

# 5. Metals of the First Transition Series

## 5.1 General Properties

Metals of the first transition series include: scandium (atomic number 21), titanium (22), vanadium (23), chromium (24), manganese (25), iron (26), cobalt (27), nickel (28), copper (29) and zinc (30). With the exception of zinc, the elements are packed in face-centred cubic, hexagonal close-packing and body-centred cubic.

### Physical Properties

- (a) They are good conductors of heat and electricity. Copper is the best conductor of the series. Electrical conductivity of the metals is associated with delocalised electrons in the metal crystal lattice.
- (b) The metals are hard, and have high melting points because they use their 4s and 3d electrons for metallic bonding. Copper and iron are fairly soft, malleable and ductile whereas chromium and vanadium are much harder and more brittle. The difference in behaviour lies in their crystalline structures.
- (c) Some of the metals are attracted by a magnetic field. This is due to unpaired electrons in their shells. Iron, cobalt and manganese are ferromagnetic. Ferromagnetic metals remain permanently magnetized after magnetization.

### Chemical Properties

- (a) **Oxidation states:** The metals exhibit variable oxidation states. This characteristic is a reflection of the fact that the 3d electrons as well as the 4s electrons participate in oxidation and reduction. Table 5.1 shows the different oxidation states of the first row transition metals.

**TABLE 5.1: Oxidation states of the first row transition elements**

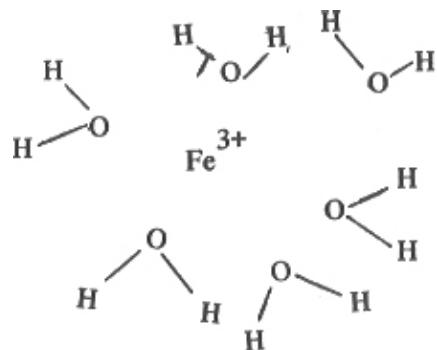
Element	Oxidation states
Sc	+3
Ti	+2, +3, +4
V	+2, +3, +4, +5
Cr	+2, +3, +6
Mn	+1, +2, +3, +4, +6, +7
Fe	+2, +3, +6
Cobalt	+2, +3, +4

Nickel	<b>+2, +3, +4</b>
Copper	<b>+1, +2</b>
Zinc	<b>+2</b>

Oxidation numbers in bold are more common and stable.

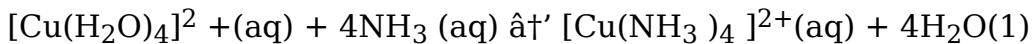
- (b) **Complex ion formation:** The metals of the first transition series form many complex ions because they have vacant d-orbitals for accepting lone pairs of electrons. The small size of the transition metal ions also enhances their ability to form complex ions.

Their ions are usually hydrated in aqueous solution. For example, iron(III) ion is surrounded by six water molecules arranged octahedrally round the ion to give  $[\text{Fe}(\text{H}_2\text{O})_6]^{3+}$

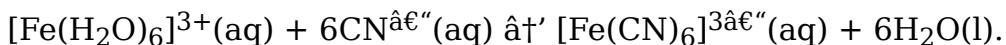


The water molecules surrounding the central iron ion are known as ligands. The charge clouds of these water molecules exert some influence on the structure of the central ion. Copper(II) ion in aqueous solution is also surrounded by four water molecules:  $[\text{Cu}(\text{H}_2\text{O})_4]^{2+}$ . Other ions and molecules like  $\text{CN}^-$  and  $\text{NH}_3$  can also act as ligands.

The water molecules surrounding a central atom can be replaced by other ligands such as  $\text{CN}^-$  and  $\text{NH}_3$ . Thus, we say that the water molecules undergo ligand exchange.



The complex ion formed in the case of copper(II) ion occurs when aqueous ammonia is added in excess to a solution of a Cu(II) ion. In similar manner, the water molecules surrounding the iron ion can be replaced.



- (c) **Coloured ions:** The transition metal ions in aqueous solution show a variety of colours. Copper(II) ions in aqueous solution or as the crystalline hydrated salts are light blue. Anhydrous copper(II) ions are white. Manganese(VII) ions are purple, and manganese(II) ions are colourless. Heptaoxochromate(VI) ions are orange, while  $\text{Cr}^{3+}$  is green. Hexacyanoferrate(II) ion, Hexacyanoferrate(III) ion, and diamine silver(II) ion are also coloured. It is the hydrated cation that is responsible for the colour of these compounds. The complex ions have some internal covalent bonding because the ligands possess lone pairs of electrons which they

use for bonding.

The colour of the transition metal ions arise as a result of their possession of unpaired d<sup>5</sup> electrons. These electrons are capable of absorbing small quantities of energy from the sun's spectrum. When all colours are present in the sun's spectrum, it appears colourless. When some of the colours are absorbed, only those not absorbed are reflected. The absorbing transition metal ion therefore assumes the colour of those colours reflected.

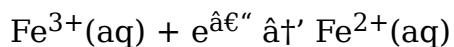
(d) **Catalytic actions:** Transition metals and their compounds are often used as catalysts. Table 5.2 shows some of the transition metals and their compounds and the reactions they catalyse.

**Table 5.2 : Transition metals and their compounds which are used as catalyst.**

Transition metal/compound	Reaction catalysed.
(i) MnO <sub>2</sub>	$2\text{KClO}_3 \xrightarrow{\text{heat}} 2\text{KCl} + 3\text{O}_2$ $2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2$
(ii) V <sub>2</sub> O <sub>5</sub> or Pt	$2\text{SO}_2 + \text{O}_2 \xrightleftharpoons{\text{Pt}} 2\text{SO}_3$
(iii) Fe (finely divided)	$\text{N}_2 + 3\text{H}_2 \xrightleftharpoons{\text{Fe}} 2\text{NH}_3$
(iv) Pt or Pd or Ni	$\text{C} = \text{C} + \text{H}_2 \xrightleftharpoons{\text{Ni}} \text{H}_2\text{C} = \text{C} = \text{H}$

These catalytic actions are due to the fact that:

(i) the transition metal ions readily undergo reduction/oxidation.

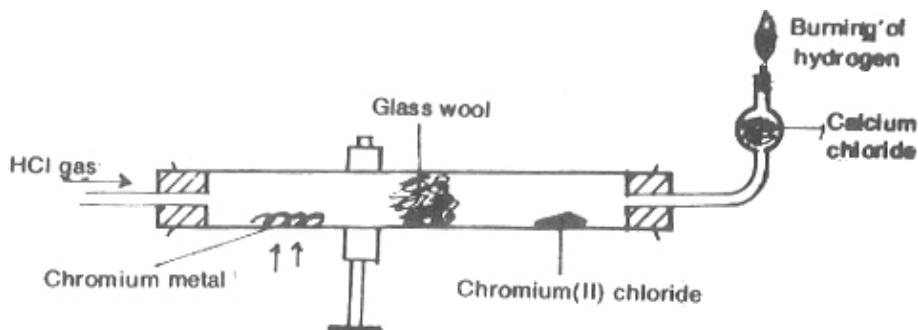


In the catalytic actions of these transition metals and their compounds they donate or accept electrons from the reactants. This speeds up the chemical reaction.

(ii) the metals readily adsorb gases on their surfaces thereby making the gas available for reaction.

*Experiment 5.1: (Demonstration): Investigating the oxidation states of chromium.*

(i) Put chromium metal in a hard boiling-tube and heat. Pass dry hydrogen chloride gas over the heated chromium.



*Figure 5.1: Preparation of chromium(II) chloride.*

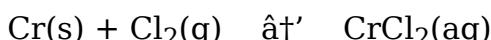
(ii) Put some chromium(III) chloride solution in a test-tube. Add sodium

trioxosulphate(IV) solution to it, and warm. What do you observe?

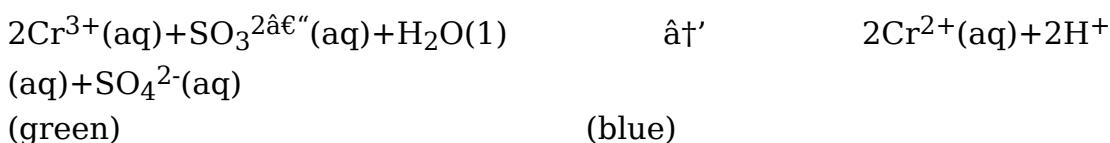
- (iii) Put potassium heptaoxodichromate(VI) solution in a test-tube. Add sodium trioxosulphate(IV) solution to it followed by dilute tetraoxosulphate(VI) acid and warm. Does the colour change from orange to green to blue?
- (iv) Put some potassium heptaoxodiochromate(VI) solution in a test-tube. Add zinc dust or magnesium ribbon to it followed by dilute hydrochloric acid. Observe the colour change from orange to green to blue.
- (v) Decant the blue solution obtained from (iv) and leave it in the open. Does the colour change back to green?

