9 Metals

Content

Learning outcomes

Candidates should be able to:

9.1 Properties of metals

(a) describe the general physical properties of metals (as solids having high melting and boiling points; malleable; good conductors of heat and electricity) in terms of their structure

(b) describe alloys as a mixture of a metal with another element, e.g. brass; stainless steel

(c) identify representations of metals and alloys from diagrams of structures

(d) explain why alloys have different physical properties from their constituent elements

9.2 Reactivity series

(a) place in order of reactivity: aluminium (see also 9.5(b)) calcium, copper, (hydrogen), iron, lead, magnesium, potassium, silver, sodium and zinc by reference to

(i) the reactions, if any, of the metals with water, oxygen, steam and dilute hydrochloric acid,

(ii) the reduction, if any, of their oxides by carbon and/or by hydrogen

(b) describe the reactivity series as related to the tendency of a metal to form its positive ion, illustrated by its reaction with

(i) the aqueous ions of the other listed metals

(ii) the oxides of the other listed metals

(c) deduce the order of reactivity from a given set of experimental results

(d) describe the action of heat on the carbonates of the listed metals and relate thermal stability to the reactivity series

9.3 Extraction of metals

(a) describe the ease of obtaining metals from their ores by relating the elements to their positions in the reactivity series

(b) describe metal ores as a finite resource and hence the need to recycle metals

(c) discuss the social, economic and environmental advantages and disadvantages of recycling metals, e.g. aluminium and copper

9.4 Iron

(a) describe and explain the essential reactions in the extraction of iron using haematite, limestone and coke in the blast furnace

(b) describe steels as alloys which are a mixture of iron with carbon or other metals and how controlled use of these additives changes the properties of the iron, e.g. high carbon steels are strong but brittle whereas low carbon steels are softer and more easily shaped

(c) state the uses of mild steel (e.g. car bodies; machinery) and stainless steel (e.g. chemical plant; cutlery; surgical instruments)

(d) describe the essential conditions for the corrosion (rusting) of iron as the presence of oxygen and water; prevention of rusting can be achieved by placing a barrier around the metal (e.g. painting; greasing; plastic coating; galvanising)

(e) describe the sacrificial protection of iron by a more reactive metal in terms of the reactivity series where the more reactive metal corrodes preferentially (e.g. underwater pipes have a piece of magnesium attached to them)

9.5 Aluminium

(a) outline the manufacture of aluminium from pure aluminium oxide dissolved in cryolite (starting materials and essential conditions, including identity of electrodes should be given together with equations for the electrode reactions but no technical details or diagrams are required)

(b) explain the apparent lack of reactivity of aluminium

(c) state the uses of aluminium and relate the uses to the properties of this metal and its alloys, e.g. the manufacture of aircraft; food containers; electrical cables

Tuesday 6th October, 2011

Metals and their reactivity

Reactivity Orders

Cold Water < Steam < dilute Acid

Metals	Cold Water	Steam	Dilute Acid
W	✓	✓	✓
X	X	✓	✓
Y	X	X	✓
Z	X	X	X

◎ W > X > Y > Z ◎

→ W = K, Na, Ca, Mg

→ X = Al, Zn, Fe

→ Y = Pb

→ Z = Cu, Hg, Ag, Au, Pt

Stability of metal compounds

K (more reactive) less stable K^+ (less reactive) more stable. K_2^+O (less reactive) more stable

Ag (less reactive) more stable Ag^+ (more reactive) less stable. Ag_2^+O (more reactive) less stable

$Ag^+ + \bar{e} \longrightarrow Ag$

Reactive — Gain of electrons — Less reactive

X^+ & Ag^+ can not exist individually. It means they are present in compounds.

Reactivity Order

Cold water < Steam < dilute Acid

Metals	Cold Water	Steam	Dilute Acid
K	React to form Metal Hydroxide + H_2	React to form Metal Oxide + H_2	React to form salt + H_2
Na	$K + H_2O_{(g)} \rightarrow KOH + H_{2(g)}$	$K + H_2O_{(g)} \rightarrow K_2O + H_2$	NOTE: K, Na, Ca react violently with acids, their reaction must be avoided
Ca	$Mg + H_2O_{(g)} \rightarrow Mg(OH)_2 + H_{2(g)}$	$Mg + H_2O_{(g)} \rightarrow MgO + H_2$	$Mg + HCl \rightarrow MgCl_2 + H_2$
Mg		$Al + H_2O_{(g)} \rightarrow Al_2O_3 + H_2$	$Al + HCl \rightarrow AlCl_3 + H_2$
Al		$Zn + H_2O_{(g)} \xrightarrow{Heat} ZnO + H_2$	$Fe + HCl \rightarrow FeCl_2 + H_2$
Zn			$Pb + HNO_3 \rightarrow Pb(NO_3)_2 + H_2$
Fe			
Pb			
H			
Cu			
Hg			
Ag			
Au			
Pt			

Don't react with water, steam and dilute acids as they are less reactive than hydrogen & are unable to displace it.

⇒ Water mean cold water in past papers.

⇒ Iron reacts with acids to form salts in which its valency is +2.

