonitrate(V) salt in a test-tube. Add water to dissolve it, then add sodium hydroxide solution in drops until a white gelatinous precipitate appears.

Divide the precipitate into two portions. To the first portion, add excess sodium hydroxide solution until the precipitate dissolves. To the second portion, add dilute tetraoxosulphate(VI) acid.

The white gelatinous precipitate dissolves in excess sodium hydroxide solution and it also dissolves in acid. Thus, we can say that aluminium hydroxide neutralises acids and alkalis. That is, it is amphoteric.



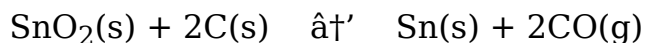
## 4.12 Tin

Tin exhibits oxidation states of +2 and +4. The important ore of tin is cassiterite,  $\text{SnO}_2$ . Cassiterite is an important mineral in Nigeria. It occurs in Jos, Plateau State. It also occurs in Malaya.

### Extraction of Tin

Tin ores are usually very lean. That is, they contain only very small percentages of tin. (1 – 2%). The oxide ore is concentrated by hydraulic washing. Since the ore is much denser than the earthy impurities, these impurities get carried away by water, leaving the ore behind. The concentrated ore is then roasted to convert sulphate and arsenic impurities to their oxides. These volatile oxides escape, leaving tin ore which contains 50 – 70% tin behind.

Reduction of tin(IV) oxide is achieved by heating with coke or coal in a furnace.



The furnace may be a horizontal one, a reverberatory one or a vertical blast furnace. From the furnace, tin often enters the slag which may be acidic or basic. The metal can combine with either due to the amphoteric nature of tin. The slag has to be remelted and heated with more coke to extract tin from it.

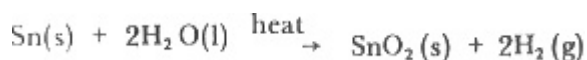
Crude tin is run off from the furnace in its molten form. It is purified by stirring the molten tin with green wooden poles and aerating it. While exposed to air, many impurities still present in the melt get oxidised in preference to tin which is difficult to oxidise. The oxide impurities rise to the surface of the molten tin from where they are removed. The pure tin left behind is run into moulds and cast as ingots.

*Experiment 4.13: Investigating the Reactions of Tin with common Reagents.*

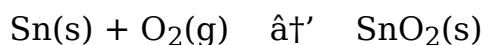
- (i) Use a crucible tong to hold a piece of tin. Burn the piece of tin in air. What do you observe?
- (ii) Repeat the combustion reaction by burning it in pure oxygen.
- (iii) Add a piece of tin to water in a beaker and leave it for some hours. Does it corrode?
- (iv) Add dilute hydrochloric acid to tin foil in a test-tube. Test any gas given off with a burning splint.

**Properties of Tin:**

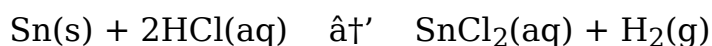
- (a) It is a silvery-white metal.
- (b) It is malleable and has a high metallic lustre.
- (c) It has a low melting point.
- (d) It exhibits allotropy. The allotropes are white tin and grey tin.
- (e) It reacts with steam when hot to form tin(IV) oxide and hydrogen.



- (f) At high temperature, tin burns in air or oxygen to form tin(IV) oxide.

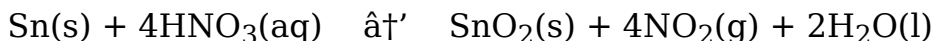


- (g) It does not corrode in air or water, hence it is used for plating.
- (h) It liberates hydrogen from hydrogen chloride acid(hydrochloric acid). Tin is slightly above hydrogen in the activity series.



Dilute acids react slowly with tin with the evolution of hydrogen.

- (i) With concentrated trioxonitrate(V) acid, tin forms tin(IV) oxide.



- (j) Tin reacts with excess chlorine gas to form tin(IV) chloride.



- (k) It reacts with concentrated tetraoxosulphate(VI) acid to form  $\text{Sn(SO}_4)_2$ .

- (l) It reacts with hot concentrated alkalis to form stannates,  $\text{Sn(OH)}_4^{2-}$ , and hydrogen.