Experiment 5.1 shows the changes in colour as the oxidation state of chromium changes:

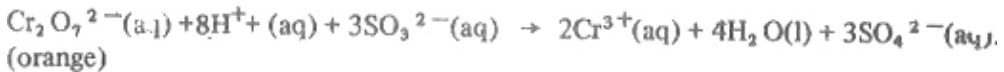
- (i) shows the oxidation of chromium metal to chromium(II)



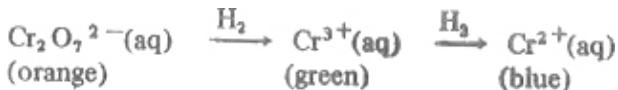
- (ii) shows the reduction of chromium(III) ion to chromium(II)



- (iii) shows the reduction of  $\text{Cr}^6$  to  $\text{Cr}^{3+}$  brought about by sulphur(IV) oxide, a reducing agent.



- (iv) the reduction here is by active hydrogen. First it reduces  $\text{Cr}^{6+}$  to  $\text{Cr}^{3+}$ , then further to  $\text{Cr}^{2+}$



- (v) the change in colour here is due to the atmospheric oxidation of  $\text{Cr}^{2+}$  back to  $\text{Cr}^{3+}$ .

## 5.2 Compounds of Transition Metals

### Chlorides:

Most of the chlorides show some typical covalent character. Most of their solid chlorides vapourize on heating, and are readily hydrolysed in aqueous solution. They are often soluble in organic solvents. In general, the higher oxidation states of the transition metals form chlorides which are covalent in character while the lower oxidation states are ionic.

The chlorides of iron and copper form complex ions. Chromium(II) chloride is a powerful reducing agent. It is also readily oxidised to chromium(III) chloride. Chromium(III) chloride has some ionic character.

## **Oxides:**

Higher oxidation state oxides are acidic. For example, chromium(VI) oxide forms chromium(III) compounds in alkaline solutions. Chromium(VI) oxide is soluble in both alkalis and acids.  $\text{Cr}(\text{OH})_3$  also dissolves in both acidic and alkaline solution, that is it is amphoteric.

## **Tetraoxosulphate(VI) Salts:**

The transition metals form stable tetraoxosulphate(VI) salts. With potassium, sodium and ammonium tetraoxosulphate(VI) salts, they form alums.

## **Trioxonitrate(V) salts:**

They all form trioxonitrate (V) salts which are decomposed by heat.



## **Trioxocarbonate(IV) Salts:**

Not all the transition metals form trioxocarbonate(IV) salts. All the trioxocarbonate (IV) salts are decomposed by heat.

**TABLE 5.3: Colours of some transition metal ions**

<b>Transition Metal</b>	<b>Ions</b>	<b>Colour of Ion</b>
Copper	$\text{Cu}^+$	White
	$\text{Cu}^{2+}$	Blue
Iron	$\text{Fe}^{2+}$	Green
	$\text{Fe}^{3+}$	Orange/Yellow
Chromium	$\text{Cr}^{2+}$	Blue
	$\text{Cr}^{3+}$	Green
	$\text{Cr}^{6+}$	Orange

## **5.3 Copper**

The name of copper is derived from the latin word ‘cuprum’ which was derived from Cyprum. Cyprum refers to the island of Cyprus. Copper materials which are several hundred years old have been found in Egypt. Its early production was due to its existence in the free state.

On exposure to the atmosphere, copper forms an oxide/sulphide film on its surface, thus rendering it ‘passive’.

Copper occurs as copper pyrites,  $\text{CuFeS}_2$ ; copper glance,  $\text{Cu}_2\text{S}$ ; malachite,  $\text{CuCO}_3\cdot\text{CuO}$ ; and cuprite,  $\text{Cu}_2\text{O}$ . Copper is also found as the metal in the Lake Superior region (USA), Arizona (USA) and New Mexico.

## **Extraction of Copper**

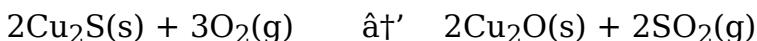
The extraction of copper is carried out in three stages: “concentration, extraction and refining”:

- (a) **Oil flotation (or concentration):** This first stage has been fully

described earlier in Chapter 4 of this book.

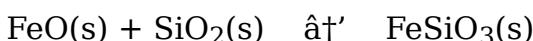
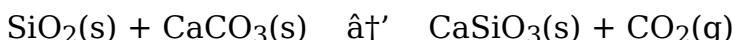
(b) **Extraction:** This involves three stages :-

(i) *Roasting:* After the concentration of the ore, it is strongly heated in air to convert part of the ore to the copper(I) oxide.



Some of the copper sulphide remains unchanged.

(ii) *Smelting:* The roasted ore is heated with limestone,  $\text{CaCO}_3$ , in a vertical cylindrical furnace. This helps to remove silica and iron oxides as slag which floats on top of the melted copper oxide – sulphide mixture.



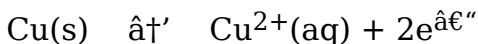
(iii) *Reduction of the copper(I) sulphide – oxide mixture:* The mixture is strongly heated in a Bessemer converter with a hot blast of air. The copper oxide and sulphide are converted to metallic copper.



The molten copper is poured out into slabs for cooling and solidification. The copper obtained is called ‘blister copper’ because the gases escaping from the copper as it solidifies give the metal a blistered appearance.

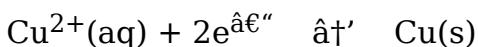
(c) Refining of copper: This is done by the electrolysis of copper(II) tetraoxosulphate(VI) solution containing some dilute tetraoxosulphate(VI) acid as the electrolyte. Impure copper, i.e., ‘buster’ copper is made into bars and rods are used as the anode. Pure copper is made the cathode.

*Chemistry of the process of electrolytic purification:* Copper from the impure copper anode goes into solution as copper ions by each copper atom losing two electrons:



The impure copper anode becomes reduced in size.

Copper ions in solution migrate to the cathode. Each copper ion gains two electrons, to become a copper atom. Copper is thus deposited at the cathode.



The impurities collect under the anode at the bottom of the electrolytic cell as sludge. Precious metals like gold, silver and platinum which are associated with copper ores are not discharged during the electrolysis. So, they collect in the sludge where they are later recovered.

*Experiment 5.2: Investigating the properties of copper.*

- (i) Examine a piece of copper plate. What is the colour?
- (ii) Place some copper turnings in a test-tube and add dilute trioxonitrate(V) acid to it. Does the copper dissolve? Repeat, using dilute hydrogen chloride acid (hydrochloric acid).
- (iii) Clean the surface of a piece of copper plate with sand-paper. Expose it to the atmosphere for a few hours. Does the surface tarnish?

### **Properties of Copper:**

- (a) Copper metal has a characteristic reddish-brown colour.
- (b) It has a density of 8.9 g/cm<sup>3</sup> and a melting point of 1083°C.
- (c) Copper resists atmospheric corrosion when exposed to air except for superficial tarnishing owing to the formation of a thin surface layer of copper(II) oxide/sulphide.
- (d) Copper is a good conductor of heat and electricity.
- (e) Copper does not react with non-oxidising acids such as hydrochloric acid. It reacts with:
  - (i) dilute trioxonitrate(V) acid;
  - $3\text{Cu(s)} + 8\text{HNO}_3\text{(aq)} \rightarrow 3\text{Cu(NO}_3)_2\text{(aq)} + 2\text{NO(g)} + 4\text{H}_2\text{O(l)}$
  - (ii) concentrated trioxonitrate(V) acid.
  - $\text{Cu(aq)} + 4\text{HNO}_3\text{(aq)} \rightarrow \text{Cu(NO}_3)_2\text{(aq)} + 2\text{NO}_2\text{(g)} + 2\text{H}_2\text{O(l)}$
  - (iii) hot, concentrated tetraoxosulphate(VI) acid.
  - $\text{Cu(s)} + 2\text{H}_2\text{SO}_4\text{(aq)} \rightarrow \text{CuSO}_4\text{(aq)} + \text{SO}_2\text{(g)} + 2\text{H}_2\text{O(l)}$

### **Uses of Copper**

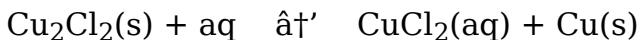
- (i) It is used as an electrical conductor in electrical wiring.
- (ii) It is used in making kettles, cups, pans and brewery stills.
- (iii) It is used in making roofing sheets.
- (iv) It is used for making plates for printing books.
- (v) It is used extensively in making alloys. Alloys of copper include: Bronze (Cu Sn, Zn), Brass (Cu, Zn), Aluminium bronze(Cu, Al), and German silver (Cu, Zn, and Ni).

### **Compounds of Copper:**

Copper exhibits two oxidation states, +1 and +2. The +2 oxidation state compounds are the more important. Covalent copper(I) compounds are more stable than covalent copper(II) compounds. Thus, on heating, copper(II) chloride and copper(II) bromide decompose, yielding the corresponding copper(I) compounds.

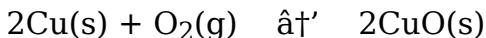


However, copper(I) ion does not exist in aqueous solution. It rapidly changes to copper(II) ion and metallic copper.

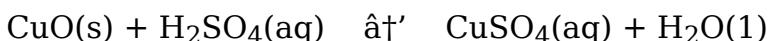
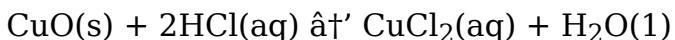
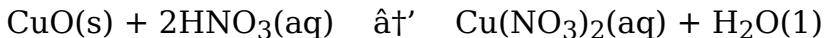


## **Copper(II) Oxide, CuO**

Copper(II) oxide is a black solid. It is prepared by heating copper in air or oxygen.