SLM Chlorides

⇒ A protective layer of $PbCl_2$ will protect Pb and prevent it from reacting with HCl ($Pb(NO_3)_2$ is soluble so the reaction of Pb with HNO_3 will continue.

HCl is in excess!

STABILITY (03) OF METAL COMPOUNDS

⇒ X, Y and Z are groups of unknown Metals

$XO \xrightarrow{Heat} X + O_2$ (Reduction - removal of O_2)

$YO \xrightarrow{Heat}$ No reaction

Reducing agent → $YO + C \xrightarrow{Heat} Y + CO$

$ZO \xrightarrow{Heat}$ No reaction

$ZO + C \xrightarrow{Heat}$ No reaction

$Z^{+2}O^{-2} \xrightarrow[\text{Electricity/ Electrolysis}]{Heat} Z + O_2$

Anode

$2O^{-2} \longrightarrow O_2 + 4e^-$

Cathode

$Z^{+2} + 2e^- \longrightarrow Z$

Reactivity Order

$XO > YO > ZO$

$X < Y < Z$

⇒ ZO is the most stable oxide because it is decomposed by the strongest method! Simple heating & reducing agent cannot decompose ZO.

Stability Order

$XO < YO < ZO$

$X > Y > Z$

In some of the past papers:-

$PbO + H_2 \xrightarrow{Heat} Pb + H_2O$

⇒ "C is b/w Zn & Al"

Reactivity is very close. A lot of heat can break PbO to form Pb & H_2O! But examiner will tell when reaction takes place!

→ Reactivity $\propto \frac{1}{\text{Stability}}$

→ More reactive metals form less reactive ions and vice versa

→ More reactive metals form less reactive and more stable compounds and vice versa. (More Stable).

Stability of Metal Compounds:-

Metals	Oxides	Hydroxides	Carbonates
K Na	Can only be reduced by electrolysis.	Can't be decomposed by heating	Can't be decomposed by heating.
Ca Mg Al	$Al_2O_3 \xrightarrow[\text{Electricity}]{\text{Heat}} Al + O_2$ (Cathode, Anode)	Can be decomposed to give metal oxide + Steam	Can be decomposed by heating to give metal oxide + CO_2.
Zn Fe Pb H Cu	Can be reduced by heating in the presence of a reducing agent e.g. Carbon (C), whereas oxides of copper can also be reduced by hydrogen $ZnO + C \xrightarrow{\text{Heat}} Zn + CO$. $CuO + H_2 \xrightarrow{\text{Heat}} Cu + H_2O(g)$. $CuO + C \xrightarrow{\text{Heat}} Cu + CO$	$Ca(OH)_2 \xrightarrow{\text{Heat}} CaO + H_2O$ $Cu(OH)_2 \xrightarrow{\text{Heat}} CuO + H_2O$	$CaCO_3 \xrightarrow{\text{Heat}} CaO + CO_2$ $CuCO_3 \xrightarrow{\text{Heat}} CuO + CO_2$
Hg Ag Au Pt	Can be reduced simple heating $Ag_2O \xrightarrow{\text{Heat}} Ag + O_2$	Unstable, don't exist!	Unstable, don't exist!

$PbO + C \xrightarrow{\text{Heat}} Pb + CO$
$Fe_2O_3 + C$

Displacement Reactions of Metals

(a) Write simple equation where possible. (01)

(B)

$Zn_{(s)} + CuSO_{4(aq)} \longrightarrow ZnSO_{4(aq)} + Cu_{(s)}$

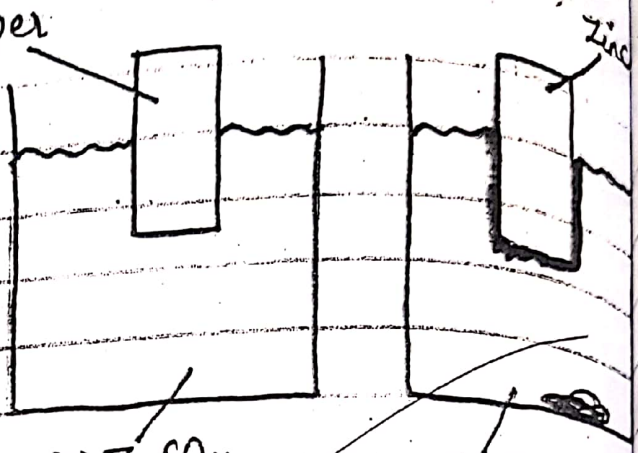
Blue → Colourless

(b) Write the ionic equation for this equation (02).

$Zn_{(s)} + Cu^{+2}_{(aq)} \longrightarrow Zn^{+2}_{(aq)} + Cu_{(s)}$

Pink

SO_4^{-2} is not undergoing a chemical change. It is a spectator ion so it is cancelled out!



