THE REACTIVITY SERIES

Metals react with other substances at different rates e.g. the reaction of metals with acids:

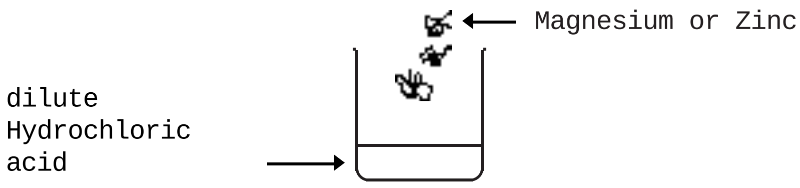
$$Mg + 2\,H^+Cl^- \xrightarrow{\text{fast}} Mg^{2+}(Cl^-)_2 + H_2$$

Magnesium chloride + Hydrogen gas

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

Zinc(II) chloride + Hydrogen gas

Experiment : Compare the reactions of Magnesium and Zinc with dilute Hydrochloric acid.



We note that Magnesium reacts much more quickly than Zinc. Magnesium is more reactive than Zinc.

In both reactions, the metal atoms LOSE electrons. Magnesium loses electrons more easily than Zinc.

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

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

A reactive metal is one which loses electrons easily.

Metals can be placed in order of reactivity (most reactive at the top):

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

This order of reactivity is called the 'electrochemical series' or 'reactivity series'.

More evidence for the above order of reactivity can be obtained by the reaction of metals with Water and with Oxygen.

REACTION OF METALS WITH WATER

Example 1 : Magnesium

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

Example 2 : Zinc

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

REACTION OF METALS WITH OXYGEN

Example 1 : Magnesium

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

Example 2 : Zinc

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

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.

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

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

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

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

i.e.

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

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

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

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

PROBLEM : 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 : 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^{+}

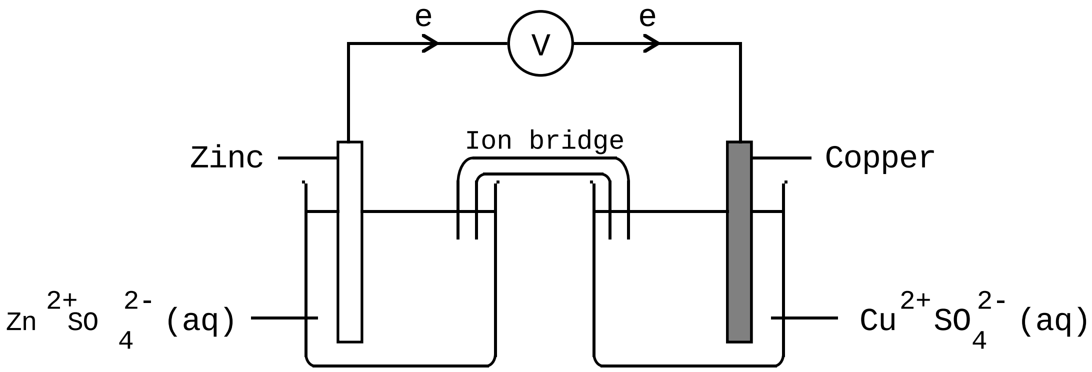
This is why metals below H in the reactivity series cannot displace Hydrogen from acids.

PROBLEM : Will the following reaction occur :

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

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

We can show, using a **CELL**, that electrons DO 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.

The ion bridge is a piece of filter paper soaked in Potassium nitrate solution. It allows ions to move from one beaker to the other.

LOSS of electrons is referred to as **OXIDATION**. The Zinc is oxidised:

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

GAIN of electrons is referred to as **REDUCTION**. The Copper ion is reduced:

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

Notice that **<u>RED</u>**uction and **<u>OX</u>**idation always happen together - a **REDOX** reaction.

