

AQA Chemistry A-Level

3.1.11: Electrode Potentials and Cells Detailed Notes



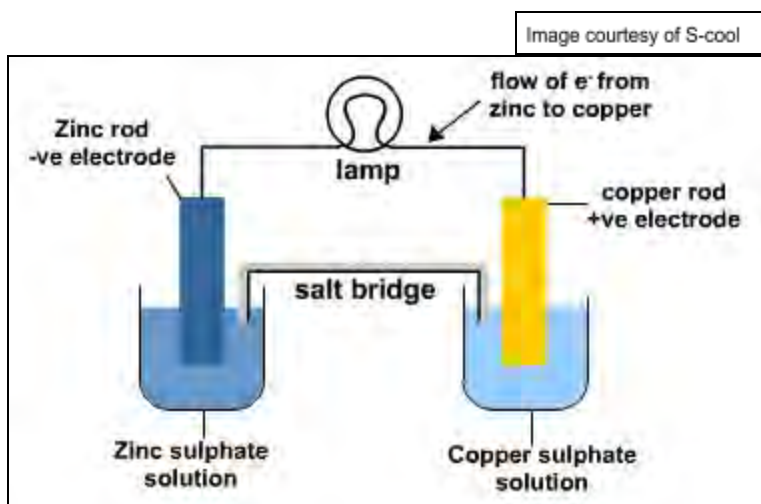


3.1.11.1 - Electrode Potentials and Cells

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.



Each solution is a **half-cell** which make 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.

Standard Cell Representation

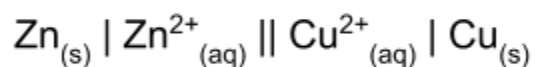
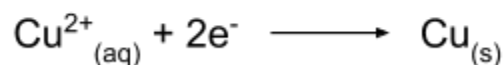
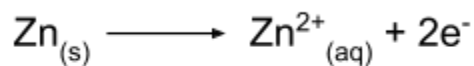
Cells are represented in a simplified way so that they don't have to be drawn out each time. This representation has **specific rules** to help show the reactions that occur:

- 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:

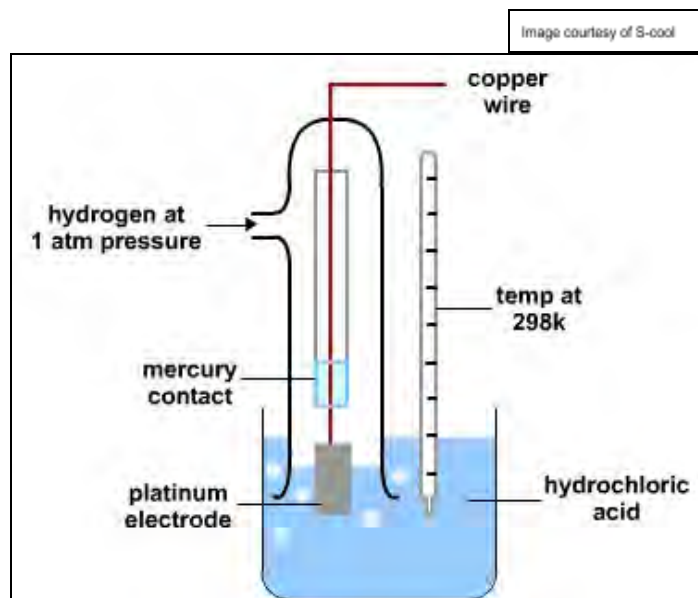


Standard Hydrogen Electrode (SHE)

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 mol dm⁻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.





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.

Positive potentials mean the substances are more easily **oxidised** and will **lose electrons**.

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

Calculating Cell Emf

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.