No observation ☆ ⇒ Zinc is not a transition me as it does not have variable oxidation state does not form coloured compo is not used as a catalyst in any reaction

It's compound is white as a solid & colourless in solutio

(c) Write observations if any. (03)

1. Blue solution will fade away (turn colourless)
2. Pink deposit at the bottom of the beaker or over zinc plate
3. Zinc plate decreases in size and mass

← 100% copper do not deposit ove the plate

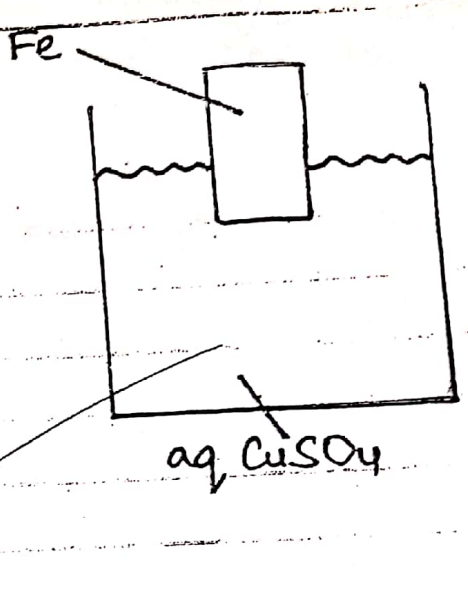
We remove copp to see the new si of the plate

(1) Equation for the Reaction

$Fe + CuSO_4 \longrightarrow FeSO_4 + Cu$

(s) (aq) (aq) (s)

(Blue) (Green)



Ionic Equations

$Fe + Cu^{+2} \longrightarrow Fe^{+2} + Cu$

(s) (aq) (aq) (s)

Observations

① → Iron plate decreases in size and mass.

② → Blue colour of the solution turns green.

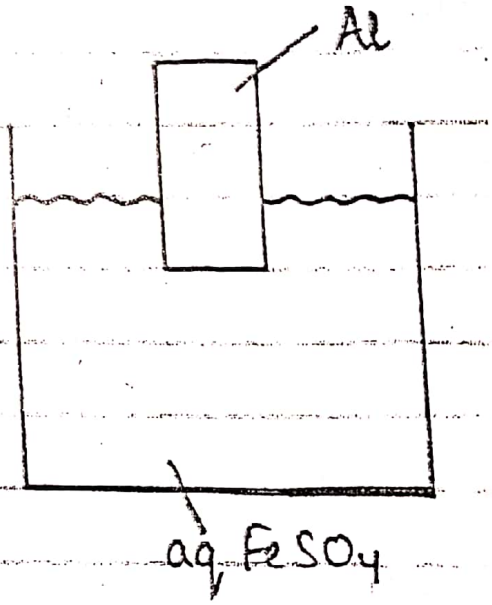
③ → Pink deposit is seen at the bottom of the beaker or over the iron plate

④ → The container becomes warm.

(2) Equation for the Reaction

$2Al + 3FeSO_4 \longrightarrow Al_2(SO_4)_3 + 3Fe$

(s) (aq) (aq) (s)



Ionic Equation

$2Al + 3Fe^{+2} \longrightarrow 2Al^{+3} + 3Fe$

(s) (aq) (aq) (s)

Observations

→ The green colour of the solution fades away.

→ The Aluminium plate decreases in size and mass.

→ The container becomes warm.

→ Precipitates of Fe formed at the bottom of the beaker.

(08)

→ $Cu_{(s)} + HCl_{(aq)} \longrightarrow$ No reaction!

→ $Zn_{(s)} + CuSO_{4\,(aq)} \longrightarrow ZnSO_{4\,(aq)} + Cu_{(s)}$

→ $Cu_{(s)} + ZnSO_{4\,(aq)} \longrightarrow$ No reaction!

→ $Al_{(s)} + CuSO_{4\,(aq)} \longrightarrow Al_2(SO_4)_{3\,(aq)} + Cu_{(s)}$

→ $Mg_{(s)} + 2H_2O_{(l)} \longrightarrow Mg(OH)_{2\,(aq)} + H_{2\,(g)}$

→ $Mg_{(s)} + 2HCl_{(aq)} \longrightarrow MgCl_{2\,(aq)} + H_{2\,(g)}$

→ $Fe_{(s)} + 2HCl_{(aq)} \longrightarrow FeCl_{2\,(aq)} + H_{2\,(g)}$

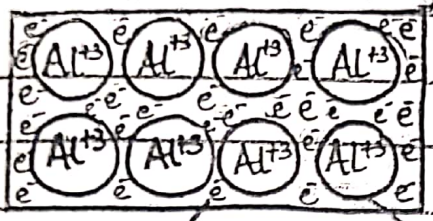
⇒ When iron reacts with an acid, it forms Fe^{+2} ions meaning its oxidation number is '+2' not '+3'!

Friday 13th October, 2017

Metals

Metals and their properties

Metallic Bonding



→ Minimum two rows & each row must have minimum 4 positive ions

Metallic Bonding

In metals, positive ions are surrounded by a sea of delocalised electrons.

Q) Explain in terms of metallic bonding, why Aluminium is a good conductor of electricity? (02)

Ans In Aluminium positive ions are surrounded by a sea of electron. Aluminium conducts electricity due to the presence of freely moving (mobile/delocalised) electrons.

Note :-

Metallic Bonding is stronger than ionic bonding which is stronger than covalent bonding which is stronger than Van der Waals forces.

