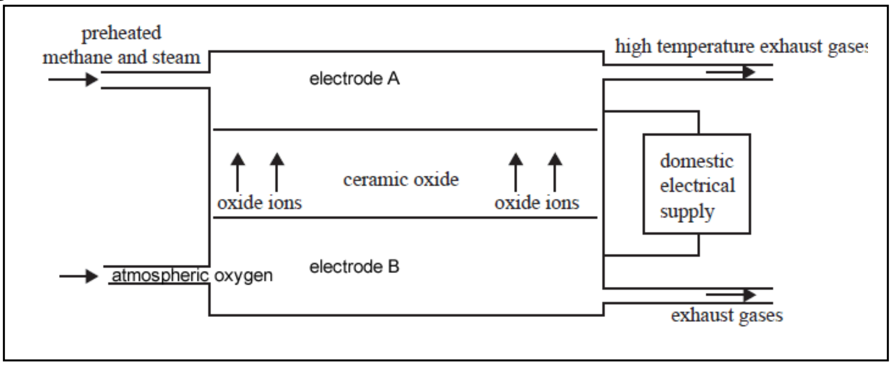
Redox reactions – Revision galvanic cells and fuel cells

Lesson 7

Revise fuel cells by visiting the link below.

www.dynamicscience.com.au/tester/solutions1/chemistry/redox/fuelcl.html

1) A fuel cell uses a solid oxide electrolyte to generate electrical energy, as shown in the diagram below.



Combustion of methane drives the fuel cell. One of the half equations is given below.

$$CH_4(g) + 4O^{-2}(g) \rightarrow CO_2(g) + 2H_2O(g) + 8e$$

a) At which electrode does the given half reaction, above, take place?
Oxidation takes place at the anode. This will take place at electrode A

b) Give the other half equation *$O_2(g) + 4e \rightarrow 2O^{2-}(g)$*

c) Give the overall equation *$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$*

d) Label the anode and cathode *Electrode A is the anode, Electrode B the cathode*

e) Label the direction of electron flow. *From anode to cathode*

f) Although a fuel cell is a galvanic cell it differs markedly from other galvanic cells. Compare fuel cells with other galvanic cells by labelling the following statements as true or false

i. Fuel cells can be recharged in a similar way to secondary cells. *False*
Products are constantly removed.

ii. Electrodes used in primary cells and secondary cells are similar to the electrodes used in fuel cells. *False*
Electrodes in a fuel cell act as catalysts and are porous.

iii. Fuel cells and all other galvanic cells transform chemical energy into electrical energy *True*

iv. Oxidation occurs at the anode of fuel cell, primary and secondary cells. *True*

v. Fuel cells deliver a constant voltage during their operation as compared to other galvanic cells which reduce in voltage as they discharge *True*

vi. The products of all galvanic cells, including fuel cells, must remain in contact with the electrodes so they can be recharged. *False*
Products are constantly removed.

vii. The anode in fuel cell is positive whereas the anode in other galvanic cell is negative *False*

Like all galvanic cells, the anode is negative and where oxidation takes place.

viii. Electrodes in fuel cells act as catalysts for the oxidation and reduction reactions, whereas electrodes in other galvanic cells do not. *True*

ix. Fuel cells represent a cheap alternative to the supply of electrical energy *False*

Fuel cells are expensive.

g) Assuming this fuel cell is 75.0% efficient in converting chemical energy into electrical energy and that methane is supplied at the rate of 44.50 litres per second at a pressure of 1 atm at 25°C, calculate the following, to the right number of significant figures.

i. Mol of CH_4 consumed every second.

n = PV/RT

=> n = (101.3 X 44.5) / (8.31 X 298)

=> n = 1.82 mol

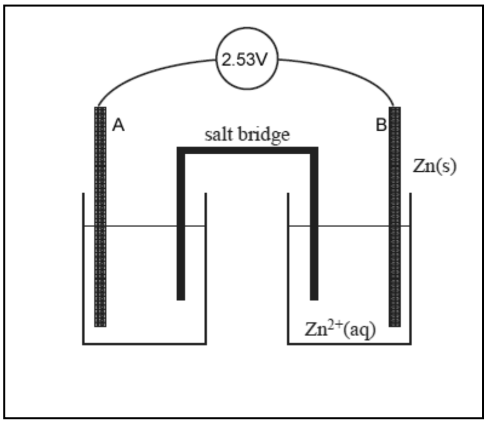
ii. Total, theoretical, heat energy available from the combustion of methane, in kJ, every second.

