Topic 10/11 : Metals and Making Electricity

Revised April 1995

The uses of metals depend on their properties :

Examples :

1. Copper is used in electrical wiring because it is a good conductor of electricity It is also used in boilers because it is a good conductor of heat.
2. Aluminium is used in building aircraft because of its strength and low density.
3. Gold can be beaten into thin sheets because of its malleability.

The Reactivity Series (also called 'The Electrochemical Series')

This series arranges the metals in order of chemical reactivity (most reactive at the top).

Metal	Symbol
Potassium	K
Sodium	Na
Lithium	Li
Calcium	Ca
Magnesium	Mg
Aluminium	Al
Zinc	Zn
Iron	Fe
Tin	Sn
Lead	Pb
Hydrogen	H
Copper	Cu
Mercury	Hg
Silver	Ag
Gold	Au
Platinum	Pt

Reactive metals are metals which lose electrons easily

Evidence for a reactivity series

1. Reaction of metals with dilute acids (e.g. Hydrochloric acid)

Example 1 : Magnesium

$$Mg + 2\,H^{+}Cl^{-} \xrightarrow{\text{fast}} Mg^{2+}(Cl^{-})_2 + H_2$$

By studying this equation we can see that the chloride ions are not involved in the reaction. They are called 'Spectator Ions'. We will remove them from the equation :

$$Mg + 2\,H^{+} \rightarrow Mg^{2+} + H_2$$

Now we can see that the reaction involves :

(a) Loss of electrons from the Magnesium (Oxidation) :

$$Mg \xrightarrow{\text{fast}} Mg^{2+} + 2e$$

(b) Gain of electrons by the Hydrogen ions (Reduction) :

$$2\,H^{+} \quad 2e \rightarrow H_2$$

Note 1. The above two equations are called Ion-electron half-equations.

Note 2. **RED**uction and **OX**idation always go on together (**REDOX**)

Example 2 : Zinc

$$Zn + 2\,H^{+}Cl^{-} \xrightarrow{\text{slow}} Zn^{2+}(Cl^{-})_2 + H_2$$

(a) Zn loses electrons (Zn is oxidised) :

$$Zn \xrightarrow{\text{slow}} Zn^{2+} + 2e$$

(b) H^{+} gains electrons (H^{+} is reduced) :

$$2\,H^{+} \quad 2e \rightarrow H_2$$

Since Mg is more reactive than Zn, Mg loses electrons more readily and so the reaction is faster.

2. Reaction of metals with Water.

Remember that, due to its ionisation, Water contains H^+

$$H_2O \rightleftharpoons H^+ + OH^-$$

The metals react with the H+ ions in the Water.

Example 1 : Magnesium

$$Mg + H_2O \xrightarrow{\text{fast}} Mg^{2+}(OH^-)_2 + H_2$$

Ion-electron half-equations :

fast

$$Mg \rightarrow Mg^{2+} + 2e$$
$$2H^+ + 2e \rightarrow H_2$$

Example 2 : Zinc

$$Zn + H_2O \xrightarrow{\text{slow}} Zn^{2+}(OH^-)_2 + H_2$$

Ion-electron half-equations :

slow

$$Zn \rightarrow Zn^{2+} + 2e$$
$$2H^+ + 2e \rightarrow H_2$$

Again, the reaction of Mg is faster because it loses electrons more easily.

3. Reaction of metals with Oxygen

Example 1 : Magnesium

$$Mg + O_2 \xrightarrow{\text{fast}} Mg^{2+}O^{2-}$$

Ion-electron half-equations :

fast

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

Example 2 : Zinc

$$Zn + O_2 \xrightarrow{\text{slow}} Zn^{2+}O^{2-}$$

Ion-electron half-equations :

$$Zn \xrightarrow{\text{slow}} Zn^{2+} + 2e$$

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

Again, the reaction of Mg is faster because it loses electrons more easily.