It reacts with acids to form the corresponding copper(II) salt.



All the resulting copper(II) salts are blue in colour in solution.

Copper II oxide does not decompose on heating and it is insoluble in water.

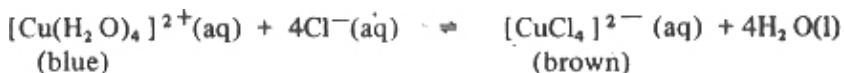
## Copper(I) Oxide, Cu<sub>2</sub>O

Copper(I) oxide has a reddish appearance. Copper(I) salts are unstable in solution. They decompose to form copper(II) salts and copper.

## **Copper(II) Chloride, CuCl<sub>2</sub>**

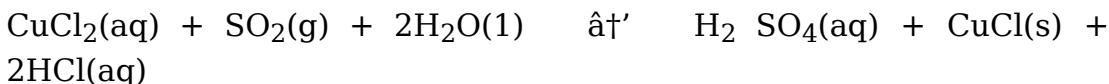
The direct combination of copper and chlorine (by heating copper in a stream of dry chlorine gas) yields anhydrous copper(II) chloride. This solid is brownish-yellow in colour. When prepared by heating copper(II) oxide with dilute hydrogen chloride acid (hydrochloric acid), the solution is filtered and concentrated by evaporation to obtain green crystals of copper(II) chloride dihydrate,  $\text{CuCl}_2\text{H}_2\text{O}$ .

A dilute solution of the solid in water is blue in colour, whereas its concentrated solution is brown. The difference in colour is due to changes occurring in the nature of the complexing agent.



## **Copper(I) Chloride, CuCl**

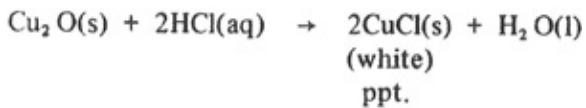
This is prepared by passing sulphur(IV) oxide into copper(II) chloride solution. Copper(I) chloride is precipitated.



Copper(I) trioxonitrate(V) and copper(I) tetraoxosulphate(VI) salts are unstable in aqueous solution.

## *Experiment 5.3: Investigating the Reactions of Copper(I) Oxide*

- (i) Put copper(I) oxide in a beaker. Add concentrated hydrogen chloride acid (hydrochloric acid) to it until a deep brown solution is observed. Pour the solution into water to obtain a white precipitate of copper(I) chloride.



- (ii) Put some copper(II) tetraoxosulphate(VI) solution in a test-tube. Add aqueous ammonia until the precipitate becomes soluble. Boil the solution with glucose solution. What is the colour of the precipitate? (Red precipitate of copper(I) oxide is obtained).

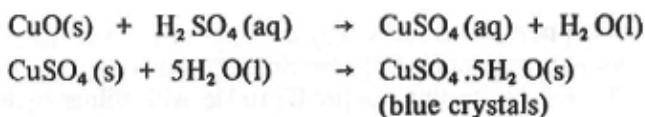
Glucose reduces the copper(II) ion to copper(I) ion. Since copper(I) oxide is insoluble, it is stable. This is the test for reducing sugars.

### **Copper(II) Tetraoxosulphate(VI), CuSO<sub>4</sub> .5H<sub>2</sub>O**

Copper(II) tetraoxosulphate(VI) crystals CuSO<sub>4</sub>.5H<sub>2</sub>O, are blue in colour, hence the crystal is called blue vitriol.

#### *Experiment 5.4: Preparation of Copper(II) Tetraoxosulphate(VI) Crystals:*

Add some copper(II) oxide to 2 M tetraoxosulphate(VI) acid solution in an evaporating dish. The solid copper(II) oxide dissolves into a blue solution. Filter off excess copper(II) oxide. Heat the filtrate to concentrate. Test for saturation periodically until a test sample gives crystals on cooling. Cool the whole solution to obtain crystals. Filter off the crystals, wash and dry between two filter papers.

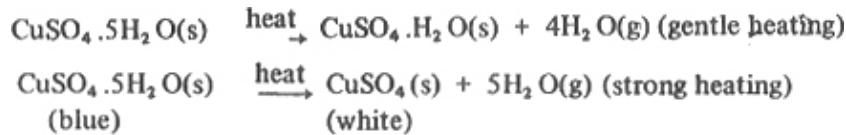


#### *Experiment 5.5: Investigating the Properties of Copper(II) Tetraoxosulphate(VI)*

- Put some copper(II) tetraoxosulphate(VI) crystals in a test-tube and heat strongly. Does the colour of the crystals change?
- Put some copper(II) tetraoxosulphate(VI) solution in a test-tube. Add 2M sodium hydroxide solution in drops and later in excess. Is there any precipitate formed? What is the colour of the precipitate?
- To another portion of copper(II) tetraoxosulphate(VI) solution, add dilute aqueous ammonia in drops. Is there any precipitate formed? What is the colour of the precipitate? Add excess dilute aqueous ammonia. Does the precipitate dissolve? What is the colour of the solution, if any?
- Put some copper(II) tetraoxosulphate(VI) solution in a test-tube. Add potassium iodide solution to it. Identify any gas given off by testing with starch paper.

### **Properties of Copper(II) Tetraoxosulphate(VI):**

- On heating copper(II) tetraoxosulphate(VI) crystals, the colour changes from blue to white. This is because it loses its water of crystallization

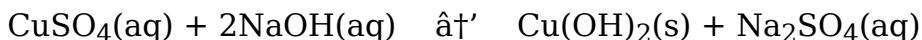


- (b) If the white powder of copper(II) tetraoxosulphate(VI) is heated very strongly, it decomposes to form copper(II) oxide and sulphur(VI) oxide.

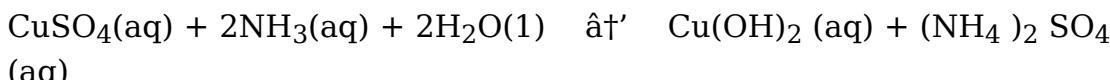


Anhydrous copper(II) tetraoxosulphate(VI) which is white, turns blue when water is added to it. It is therefore used to test for water.

- (c) If sodium hydroxide solution is added to copper(II) tetraoxosulphate(VI) solution in a test-tube, a blue precipitate is formed. The precipitate is insoluble in excess sodium hydroxide.

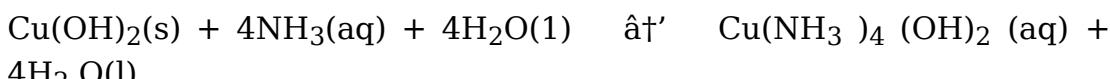


- (d) By adding aqueous ammonia in drops to a solution of copper(II) tetraoxosulphate(VI), a blue precipitate is first formed:



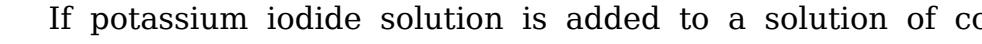
(blue ppt)

On adding excess aqueous ammonia, the blue precipitate dissolves, forming a deep blue solution.



(Tetramine Copper(II) hydroxide)

- The deep blue colour is due to the presence of tetramine copper(II) ion.



tetraoxosulphate(VI), copper(I) iodide is formed.

## Iodine is a

**5.4 Iron** Iron is a  $\hat{G}^{\text{S1}}$  and  $\hat{G}^{\text{TM}}$  metal. It has been extensively studied due to its biological importance.

The names given to iron by the ancient Egyptians, Sumerian and Hebrew signify *âfœ*the metal from heavenâœ. This indicates an early

About 5% of the earth's crust is made up of iron. The interior of the earth contains a large proportion of iron. Although iron compounds may be found widely dispersed in the soil and rocks, this does not mean that iron

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Iron is not found free in nature. It occurs as haematite ( $\text{Fe}_2\text{O}_3$ ), magnetite ( $\text{Fe}_3\text{O}_4$ ), limonite ( $2\text{Fe}_2\text{O}_3 \cdot 3\text{H}_2\text{O}$ ), siderite ( $\text{FeCO}_3$ ), and iron pyrites ( $\text{FeS}_2$ ). The most important ore of iron, from which it is mainly extracted, is haematite. Haematite is dark grey.

Iron ore deposits occur in Ajaokuta and other parts of Nigeria.

## Extraction

Apart from the iron ore, other raw materials for the extraction of iron are coal (or coke) and limestone. The site of an extraction plant is determined by the availability of ore in large deposits, chemical and physical states of the ore and the costs of transportation of raw materials and finished product.

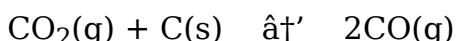
The iron ore is usually reduced in a blast furnace with coke.

The furnace is charged from the top with alternate layers of iron ore, (haematite,  $\text{Fe}_2\text{O}_3$ ), coke and limestone. Preheated air (hot blast of air) is forced into the bottom of the furnace through tubes called tuyeres. There is a space at the bottom for the collection of the molten iron. The working temperature is usually very high.

**Chemistry of the extraction:** Carbon in the hot coke is converted to carbon(IV) oxide by the oxygen in the blast of hot air at the bottom of the furnace. This reaction is highly exothermic and raises the furnace temperature to about  $1500^\circ\text{C}$



As the carbon(IV) oxide rises up the furnace, it comes in contact with more red hot coke and less oxygen and is therefore reduced by the red hot coke to carbon(II) oxide. This reaction is endothermic.