### Uses of Tin

- (i) It is used in tin-plating to protect iron against corrosion. The tin-plated iron is used to manufacture tin cans for packing food.
- (ii) It is used in making alloys such as Solder (50% Sn and 50% Pb), Bronze (33% Sn and 67% Cu), and Tin foil (88% Tin, 4% Cu, 7.5% P, 0.5% Sb).

### Tin(IV) Chloride, $\text{SnCl}_4$

Tin(IV) chloride is prepared by passing chlorine over heated tin metal in a hard glass tube.

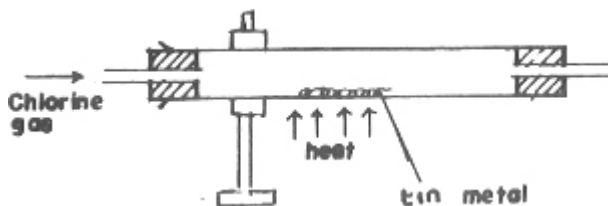
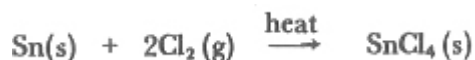


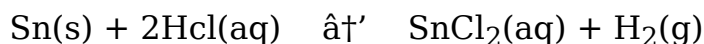
Figure 4.16: Preparation of Tin(IV) chloride.



Tin(IV) chloride is also formed during the process of recovering tin from tin cans, as an intermediate compound. It is used as a mordant in dyeing.

### Tin(II) Chloride Dihydrate, $\text{SnCl}_2 \cdot 2\text{H}_2\text{O}$

This is prepared by the action of dilute hydrogen chloride acid (hydrochloric acid) on tin.



On concentrating the solution and cooling, tin(II) chloride dihydrate,  $\text{SnCl}_2 \cdot 2\text{H}_2\text{O}$  is formed.

Tin(II) chloride is used as a reducing agent and as a mordant in

dyeing and in weighting silk. Silk impregnated with tin(II) chloride undergoes hydrolysis and oxidation and acquires a heavy luxurious appearance.

## CHAPTER SUMMARY

â€¢ Silver, gold and copper are found in the free state, the other metals are found as:

- (i) oxide ores, e.g  $\text{Fe}_2\text{O}_3$ ,  $\text{Al}_2\text{O}_3$ ,  $\text{SnO}_2$
- (ii) halide ores, e.g  $\text{NaCl}$ ,  $\text{KCl}$ ,  $\text{MgCl}_2$ ;
- (iii) sulphide ores, e.g  $\text{ZnS}$ ,  $\text{FeS}$ ,  $\text{PbS}$ ,  $\text{CuS}$ ;
- (iv) trioxocarbonate(IV) ores, e.g.  $\text{CaCO}_3$ ,  $\text{MgCO}_3$ ,  $\text{ZnCO}_3$ ; and
- (v) tetraoxosulphate(VI) ores, e.g  $\text{CaSO}_4$ .

â€¢ Extraction of the metals from their ores generally involves three steps: (i) concentration of the ore, (ii) reduction, and (iii) further purification (refining) of the product.

â€¢ Concentration of the ore may be done by oil floatation, magnetic separation, or hydraulic washing.

â€¢ Reduction of the concentrated ore may be done by electrolytic process or by the use of a reducing agent such as carbon (coke) or carbon(II) oxide.

## Alkali Metals

â€¢ The alkali metals have relatively low densities, and low melting and boiling points when compared with other metals.

â€¢ Extraction of the metals is by electrolysis of their fused halide ores.

â€¢ They are all very reactive, their reactivities increase down the group.

â€¢ They react violently with cold water and with dilute mineral acids, liberating hydrogen.

â€¢ They tarnish in air due to oxide formation.

â€¢ Their oxides and hydroxides are soluble in water, forming alkaline solutions

â€¢ They also form peroxides which yield hydrogen peroxide in water.

â€¢ All their common salts are soluble in water.

â€¢ They form salt-like hydrides in which hydrogen carries a negative charge

â€¢ The important compounds of sodium include:

- (i) sodium hydroxide which is made by the electrolysis of brine,
- (ii) sodium chloride which is found in nature,
- (iii) sodium trioxocarbonate(IV) which is made by the Solvay-ammonia process.