We can obtain the overall REDOX equation by adding the above two equations together:

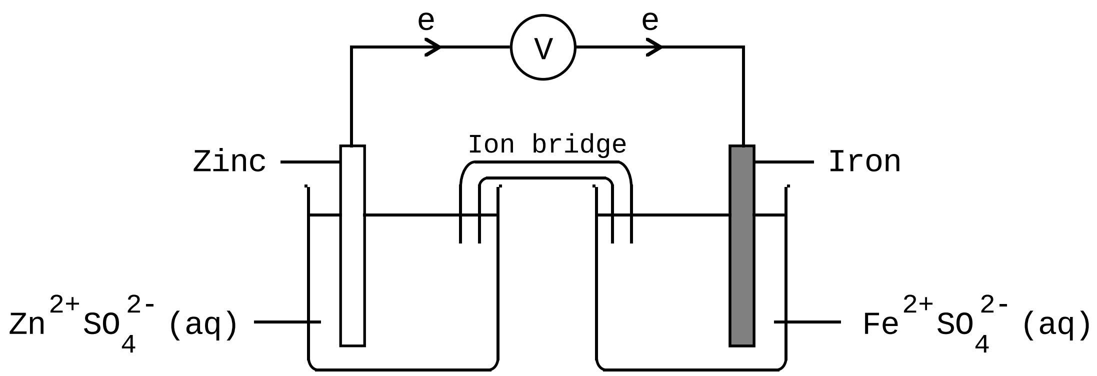
$$Zn \quad \rightarrow \quad Zn^{2+} \quad + \quad \cancel{2e}$$

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

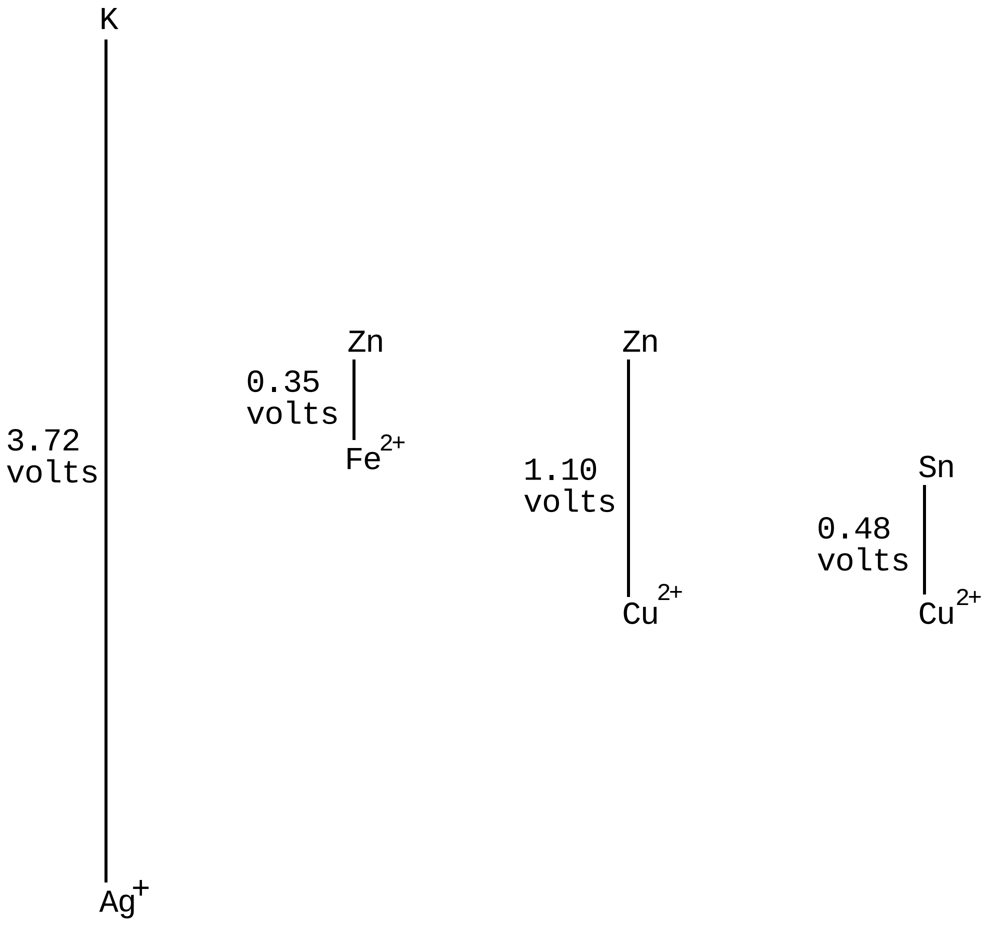
$$Zn \quad + \quad Cu^{2+} \quad \rightarrow \quad Zn^{2+} \quad + \quad Cu$$

The Zn/Cu^{2+} cell has a voltage of 1.10 volts.

