Summary: redox reactions and electrode potentials

Metal ion-metal half-cells

We can set up a simple half-cell by using a strip of metal dipping into a solution of metal ions. For example, the copper-zinc cell consists of two half-cells (Figure 6).

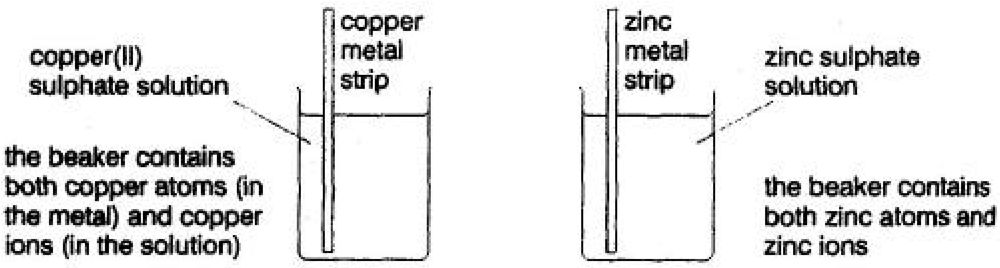
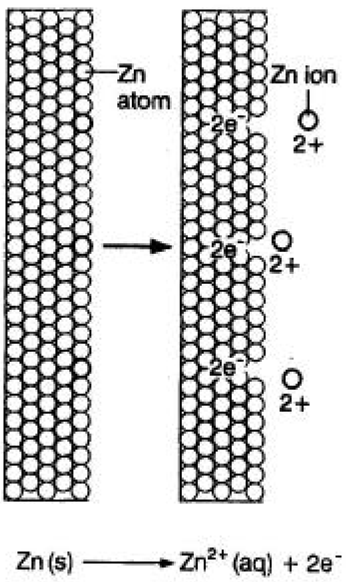


Figure 6 The copper and zinc half-cells

Each of these half-cells has its own electrode potential. Take the zinc half-cell, for example. The Zn atoms in the zinc strip form Zn^{2+} ions by releasing electrons:

$Zn(s) \rightarrow Zn^{2}(aq) + 2e^{-}$

Figure 7 Zinc atoms form zinc ions, releasing electrons and setting up a potential difference



The electrons released make the Zn scrip negatively charged relative to the solution, so there is a potential difference between the zinc strip and the solution. The Zn^{2+} ions in the solution accept electrons, re-forming Zn atoms:

$Zn^{2+}(aq) + 2e^{-} \rightarrow Zn(s)$

When Zn^{2+} ions are turning back to Zn as fast as they are being formed an equilibrium is set up:

$Zn^{2+}(aq) + 2e^{-} \leftrightarrow Zn(s)$

For a general metal, M:

$M^{2+}(aq) + 2e^{-} \leftrightarrow M(s)$

The position of this equilibrium determines the size of the potential difference (the electrode potential) between the metal strip and the solution of metal ions. The further to the right the equilibrium lies, the greater the tendency of the electrode to accept electrons, and the more positive the electrode potential.
When we put two half-cells together, the one with the more positive potential will become the positive terminal of the cell, and the other one will become the negative terminal.