TOPIC 13 Metals

Objectives

Candidates should be able to:

(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
(c) identify representations of metals and alloys from diagrams of structures
(d) explain why alloys have different physical properties to their constituent elements
(e) place in order of reactivity calcium, copper, (hydrogen), iron, lead, magnesium, potassium, silver, sodium and zinc by reference to
  (i) the reactions, if any, of the metals with water, steam and dilute hydrochloric acid,
  (ii) the reduction, if any, of their oxides by carbon and/or by hydrogen
(f) 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
(g) deduce the order of reactivity from a given set of experimental results
(h) describe the action of heat on the carbonates of the listed metals and relate thermal stability to the reactivity series
(i) describe the ease of obtaining metals from their ores by relating the elements to their positions in the reactivity series
(j) describe and explain the essential reactions in the extraction of iron using haematite, limestone and coke in the blast furnace
(k) 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
(l) state the uses of mild steel
(m) 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
(n) 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

1. **Physical Properties of Metals**

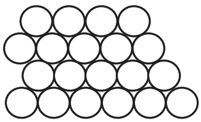
   Metals usually have high densities, melting points and boiling points. Some exceptions would be Group I metals (some are less dense than water) and mercury (which is a liquid at room temperature and pressure). Metals are good conductors of heat and electricity, and are often shiny, ductile and malleable.

2. **Alloys**

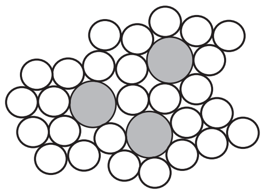
   Pure metals are usually not widely used as they are soft and may corrode easily, therefore alloys are used instead. An alloy is a mixture of a metal and one or more elements, which may be metal or non-metal.

   A pure metal is soft due to the regular arrangement of atoms in the metal lattice. The atoms are arranged in neat layers which slide past each other easily when a force is applied.

   In an alloy however, the arrangement of atoms is disrupted by the presence of atoms of different sizes. This prevents the layers of atoms from sliding easily, making the alloy harder than the pure metal.



pure metal



alloy

   Alloying metals helps to change properties to make it more suitable for a particular use. For instance, an alloy of iron and chromium has greater resistance to rusting compared to pure iron.

3. The Reactivity Series

The reactivity series arranges metals in order of reactivity. Metals that are more reactive have a higher tendency of forming ions compared to metals that are less reactive.

most reactive ↑	Potassium	(K)
	Sodium	(Na)
	Calcium	(Ca)
	Magnesium	(Mg)
	Aluminium	(A*l*)
	(Carbon)	(C)
	Zinc	(Zn)
	Iron	(Fe)
	Lead	(Pb)
	(Hydrogen)	(H)
	Copper	(Cu)
	Silver	(Ag)
least reactive	Gold	(Au)

4. Displacement Reactions of Metals

Displacement reaction takes place when a more reactive metal is placed in the salt solution of a less reactive metal. Since the more reactive metal has a higher tendency to form ions, it displaces the less reactive metal from its salt.

For example, when copper metal is placed in a solution of silver nitrate, copper displaces silver from the silver nitrate solution.

copper metal + silver nitrate → copper nitrate + silver

$Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$

However, no reaction occurs when a less reactive metal is placed in the salt solution of a more reactive metal. No change is seen when copper metal is placed in magnesium sulfate solution since magnesium is more reactive than copper.

5. Reactions of Metals with Water

Metal	Observations	Equation
Potassium	Reacts violently with cold water. Hydrogen gas catches fire and explodes.	Reaction with cold water: $2K(s) + 2H_2O(l) \rightarrow$ $2KOH(aq) + H_2(g)$
Sodium	Reacts violently with cold water. Hydrogen gas may catch fire.	Reaction with cold water: $2Na(s) + 2H_2O(l) \rightarrow$ $2NaOH(aq) + H_2(g)$
Calcium	Reacts moderately with cold water.	Reaction with cold water: $Ca(s) + 2H_2O(l) \rightarrow$ $Ca(OH)_2(aq) + H_2(g)$
Magnesium	Reacts slowly with cold water. Hot magnesium reacts violently with steam and burns with a white glow.	Reaction with cold water: $Mg(s) + 2H_2O(l) \rightarrow$ $Mg(OH)_2(aq) + H_2(g)$ Reaction with steam: $Mg(s) + H_2O(g) \rightarrow$ $MgO(s) + H_2(g)$
Aluminium	Reacts readily with steam. Reaction slows down due to the formation of a protective oxide layer.	Reaction with steam: $2Al(s) + 3H_2O(g) \rightarrow$ $Al_2O_3(s) + 3H_2(g)$
Zinc	Hot zinc reacts readily with steam. Zinc oxide produced is yellow when hot and white when cold.	Reaction with steam: $Zn(s) + H_2O(g) \rightarrow$ $ZnO(s) + H_2(g)$
Iron	Hot iron reacts slowly with steam.	Reaction with steam: $3Fe(s) + 4H_2O(g) \rightarrow$ $Fe_3O_4(s) + 4H_2(g)$
Lead	No reaction	-
Copper	No reaction	-
Silver	No reaction	-
Gold	No reaction	-

6. **Reactions of Metals with Dilute Hydrochloric Acid**

Metal	Observations	Equation
Potassium	Reacts violently	$2K(s) + 2HCl(aq) \rightarrow 2KCl(aq) + H_2(g)$
Sodium	Reacts violently	$2Na(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2(g)$
Calcium	Reacts violently	$Ca(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$
Magnesium	Reacts readily	$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
Aluminium	Reacts readily	$2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$
Zinc	Reacts moderately fast	$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
Iron	Reacts slowly	$Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$
Lead	No reaction	-
Copper	No reaction	-
Silver	No reaction	-
Gold	No reaction	-

Reactions of metals with dilute hydrochloric acid can be seen as the displacement of hydrogen in the acid by a more reactive metal.

