AQA Chemistry A-level

3.1.11: Electrode Potentials and Cells

Detailed Notes

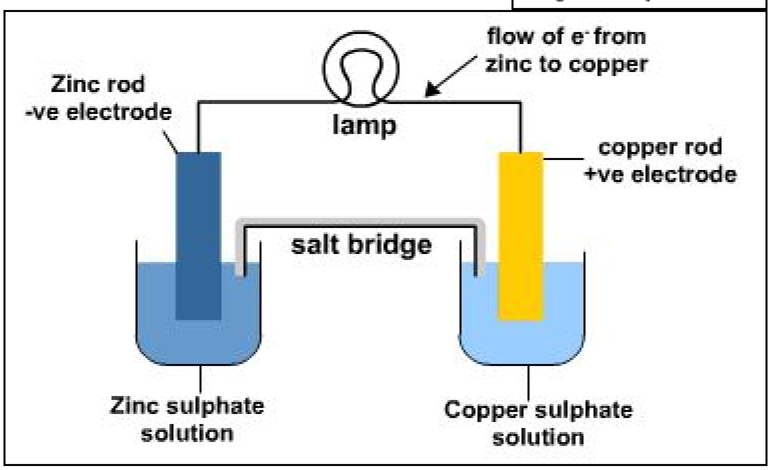
3.1.11.1 - Electrode Potentials and Cells

Electrochemical cells use **redox reactions** as the **electron transfer** between products creates a flow of electrons. This flow of charged particles is an **electrical current** which flows between **electrodes** in the cell. A **potential difference** is produced between the two electrodes which can be measured.

Electrochemical Cells

Most electrochemical cells consist of **two solutions with metal electrodes** and a **salt bridge**. A salt bridge is a tube of **unreactive ions** that can move between the solutions to carry the flow of charge but will not interfere with the reaction.

Image courtesy of S-cool



Each solution is a **half-cell** which makes up the full chemical cell. These half-cells have a **cell potential** which indicates how it will react, either as an oxidation or reduction reaction.

Conventional Cell Representation

Cells are represented in a simplified way so that they don't have to be drawn out each time. This representation has **specific rules** to help show the reactions that occur:

- The half-cell with the **most negative** potential goes on the **left**.
- The **most oxidised** species from each half-cell goes **next to the salt bridge**.
- A salt bridge is shown using a **double line**.
- Always include **state symbols**.

Example:

$$Zn_{(s)} \longrightarrow Zn^{2+}_{(aq)} + 2e^-$$

$$Cu^{2+}_{(aq)} + 2e^- \longrightarrow Cu_{(s)}$$

$$\boxed{Zn_{(s)} \mid Zn^{2+}_{(aq)} \parallel Cu^{2+}_{(aq)} \mid Cu_{(s)}}$$

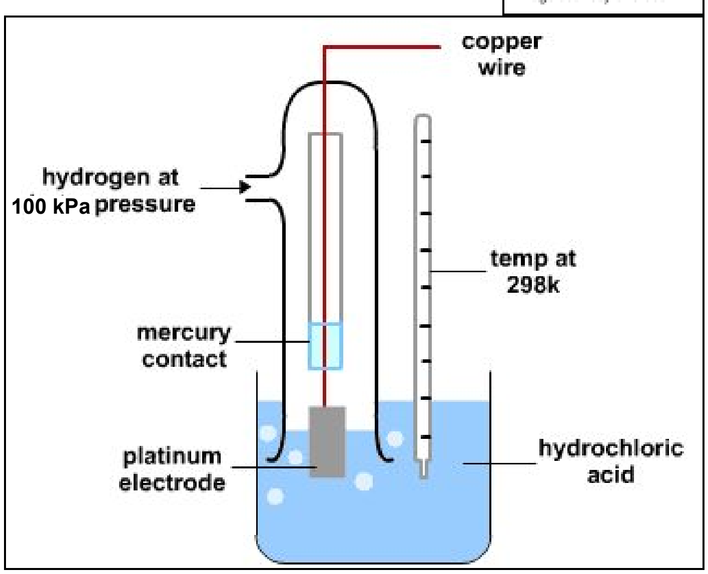
Standard Hydrogen Electrode (SHE)

The standard hydrogen electrode is the **measuring standard** for half-cell potentials. It has a cell potential of **0.00V**, measured under **standard conditions**. These conditions are:

- Solutions of **1.0 moldm^{-3}** concentration
- A temperature of **298K**
- **100 kPa** pressure

The cell consists of **Hydrochloric acid, Hydrogen gas** and uses **Platinum electrodes**. These are very useful as they are **metallic**, so will conduct electricity, but are also **inert** so will not interfere with the reaction.

Image courtesy of S-cool



Cell Potentials

If measured under **standard conditions**, cell potentials are measured compared to the **SHE** to give a numerical value for the half-cell potential.

Negative potentials mean the substances are more easily **oxidised** and will **lose electrons**. **Positive** potentials mean the substances are more easily **reduced** and will **gain electrons** to become more stable.