The hot carbon(II) oxide rises up the tower and comes in contact with the hot iron(III) oxide, coke and limestone going down from the top. At a temperature of about  $600^\circ\text{C}$ , the iron(III) oxide is reduced by carbon(II) oxide to iron(II) oxide.

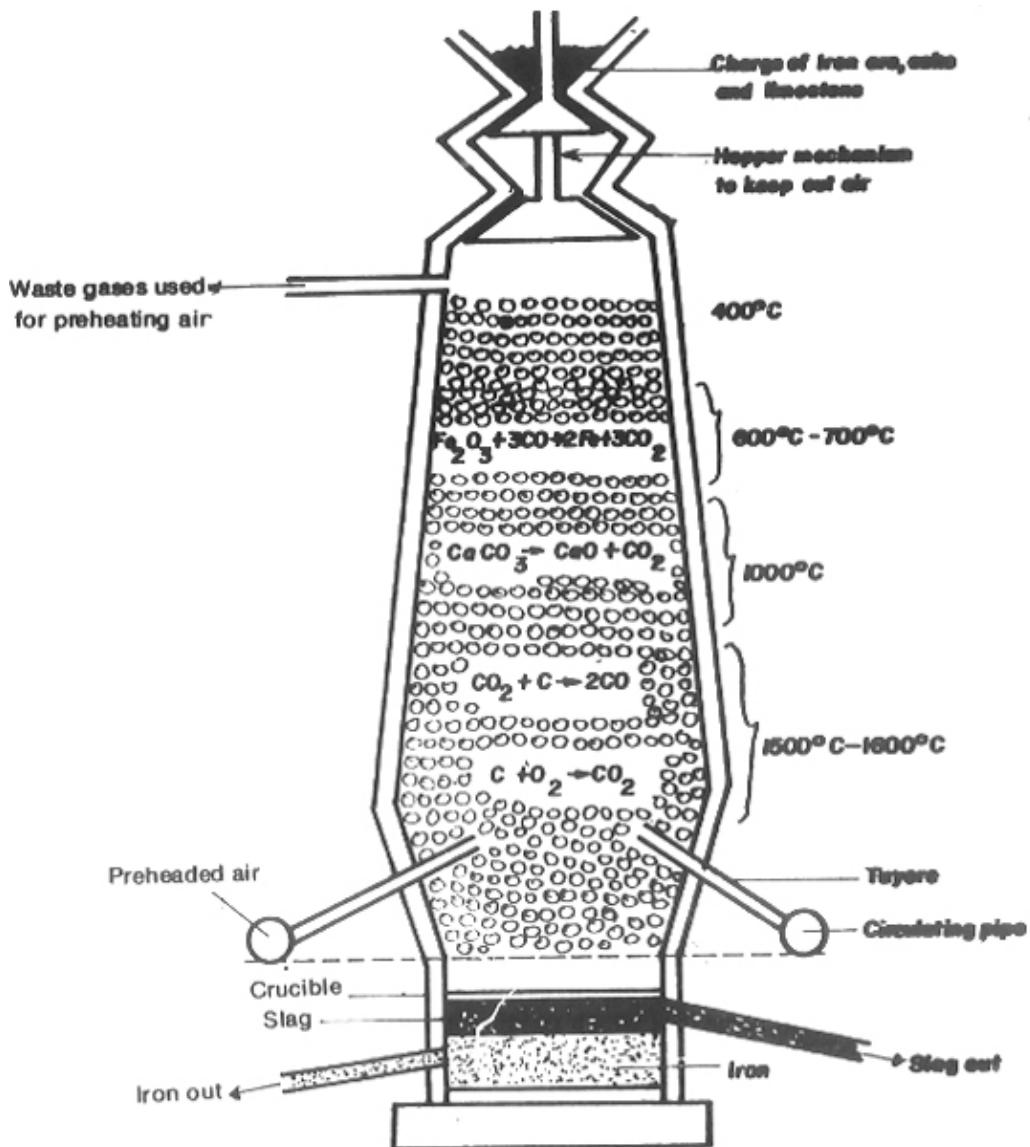
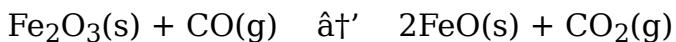
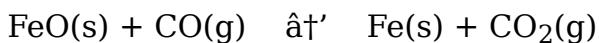


Figure 5.3: The blast furnace for the extraction of iron



As the iron(II) oxide moves down the tower and as the charge continues, the temperature increases. Iron(II) oxide is further reduced by carbon(II) oxide to iron.

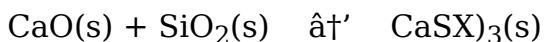


The reduced iron is melted by the high temperature of the lower region of the blast furnace. The molten iron is then tapped off at the bottom of the blast furnace periodically.

Decomposition of limestone also occurs, producing carbon(IV) oxide, for further reduction to carbon(II) oxide by coke.



The calcium oxide produced by the decomposition helps to remove sand by reacting with it to form calcium trioxosilicate(IV).



The calcium trioxosilicate(IV) and other impurities which are melted by the high temperature form slag which floats on top of the molten iron.

Calcium trioxosilicate(IV) is used in road building and for making cement.

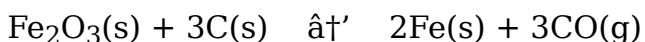
## Types of Iron

**Pig iron:** The iron from the blast furnace is called pig iron. It is poured into large containers in order to solidify. Pig iron is impure. It contains about 2 – 5% of carbon and some other impurities such as manganese, sulphur and silicon.

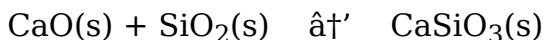
Pig iron is hard because of the impurities it contains. It cannot be worked by hammering. It is used for making materials which will stand up to shocks, like stoves, railings and cylinder blocks of engines.

**Cast iron:** Cast iron is remelted pig iron mixed with steel scrap and allowed to cool in moulds. The impurities in the original pig iron are still present in cast iron. Therefore, it has the same properties as pig iron and is employed for the same purpose.

**Wrought iron:** This is the purest form of commercial iron. It is made by oxidizing most of the impurities in pig or cast iron in a reverberatory furnace. The furnace is lined with haematite which oxidizes the carbon impurity to carbon(II) oxide.



Other impurities such as sulphur, silicon and phosphorus are also oxidized. Limestone is added to the molten pig iron. The oxides of impurities form slag with calcium oxide resulting from the decomposition of limestone.



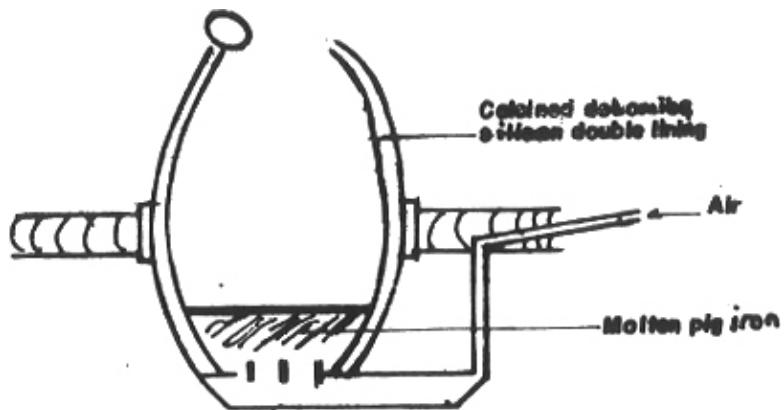
The molten, purified iron and the slag are tapped off at different points of the furnace. Wrought iron produced in this way contains less than 0.2% of carbon. It is soft and malleable and melts at about  $1400^{\circ}\text{C}$ . It can be welded, and so is used for making chains, nails, bolts and framework for many purposes.

## Steel

Steel is an alloy of iron containing some other metals and a small quantity of carbon (about 0.15 to 1.5%). The properties of steel depend largely on the percentage of carbon present and how the carbon is mixed with the iron. Steel is the most commonly used form of iron.

## How to produce different types of steel:

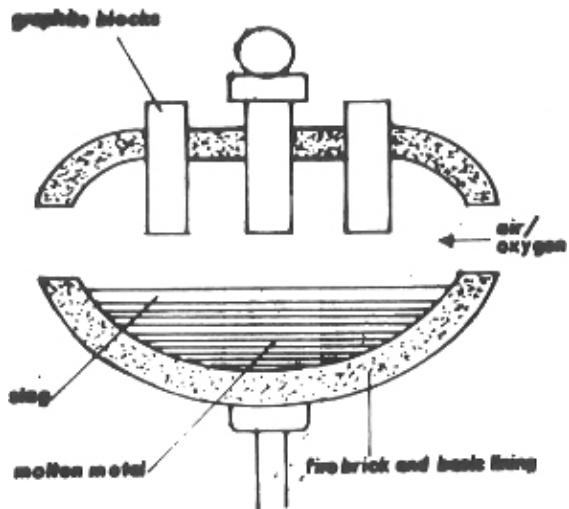
*The Bessemer Process.*



*Figure 5.4: The Bessemer converter*

Bessemer converter is a large egg-shaped vessel rotating on a horizontal axis. Molten pig iron is fed into it and a strong hot blast of air is blown in. Sulphur, carbon, and other impurities are oxidized to their volatile oxides. Silicon and manganese react with the calcium oxide and magnesium oxide (calcined dolomite) in the lining of the furnace to form slag. The slag is easily removed. Appropriate amounts of carbon and manganese are then added to produce the required type of steel.