=> 1.82 X 8.90 X 10^2 = 1.62 X 10^3 kJ

iii. Electrical energy, in kJ, produced every second.

=> 1.62 X 10^3 kJ X 0.75 = 1220 kJ

2) Consider the diagram of a galvanic cell shown on the right operating under standard conditions.



a) What is the half cell on the left composed of?

H_2O_2(aq) in an acidified solution

b) In which direction are electrons flowing?

From electrode B to electrode A

c) What is electrode A composed of?

Carbon or platinum

d) What properties should electrode A have?

Conduct electricity and be inert.

e) Identify the

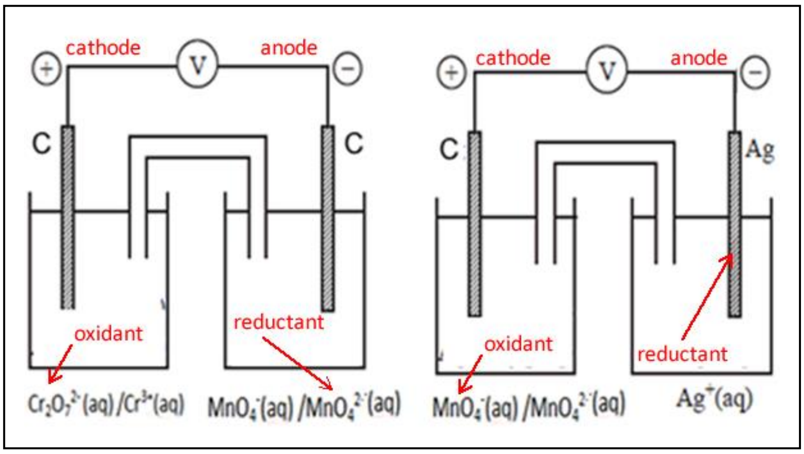
- oxidant *- H_2O_2*
- reductant *– Zinc metal*

f) As the cell discharges label the following

- direction of cation flow *– from the salt bridge towards electrode A*
- direction of anion flow *– from the slat bridge towards electrode B*
- anode *– electrode B*

- cathode *– electrode A*
- polarity of electrodes. *– Electrode B is negative, electrode A is positive.*

3) Consider the two galvanic cell shown below



a) Place the following half equations in the order they would be found on an E° table.

$MnO_4^-(aq) + e^- \rightleftharpoons MnO_4^{2-}(aq)$ ----------------E°

$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$ ----------------E°2

$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightleftharpoons 2Cr^{3+}(aq) + 7\ H_2O(l)$ ----------------E

E°3
E°1
E°2
In order to undergo a spontaneous reaction, the oxidant must be above the reductant on the E° table.

b) The lithium button cell, used to power watches and calculators, is a primary cell containing lithium metal. The lithium ion cell is a secondary cell that is used to power laptop computers.
a. What is the difference between a primary and secondary cell?

Secondary cells can be recharged and primary cells are used once only. Products adhere to the electrodes of a secondary cell whereas the products of a primary cell migrate away from the electrodes.

c) By referring to information provided in the Data Book, give one reason why lithium is used as a reactant in these galvanic cells.
It is a strong and light reductant able to carry more charge per gram than most other reductants.

d) Some early lithium metal batteries exploded when exposed to water. Explain why, using a balanced equation, including states, for the reaction between lithium metal and water.

