Focus Idea 4 – Electrochemistry: *Oxidation-reduction reactions are increasingly important as a source of energy.*

Prior learning: Preliminary modules 8.3.2.
Background: Reactions of metals usually require the transfer of electrons. Metals can be arranged in an activity series from most active to least active:

MOST ⟵⟶ LEAST
K Na Li Ba Ca Mg Al Cr Zn Fe Co Ni Sn Pb Cu Hg Ag Pt Au

It is recommended that you become familiar with the names for the metals signified by the symbols listed. A Periodic Table that gives the names as well as symbols will be provided in the HSC examination.

<u>explain</u> **the displacement of metals from solution in terms of transfer of electrons**

- More active metals will displace less active metal ions from solution in an oxidation-reduction reaction.
- When an active metal is placed in a solution containing ions of a less active metal, the active metal displaces the less active metal from solution. This occurs because a more active metal atom loses one or more electrons and becomes a positive ion. The electrons lost are transferred to the ions of the less active metal, resulting in them becoming metal atoms.

  For example, if an iron nail is placed in a solution of blue copper (II) salt, some of the iron nails dissolves.

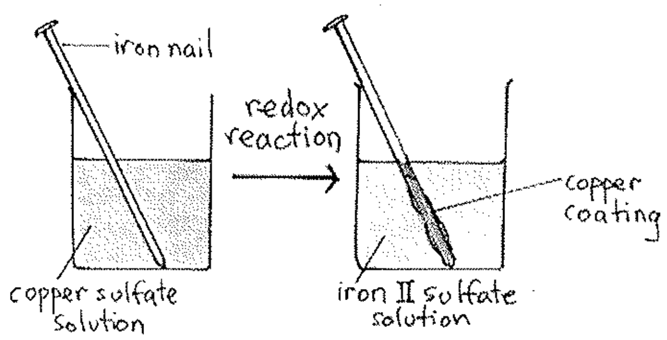
  $Fe \rightarrow Fe^{2+} + 2e^-$ (oxidation)

  At the same time, the blue colour of Cu^{2+} ions disappears and a dark copper coating appears on the nail surface.

  $Cu^{2+} + 2e^- \rightarrow Cu$ (reduction)

  The overall reaction is:

  $Fe + Cu^{2+} \rightarrow Fe^{2+} + Cu$ (oxidation-reduction)



The electrons lost by iron atoms undergoing oxidation are used to reduce copper (II) ions to copper atoms. Oxidation–reduction reactions (also called *redox* reactions) involve transfer of electrons.

<u>identify</u> **the relationship between displacement of metal ions in solution by other metals to the relative activity of metals**

- If a metal is higher in the activity series, the metal atoms will react when put in a solution of ions of a metal that is lower in the activity series. The less active metal ions are displaced from solution as they form atoms.

- In reacting, the more active metal atom (M) changes to a metal ion (M^{n+}) by losing one or more electrons to form a cation.

Summary: $M \rightarrow M^{n+} + ne^-$

- Metal reactions can be related to the activity series. For example:
  - the metals from **K** to **Pb** react with dilute acids releasing hydrogen gas
  - the metals from **K** to **Mg** react with liquid water
  - the metals from **Al** to **Ni** require water to be in the form of steam before reacting.

(information taken from HSC Online Chemistry website)

Practice Problems

1. Write a balanced net ionic equation for the reaction of zinc with aqueous iron(II) chloride. Include the physical states of the reactants and products.
2. Write a balanced net ionic equation for each reaction, including physical states.
   (a) magnesium with aqueous aluminum sulfate
   (b) a solution of silver nitrate with metallic cadmium
3. Identify the reactant oxidized and the reactant reduced in each reaction in question 2.
4. Identify the oxidizing agent and the reducing agent in each reaction in question 2.

Practice Problems

5. Write balanced half-reactions from the net ionic equation for the reaction between solid aluminum and aqueous iron(III) sulfate. The sulfate ions are spectator ions, and are not included.
   $Al_{(s)} + Fe^{3+}_{(aq)} \rightarrow Al^{3+}_{(aq)} + Fe_{(s)}$
6. Write balanced half-reactions from the following net ionic equations.
   (a) $Fe_{(s)} + Cu^{2+}_{(aq)} \rightarrow Fe^{2+}_{(aq)} + Cu_{(s)}$
   (b) $Cd_{(s)} + 2Ag^{+}_{(aq)} \rightarrow Cd^{2+}_{(aq)} + 2Ag_{(s)}$
7. Write balanced half-reactions for each of the following reactions.
   (a) $Sn_{(s)} + PbCl_{2(aq)} \rightarrow SnCl_{2(aq)} + Pb_{(s)}$
   (b) $Au(NO_3)_{3(aq)} + 3Ag_{(s)} \rightarrow 3AgNO_{3(aq)} + Au_{(s)}$
   (c) $3Zn_{(s)} + Fe_2(SO_4)_{3(aq)} \rightarrow 3ZnSO_{4(aq)} + 2Fe_{(s)}$
8. Write the net ionic equation and the half-reactions for the disproportionation of mercury(I) ions in aqueous solution to give liquid mercury and aqueous mercury(II) ions. Assume that mercury(I) ions exist in solution as Hg_2^{2+}.