Properties of Metals:

Practice questions

Complete the following equations. You can write "no reaction" where reaction is not possible

(1) $Na + H_2O_{(l)} \rightarrow$

(2) $K + H_2O_{(g)} \rightarrow$

(3) $Mg + H_2O_{(g)} \rightarrow$

(4) $Mg + H_2O_{(l)} \rightarrow$

(5) $Pb + H_2O_{(g)} \rightarrow$

(6) $Al + H_2O_{(g)} \rightarrow$

(7) $Fe + HCl_{(aq)} \rightarrow$

(8) $Cu + HCl_{(aq)} \rightarrow$

(9) $Na_2O \xrightarrow{Heat}$

(10) $K_2O \xrightarrow{Heat/Electricity}$

(11) $Al_2O_3 + C \xrightarrow{Heat}$

(12) $Al_2O_3 + Heat \xrightarrow{Electricity}$

(13) $Fe_2O_3 + H_2 \xrightarrow{Heat}$

(14) $Fe_2O_3 + C \xrightarrow{Heat}$

(15) $CuO + H_2 \xrightarrow{Heat}$

(16) $Na_2CO_3 \xrightarrow{Heat}$

(17) $K_2CO_3 \xrightarrow{Heat}$

(18) $CaCO_3 \xrightarrow{"}$

(19) $CuCO_3 \xrightarrow{"}$

(20) $NaOH \xrightarrow{"}$

Colour of metals and their compounds

11th October, 2021

Lab Activity

To observe:

Colours of metals & their compounds and their displacement reactions.

Metals & their appearance

* Copper Metal : Pinkish Brown
  CuO : Black

* Magnesium Metal : Grey
  MgO : White

* Zinc Metal : Grey
  ZnO : White
  $ZnSO_4$: white

* Aluminium Metal : Silvery white
  $Al(NO_3)_3$: White

* Iron Metal : Silvery white
  Fe_2O_3 : Brown

NOTE:
Iron seemed brown because of the rust formed over it. From some parts it seemed grey, no silvery white because it was not fresh iron. Rubbing iron by sand paper looking at freshly cut iron, will observe a silvery white colour of iron.

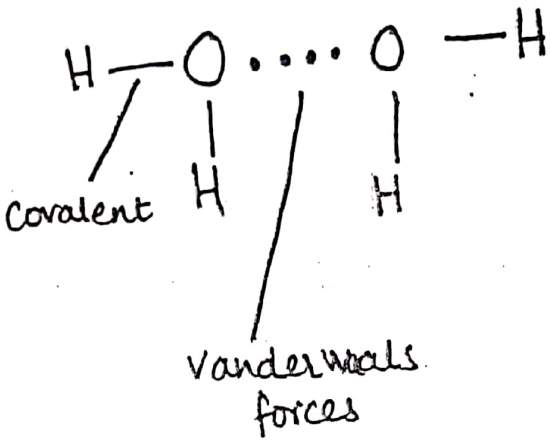
Reactions:

$$Al_{(s)} + 6HCl_{(aq)} \longrightarrow 2AlCl_{3(aq)} + 3H_{2(g)}$$

$$Zn_{(s)} + 2HCl_{(aq)} \longrightarrow ZnCl_{2(aq)} + H_{2(g)}$$

⇒ According to the reactivity series, Aluminium is more reactive than zinc but when both are made to react with HCl, zinc will appear to react faster. This is due to the layer of Al_2O_3 (insoluble) over Aluminium that prevents it from reacting at the start.

⟹ In metals, positive ions do not move, they vibrate only; they do not conduct electricity or heat. The electrons are responsible for conduction of heat & electricity.



⟹ Metals have high melting points and densities due to strong electrostatic forces b/w opposite charges, which need high energy to break.

Stronger bonding! ↑

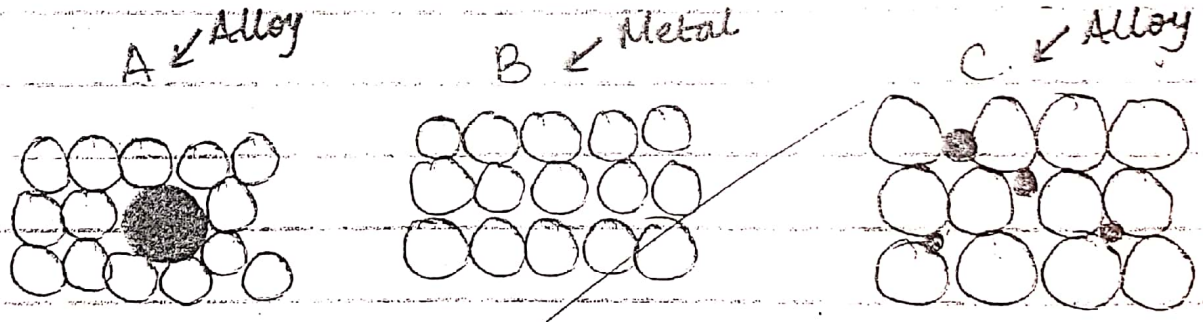
Metallic Bonding
Ionic Bonding
Covalent Bonding
Van der Waals Force

(2) Metals have high density
(3) Metals are good conductors of electricity and heat due to the presence of freely moving electrons.
(4) Metals are malleable and ductile as layers of atoms slide over each other when force is applied. (In ionic compounds, same charges repel each other).
(5) They are magnetic in nature (only transition metals).
(6) Metals are sonorous (Make a ringing sound when struck).

