Friday Worksheet **Name: ………………**

Fuel cells worksheet 3

1) Indicate if the following statements are true or false for both fuel cells and rechargeable cells. Give an explanation

a) Reactants and products are stored within the cell.

False. Fuel cells have reactants continuously fed and products removed.

b) Reaction products are continuously removed from the cell.

False . Rechargeable cells have the products of their reactions in close proximity to the electrodes so they can be recharged. Reactatns and prodcuts are contained within the rechargeable cell.

c) Electrons pass from the oxidant to the cathode as electricity is produced.

False. At the cathode, positive electrode, electrons pass from the cathode to the strongest oxidant present in both rechargeable cells and fuel cells, not from the oxidant to the cathode as stated.

d) Electrical energy is converted to chemical energy as the cell is recharged.

False. Fuel cells do not get recharged.

2) Methane can be obtained from natural gas deposits **or** as a biochemical fuel from biomass. A methane fuel cell is being considered. The overall, unbalanced reaction is given below

$CH_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$

a) What species is being oxidised?

Carbon in CH_4 has an oxidation state of -4 and in CO_2 has an oxidation state of +4

b) Give the equation for the half-reaction that occurs at the anode.

Oxidation occurs at the anode

=> $CH_4(g) \rightarrow CO_2(g)$

=> balance for oxygen by adding water to the left side

=> $CH_4(g) + 2H_2O(l) \rightarrow CO_2(g)$

=> balance for hydrogens by adding H^+ to the right side

=> $CH_4(g) + 2H_2O(l) \rightarrow CO_2(g) + 8H^+(aq)$

=> balance for charge by adding electron to the most positive side.

=> $CH_4(g) + 2H_2O(l) \rightarrow CO_2(g) + 8H^+(aq) + 8e$

c) Give the equation for the half reaction that occurs at the cathode.

Reduction occurs at the cathode

$O2(g) \rightarrow H_2O(l)$

=> balance for oxygen by adding H_2O on the right side

=> $O_2(g) \rightarrow 2H_2O(l)$

=> balance for hydrogen by adding H^+ on the left side

=> $4H^+(aq) + O_2(g) \rightarrow 2H_2O(l)$

=> balance for charge by adding electrons to the left side

=> $4e + 4H^+(aq) + O_2(g) \rightarrow 2H_2O(l)$

d) Give a balanced overall equation

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$

3) Fuel cells are being developed that use fuels other than hydrogen as their energy sources. Potential fuels include ethane and ethane-1,2-diol. These fuels are reacted with oxygen gas to produce CO_2(g) and H_2O(l) using an acidic electrolyte.

a) Write the reaction occurring at the anode for:

ethane ----- ---- $C_2H_6(g) + 4H_2O(l) \rightarrow 2CO_2(g) + 14H^+(aq) + 14e$

ethane-1, 2-diol ---- $C_2H_4(OH)_2(l) + 2H_2O(l) \rightarrow 2CO_2(g) + 10H^+(aq) + 10e$

b) Which one of the two fuels would produce the greatest amount of CO_2 per coulomb of electrical charge generated?

Step 1 from the oxidation equations above derive the mol of CO_2 prodcued per mol of electrons produced

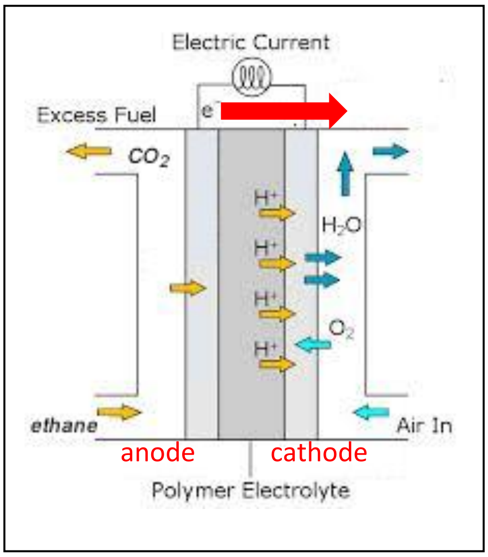
For ethane 2 mol of CO_2 : 14 mol of e = for every 7 mol of electrons 1 mol of CO_2 is produced.

For ethane-1, 2-diol it is 1 mol of CO_2 for every 5 mol of electrons.

Ethane-1, 2-diol will produce more CO_2 than ethane for the same amount of charge.

c) The ethane fuel cell operates at a voltage of 0.842 V.

i. Calculate the amount of energy, in kJ, delivered by one mol of ethane.



Earlier we studied calibration of a calorimeter and were introduced to the formula E = VIt.

Now Charge is given by the formula

Q= It.

$C_2H_6(g) + 4H_2O(l) \rightarrow 2CO_2(g) + 14H^+(aq) + 14e$

so for one mol of ethane 14 mol of electrons are given off this represents a charge of

$Q = It = 14 \times 96500 = 1.35 \times 10^6$ C

Hence energy given out by one mol of ethane is given by relationship below.