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

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

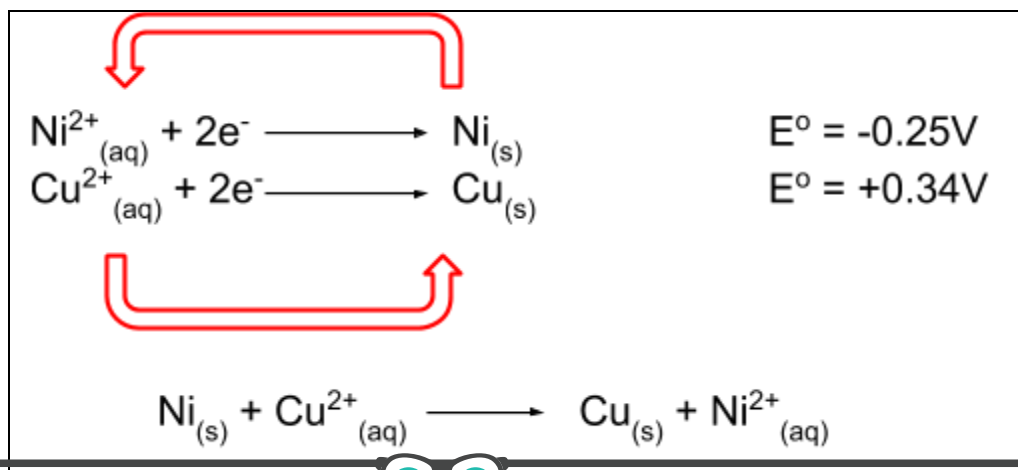
If the overall cell potential is a **positive** value, the reaction taking place is **spontaneous and favourable**. The more positive the potential, the more favourable the reaction.

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:





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

Half Reaction	Standard Potential (V)
$F_2 + 2e^- \rightleftharpoons 2F^-$	+2.87
$Pb^{4+} + 2e^- \rightleftharpoons Pb^{2+}$	+1.67
$Cl_2 + 2e^- \rightleftharpoons 2Cl^-$	+1.36
$O_2 + 4H^+ + 4e^- \rightleftharpoons 2H_2O$	+1.23
$Ag^+ + 1e^- \rightleftharpoons Ag$	+0.80
$Fe^{3+} + 1e^- \rightleftharpoons Fe^{2+}$	+0.77
$Cu^{2+} + 2e^- \rightleftharpoons Cu$	+0.34
$2H^+ + 2e^- \rightleftharpoons H_2$	0.00
$Pb^{2+} + 2e^- \rightleftharpoons Pb$	-0.13
$Fe^{2+} + 2e^- \rightleftharpoons Fe$	-0.44
$Zn^{2+} + 2e^- \rightleftharpoons Zn$	-0.76
$Al^{3+} + 3e^- \rightleftharpoons Al$	-1.66
$Mg^{2+} + 2e^- \rightleftharpoons Mg$	-2.36
$Li^+ + 1e^- \rightleftharpoons Li$	-3.05

↑ stronger oxidizing agent (red arrow pointing up)
↓ stronger reducing agent (blue arrow pointing down)

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 electron 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**.

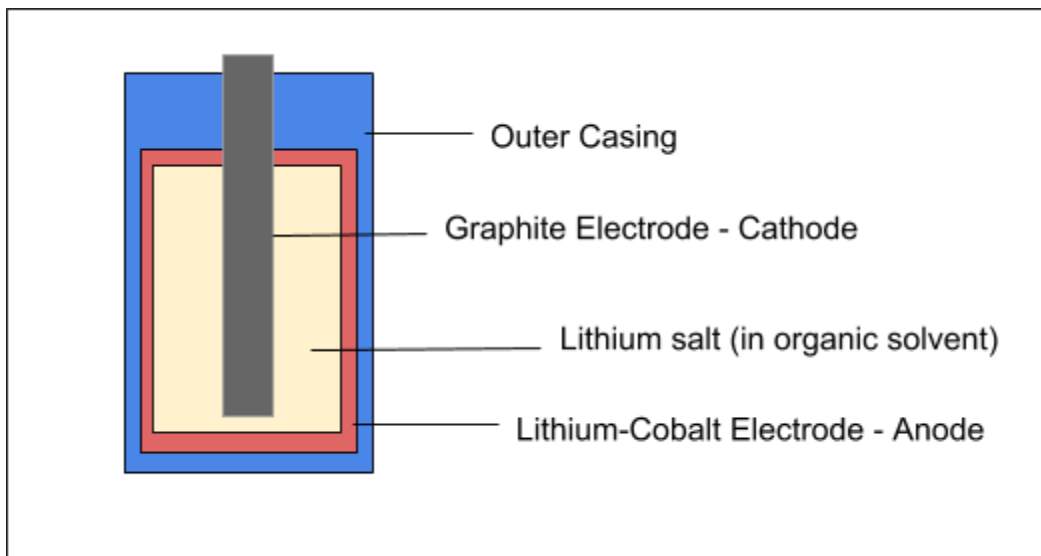
Rechargeable Cells

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.