Calculating Cell EMF (electromotive force)

Standard cell potential values are used to calculate the **overall cell EMF**. This is always done as **potential of the right of the cell minus the potential of the left** of the cell when looking at the cell representation.

$$\boxed{Emf_{(cell)} = E^{o}_{(right)} - E^{o}_{(left)}}$$

It can also be remembered as the **most positive potential minus the most negative potential**.

If the overall cell potential is a **positive** value, the reaction taking place is **spontaneous and favourable**. The more positive the potential, the more favourable the reaction.

Cell Reactions (Anticlockwise rule)

In a similar way to redox reactions, half-cell reactions can be **combined** to give the overall cell reaction. The **'anti-clockwise rule'** is a good method for ensuring the reaction is formed correctly.

1. Write the **most negative** EMF out of the pair on **top**.
2. Draw **anticlockwise arrows** around the reactions.
3. **Balance** the electrons on both sides of the reaction.
4. Write out the cell reaction.

Example:

$$Ni^{2+}_{(aq)} + 2e^- \longrightarrow Ni_{(s)} \qquad E^o = -0.25V$$

$$Cu^{2+}_{(aq)} + 2e^- \longrightarrow Cu_{(s)} \qquad E^o = +0.34V$$

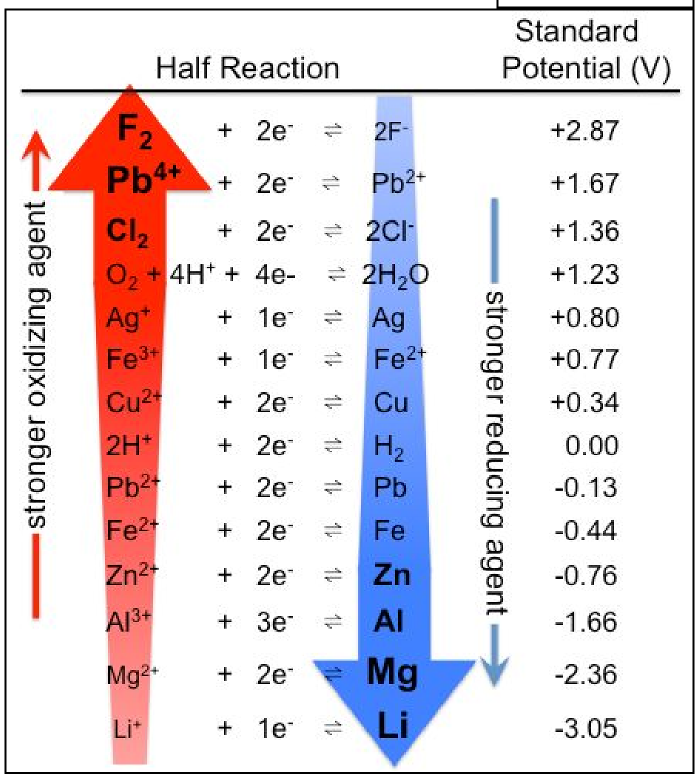
$$Ni_{(s)} + Cu^{2+}_{(aq)} \longrightarrow Cu_{(s)} + Ni^{2+}_{(aq)}$$

Oxidising and Reducing Agents

Electrode potentials that are very **positive** are better **oxidising agents** and will oxidise those species more negative than it.

Species that are very **negative** are better **reducing agents** and will reduce those less negative than it.

Image courtesy of Quora



Effects of Concentration and Pressure

Increasing the concentration of the solutions used in the electrochemical cell makes the cell EMF more **positive** as fewer electrons are produced in the reaction.

Increasing the pressure of the cell will make the cell EMF more **negative** as more electrons are produced.