$E = VIt = 0.842 \times 1.35 \times 10^6 = 1{,}137$ kJ

ii. Using the molar heat of combustion of ethane, calculate the energy efficiency of this fuel cell.

Molar heat of combustion of ethane is derived from the Data Book at 1,557 kJ/mol

=> efficiency = (1,137/1,557) X 100 = 73.0% efficient.

iii. Name one difference between the electrodes of a fuel cells and the electrodes of most galvanic cells.

Electrodes of a fuel cell act as semipermeable barriers and also act as catalysts while those of galvanic cells do not act as catalysts or as semipermeable barriers.

The electrodes of a fuel cell are not used up but elecrodes of some galvanic cells are part of the half-cell reaction.

iv. Indicate on the diagram:

- the direction of electron flow,
- cathode,
- anode.

v. What is the difference between the electrolyte in a fuel cell and the salt bridge in a galvanic cell?

The salt bridge in a galvanic cell and the electrolyte of a fuel cell allow the movement of ions between the half-cells thus preventing the build-up of positive charge at the anode, as electrons leave, and a build-up of negative charge in the cathode half-cell,

as electrons move in.
In a fuel cell the ions travel through the electrolyte to the cathode where they regain their electrons and undergo a chemical reaction with, usually oxygen, to create water. Ions in a salt bridge do not take part in reactions at the cathode or anode.

vi. Write the equation for the reaction taking place at the anode if an ethane fuel cell using an alkaline electrolyte, such as KOH paste.

$C_2H_6(g) + 4H_2O(l) \rightarrow 2CO_2(g) + 14H^+(aq) + 14e$

Remove all the H^+ ions by adding OH^- ions to both sides.

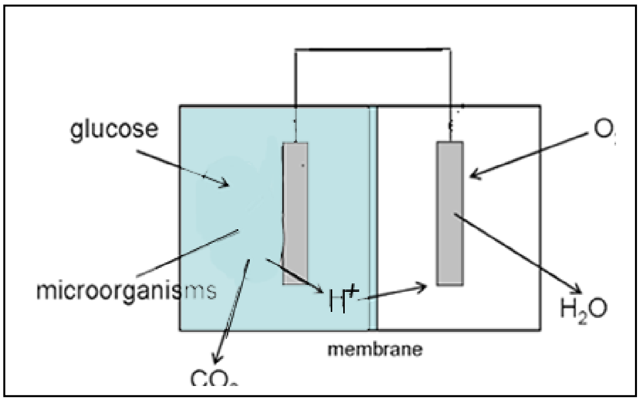
$C_2H_6(g) + 4H_2O(l) + 14OH^-(aq) \rightarrow 2CO_2(g) + 14H^+(aq) + 14OH^-(aq) + 14e$

$C_2H_6(g) + 4H_2O(l) + 14OH^-(aq) \rightarrow 2CO_2(g) + 14H_2O(l) + 14e$

Remove the H_2O molecules that appear on both sides of the equation.

$C_2H_6(g) + 14OH^-(aq) \rightarrow 2CO_2(g) + 10H_2O(l) + 14e$

d) A microbial fuel cell is a device that converts chemical energy into electrical energy by the catalytic reaction of microorganisms. The image on the right shows the schematic of a microbial fuel cell, where glucose ($C_6H_{12}O_6$)In solution is broken down by bacteria to CO_2.



i. What is the oxidation state of carbon in:

- $C_6H_{12}O_6$, ----- 0
- CO_2 ----- +4

ii. Write the equation to the reaction that occurs at the anode.

Since oxidation occurs at the anode

$C_6H_{12}O_6(aq) \Rightarrow CO_2(g)$

Balance for carbons

$C_6H_{12}O_6(aq) \Rightarrow 6CO_2(g)$

Balance for oxygen by adding H_2O to the left

$C_6H_{12}O_6(aq) + 6H_2O(l) \Rightarrow 6CO_2(g)$

Balance for hydrogen by adding H^+ to the right

$6H_2O(l) + C_6H_{12}O_6(aq) \Rightarrow 6CO_2(g) + 24H^+(aq)$

Balance for charge by adding electrons to the most positive side.

$6H_2O(l) + C_6H_{12}O_6(aq) \Rightarrow 6CO_2(g) + 24H^+(aq) + 24e$

iii. Write the equation for the reaction that occurs at the cathode.

$O_2(g) + 2H^+(aq) + 2e \Rightarrow H_2O(l)$

iv. The cell operates for 6.00 hours producing a constant current of 1.05 Amps. What amount of glucose, in grams, is consumed by the cell?

According to the $6H_2O(l) + C_6H_{12}O_6(aq) \Rightarrow 6CO_2(g) + 24H^+(aq) + 24e$ for every mol of glucose 24 mol of electrons are produced.

Step 1 find the charge delivered

=> $Q = It = 1.05 \times 6 \times 60 \times 60 = 22680$ C

Step 2 find the mol of electrons

=> $n_e = 22680/96500 = 0.235$

Step 3 find the mol of glucose

=> $0.235/24 = 9.79 \times 10^{-3}$

Step 4 find the mass of glucose = $180.2 \times 9.79 \times 10^{-3} = 1.76$ g