Displacement Reactions

The easier it is for the metal to lose electrons the more difficult it is for its positive ion to regain them e.g.

$$Zn \underset{\text{difficult}}{\overset{\text{easy}}{\rightleftharpoons}} Zn^{2+} + 2e$$

$$Cu \underset{\text{easy}}{\overset{\text{difficult}}{\rightleftharpoons}} Cu^{2+} + 2e$$

Consequently, when a piece of Zinc is placed in a solution containing Cu^{2+} ions, electrons will flow from the Zn to the Cu^{2+} :

$$Zn + Cu^{2+}SO_4^{2-} \rightarrow Zn^{2+}SO_4^{2-} + Cu$$

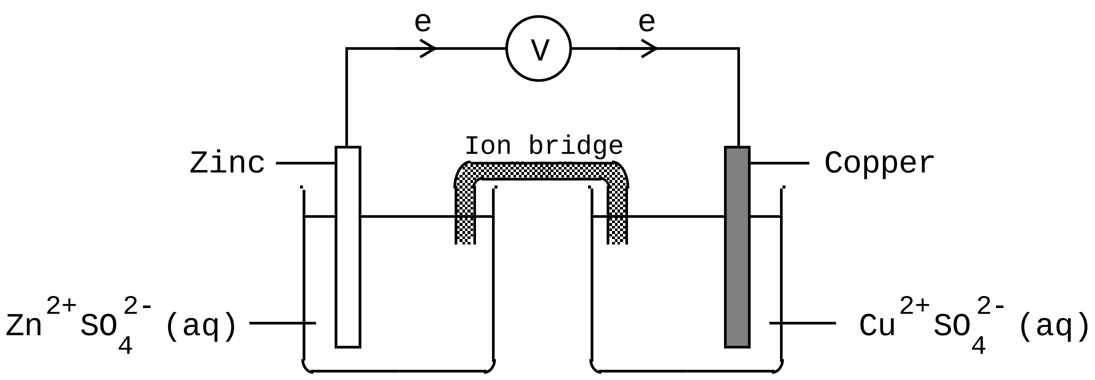
Ion-electron half-equations :

$$Zn \rightarrow Zn^{2+} + 2e$$

$$Cu^{2+} + 2e \rightarrow Cu$$

The solution becomes colourless due to the blue Cu^{2+} being replaced by the colourless Zn^{2+}. Copper metal is formed. Copper has been displaced by Zinc.

We can show, using a **CELL**, that electrons flow from Zn to Cu^{2+} :



Electrons flow from Zn to Cu^{2+} along the wires and through the voltmeter. This is an electric current (V = 1.10 V).

The Ion-bridge

This is a piece of filter paper soaked in Potassium nitrate solution. As the reaction proceeds, positive charge would build up in the left hand beaker due to the production of Zn^{2+} ions :

$$Zn \rightarrow Zn^{2+} + 2e$$

This would prevent negative electrons from leaving this beaker.

Negative charge would build up in the right hand beaker due to the removal of Cu^{2+} ions :

$$Cu^{2+} + 2e \rightarrow Cu$$

This would prevent negative electrons from entering this beaker.

The $K^+NO_3^-$ ion-bridge prevents these charges from building up : the NO_3^- ions diffuse into the left beaker cancelling the positive charge ; the K^+ ions diffuse into the right beaker cancelling the negative charge.

In a displacement reaction electrons always flow from the more reactive metal to the less reactive metal's ions.

Problem 1 : Will the following reaction occur :

$$Zn + Ag^+NO_3^- \rightarrow Zn^{2+}(NO_3^-)_2 + Ag$$

Answer : Yes, because Zn is more reactive than Ag.

Electrons ***will*** flow from Zn to Ag^+

$Zn \rightarrow Zn^{2+} + 2e$
$Ag^+ + e \rightarrow Ag$

Problem 2 : Will the following reaction occur :

$Cu + (H^+)_2SO_4^- \rightarrow Cu^{2+}SO_4^{2-} + H_2$

Answer : No, because Cu is less reactive than H.

