Fuel Cells

The principle of fuel cells

Oxygen and hydrogen, when mixed together in the presence of enough activation energy have a natural tendency to react and form water, because the Gibbs free energy of H_2O is smaller than that of the sum of H_2 and $½O_2$ added together (Hence, we don't smoke our pipes on Zeppelins!). If hydrogen and oxygen were combined directly, we would see combustion:

$H_2 + ½O_2 \leftrightarrow H_2O$

Combustion involves the direct reaction of H_2 gas with O_2. The hydrogen donates electrons to the oxygen . We say that the oxygen has been reduced and the fuel oxidised. This combustion reaction releases heat energy.

The fuel cell separates hydrogen and oxygen with a gas-impermeable electrolyte through which only ions (e.g. H^+, O^{2-}, CO_3^{2-}) can migrate. Hence two half reactions occur at the two electrodes. The type of reactions at the electrodes is determined by the type of electrolyte.

Grove's fuel cell is one of the simplest examples.

The half-reaction at the anode:

$H_2 \leftrightarrow 2H^+ + 2e^-$

The half-reaction at the cathode:

$O_2 + 4e^- + 4H^+ \leftrightarrow 2H_2O$

The net reaction is the combustion reaction:

$H_2 + ½O_2 \leftrightarrow H_2O$

Activation polarization is caused by the energy intensive activity of the making and breaking of chemical bonds: At the anode, the hydrogen molecules enter the reaction sites so that they are broken into ions and electrons. The resulting ions form bonds with the catalyst atoms and the electrons remain in the vicinity until new hydrogen molecules start bonding with the catalyst, breaking the bond between the earlier ion. The electrons migrate through the bipolar plate if the bonding energy of the ion is low enough and the ions diffuse through the electrolyte. A similar process occurs at the cathode: Oxygen molecules are broken up and react with the electrons from the anode and the protons that diffused through the electrolyte to form water. Water is then ejected as a waste product and the fuel cell runs (can supply a current), as long as fuel and oxygen is provided.

The exact reactions at the electrodes depend upon which species can be transported across the electrolyte. Fuel cells are classified according to the type of electrolyte (see Types of Fuel Cells). The most common electrolytes are permeable for protons and the reactions are as discussed above. The second most common electrolytes, found in solid oxide fuel cells (SOFCs), are permeable for oxide ions and the following half-reactions occur:

The half-reaction at the anode:

$H_2 + O^{2-} \leftrightarrow H_2O + 2e^-$

The half-reaction at the cathode:

$O_2 + 4e^- \leftrightarrow 2O^{2-}$

The net reaction is the same as before:

$H_2 + \frac{1}{2}O_2 \leftrightarrow H_2O$

A third type of electrolyte, used for molten carbonate fuel cells at high temperatures conducts carbide ions (CO_3^{2-}):

The half-reaction at the anode:

$H_2 + CO_3^{2-} \leftrightarrow H_2O + CO_2 + 2e^-$

The half-reaction at the cathode:

$½O_2 + CO_2 + 2e^- \leftrightarrow CO_3^{2-}$

The net reaction is the combustion reaction:

$H_2 + ½O_2 \leftrightarrow H_2O$

We also commonly see alkaline electrolytes, across which OH^- is the transported species. In this case the half-reactions would be:

The half-reaction at the anode:

$H_2 + 2OH^- \leftrightarrow 2H_2O + 2e^-$

The half-reaction at the cathode:

$O_2 + 4e^- + 2H_2O \leftrightarrow 4OH^-$

The net reaction is the combustion reaction:

$H_2 + ½O_2 \leftrightarrow H_2O$

Building a simple fuel cell

Collect the necessary components:

This is the most challenging part of the project, due to the difficulty many students will have in procuring the two lengths of platinum (Pt) wire required. Persist.

- Two short lengths of platinum wire – 2 cm a piece is sufficient
- A beaker
- Weak NaCl solution (table salt and tap water will suffice)

- Crocodile clips ×2
- Wires ×2
- Voltmeter
- DC power source, e.g. 6-volt battery

Step by step:

1. Fill beaker with salt solution
2. Clamp the ends of the Pt wires in the crocodile clips
3. Place the Pt wires in the solution in such a way that they will not move and so that only the Pt is in contact with the water. Plasticine (or Blu-Tac) would be helpful at this point. If the crocodile clips are in contact with the solution they will corrode quickly and taint the water.
4. Verify that the voltage across the two electrodes is zero (two wires made of the same metal have no galvanic difference between them and hence no voltage will be seen).
5. Connect each Pt wire to one of the battery's terminals and watch the gas bubbles from the electrodes.
6. Verify that the voltage across the electrodes is now equal to that of the battery, i.e. 6 V.
7. Disconnect the battery and be careful not to knock the beaker.
8. Verify that a voltage remains across the two electrodes.

This remaining voltage is due to bubbles of chlorine gas adhering to the cathode and to bubbles of hydrogen gas adhering to the anode. HCl is produced as the two atoms combine, via H^+ ions travelling through the solution from anode to cathode, to produce HCl.

Things to consider:

An ammeter will confirm that we only obtain a tiny current in the above experiment (although the voltage remains set to the value we can calculate using the Nernst equation). Factors limiting the current in this case are:

- the small surface area of the electrode - small amount of gas
- wide spacing of the electrodes and electrolyte's ionic resistance.

In production cells, the electrodes are made larger and flat so as to maximise contact area, and the thickness of the electrolyte is kept to a minimum so as to reduce resistance.

Source : http://www.doitpoms.ac.uk/tlplib/fuel-cells/principle.php