Alloys

Mixture of two or more metals are called alloys e.g; brass, bronze, stainless steel

Brass = copper + zinc
Bronze = copper + tin
Stainless steel = Iron + chromium + nicked



☆ ⇒ A and C are alloys

We removed O_2 combined with Fe and SiO_2

Overall eq: $CaCO_3 + SiO_2 \rightarrow CaSiO_3 + CO_2$

Fe's oxides can be reduced by reducing agent not by simple heating

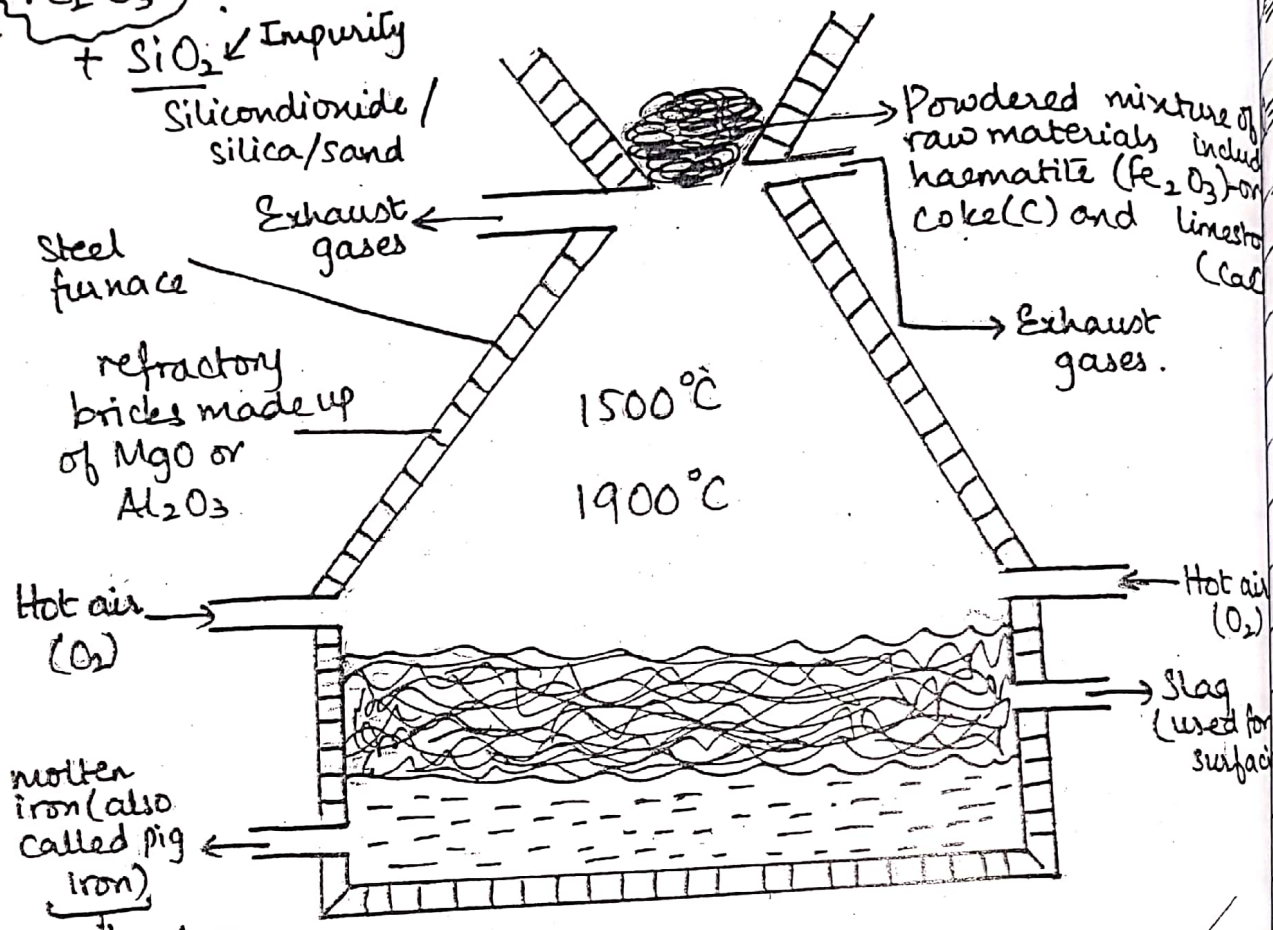
Blast Furnace!

ORE:-

Haematite

Fe_2O_3 + SiO_2 ← Impurity

Silicondioxide / silica/sand



not pure form 96% Fe, 4% C

To reduce Fe_2O_3 by removal of oxygen, to extract Fe

Reduction of the ore:

1. $C + O_2 \longrightarrow CO_2 \quad -\Delta H$ (produces a very high temp) ← coke
2. $CO_2 + C \longrightarrow CO \quad +\Delta H$ (CO is a very powerful reduc a
3. $Fe_2O_3 + CO \longrightarrow Fe_{(l)} + CO_2$
4. $Fe_2O_3 + C \longrightarrow Fe_{(l)} + CO$

If pig iron a shape, it cast iron!