The Zn/Fe^{2+} cell, shown below, has a voltage of only 0.35 volts.



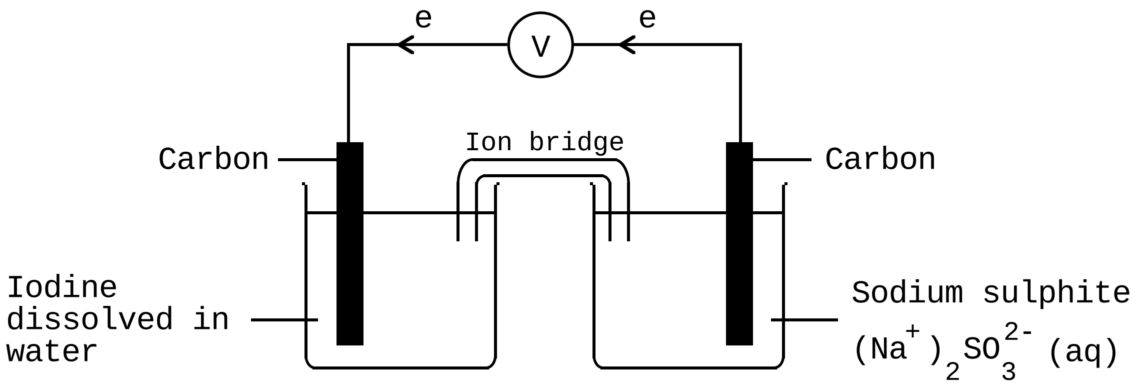
The voltage is so low because the two metals are so close together in the reactivity series. Metals at the top of the series **push** electrons forcibly ; metal ions at the bottom of the series **pull in** electrons forcibly. So a combination of a very reactive metal with an ion of a very unreactive metal will give the highest voltage e.g.



Note that, in a cell, electrons always flow from the most reactive metal to the least reactive metal's ions.

We should also note that cells without metals are possible!

EXAMPLE



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

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

* This reaction is in the SQA Data Book

The sulphite/water mixture is oxidised; the Iodine is reduced.

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 e.g. Iron ore is Iron(III) oxide

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 e.g. Silver

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

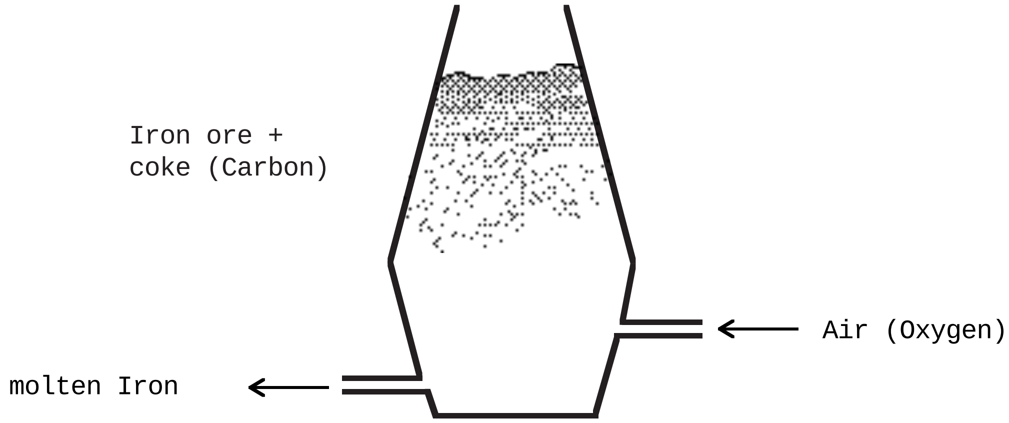
There are, however, vast deposits of the ores of **more reactive** metals. These metals are important for industry (Iron to make machinery etc) and in our daily lives (Aluminium to make cooking foil, pots and pans, window frames etc).

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

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

Since Fe^{3+} has been reduced, we refer to this reaction as 'the reduction of Iron(III) oxide'.

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



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

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

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 e.g. Aluminium

Bauxite (impure Aluminium oxide) 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 negative electrode :

$Al^{3+} + 3e \rightarrow Al$

Oxidation of the oxide ion occurs at the positive electrode :

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