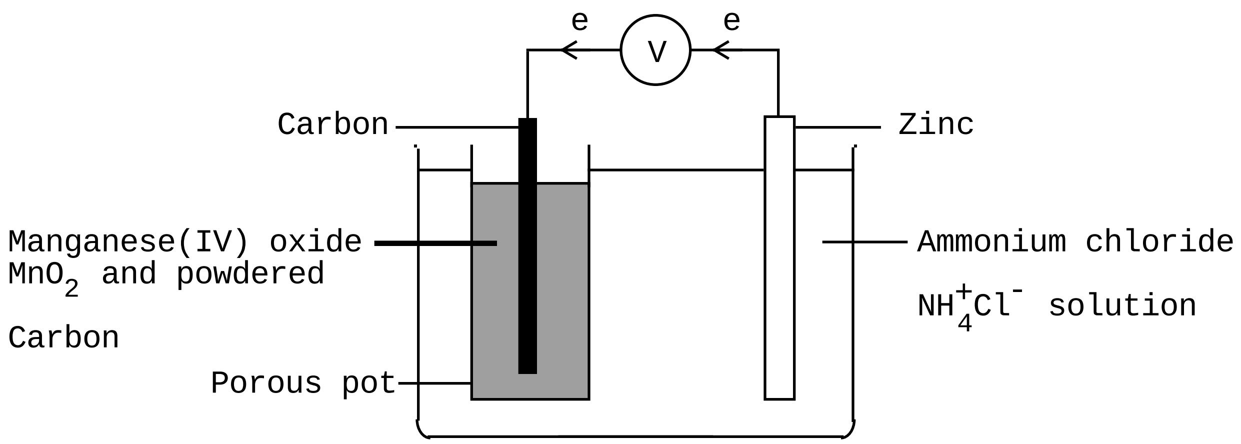
Electrons ***cannot*** flow from Cu to H^+

This is why all metals below H in the reactivity series cannot displace Hydrogen from acids.

Cells

A cell or battery is a device which makes electricity from a chemical reaction. The Zinc/Copper ion cell has been discussed ; here are some others :

Example 1 : The Leclanché cell



Electrons flow form Zinc to Manganese dioxide :

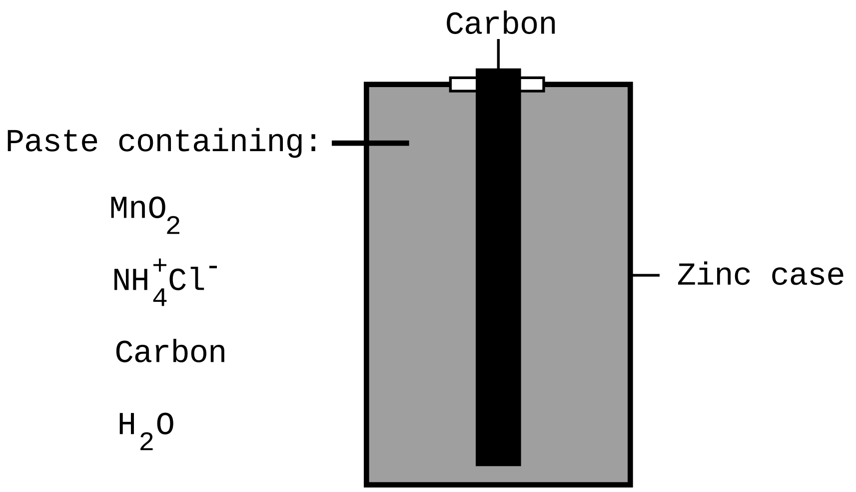
$Zn \rightarrow Zn^{2+} + 2e$ (Oxidation of Zn)
$MnO_2 + e \rightarrow$ other substances (reduction of MnO_2)

Zn is oxidised by MnO_2 - an oxidising agent.
MnO_2 is reduced by Zn - a reducing agent.

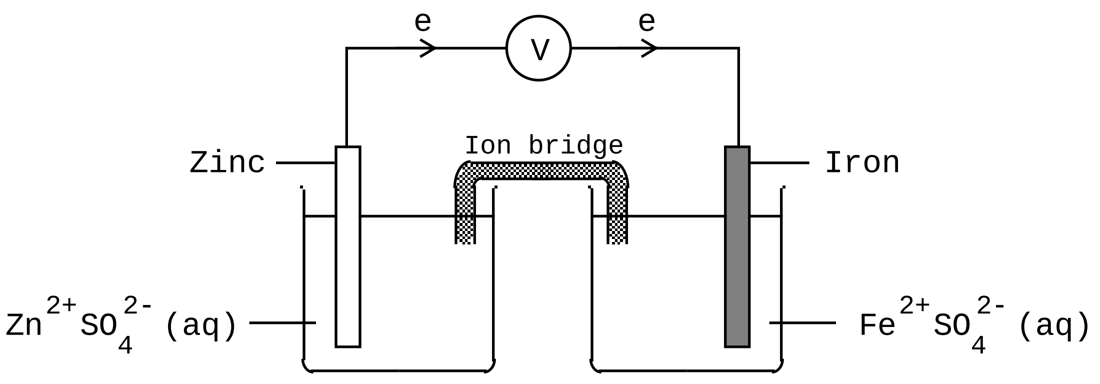
The ions in the $NH_4^+Cl^-$ electrolyte diffuse through the porous pot and prevent a build up of charge.