5. $CaCO_3 \xrightarrow{Heat} CaO + CO_2$
6. $CaO + SiO_2 \longrightarrow CaSiO_3$ (slag - calcium silicate) ← salt

(Basic Oxide) (Acidic Oxide)

Nuetralisation Reaction!

Q⇒ Why metals are converted into alloys?

Ans To increase the strength of metals.

Q⇒ Why alloys are stronger than pure metals?

Ans In alloys, there are different sized atoms, which disrupt metallic lattice (regular arrangement) which does not allow layers of atoms to slide past each other.

Ore

An ore is a naturally occuring compound of a metal along with impurities.

Hematite = Fe_2O_3

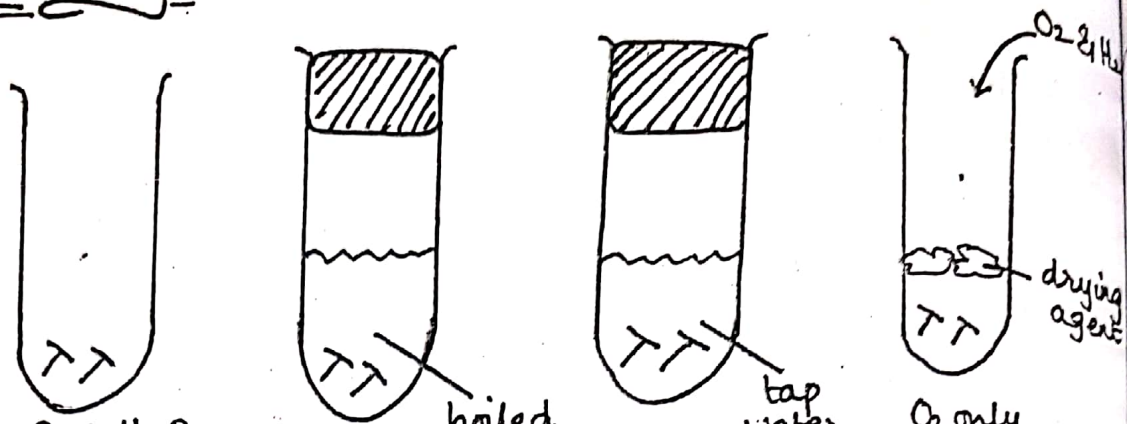
↑ Ore of iron that is in our syllabus!

Types of Iron and their properties

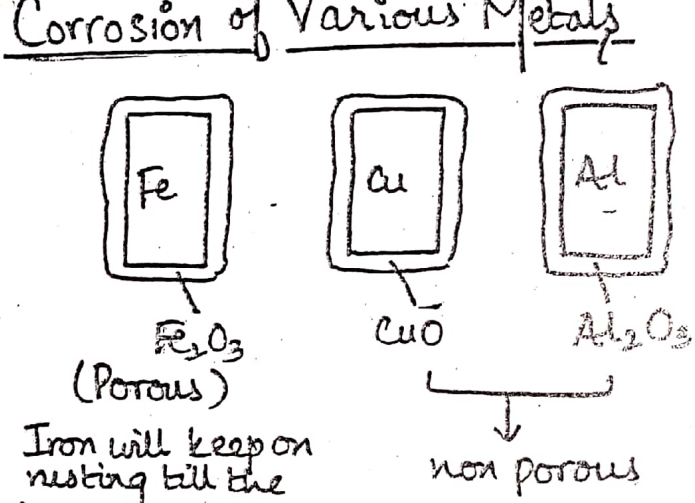
↑ carbon content makes Fe strong but brittle. It no longer remains malleable.

	Type	Composition	Properties	Uses
① →	Pig or Cast Iron	96% Fe, 4% C	Strong, very brittle	Engine Blocks
② →	Wrought Iron	100% Fe	Very malleable	Gates & Chains
③ →	Carbon Steel:			
i.	Mild Steel	0.2 to 0.5% C.	Soft and malleable	Car bodies & Machines
ii.	High Carbon Steel	0.5 to 1.5% C.	Strong & Brittle	Drills, Nails & Hammers

Rusting



Corrosion of Various Metals

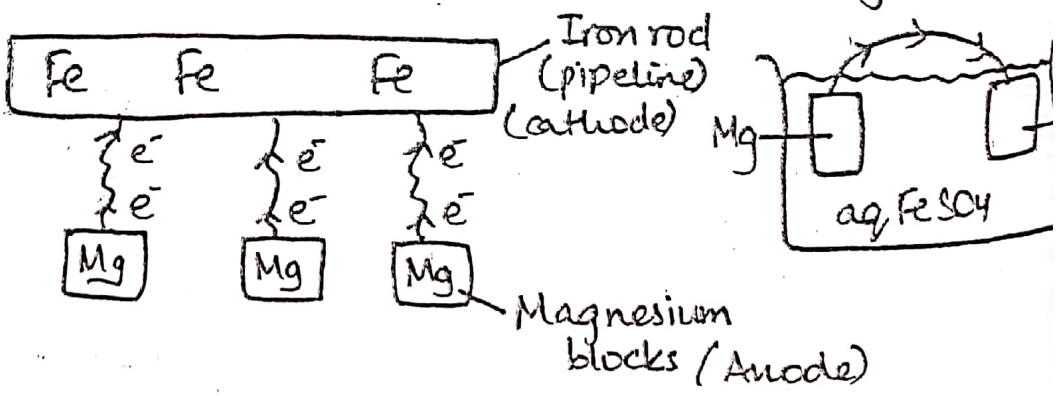


☆⇒ Only Oxygen is required for the corrosion of all metals

Iron is exception because it needs both oxygen & moisture.

