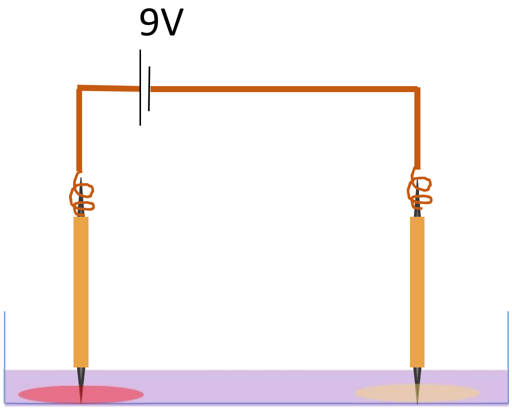
Advanced Case Study: Water Electrolysis in acidic solution



The hydrogen economy is an option for removing fossil fuels from transport and energy storage. One way of obtaining hydrogen is through splitting water. You have done this yourself with the pencils. Why isn't can't you just scale up that method? What would happen on a massive scale?

The half equations are:
Reduction reaction (water is :
Oxidation reaction (water is being stripped of electrons at one electrode):
Anode: $2H_2O_{(l)} \rightleftharpoons O_{2(g)} + 4e^- + 4H^+$ $E^o = +1.23$

Reduction reaction (water has electrons shoved onto it on the other electrode):
Cathode: $4e^- + 4H^{+\,-} \rightleftharpoons 2H_{2(g)}$ $E^o = 0.00$ V

Is this reaction spontaneous? What would you expect? Use the method your teacher has shown you to work out cell potentials (eg. Reduction minus Oxidation)

E_{cell} = (0.00V) – (+1.23 V) = -1.23 V

To work out whether a reaction is spontaneous in electrochemistry, the following equation is used to relate cell potential to Gibb's Free energy. Remember, if Gibbs Free energy is negative, a reaction is spontaneous.

$$\Delta G = -nFE$$

n = number of electrons involved in the reaction (in this case 4)
F = Faraday's constant, a constant that converts between Coulombs and Moles)
This reaction is non-spontaneous, which is what we expected as we need to provide a power source for this reaction.

In a fuel cell, the reverse of the above reactions are taking place, so flip the signs:

E_{cell} = (+0.83 V) – (-0.40 V) = 1.23 V

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