Example 2 : The Dry Cell (the every-day 'battery')

This is a more convenient Leclanché cell :



Example 3 :



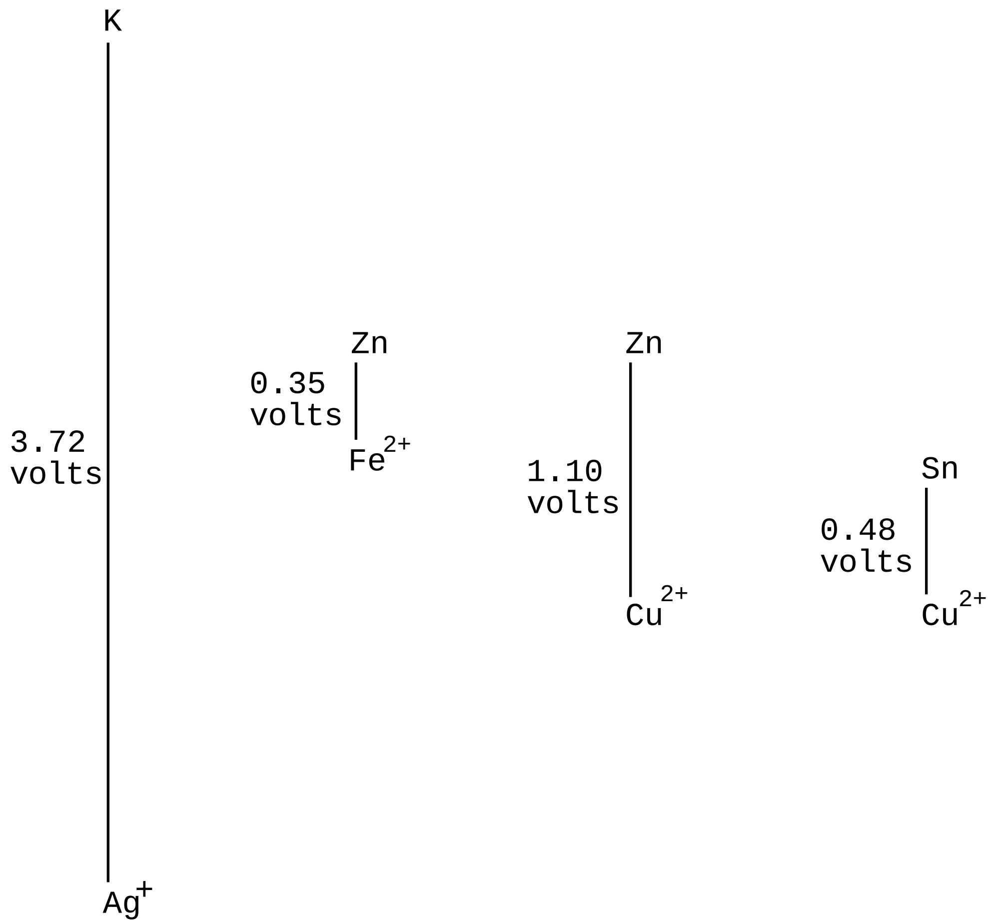
The Zinc loses electrons ; the Fe^{2+} ions gain them :