(b) *The Open-hearth process*



*Figure 5.5: The Open-hearth converter*

Pig iron, scrap steel and iron(III) oxide are heated in a shallow rectangular blast furnace. Dolomite or limestone is used to line the hearth and impurities such as sulphur and carbon are oxidized to gaseous products. Calculated amount of scrap metal is used in proportion to the pig iron to give the required steel.

The Open hearth process is slower than the Bessemer process, but it is easier to control.

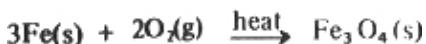
(c) *The Linz-Donawitz (Lâ€"D) process:* This is the most modern process for producing steel. A blast of oxygen is passed over the molten iron, causing the immediate burning of carbon and sulphur to carbon(IV)

oxide and sulphur(IV) oxide respectively. Powdered lime is added so that silicon(IV) oxide reacts with it to form calcium trioxosilicate(IV) (slag).

The Lâ€"D process is faster and equal in purity to the Open-hearth process and does not require large quantities of scrap steel.

## Properties and Reactions of Iron

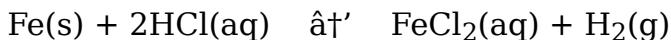
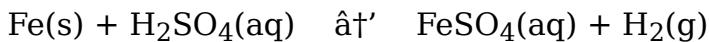
- (a) It has a high melting point of  $1535^{\circ}\text{C}$  and a high density of  $7.878 \text{ g cm}^{-3}$ .
- (b) Pure iron is grey in appearance.
- (c) Iron burns in air and in oxygen at very high temperatures to form triiron tetroxide.



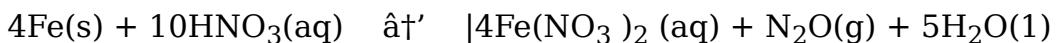
- (d) Red hot iron reacts with steam to form triiron tetraoxide.



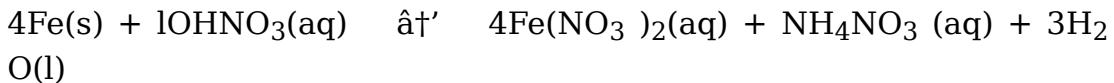
- (e) When exposed to the atmosphere, iron rusts. Iron rust is hydrated iron(III) oxide,  $2\text{Fe}_2\text{O}_3 \cdot X\text{H}_2\text{O}$ . The composition of iron rust is variable and it often contains some iron(II) trioxocarbonate(IV).
- (f) Iron reacts with dilute tetraoxosulphate(VI) acid, and hydrogen chloride acid (hydrochloric acid) to form the corresponding salts. Hydrogen gas is also liberated.



- (g) Dilute trioxonitrate(V) acid does not react with iron to produce hydrogen. This is because trioxonitrate(V) acid is a strong oxidizing agent.



With very dilute trioxonitrate(V) acid, iron forms ammonium trioxonitrate(V).



Concentrated trioxonitrate(V) acid forms an oxide film on the iron and this renders it passive.

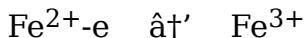
## Uses

- (i) Cast iron is used for making pipes, lamp parts, railings, retort stands and heavy bases for appliances.
- (ii) Steel is used to make stainless steel for cutlery, machinery, steel plates, steel rods and tubes.
- (iii) Wrought iron is used for making iron gates, chairs, chains, etc.

(N.B: Since industrialization depends to a large extent on the availability of iron and steel, many African countries are investing heavily in iron and steel industries. Nigeria has established a large iron and steel industry at Aja-okuta Steel rolling mills and plants have also been established in Katsina, Oshogbo and Aladja.)

## Compounds of Iron

Iron exhibits oxidation states of +2 and +3. Iron(II) compounds are generally green while iron(III) compounds are orange or reddish-brown in colour. Iron(II) and iron(III) ions are stable. Iron(II) ions are easily oxidised to iron(III) ions.

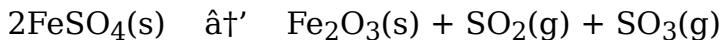


### Experiments 5.7: Investigating the Reactions of Iron(II) Tetraoxosulphate(VI).

- (i) Expose some crystals of iron(II) tetraoxosulphate(VI) heptahydrate to the atmosphere for some hours. Do you observe any change in colour?
- (ii) Put some water in a beaker. Add dilute tetraoxosulphate(VI) acid to it. Later, add iron(II) tetraoxosulphate(VI) crystals. Do the crystals dissolve? Why do you add dilute tetraoxosulphate(VI) acid?
- (iii) Divide the solution from (ii) into five portions.
  - (a) To the first portion, add dilute sodium hydroxide solution.
  - (b) To the second portion, add aqueous ammonia.
  - (c) To the third portion, add a few drops of potassium hexacyanoferate(II).
  - (d) To the fourth portion, add a few drops of ammonium thiocyanate solution.
  - (e) To the fifth portion, add potassium hexacyanoferrate(III) solution. Record your observations in each case.
- (iv) Place some crystals of iron(II) tetraoxosulphate(VI) in an evaporating dish and heat. Identify the gas given off using cobalt(II) chloride paper (the colour is pink when wet, and blue when dry).

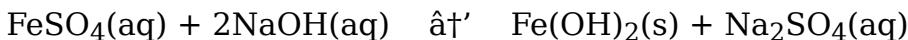
## Properties of iron(II) tetraoxosulphate(VI)

- (a) Iron(II) tetraoxosulphate(VI) crystals decompose in three stages when the heating is gradually increased to a high temperature.

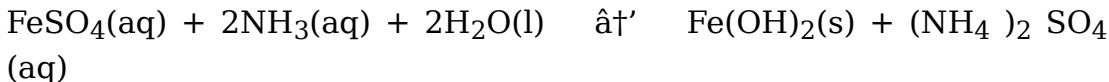


The iron(II) ions are oxidized to iron(III) ions on heating in air.

- (b) A dirty green precipitate of iron(II) hydroxide is formed when sodium hydroxide is added to a solution of iron(II) tetraoxosulphate(VI).



With aqueous ammonia, iron(II) hydroxide is formed as a green gelatinous precipitate.



In the presence of oxygen in air, the green precipitate is oxidised to reddish-brown precipitate of iron(III) hydroxide.

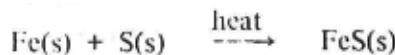


- (c) In preparing a solution of iron(II) tetraoxosulphate(VI), dilute tetraoxosulphate(VI) acid is added to prevent hydrolysis.
- (d) When potassium tetraoxomanganate(VII) solution is added to iron(II) tetraoxosulphate(VI) solution in the presence of dilute tetraoxosulphate(VI) acid, the colour of the  $\text{KMnO}_4$  changes from purple to colourless. The iron(II) tetraoxosulphate(VI) behaves as a reducing agent, while potassium tetraoxomanganate(VII) behaves as an oxidising agent. Iron(II) ion is oxidised to iron(III) ion.
- (e) With potassium hexacyanoferrate(II) solution, a pale blue precipitate is formed when iron(II) tetraoxosulphate(VI) solution is added to it.

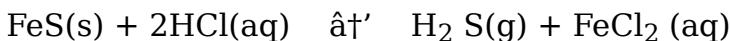
With potassium hexacyanoferrate(III), it gives a blue precipitate. Identification of iron(II) ions in solution is based on the above properties.

## Iron(II) Sulphide, $\text{FeS}$

This is prepared by strongly heating a mixture of iron filings and sulphur in equimolar proportions.



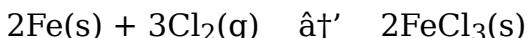
Iron(II) sulphide is a greyish-black insoluble solid which occurs as the mineral-iron pyrite. It reacts with dilute acids to liberate hydrogen sulphide.



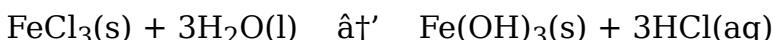
Iron(II) sulphide is used to generate hydrogen sulphide in the laboratory. The mineral, iron pyrite, is an important source of sulphur.

## Iron(III) Chloride, $\text{FeCl}_3$

It is prepared by passing dry chlorine gas over heated iron filings.



The apparatus is similar to that used for preparing iron(II) chloride. Iron(III) chloride is a dark brown solid which is deliquescent. It forms a yellowish solution which is acidic due to hydrolysis.



### Experiment 5.8: Investigating the Properties of Iron(HI) Salts:

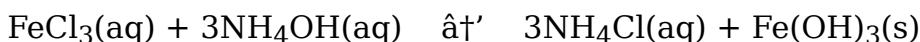
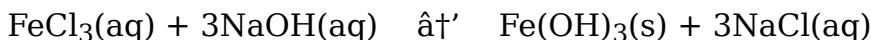
Carry out the following instructions and record your observations in each

case.