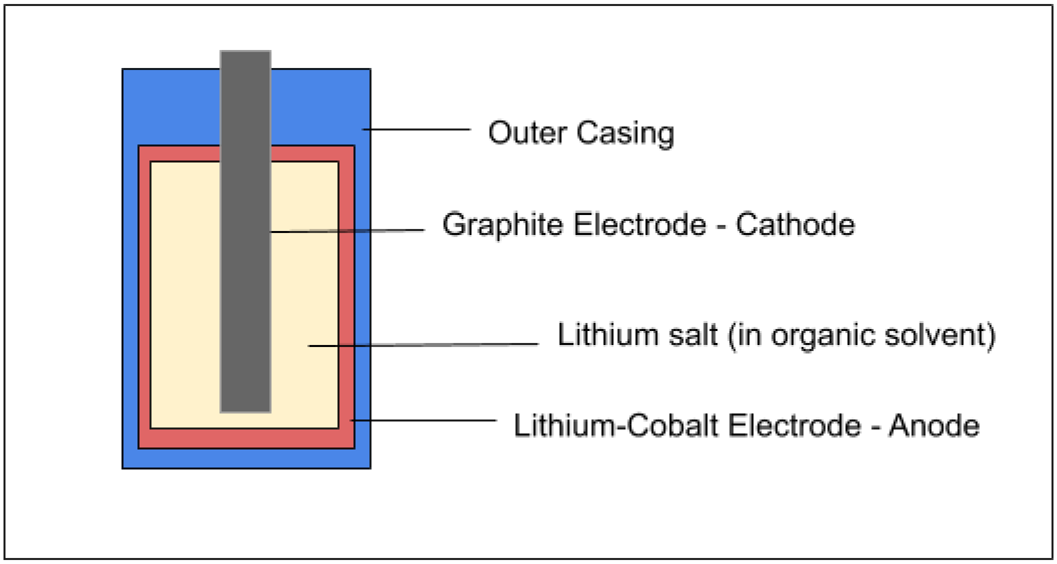
3.1.11.2 - Commercial Cells

Electrochemical cells can be a useful **source of energy for commercial use**. They can be produced to be **non-rechargeable, rechargeable or fuel cells**.

Rechargeable Cells

The reaction that takes place within a rechargeable cell is a **reversible reaction** meaning the reactants can reform. Therefore the cell can be 'reformed' meaning it is a rechargeable cell.

Lithium ion cells are commonly used as rechargeable batteries in phones, laptops and cars. They consist of a **Lithium Cobalt Oxide electrode** and a **Graphite (Carbon) electrode**. An electrolyte of a **Lithium salt** in an organic solvent is used to carry the flow of charge.



The half-cell equations for the equations can be **combined** to give the full cell equation:

Negative Electrode: $Li \longleftrightarrow Li^+ + e^-$

Positive Electrode: $e^- + Li^+ + CoO_2 \longleftrightarrow Li^+[CoO_2]^-$

$$Li + Li^+ + CoO_2 \longleftrightarrow Li^+ + Li^+[CoO_2]^-$$

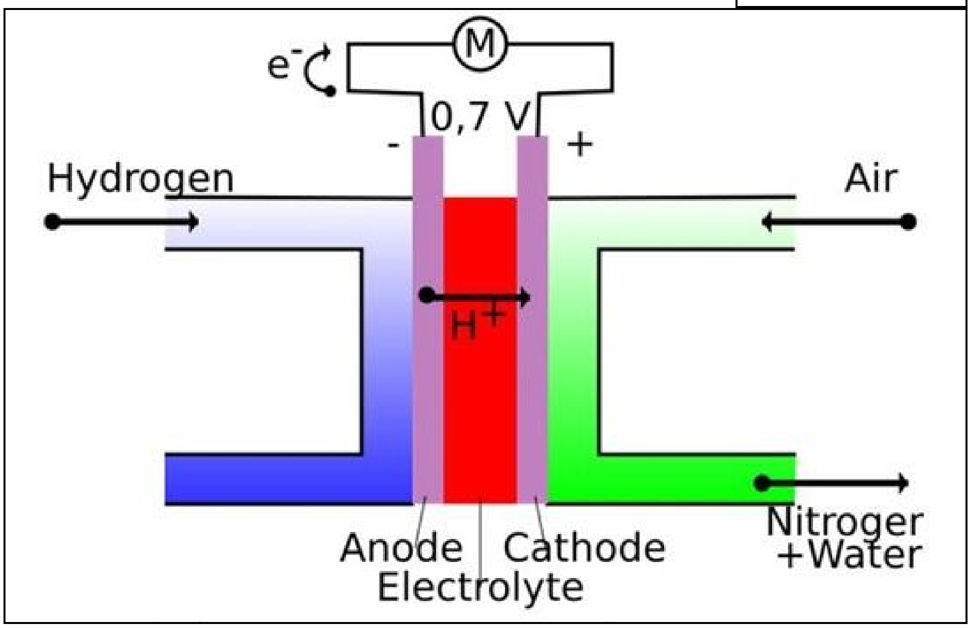
In order to be recharged, a **current has to be applied** over the cell which forces electrons to move in the opposite direction. This causes the reaction to reverse, recharging the cell.

Non-rechargeable cells are not able to do this as the reactions used are **impossible to reverse**.

Fuel Cells

This type of electrochemical cell is used to generate an electrical current without needing to be recharged. The most common type of fuel cell is the **Hydrogen fuel cell**, which uses a **continuous supply** of Hydrogen and Oxygen from air to generate a **continuous current**.

Image courtesy of wiremea.com



The reaction that takes place produces **water** as the only waste product, meaning the Hydrogen fuel cell is seen as being much more **environmentally friendly**.

Image courtesy of In Depth Tutorials and Information

Anode : $H_2(g) + O^{2-} = H_2O(g) + 2e^-$

Cathode : $2e^- + \frac{1}{2}O_2(g) = O^{2-}$

and the overall cell reaction becomes

$$H_2(g) + \frac{1}{2}O_2(g) = H_2O(g).$$

The downsides to Hydrogen fuel cells include the **high flammability of Hydrogen** and that they are **expensive to produce** meaning they are not yet used too commonly.