$Zn \rightarrow Zn^{2+} + 2e$

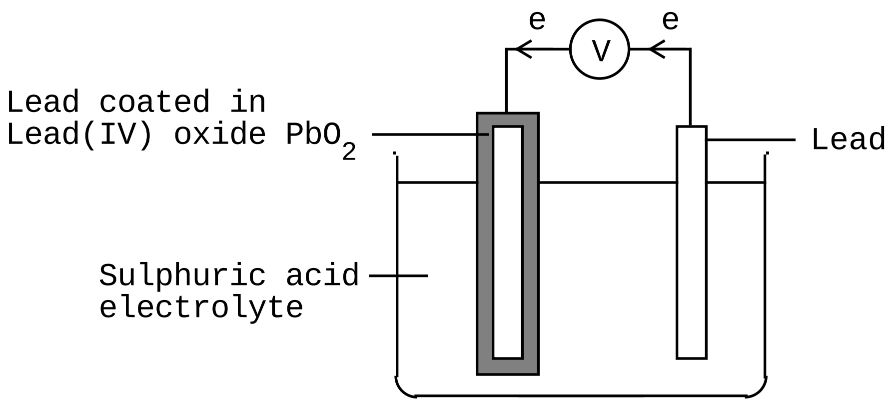
$Fe^{2+} + 2e \rightarrow Fe$

The cell voltage is only 0.35 volts. This is because the two metals are so close together in the reactivity series. Metals at the top of the series **push** electrons forcibly ; metal ions at the bottom of the series **pull in** electrons forcibly.

So a combination of a reactive metal with an ion of an unreactive metal will give the highest voltage e.g.



Example 4 : The car battery



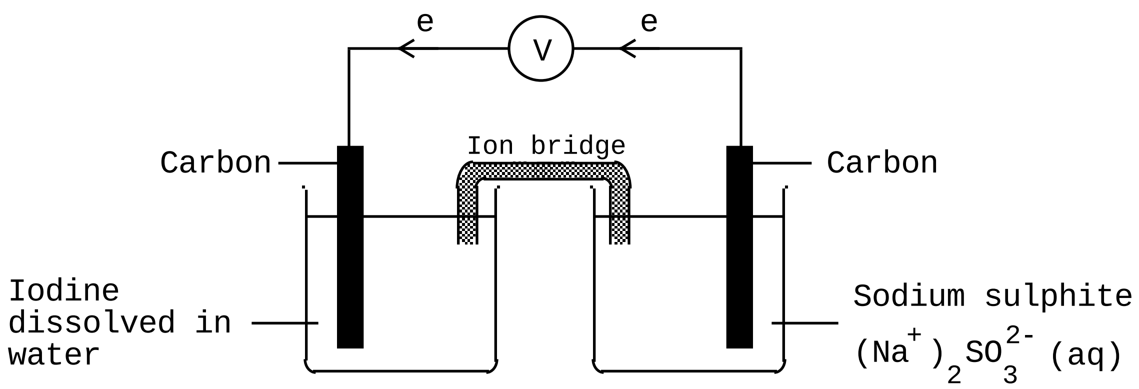
Electrons flow from Lead to Lead(IV) oxide (V = 2.03 volts)

$Pb \rightarrow Pb^{2+} + 2e$

$PbO_2 + e \rightarrow$ other substances

Most cells run down when their chemicals are used up and have to be replaced. The Lead/acid cell can be recharged : if the two Lead plates are connected to a power supply Lead reforms at the cathode (-ve) and Lead(IV) oxide at the anode (+ve).

Example 5 :



Electrons flow from Sodium sulphite to Iodine (V = 0.37 volts) :

*	SO_3^{2-}	+	H_2O	->	SO_4^{2-}	+	$2H^+$	+	2e
	I_2	+	2e	->	$2I^-$				

(* reduction by the Sulphite ion : see Data Book)

Cells have two advantages over mains electricity : they can be carried around and they are safer (lower voltages). They have two disadvantages : when they run down they have to be replaced and they are expensive !

Extraction of metals from their ores

Only the very unreactive elements (Hg - Pt) occur uncombined in nature. The more reactive metals have combined with other elements and, therefore, occur in compounds. These metal compounds are called **ORES**.

Example : Iron ore is Iron(III) oxide $(Fe^{3+})_2(O^{2-})_3$

The more reactive the metal the stronger the bonds in the ore and the more difficult the extraction of the metal from its ore.
Unreactive metals (Hg - Pt) were the easiest to obtain and therefore the first to be discovered. They can be obtained from their ores by heating alone.