Sacrificial Protection

Electrons travel from m reactive metal to the le reactive metal, thus pre it from rusting.



Stainless Steel	Fe, Cr, Ni	Resistant to corrosion	Cutlery, Surgical Instruments & Chemical Plants

Rusting of Iron

$$Fe + O_2 + H_2O \longrightarrow \underbrace{Fe_2O_3 \cdot x H_2O}_{\text{Rust}}$$

Oxidation of iron in two ways :-
① Gain of O_2
② Loss of electrons

$$Fe \longrightarrow Fe^{+3} + 3e^- \quad \text{(Oxidation of Iron)}$$

When iron is exposed to air, it reacts with oxygen & moisture to form a porous, flaky, brown layer over it. This process is known as rusting and the layer is called rust.

Methods of preventing rusting

- Painting
- Oiling and greasing
- Plastic Coating
- Galvanising (zinc coating)
- Tining (tin coating)
- Electroplating
- Alloying (e.g stainless steel)
- Sacrificial Protection

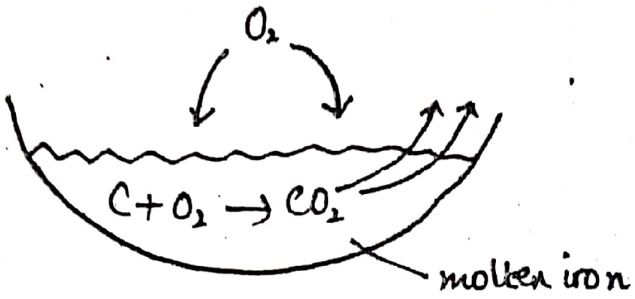
Less reactive metal is the cathode → It is prevented from oxidation.

Sacrificial Protection (Cathodic protection)

→ This method is used to protect underground pipelines of iron and ships from rusting.

→ Magnesium being more reactive than iron, will preferably

Conversion of Iron to Steel



⇒ Pig iron can be converted into pure iron or steel by blowing pre-heated oxygen with high pressure into molten iron which converts the carbon to carbondioxide gas, which escapes.

be oxidised by loosing electrons, thus prevents iron from rusting by not allowing it to get oxidised.

Extraction of Aluminium

Ores of Aluminium

→ Bauxite (or Alumina) — Al_2O_3 ← We will extract Aluminium from this ore.

→ Cryolite — Na_3AlF_6

Al_2O_3 → Basic Oxide
Insoluble in water
so for electrolysis, it must be molten

Two stages of extraction

→ Melting of ore (Al_2O_3 — mp is 2000°C) → Process is very expensive due to both these stages:-
① Electrolysis is a very expensive method.
② High temp of 2000°C has to be maintained

→ Electrolysis of molten Al_2O_3

⇒ How to decrease the melting point of Al_2O_3?

⇒ We will decrease the melting point of Al_2O_3 by adding impurities. This impurity needs to be wisely selected. The impurity selected is cryolite. When large amounts (70-80%) of cryolite are added to bauxite the melting point will drop from 2000°C to 800°C.

① ⇒ Solid cryolite & bauxite are mixed together. Then they are heated to melt the mixture

⇒ What is the purpose of adding cryolite in bauxite? (02)

→ To decrease the melting point of Al_2O_3 (or the mixture)

→ It acts as a solvent to dissolve Al_2O_3

$$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)} - \Delta H$$

Ad. of this reaction

① → Uses up oxygen so it does not react with molten Aluminium.

② → Reaction provides heat to keep the mixture in its molten state (800°C temp maintained) for electrolysis to continue.

Also reasons for not using platinum electrodes even though they wont need replacing!