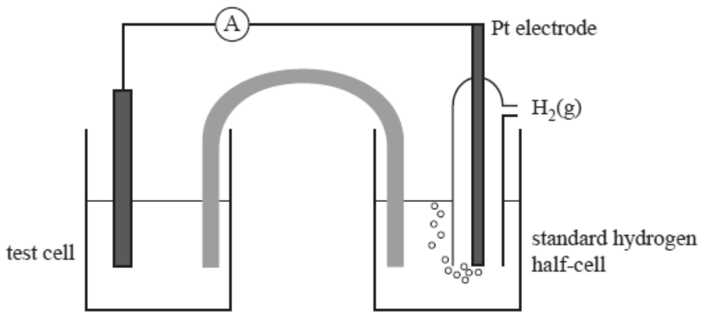
Lithium reacts with water to produce hydrogen gas.

$2Li(s) + 2H_2O(l) \rightarrow H_2(g) + 2LiOH(aq)$

4) In a problem-solving activity a student is given the following information regarding three half-equations. However, although the three numerical values of E^0 are correct, they have been incorrectly assigned to the three half-equations

Half-equation	$\mathbf{E^0}$
$AgCl(s) + e \rightleftharpoons Ag(s) + Cl^-(aq)$	–0.40 V
$Cd^{2+}(aq) + 2e \rightleftharpoons Cd(s)$	–0.36 V
$PbSO_4(s) + 2e \rightleftharpoons Pb(s) + SO_4^{2-}(aq)$	+0.22 V

The objective of this task is to correctly assign the E^0 values to the corresponding half-equation shown on the right. To do this, the student constructs standard half-cells for each of the above half-reactions. These half-cells are connected, one at a time, to a standard hydrogen half-cell as indicated in the diagram below.



The following observations were made either during or after the electrochemical cell discharged electricity for several minutes.

Experiment	**Half-cell reaction being investigated**	**Experimental notes**
1	$AgCl(s) + e \rightleftharpoons Ag(s) + Cl^-(aq)$	Electron flow was detected passing from the standard hydrogen half-cell to the half-cell containing the silver electrode.
2	$Cd^{2+}(aq) + 2e \rightleftharpoons Cd(s)$	The mass of the cadmium electrode decreased.
3	$PbSO_4(s) + 2e \rightleftharpoons Pb(s) + SO_4^{2-}(aq)$	The pH of the solution in the standard hydrogen half-cell increased.

According to the table above we can assign the hal-cells to their relative positions on the Eo table as shown below.

$AgCl(s) + e \Leftrightarrow Ag(s) + Cl^-(aq)$ ---------0.22

$2H^+(aq) + 2e \Leftrightarrow H_2(g)$ --------- 0.00

$PbSO_4(s) + 2e \Leftrightarrow Pb(s) + SO_4^{-2}(aq)$ --------- ?

$Cd^{2+} + 2e \Leftrightarrow Cd(s)$ --------- ?

a) The above information can only be used to assign one of the E^0 values to its corresponding half-equation. Identify this half-equation by placing the correct E^0 value next to its corresponding half-equation in the table on the right.

Half-equation	E^0
$AgCl(s) + e \rightleftharpoons Ag(s) + Cl^-(aq)$	+0.22V
$Cd^{2+}(aq) + 2e \rightleftharpoons Cd(s)$	
$PbSO_4(s) + 2e \rightleftharpoons Pb(s) + SO_4^{2-}(aq)$	

b) Explain why the other two E^0 values cannot be correctly assigned to their half-equations

Although we know they appear below the H^+/H_2 half-cell we do not have any information to assign them to their positions relative to each other.

c) Explain why the pH of the solution in the standard hydrogen half-cell increased in experiment 3.

$2H^+(aq) + 2e \Leftrightarrow H_2(g)$ --------- 0.00

$PbSO_4(s) + 2e \Leftrightarrow Pb(s) + SO_4^{-2}(aq)$ --------- - ?

The overall reaction occurred using up H^+ ions decreasing $[H^+]$ and hence increasing the pH

$PbSO_4(s) + 2H^+(aq) \Leftrightarrow Pb(s) + SO_4^{-2}(aq) + H_2(g)$