- (i) Add aqueous ammonia to iron(III) chloride solution.
- (ii) Add aqueous sodium hydroxide solution to a solution of iron(III) chloride in a test-tube.
- (iii) Add sodium trioxosulphate(IV),  $\text{Na}_2\text{SO}_3$ , to iron(III) chloride solution followed by dilute tetraoxosulphate(VI) acid.
- (iv) Add potassium tetraoxomanganate(VII) to iron(III) chloride solution in the presence of dilute tetraoxosulphate(VI) acid.
- (v) Add drops of potassium hexacyanoferrate(II) solution to a solution of iron(III) chloride in a test-tube.
- (vi) Add drops of potassium hexacyanoferrate(III) to a solution of iron (III) chloride in a test-tube.

### **Properties of Iron(III) Salts**

- (a) Iron(III) chloride solution gives a reddish-brown gelatinous precipitate on adding either sodium hydroxide solution or aqueous ammonia.



- (b) When sodium trioxosulphate(IV) is added to iron(III) tetraoxosulphate(VI) in the presence of dilute tetraoxosulphate(VI) acid, the sodium trioxosulphate(IV) solution reacts with the acid to give off sulphur(IV) oxide which reduces iron(III) ion in iron(III) chloride to iron(II) ion.



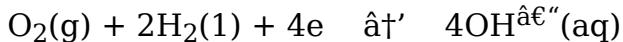
- (c) No visible reaction occurs when iron(III) salt is added to acidified potassium tetraoxomanganate(VII) solution since they are both oxidizing agents.
- (d) Iron(III) chloride gives a deep blue (Prussian blue) precipitate with potassium hexacyanoferrate(II). No precipitate is formed when potassium hexa-cyanoferrate(III) is added to iron(III) salt. Rather, a brown solution is formed.
- (e) A deep red colouration occurs when iron(III) salt solution is added to potassium thiocyanate.

### **Corrosion of Metals**

Metals are known for their metallic lustre (shining appearance). However, when metals are exposed to the atmosphere and they come in contact with moisture and oxygen of the air, they may become less shiny. The surface of the metal may wear out by reacting with the components or impurities in air. The metal is then said to be corroded. In the case of iron, the process is known as rusting.

## Rusting as an Electrolytic Process

Rusting of iron occurs in the presence of oxygen (from air), water and some other compounds such as carbon(IV) oxide, acids or salts which form electrolytic solutions in water. In our humid climate, iron is usually moist. The water making it moist usually contains dissolved oxygen. The concentration of the dissolved oxygen is higher at the circumference of the drop of water than at its centre. Since oxygen readily consumes electrons to form hydroxide ions, the following cathodic reaction occurs at the circumference of the drop of water:



Electrons for this reaction are pulled out from iron atoms at the centre of the drop of water where oxygen concentration is less<sup>TM</sup>



This reaction occurring at the centre of the drop is the anodic reaction. The electrons flow towards the circumference.

In the electrolyte, the  $\text{Fe}^{2+}$  ions move to the cathode while the  $\text{OH}^-$  ions move to the anode. They unite and form iron(II) hydroxide precipitate. This is eventually oxidised to hydrated iron(III) oxide (rust).

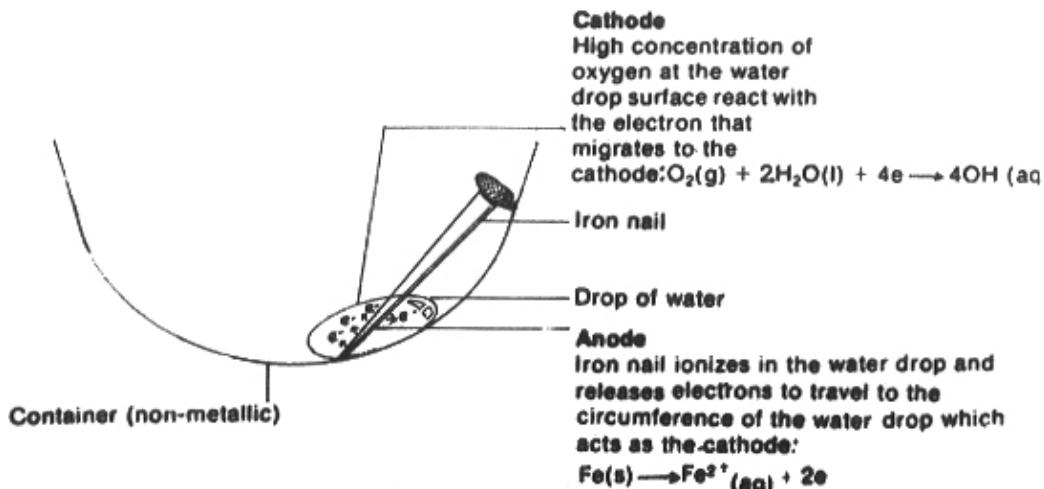
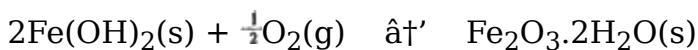


Figure 5.7 Rusting of iron

### Experiment 5.9: Investigating the Corrosion of Iron Nails in Agar Mixture

Boil about 150 cm<sup>3</sup> of water in a beaker. Remove the burner and add 1.5 g of agar powder to it. Stir continuously and heat until the agar is evenly dispersed. Add 5 drops of 0.1 M potassium hexacyanoferrate(III) solution and 3 drops of 0.1% phenolphthalein solution.

Place a nail in a watch glass. Wind round a second nail, a piece of copper wire so that the nail is in contact with the wire. Place it on one side of the watch glass. Wind a strip of zinc metal round the third nail and

place it in the second watch glass.

Cover the nails with the warm agar mixture. Leave the watch glasses overnight. What do you observe with the nails? What do you think would happen if:

- (i) the galvanised iron is scratched?
- (ii) the copper plated iron is scratched?

The result of the experiment shows that any metal, higher than iron in the redox series can protect it from corrosion (rusting). When galvanized iron is scratched, the zinc is preferentially ionized/corroded because zinc is higher in the electrochemical series. Thus, it protects the iron. However, if copper-plated iron is scratched, iron is preferentially ionized/corroded, thus iron is no longer protected.

### **Experiment 5.10: Investigating the conditions for rusting:**

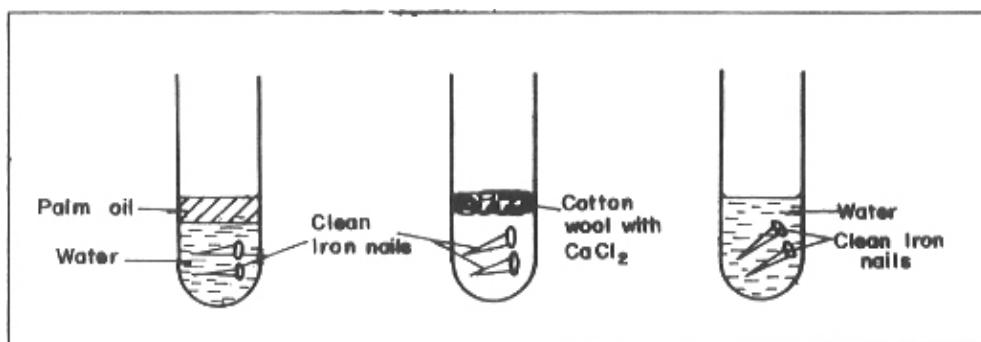


Figure 5.8: Conditions for rusting.

Place three clean iron nails in separate test-tubes. Label the test-tubes 1, 2 and 3. To the first test-tube, add recently boiled water to cover the iron nails. Put palm oil on top of the boiled water immediately it is poured in the test-tube to exclude air. By boiling, air is removed from water. In the second test-tube, put some cotton wool. Add anhydrous calcium chloride in order to absorb water vapour from the air. To the third test-tube, add some water. Leave the experiment for a few days. At the end of this period, observe the iron nails and record your observations. In which test-tube was there no rusting? Why?

### **Prevention of Rusting:**

Rusting can be prevented by oiling/greasing and by painting iron construction works or tools. These prevent water from getting to the surface of the metal. Another method is by covering the surfaces with zinc (galvanizing) or with tin (tin-plating). Tin-plate is used in the canning industry for food preservation. Galvanized iron is not so used. The reason is because when the tin-coated iron is damaged, iron goes into solution in preference to tin. Poisonous tin compounds will therefore not pass into solution. On the other hand, if galvanized iron is damaged, exposing the iron, zinc goes into solution, and its poisonous compounds will thus render the food unwholesome.

For other purposes, galvanization is a better protective method for iron than tin-plating. When the zinc surface of galvanized iron is scratched, exposing iron, zinc goes into solution, providing electrons for the electrolytic reaction that occurs. That is, zinc is "sacrificed" for the iron which it protects. On the other hand, if tin-plated iron is scratched exposing the iron surface, it is iron which goes into solution to provide electrons for the electrolytic reaction. The iron is thus no longer protected.

**Cathodic protection:** The principle of sacrificial protection is employed in cathodic protection of iron from rusting. This is applied in underground pipes, tanks, bridges and even boats. Magnesium anodes are attached at intervals along the pipe or bridge. Being more electropositive than iron, the magnesium anode goes into solution to provide the electrons for the electrolytic reaction instead of iron.