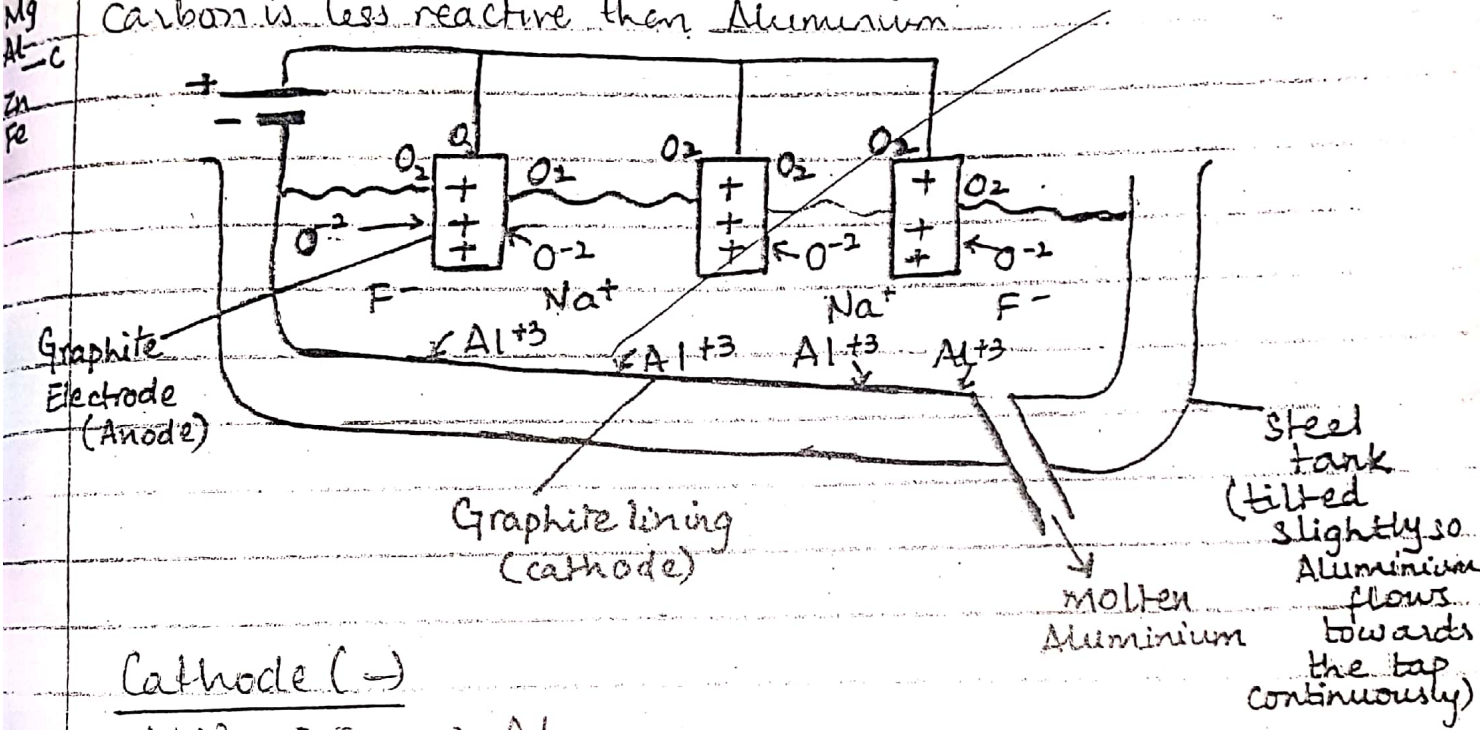
Electrolysis of molten mixture of Al_2O_3 and cryolite

$Al_2O_3 + C \xrightarrow{Heat}$ No reaction

$Al_2O_3 \xrightarrow{Heat}$ No reaction

Activity series
K
Na
Ca
Mg
Al
C
Zn
Fe

Aluminium Oxide cannot be reduced with carbon because carbon is less reactive than Aluminium



Cathode (−)

$Al^{+3}_{(l)} + 3e^- \longrightarrow Al_{(l)}$

Anode (−)

$E > 0$

$F^- < O^{-2}$ Fluorine is more reactive than oxygen but fluoride ion is less reactive than oxide ion.

O^{-2} is more reactive so it has a greater ability to discharge.

$2O^{-2}_{(l)} \longrightarrow O_{2(g)} + 4e^-$

$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)} - \Delta H$

From graphite anode

Pakistan	K
steel	Na
covers	Ca
many	Mg
Acres	Al
Zameen	Zn
its	Fe
loss	Pb
has	H
caused	Cu
many	Hg
serious	Ag
Govt	Au
Problems	Pt

(⇒) Why graphite anodes are replaced periodically (after regular intervals)?

(1)→ Oxygen produced at the anode reacts with carbon atoms in graphite anode to form CO_2 gas (Due to the high temp. in the surroundings, this reaction is more likely to take place). Thus the anode is used up.

NOTE: This reaction is exothermic & heat produced keeps the mixture in its molten state.

Uses of Aluminium along with reasons!

(1)→ It is used to make window panes because it is resistant to corrosion.

(2)→ It is used to make overhead cables & wires because it is good conductor of electricity.

(If asked in comparison with copper, it is because Aluminium has a lesser density than copper).

(3)→ Aluminium alloys are used to make aircraft bodies because Aluminium alloys have lower densities and are stronger.

(4)→ Aluminium is used for making kitchen foil and is also used in packaging because it is very malleable.

(5)→ Aluminium is used for making petrol and milk tanks because it is a good reflector of light (but it is not good for milk + petrol to come in contact with it as it reacts at high temperatures).

(6)→ It is used for making cooking utensils because it is a goo

conductor of heat.

Recycling of Metals:

Advantages (Environmental Ad.)

- It will reduce litter problem.
- It decreases the need for landfill sites.
- More land will be available for farming, as less metal will be extracted (Extraction is done by mining & that damages land)
- Recycling decreases the release of pollutant gases (released during extraction)

Advantages (Economic Ad.)

- Recycled metal is cheaper than the metal extracted from its ore.

Advantages (Social Ad.)

- Recycling helps conserve natural resources of metals for future generations.

Disadvantages

- Recycling needs man power, transportation for collection and its storage needs space. This causes wastage of time, money and manpower.
- Recycled metal is lower in quality than the metal extracted from its ore.