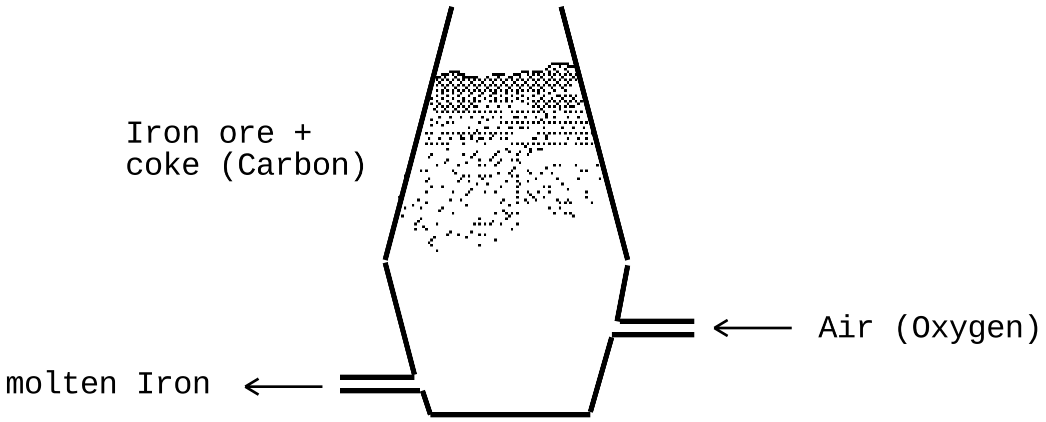
Example : $(Ag^+)_2O^{2-} \xrightarrow{\text{heat}} Ag + O_2$

There are, however, vast deposits of the ores of **more reactive** metals. These metals are important for industry (Iron to make machinery etc) and in our daily lives (Aluminium to make cooking foil, pots and pans, window frames etc).

The more reactive metals (Fe - Cu) can be obtained by heating with Carbon monoxide which pulls the Oxygen away from the metal e.g.

$$(Fe^{3+})_2(O^{2-})_3 \quad + \quad CO \quad \xrightarrow{\text{heat}} \quad Fe \quad + \quad CO_2$$

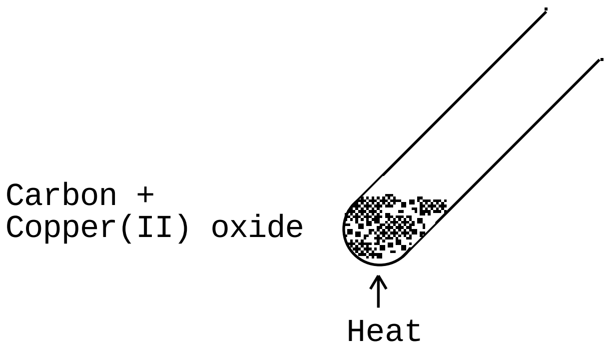
This reaction is used in the blast furnace to produce Iron :



The Carbon monoxide is formed by reaction of Carbon with Oxygen :

$$2\ C \quad + \quad O_2 \quad \rightarrow \quad 2\ CO$$

<u>Experiment</u> : To obtain Copper from Copper(II) oxide



Carbon monoxide is produced by reaction of Carbon with Oxygen in the air. The Carbon monoxide then reacts with the Copper(II) oxide :

$$C \quad + \quad O_2 \quad \rightarrow \quad CO$$

$$Cu^{2+}O^{2-} \quad + \quad CO \quad \rightarrow \quad Cu \quad + \quad CO_2$$

When the contents of the test tube are shaken out, pieces of Copper are seen.

In all these examples the metal is obtained by **REDUCTION** of the metal ion.

Ores of the very reactive metals (K - Zn) are stable to heat. The only way to obtain the metal is by electrolysis of the molten ore.

Example : Extraction of Aluminium from bauxite
(impure Aluminium oxide)

The bauxite is first treated with Sodium hydroxide to remove Iron impurities and leave pure Aluminium oxide.

This oxide has such a high melting point (2050 ^{0}C) that cryolite $(Na^{+})_3AlF_6^{3-}$ is added to lower the melting point to about 900 ^{0}C.