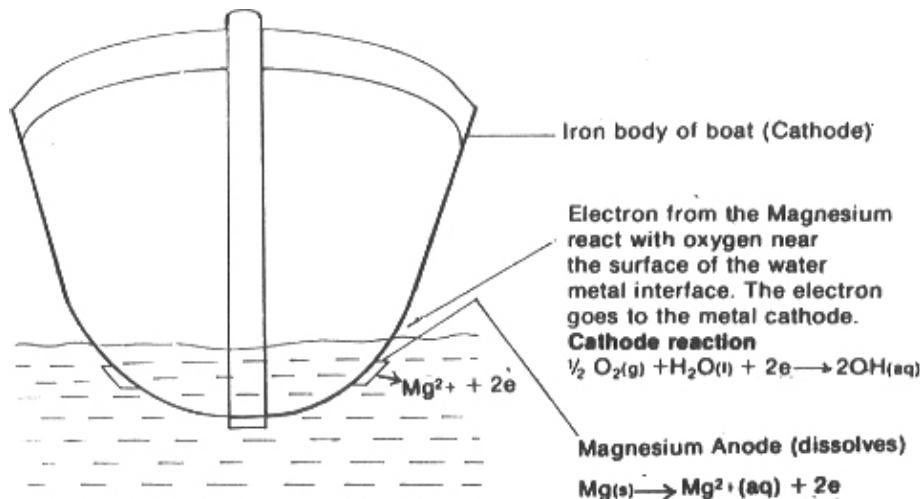
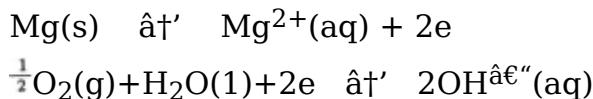


Figure 5.9: Cathodic protection of iron.

#### Experiment 5.11: To determine the fraction of Air used in Rusting

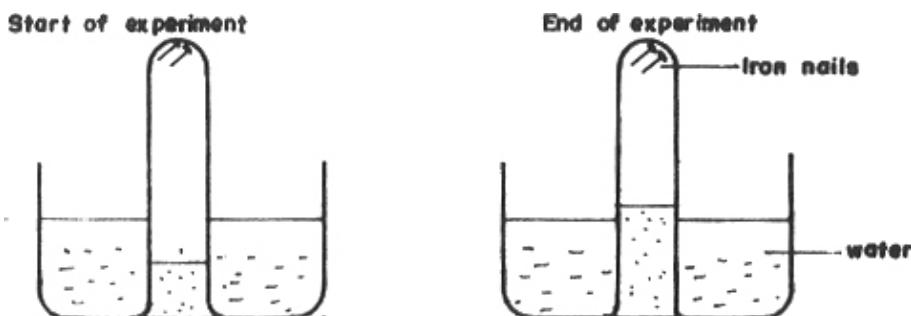


Figure 5.10: Fraction of air used in rusting.

Put some water in a trough. Wet the inside of a test-tube. Sprinkle some iron filings inside the test-tube. Allow the filings to stick to the inside of the test-tube. Invert the test-tube over the water in the bowl. Adjust the

level of water inside. Leave for five days to allow rusting to take place.

As rusting takes place, water rises inside the test-tube up to about  $\frac{1}{3}$  of the original empty space in the test-tube. This shows that the fraction of air used in rusting is about  $\frac{1}{3}$  of air, which corresponds to the fraction of oxygen in air.

From the experiments, you would have observed that rusting took place in the presence of air (oxygen) and water. Salt water also aids rusting.

**Anodic Protection:** Corrosion control may be achieved either by using an inactive substance or with paint. Metals which do not corrode rapidly are often used to coat steel. These include zinc, tin, copper, lead, silver, etc. Zinc is used to coat steel because the zinc becomes coated with either an oxide layer or trioxocarbonate(IV) on exposure to air.

Galvanised iron is steel coated thinly with zinc either by dipping the steel in molten zinc or by electroplating. Tin is very effective for plating. Canning industry makes use of tin-plated steel in containers for food.

Paints and lacquers are often used to protect metals from corrosion. The paints and lacquers protect the metal by mechanically preventing moisture and air from coming into contact with the metal.

To prevent corrosion of aluminium, anodised aluminium is used. Aluminium is given a heavy coat of aluminium oxide which becomes the anode of an electrolytic cell.

### **Economic importance of corrosion:**

A wide variety of domestic, industrial and other appliances are made of iron and other metals which may eventually undergo rusting if not adequately protected by anti-corrosion devices. Examples of such materials include: water tanks, oil tanks, chairs, tables, beds, rods, pipes, spoons, pots, plates, car bodies and parts, bridges, etc. Corrosion may lead to temporary or permanent damage of such appliances. The huge cost of maintenance, repairs and replacement arising from corrosion, underlines the economic importance of corrosion.

## **5.5 Alloys**

Alloys are mixtures of two or more metals combined in molten or solution state to produce a substance with more desirable properties on solidification. Alloys are generally preferred to pure metals for certain purposes.

### **Advantages of Alloys over Pure Metals**

For construction purposes, it is better to use an alloy of aluminium rather than the pure metal. This is because aluminium cannot be machined easily as it sticks to cutting tools. For example, duralumin (alloy of aluminium, copper, magnesium, and manganese) is light and has high tensile strength, but does not stick to cutting machines.

Aluminium bronze (an alloy of aluminium and copper) has resistance to chemical corrosion and mechanical wear. Hence, it is preferred to aluminium in the manufacture of engine parts.

Steel, an alloy of iron, carbon, and traces of sulphur, nickel, tungsten, vanadium, copper, silicon or phosphorus, possesses a much higher tensile strength than iron. Steel is used when high resistance to reagents, coupled with toughness, **is** required.

An alloy of two or more metals has a wide range of properties: lower melting point than the original metals, resistance to corrosion, higher tensile strength than the original metals, and less malleable and ductile. By alloying, the fundamental properties of metals are improved.

**TABLE 5.4: Alloys**

Alloys	%Composition	Properties	Uses
Stainless Steel	Cr 10-20; Ni 8-20; Fe 80-60	Does not rust or corrode.	In chemical and food processing industry.
Armour plate	Ni 1-4; Cr 0.5-2; Fe 95-98	Very hard. Has high tensile strength.	Used for manufacturing armour plates and wires.
Tool Steel	Cr 2-4; Mo 6-7; Fe 90-96.	Very hard.	Used for making chisels, hammers etc.
Aluminium brass	Cu-90; Al-10.	Resists chemical wear.	Used to make castings.
Coinage brass	Cu-95; Sn-4, Zn-1.	Resists chemical wear.	Coins and medals.
Common brass	Cu-66.6; Zn-33.4.		
Dentist amalgam	Hg-70; Cu-30.		Dental filling.
Coinage Gold	Au-90; Cu-10.	Does not corrode	Coinage
Dental Gold	Au-58; Cu-14.28; Ag-28.14	Does not corrode	Dental cap.
Lead battery plates	Pb-94; Sb-6.		Storage batteries
Type metal	Pb-58; Sn-26; Sb-15; Cu-1.		Printing industry.
Solder lead	Pb-67; Sn-33.		For joining metals
Mangahum	Al 70-90; Mg 30-10		Castings and scientific apparatus
Manganin	Cu; Mn; Ni	Has constant resistance.	Electrical apparatus
Silver coinage	Ag-90; Cu-10.	Does not corrode easily.	Coinage.
Duralumin	Al, Mg, Cu, Mn.	High tensile strength and, light weight.	
Pewter	Contains lead.		
German Silver	Cu; Zn, Ni.		
Copper Coinage	Cu, Sn.		

## 5.6 Summary of Chemistry of Metals

**TABLE 5.5 : Metallic ores and methods of extraction**

Metal	Reactivity	Most Common ore	Method of extraction
K Na		Chlorides	Electrolysis of fused $\text{Cl}^-$
Ca	Very reactive	Chlorides	
Mg		Trioxocarbonates(IV)	Electrolysis of fused $\text{Cl}^-$
Al		Oxides	Electrolysis of $\text{O}^{2-}$
Zn			
Fe	Moderately reactive	Oxides	Electrolysis of oxide
Sn		Trioxocarbonates(IV)	Roasting of $\text{CO}_3^{2-}$ and $\text{S}^{2-}$
Pb		Oxides	to oxide and reduction by carbon or CO.
Cu			
Hg		Sulphides	Roasting in air.
Ag	Least reactive		Heating in air.
Au		Free element	Used as free elements

**TABLE 5.6: Properties of metals according to position in the electrochemical series**

Metal	Reaction with oxygen	Reaction with water	Reaction with acids	Reduction of oxides by hydrogen	Action of heat on oxides	Reaction of water with oxides	Character of hydroxide
K Na		Decompose cold water.	Attacked by dilute acids.	Not reduced.		Oxides from hydroxides.	Soluble in water to form alkaline solution.
Ca	Burn in air or oxygen.			Not reduced.	Stable when heated.		
Mg		Decompose steam at red heat.	(Al is only attacked by dilute HCl.)			Oxides do not react with water.	Insoluble hydroxides.
Al							
Zn							
Fe							
Pb	Oxidized when heated in air.	Do not decompose water or steam at red heat.	Attacked by oxidizing acids	Reduced			
Cu							
Hg							
Ag	Unaffected		Not attacked.	Reduced	Decompose.		Hydroxides not formed.