While lead is higher than hydrogen in the series, it does not react with dilute hydrochloric acid due to the formation of an insoluble layer of lead(II) chloride. The salt acts as a protective layer and prevents the acid from reacting further with the metal.

7. **Extraction of Metals**

Metals are usually found in nature as ores, which mainly consist of metal oxides. The extraction of a metal from its ore depends on its reactivity. A more reactive metal usually requires tougher methods of extraction compared to a less reactive metal.

Zinc and metals lying below it in the reactivity series can be extracted from their oxides through heating with carbon.

Aluminium and other metals above it in the reactivity series form very stable oxides that are not easily reduced. They can only be extracted from their ores through electrolysis of their molten oxides.

8. **Thermal Stability of Metal Carbonates**

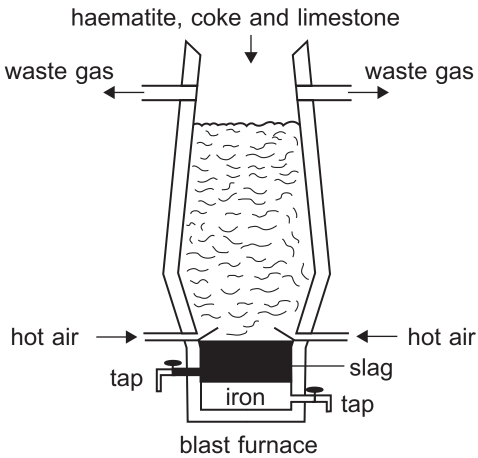
Reactive metals form very stable carbonates which do not decompose easily upon heating. On the other hand, the carbonates of metals which are less reactive are easily decomposed by heat.

Carbonates of potassium and sodium are thermally stable since these metals are found high in the reactivity series. Carbonates of calcium, magnesium, zinc, iron, lead and copper decompose upon heating to form metal oxide and carbon dioxide.

Silver carbonate is the least stable since silver metal is the least reactive. It decomposes completely into silver metal and carbon dioxide.

9. Extraction of Iron

Iron is extracted from its ore, haematite (contains iron(III) oxide, Fe_2O_3), by heating with carbon. Haematite, coke (mainly carbon) and limestone (calcium carbonate, $CaCO_3$) are loaded at the top of the blast furnace while hot air is introduced at the bottom of the furnace.



1. Coke is oxidised by oxygen in the hot air in an exothermic reaction.

   carbon + oxygen → carbon dioxide
   $C(s) + O_2(g) \rightarrow CO_2(g)$

   Carbon dioxide then further reacts with carbon to produce carbon monoxide.

   carbon dioxide + carbon → carbon monoxide
   $CO_2(g) + C(s) \rightarrow 2CO(g)$

2. Carbon monoxide reduces iron(III) oxide in haematite to molten iron. Since iron has high density, it sinks to the bottom of the furnace.

   iron(III) oxide + carbon monoxide → molten iron + carbon dioxide
   $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(l) + 3CO_2(g)$

3. Limestone undergoes thermal decomposition.

   calcium carbonate → calcium oxide + carbon dioxide
   $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

   Silicon dioxide, an acidic impurity, is removed by reacting with calcium oxide, which is basic in nature.

   calcium oxide + silicon dioxide → calcium silicate (slag)
   $CaO(s) + SiO_2(s) \rightarrow CaSiO_3(l)$

   The reaction forms slag, which floats on top of molten iron. The molten iron and slag are tapped off separately at the bottom of the furnace.

10. Steel

Steel is an alloy of iron and carbon. The properties of steel can be further altered by the addition of other metals and controlling the amounts of these components.

Molten iron produced from the blast furnace is known as pig iron or cast iron and contains many impurities. These impurities can be removed by introducing oxygen through the molten iron. Removal of these impurities leaves behind pure iron, which is also known as wrought iron.

Different percentages of carbon may be added to wrought iron to form steel. Low-carbon steel (or mild steel) is hard and malleable and can be used to make car bodies. High-carbon steel is harder but more brittle than low-carbon steel and is used to make cutting tools.

Stainless steel is a mixture of iron, carbon, chromium and nickel. Addition of these elements makes it more resistant to corrosion and hence, makes it suitable for use as cutlery or surgical tools.

11. Rusting of Iron

Iron corrodes in the presence of water and oxygen to form rust (hydrated iron(III) oxide).

iron + oxygen + water → hydrated iron(III) oxide
$4Fe(s) + 3O_2(g) + 2H_2O(l) \rightarrow 2Fe_2O_3.H_2O(s)$

Rusting can be prevented by painting or covering the metal with a layer of oil. This protects iron from being exposed to oxygen and water.

Sacrificial protection can be used to prevent rusting. A more reactive metal is used as the sacrificial metal and corrodes in place of iron. This is usually done by attaching a block of magnesium or zinc to the iron, or galvanisation, where iron is coated with zinc.

Plating iron with zinc also protects the metal from direct contact with water and oxygen.