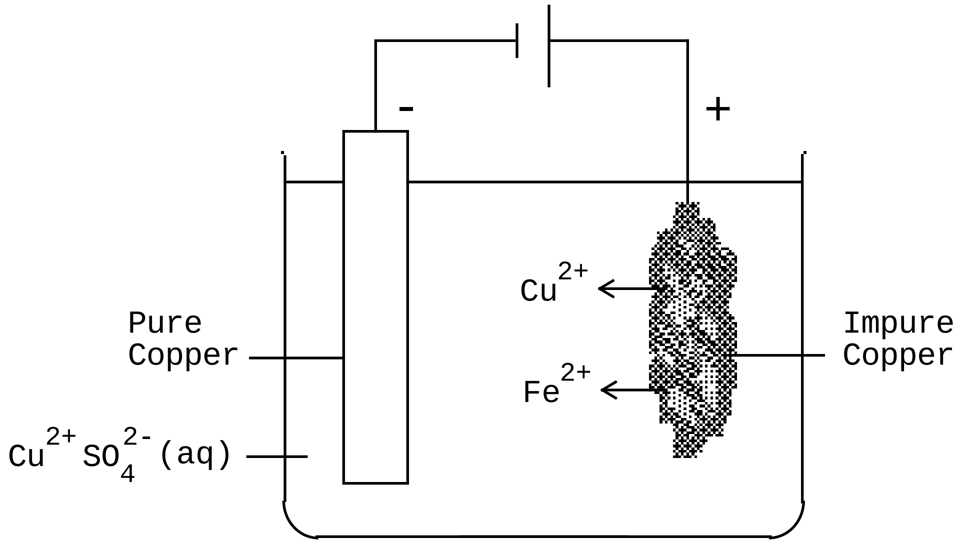
Reduction of the Aluminium ion occurs at the cathode :

Al^{3+} + 3e -> Al

It is essential that metals are recycled (used again) since these ores will not last forever.

Purification of metals by electrolysis

Example : impure Copper (main impurity is Iron)



The anode draws electrons away from Cu and Fe.

Cu -> Cu^{2+} + 2e

Fe -> Fe^{2+} + 2e

Both the Cu^{2+} and Fe^{2+} ions move to the Copper cathode. Cu^{2+} gain electrons more easily than Fe^{2+} (Cu is less reactive than Fe) so Cu^{2+} ions are reduced at the Copper cathode to form a layer of pure Copper :

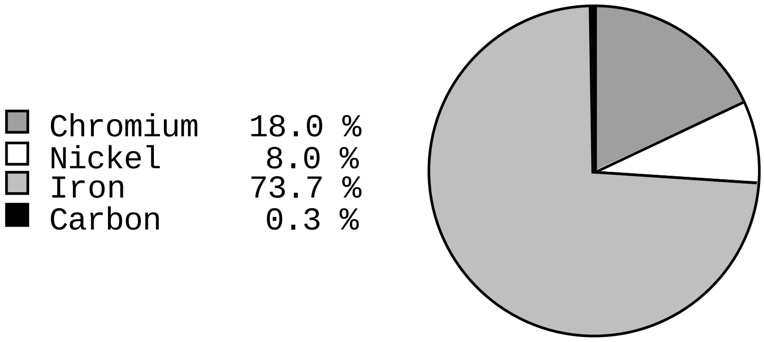
Cu^{2+} + 2e -> Cu

The Fe^{2+} ions remain in solution.

Alloys

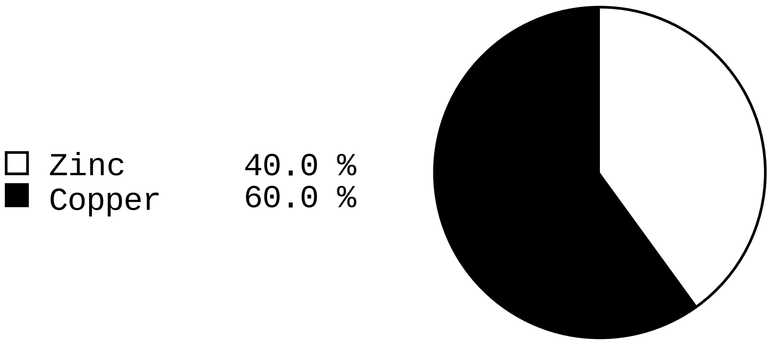
When other elements are added to a metal to improve its properties the mixture is called an **ALLOY**.

Example 1 : stainless steel



The properties of an alloy vary with its composition e.g. the more Carbon in steel the harder it becomes.
Unlike Iron, stainless steel does not corrode. It is therefore used for making cutlery, sinks etc.

Example 2 : brass



Brass is stronger than either Copper or Zinc and does not corrode. It is used to make shell cases, electric plugs etc.