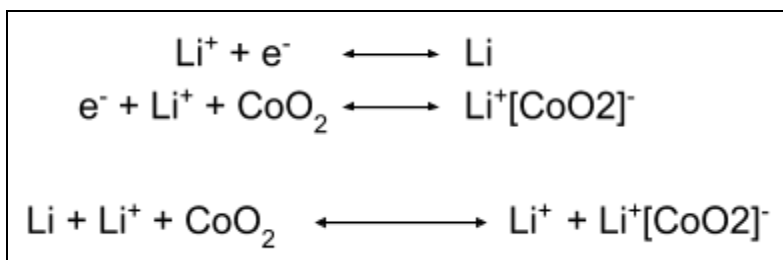




Lithium ion cells are commonly used as rechargeable batteries in phones, laptops and cars. They consist of a **lithium cobalt oxide electrode** and a **graphite (carbon) electrode**. An electrolyte of a **lithium salt** in an organic solvent is used to carry the flow of charge.



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



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.

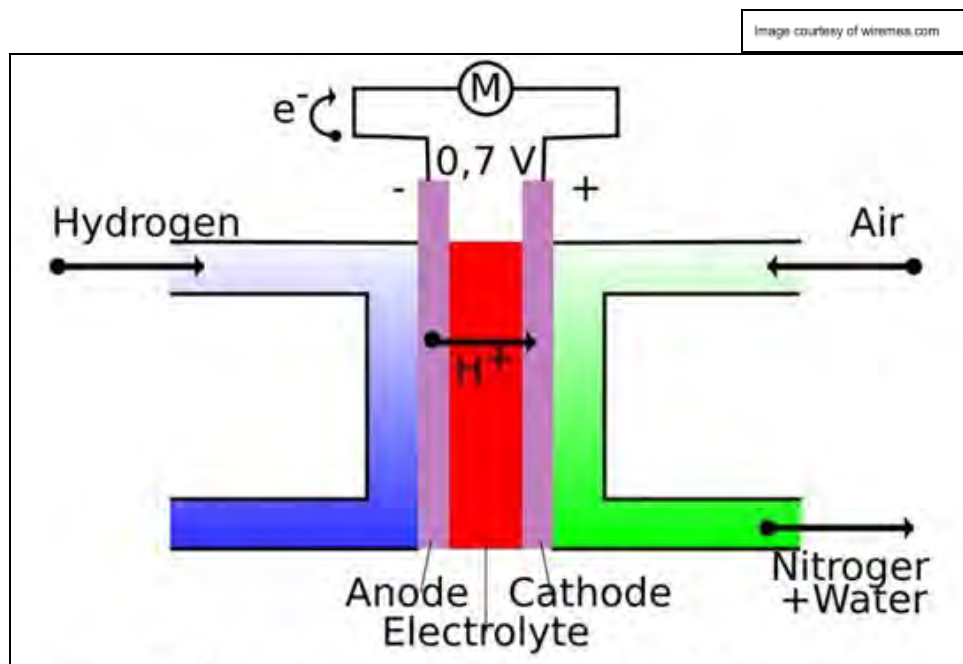
Non-rechargeable cells are not able to do this as the reactions used are **impossible to reverse**.





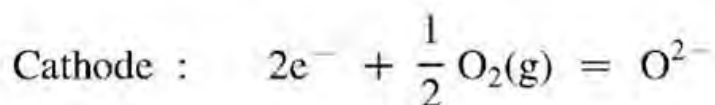
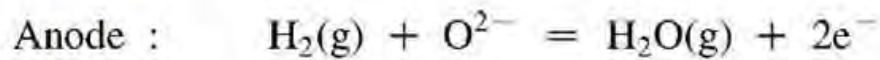
Fuel Cells

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 the air to generate a **continuous current**.

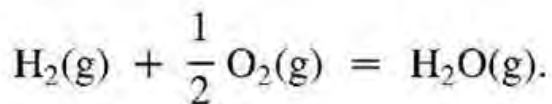


The reaction that takes place produces **water** as the only waste product, meaning the hydrogen fuel cell is seen as being much more **environmentally friendly**.

Image courtesy of In Depth Tutorials and Information



and the overall cell reaction becomes



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.