â€¢ Sodium hydroxide is used for making soap, paper and the manufacture of other chemicals.

â€¢ Sodium chloride is used for making sodium hydroxide and sodium trioxocarbonate(IV).

â€¢ Sodium trioxocarbonate(IV) is used for making glass, washing powder, and for softening hard water.

## **Alkaline Earth Metals**

â€¢ The alkaline earth metals are extracted by electrolysis of their fused halides.

â€¢ They are not as reactive as the alkali metals.

â€¢ They react gently with cold water, liberating hydrogen.

â€¢ Their oxides and hydroxides are slightly soluble in water, producing weakly alkaline solutions.

â€¢ Their chlorides and trioxonitrate(V) salts are soluble in water, but their trioxocarbonate(IV) and tetraoxosulphate(VI) salts are insoluble.

â€¢ Their hydroxides, trioxocarbonates(IV) and trioxosulphates(IV) are decomposed by heat to the oxides.

â€¢ Important compounds of calcium include:

- (i) calcium oxide (quicklime), which is made in lime kilns by decomposition of calcium trioxocarbonate(IV) (limestone),
- (ii) calcium hydroxide which is made by slaking quicklime,
- (iii) calcium trioxocarbonate(IV) which is found in nature, and
- (iv) bleaching powder, which is made by bubbling chlorine into slaked lime,  $\text{Ca(OH)}_2$

## **Groups III and IV Metals**

â€¢ Aluminium is by far the most important member of group III.

â€¢ Aluminium is extracted from its bauxite ore by purifying the ore before electrolysis of the pure oxide dissolved in molten cryolite.

â€¢ Aluminium burns in air to form both aluminium oxide and aluminium nitride.

â€¢ Aluminium is active enough to liberate  $\text{H}_2$  from acids but not from water.

- â€¢ Aluminium is not attacked by trioxonitrate(V) acid due to a protective oxide formation.
- â€¢ It has a high affinity for oxygen and so aluminium is used in the thermit process.
- â€¢ Aluminium oxide and aluminium hydroxide are amphoteric.
- â€¢ Aluminium is used for making cooking utensils, overhead electrical cables and light alloys for aircraft construction.
- â€¢ Tin is a group IV metal.
- â€¢ It is extracted from its oxide ore, cassiterite,  $\text{SnO}_2$ .
- â€¢ Concentration of the ore is by washing, reduction is with coke, and purification is by aeration.
- â€¢ Tin does not corrode in air or in water.
- â€¢ Tin liberates hydrogen from non-oxidizing acids but forms tin(IV) oxide with concentrated  $\text{HNO}_3$ .
- â€¢ Tin finds use in plating, to protect iron from rusting, and in making alloys such as solder, type metal and bronze.

## ASSESSMENT

1. How is aluminium obtained from crude bauxite? Why is aluminium used in (i) cooking-utensils, (ii) aircraft industry?  
Why is aluminium oxide regarded as an amphoteric oxide? Write equations to illustrate the reaction between aluminium and sodium hydroxide.
2. What is the reaction of sodium hydroxide with solutions of :  
(a) Copper(II) ions,  
(b) Iron(III) ions  
(c) Zinc ions,  
(d) Aluminium ions.
3. (a) List three physical and three chemical properties of metals as compared to non-metals.  
(b) Name three methods by which you would prevent rusting?  
Write the formula for iron rust. What conditions are necessary for rusting to take place?
4. What are alloys? Give two examples of alloys and their composition.
5. (a) How do aluminium and calcium react with:  
(i) water,  
(ii) oxygen?  
(b) What is the reaction if any, when



- (i) a rod of zinc is placed in a solution of iron(II) tetraoxosulphate(VI).
- (ii) a rod of copper is placed in a solution of copper(II) tetraoxosulphate(VI).
- (iii) a zinc rod is placed in copper(II) trioxonitrate(V). Give a brief explanation of your answer in each case.

(WAEC)

6. Describe an electrolytic method for the production of sodium hydroxide. Give the chemistry of the reactions involved in the process. How does sodium hydroxide react with chlorine?

(WAEC)