Fuel Cells

Electrode potentials part 1
Electrode potentials part 2

Fuel cells are simply real-life applications of the redox reactions that you have been learning i.e. they produce electricity via a redox reaction.

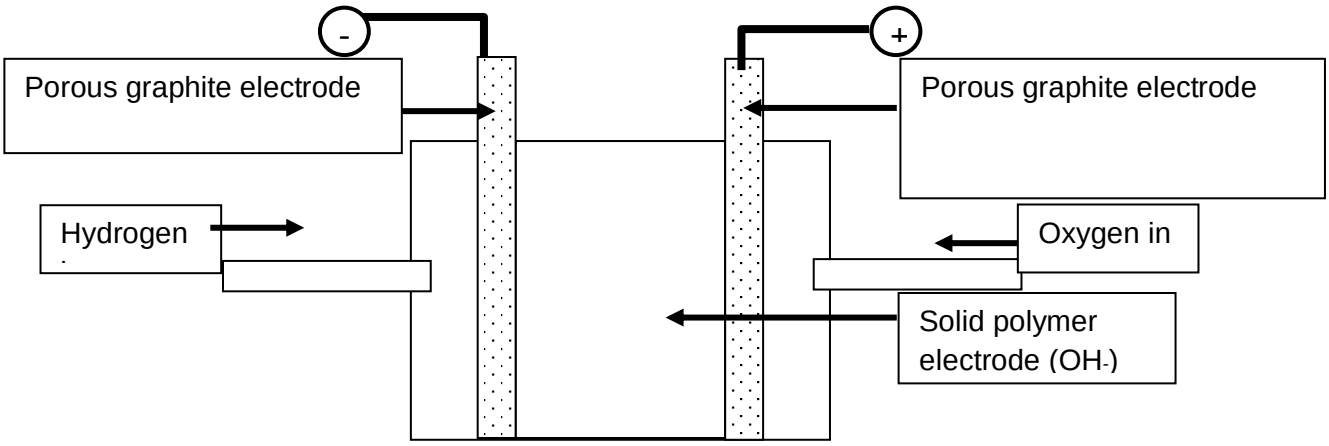
Typical examples of fuels used are hydrogen, ethanol and methanol, which are constantly supplied to the cell.

The cell consists of two electrodes and a membrane that allows ions to pass between the electrodes and an electrolyte.

Hydrogen-Oxygen Fuel Cell

Hydrogen (the fuel) and oxygen are added to the cell. A redox reaction occurs to **produce water** and electricity. These reactions can take place under **acidic or alkaline** conditions.

✓ The specification ONLY mentions alkaline fuel cells (used commercially) so you can ignore anything to do with acidic conditions.



The alkali is the electrolyte to allow the passage of ions between the two electrodes.

The electrodes used are made from porous platinum rather than platinum rods as they will have a higher surface area.

✓ just think of this as a normal cell with the two beakers, electrodes and salt bridge. It just looks a bit different but has all the same parts.

Half-Equations

Below are the half and full equation for the reactions that occur under alkaline conditions:

Alkaline

Anode: $H_2 + 2OH^- \rightarrow 2H_2O + 2e^-$ Cathode: $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$

Overall: $2H_2 + O_2 \rightarrow 2H_2O$

I would recommend knowing these half-equations or at least have a method to work them out. They often ask for them in exams.

✓ Alkaline half-equations ones are more difficult to work out as you are used to adding H_2O and H^+ to balance half equations. Click the link for more on balancing alkaline half-equations.

Advantages

- the hydrogen fuel cell produces **only water**, whereas fossil fuels produce **CO_2** gas.
- the hydrogen fuel cell is **more efficient** than directly combusting hydrogen to make water.
- the electricity is released in a **more controlled** manner from the hydrogen fuel cell and minimizes heat loss.

Disadvantages

Hydrogen fuel cell versus a rechargeable battery:

- hydrogen is difficult to transport
- hydrogen is highly flammable
- fuel cell needs a constant supply of hydrogen
- hydrogen is made from burning fossil fuels which is not environmentally friendly
- hydrogen is not as readily available as fossil fuels are non-renewable
- hydrogen fuel cell is more expensive

✓ I don't think you need much of this advantage/disadvantage stuff. It's more of an old spec thing. I just listed a few incase they throw in a sneaky 2 marker.

Rechargeable and Non-rechargeable cells

You don't need to know too much about this or memorise the types of chemicals used in these batteries (cells). It's more about applying the redox knowledge from this topic and working out whether a reaction is spontaneous.

Non-rechargeable

Should be self-explanatory! It cannot be recharged and is therefore irreversible. So this is just a normal redox reaction.

Rechargeable cells

They often use nickel-cadmium cells and lead-acid cells.

The "discharge" reaction is just the normal forward redox reaction i.e. **reverse the more negative** equation → a positive EMF value.

To "recharge" the cell, you need to do the opposite and **reverse the more positive** half equation. Don't worry why but recharging involves providing electricity to reverse the reaction. Very different from everything else in this topic!

Example

They have listed the lithium cell in the specification to demonstrate a rechargeable cell but there are plenty of others they could use in exams. The principles are the same.

Positive electrode

$$Li^+ + CoO_2 + e^- \rightarrow Li^+[CoO_2]^-$$

✓ it might not be obvious at first but the Co has gone from 4+ → 3+

Negative electrode

$$Li \rightarrow Li^+ + e^-$$

Overall Reaction

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

✓ Remember this is the "normal" forward reaction (discharge). For recharging, just reverse the overall equation.

✓ These half-equations are specifically mentioned in the specification.

Conventional Cell Representation

$$Li \mid Li^+ \mid\mid Li^+, CoO_2 \mid LiCoO_2 \mid Pt$$

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