

APPLICATIONS of ELECTROCHEMICAL CELLS

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ELECTROCHEMICAL CELLS

Electrochemical cells produce electrical energy from redox reactions and are used in many commercial applications such as mobile phones and electric vehicles. There are three types of electrochemical cell:

1. **Irreversible (non-rechargeable) cells** in which the chemical reaction can go only one way. When the reactants are used up and the reaction is over, the cell must be disposed of and cannot be used again.
2. **Rechargeable cells** involve reversible chemical reactions. When the reactants have all turned into products, the reaction can be made to go in reverse by passing an electric current through the cell. In this way the cell can be used many times.
3. In **fuel cells**, reactants are constantly flowing into the cell (and products flowing out). As long as there is a constant supply of reactants (the fuel), the cell will continuously generate an electric current.

LITHIUM CELLS

An example of a rechargeable cell is the lithium cell, in which a positive electrode, made of lithium cobalt oxide (LiCoO_2), and a negative electrode, made of carbon, are separated by a thin permeable polymer layer. When it is charged, an electric current is forced to flow from the positive electrode to the negative electrode, and lithium ions move from the positive electrode through the polymer to the negative electrode.

The half-equations at each electrode during charging are:

- Positive electrode: $\text{Li}^+[\text{CoO}_2]^- \rightarrow \text{Li}^+ + \text{CoO}_2 + \text{e}^-$
- Negative electrode: $\text{Li}^+ + \text{e}^- \rightarrow \text{Li}$

As the cell discharges, it creates an electric current with electrons flowing from the negative electrode to the positive electrode, and lithium ions moving from the negative electrode to the positive electrode. In this way the charge remains balanced (the charges of the negative electrons and positive lithium ions cancel).

It is the flow of electrons that powers devices such as laptops when they are running on a battery.

When the cell is discharging, the electrode equations are reversed compared to the ones given above (which are correct when the cell is being charged).

The electrode potentials for the lithium cell are:

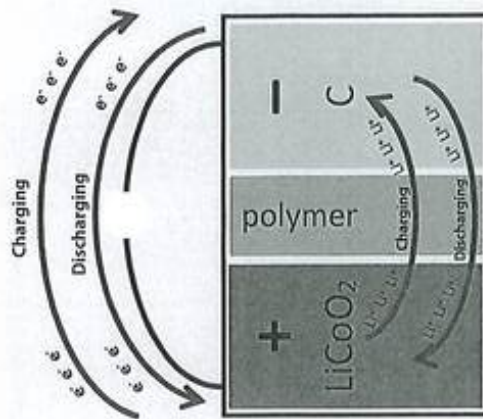
- $\text{Li}^+ + \text{CoO}_2 + \text{e}^- \rightleftharpoons \text{Li}^+[\text{CoO}_2]^-$ $E^\ominus = +1.42 \text{ V}$
- $\text{Li}^+ + \text{e}^- \rightleftharpoons \text{Li}$ $E^\ominus = -3.05 \text{ V}$

The EMF of the lithium cell can be calculated using $E^\ominus_{\text{cell}} = E^\ominus_{\text{reduced}} - E^\ominus_{\text{oxidised}}$. When discharging, Li is oxidised to Li^+ , and CoO_2 is reduced to $[\text{CoO}_2]^-$, and so the calculation is $E^\ominus_{\text{cell}} = 1.42 - (-3.05) = 4.47 \text{ V}$.

A higher voltage can be achieved by connecting several cells together in a battery.

EXAM TIP

You may be given some values for standard electrode potentials and asked to use them to calculate the EMF of a cell



Remember electrode potentials are always given for reductions

FUEL CELLS

An example of a **fuel cell** is the alkaline hydrogen-oxygen fuel cell. In this cell, hydrogen flows into the negative electrode and oxygen to the positive electrode. The half-equations occurring at each electrode are:

- Positive electrode: $\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^- \rightarrow 4\text{OH}^-$
- Negative electrode: $2\text{H}_2 + 4\text{OH}^- \rightarrow 4\text{H}_2\text{O} + 4\text{e}^-$
- Overall reaction: $\text{O}_2 + 2\text{H}_2 \rightarrow 2\text{H}_2\text{O}$

Fuel cells like this have many advantages:

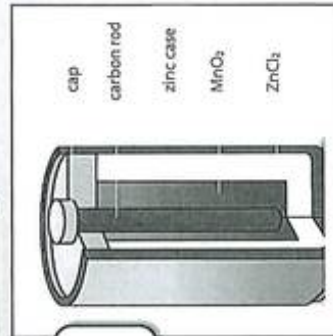
- The only product is water. Unlike burning many other fuels, this fuel cell produces no toxic chemicals or greenhouse gases.
- It does not need to be recharged. The fuel needs to be replaced when it is used up, but this is much faster than recharging a lithium cell.
- There are some disadvantages, however:
- The production of the hydrogen fuel (unless it is done using electricity from a source such as solar or wind power) can cause emission of environmentally damaging gases.
- Hydrogen is a very dangerous gas as it is highly flammable and can even be explosive. It must be stored at high pressure (otherwise it would take up an enormous amount of space), which makes it even more dangerous.

Common non-rechargeable cells are zinc-carbon cells. One type of these is shown in this diagram.

You can use the electrode potentials to work out the reaction that produced the electric current.

- $\text{Zn}^{2+} + 2\text{e}^- \rightleftharpoons \text{Zn}$ $E^\ominus = -0.76 \text{ V}$
- $2\text{MnO}_2 + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{Mn}_2\text{O}_3 + \text{H}_2\text{O}$ $E^\ominus = +0.95 \text{ V}$

Since Zn has the more negative (smaller) value for $E^\ominus$, this will be oxidised. Combining equations gives:



Electrode potentials in real cells will be slightly different as the conditions are not standard.

EXAM TIP

You may be given electrode half-equations and electrode potential values and be expected to work out the reaction that occurs when electricity is produced. This could be for any rechargeable or non-rechargeable cell.