**TABLE 5.7: Properties of compounds of metal**

Metal	Character of trioxocarbonate(IV)	Action of heat on trioxonitrate(V)	Solubility of chlorides	Solubility of sulphides
K Na Ca	Soluble and not decomposed by heat.	Trixonitrate(V) salts decompose to dioxonitrate (III)	Soluble	Soluble in water.
Mg Al Zn Fe	Insoluble and decomposed by heat. $\text{Al}_2(\text{CO}_3)_3$ does not exist	Trixonitrate(V) decompose to oxide.		Insoluble in water but soluble in dilute HCl.
Pb Cu Hg			Soluble in hot water. Soluble	Insoluble in water and dilute HCl.
Ag Au	Unstable	Trixonitrate(V) decompose to metal.	Insoluble	

TABLE 5.8: Physical Properties of metals

	K	Mg	Ca	Al	Zn	Fe
State	Solid	Solid	Solid	Solid	Solid	Solid
Appearance	White metallic lustre.	Silvery white, tarnishes in air.	Silvery, tarnishes in air.	Silvery white.	Bluish white.	Pure iron is white
Density ( $\text{g cm}^{-3}$ )	0.86	0.97	1.55	2.69	7.1	7.9
Malleability	Malleable and ductile	Malleable and ductile	Malleable and ductile	Can be rolled into foil	Malleable	Extremely malleable
Tensile Strength	Poor	Poor	Fair	Fair	High	High
Melting point	$63^{\circ}\text{C}$	$98^{\circ}\text{C}$	$850^{\circ}\text{C}$	$660^{\circ}\text{C}$	$419^{\circ}\text{C}$	$1535^{\circ}\text{C}$
Conduction of heat and electricity.	Good	Good	Good	Good	Good	Good. Ferromagnetic

Pb	Cu
Solid	Solid
Bluish white	Reddish brown.
11.3	8.95
Very malleable and ductile.	Very malleable and ductile.
Fair	Fairly high.
$327^{\circ}\text{A}$	$1080^{\circ}\text{C}$
Good	Very Good.

TABLE 5.9: Chemical Properties of metals

Metal	Reaction with Air or Oxygen	Reaction with Water	Flame colouration
Sodium	Attacked by oxygen to form sodium oxide. Oxide dissolves in moisture to form the hydroxide.  CO <sub>2</sub> from air combines, with sodium hydroxide to form sodium trioxocarbonate(IV) decahydrate.	Darts about and ignites on the water, melting to form sodium hydroxide, and hydrogen is given off	Persistent golden yellow.
Potassium	Potassium behaves like sodium. Sodium and potassium are kept under paraffin oil in the laboratory.	Potassium behaves like sodium.	Lillac-blue
Calcium	Not as reactive as sodium and potassium. Reacts to form calcium oxide.	Calcium reacts with water to form a suspension of calcium hydroxide. Hydrogen is given off.	Brick red.

## Aluminium:

A thin coat of aluminium oxide is formed on exposure to air. This prevents further attack of air. At 800°C, it burns in air to form aluminium oxide and nitride.

Aluminium is attacked by dilute hydrochloric acid and concentrated hydrochloric acid, with evolution of hydrogen. It is attacked by hot concentrated H<sub>2</sub>SO<sub>4</sub> with the evolution of sulphur(IV) oxide.

Dilute or concentrated trioxonitrate(V) acid does not react with aluminium. This is due to the fact that a thin layer of oxide is formed which prevents further attack by the acid.

Metal Reaction with Air or Oxygen  
colouration Sodium Attacked by oxygen to form sodium oxide. Oxide dissolves in moisture to form the hydroxide.

CO<sub>2</sub> from air combines, with sodium hydroxide to form sodium trioxocarbonate(IV) decahydrate. Darts about and ignites on the water, melting to form sodium hydroxide, and hydrogen is given off.

Persistent golden yellow. Potassium Potassium behaves like sodium. Sodium and potassium are kept under paraffin oil in the laboratory. Potassium behaves like sodium. Lillac-blue Calcium Not as reactive as sodium and potassium. Reacts to form calcium oxide. Calcium reacts with water to form a suspension of calcium hydroxide. Hydrogen is given off. Brick red.

**TABLE 5.10: Reactions of some common ions with reagents**

Metallic ion	Reaction with NaOH.	Reaction with aqueous ammonia	Reaction with potassium hexacyanoferrate(II) solution, $K_4Fe(CN)_6$ .	Reaction with potassium hexacyanoferrate(III) solution, $K_3Fe(CN)_6$ .
Iron(II) ion	Green gelatinous precipitate.	Green gelatinous precipitate.	Pale blue precipitate.	Dark blue precipitate.
Iron(III) ion	Reddish-brown gelatinous precipitate.	Reddish-brown gelatinous precipitate.	Dark blue precipitate.	Brown solution, no precipitate.
Copper(II) ion	Blue precipitate insoluble in excess.	Blue precipitate, soluble in excess, forming deep blue solution.	—	—
Aluminium ion	White gelatinous precipitate, soluble in excess.	White gelatinous precipitate, insoluble in excess.	—	—
Zinc ion	White gelatinous precipitate, soluble in excess.	White precipitate soluble in excess.	—	—
Calcium ion	White precipitate, insoluble in excess.	—	—	—

## 1st Row Transition Metals

These form coloured ions in solution. They have variable oxidation states, high melting points, densities and boiling points. The metallic ions are all hydrated in solution.

Copper occurs as copper pyrites. It exhibits oxidation states of +1 and +2. Most of its salts are soluble in water. Copper(II) chloride is soluble, so are  $CuSO_4$  and  $Cu(NO_3)_2$ .  $CuCO_3$  is insoluble.

Iron occurs as iron pyrites. It exhibits oxidation states of +2 and +3. Most of its salts are soluble in water. When exposed to air (oxygen) and water, it rusts, forming a brownish precipitate with the formula  $Fe_2O_3 \cdot xH_2O$ . Iron has high melting and boiling points.

Zinc occurs as zinc blend,  $ZnS$ . Zinc is extracted by first roasting to convert to oxide and then reducing the oxide using coke.

Metals corrode in air. Corrosion is a general term applied to the process in which metals change to compounds on exposure to air. In the case of iron, it is referred to as rusting. Economically, rusting or corrosion is a serious problem. A lot of waste occurs as a result of corrosion or rusting.

An alloy is a mixture of two or more metals. Metals are often alloyed with other metals or non-metals to improve on their physical properties.

## ASSESSMENT

- Starting with iron filings, describe how you would prepare the following compounds in the laboratory:

- (i) iron(III) chloride crystals,  
 (ii) iron(II) tetraoxosulphate(VI) crystals;  
 (iii) anhydrous iron(II) chloride.
2. Name three ores of iron. One of them should be commonly found in Nigeria. Indicate it.  
 Describe in detail the extraction of iron from the ore that is commonly found in Nigeria, using the blast furnace. State three uses of iron.
3. (a) Consider the following reactions:
- $$\text{Cu}^{2+}(\text{aq}) + \text{Zn}(\text{s}) \rightarrow \text{Cu}(\text{s}) + \text{Zn}^{2+}(\text{aq})$$
- $$\text{Zn}^{2+}(\text{aq}) + \text{X}(\text{s}) \rightarrow \text{Zn}(\text{s}) + \text{X}^{2+}(\text{aq})$$
- (i) What will be the reaction between X and a dilute acid?  
 (ii) Can X reduce  $\text{Cu}^{2+}$  ions? Explain.
- (b) In which of these two is iron better protected from corrosion: galvanized iron or copper-plated iron? Explain.
- (c) Iron is manufactured in a blast furnace using iron ore, coke and limestone. Write the equations for the reactions at the:
- (i) top of the furnace.  
 (ii) middle of the furnace  
 (iii) bottom of the furnace.
- (WAEC)
4. (a) Explain why transition metals:  
 (i) have high melting points,  
 (ii) have variable oxidation states,  
 (iii) exhibit paramagnetism.
- (b) (i) Describe the electrolysis of copper(II) tetraoxosulphate(VI) solution, using copper electrodes.  
 (ii) Will the colour of the copper(II) tetraoxosulphate(VI) solution change at the end of the electrolysis described in (b)(i) above?  
 Give reasons for your answer.
- (WAEC).
5. (a) Describe what happens when copper(II) is electrolysed with carbon electrodes.  
 (b) How would you purify an impure copper rod?  
 (c) During the electrolysis of a copper(II) salt, the quantity of electricity passed was 9650 coulombs. Calculate, (i) the number of mole of copper metal deposited (ii) the mass of copper deposited.  
 $1 \text{ Faraday} = 96500 \text{ coulombs}; \text{Cu} = 64$
6. (a) Consider the following reactions.
- $$\text{Cu}^{2+}(\text{aq}) + \text{Zn}(\text{s}) \rightarrow \text{Cu}(\text{s}) + \text{Zn}^{2+}(\text{aq})$$
- $$\text{Zn}^{2+}(\text{aq}) + \text{X}(\text{s}) \rightarrow \text{Zn}(\text{s}) + \text{X}^{2+}(\text{aq})$$
- (i) What will be the reaction between X and a dilute acid? (ii) Can X reduce  $\text{Cu}^{2+}$  ions? Explain.

- (b) In which of these two is iron better protected from corrosion: galvanised iron, or copper-plated iron?
- (c) Iron is manufactured in a blast furnace using iron ore, coke and limestone. Write the equations for the reaction at the:
- top of the furnace,
  - middle of the furnace, and
  - bottom of the furnace.

(WAEC)