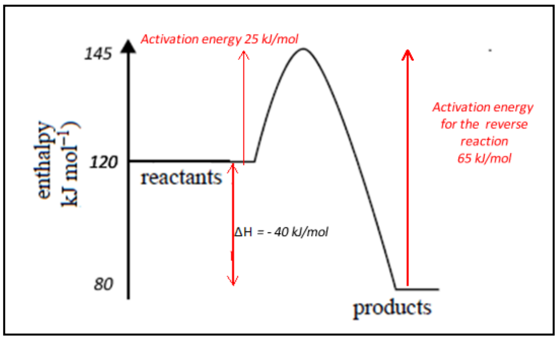
Revision Unit 3 VCE Chemistry 2018

Thermochemical equations, electrolysis, equilibrium and rates of reaction.

1) Consider the energy profile diagram of the reaction a(g) + b(g) → c(g), shown on the right.



a) Indicate on the diagram the:

- ΔH

- Activation energy for the forward reaction a(g) + b(g) → c(g)

- Activation energy for the reverse reaction c(g) → a(g) + b(g)

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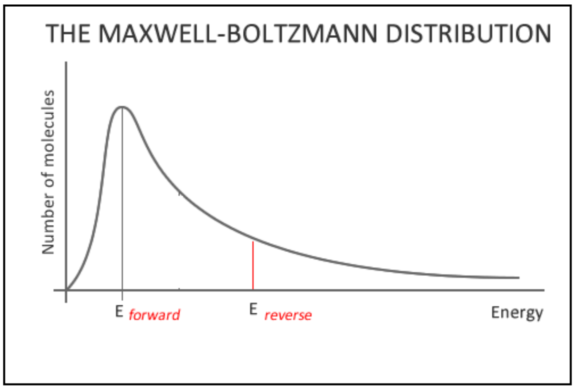
b) Which of the two reactions listed below a(g) + b(g) → c(g) or c(g) → a(g) + b(g) has the fastest rate of reaction at room temperature? Explain why.

a(g) + b(g) → c(g) has the fastest rate of reaction because it has the lowest activation energy of the two reactions and hence more molecules have the required energy to react.



c) When one mole each of a and b react what amount of energy(in kJ) is

- required in the process of bond breaking --- *25kJ*

- released in the process of bond formation --- 65kJ

d) The combustion of 0.165 grams of hexane (86.2 g/mol) takes place according to the equation $2C_6H_{14}(g) + 19O_2(g) \rightarrow 12\ CO_2(g) + 14H_2O(g)$ $\Delta H = -8316$ kJ mol^{-1} in a bomb calorimeter containing 100 mL of water at 25.0°C.

Calculate the final temperature of the water assuming no energy is lost to the surroundings.

Step 1 Calculate the mol of hexane

=> 0.165 / 86.2 = 1.914 X 10^{-3}

Step 2 Calculate the amount of energy released

=> (8316 kJ/2) X 1.914 X 10^{-3} = 7.96 kJ

Step 3 Calculate the mass of water

=> mass = 100mL X 0.997 g mL^{-1} = 99.7g

Step 4 Calculate the change in temperature

=> 7960 = 4.18 X 99.7 X ΔT

=> 7960 /(4.18 X 99.7) = ΔT = 19.1 °C

Step 5 Final temperature = 25.0 + 19.1 = 44.1°C

2) Given the following thermochemical equations

- $1/2N_2(g) + 3/2H_2(g) \rightarrow NH_3(g)$; ΔH = -46.1kJ
- $C(s) + 2H2(g) \rightarrow CH4(g)$; ΔH = -74.7kJ
- $C(s) + 1/2H2(g) + 1/2\ N2(g) \rightarrow HCN(g)$; ΔH = +135.2kJ

a) Find the ΔH for the reaction:

$CH_4(g) + NH_3(g) \rightleftarrows HCN(g) + 3H_2(g)$ ΔH = +256kJ mol⁻

b) This reaction took place in a bomb calorimeter, where 16.00 grams of CH_4 was placed in a 2.00 litre vessels with 28.05 grams of ammonia and allowed to reach equilibrium at a temperature of 20.0°C . A total of 83.5 kJ of energy was absorbed from the surroundings.

Assuming no heat was lost from the system calculate, at equilibrium, the:

Step 1 calulate the mol of CH_4 and NH_3 added.

=> $n_{methane}$ = 16.00 / 16.0 = 1.00 mol

=> $n_{ammonia}$ = 28.05 / 17.0 = 1.65 mol

Step 2 calculate the mol of HCN produced from the release of 83.5 kJ of energy.

=> n_{HCN} = 83.5 / 256 = 0.326 mol

- ⇨ Mol of CH_4 and the $[CH_4]$

  $n_{methane}$ = 1.00 – 0.326 = 0.497 mol => $[CH_4]$ = 0.497 / 2.00 = 0.249 M

- ⇨ Mol of NH_3 and the $[NH_3]$

  $n_{ammonia}$ = 1.65 – 0.326 = 1.32 mol => $[CH_4]$ = 1.32 / 2.00 = 0.662 M

- ⇨ Mol of HCN and the [HCN]

  n_{HCN} =0.326 => [HCN] = 0.326/ 2.00 = 0.163 M

- ⇨ Mol of H_2 and the $[H_2]$

  $n_{Hydrogen}$ = 0.326 X 3= 0.978 => [HCN] = 0.978/ 2.00 = 0.489 M

c) Calculate the k_c value for the reaction $CH_4(g) + NH_3(g) \rightleftarrows HCN(g) + 3H_2(g)$ at the given temperature in b) above.

$[HCN]\ [H_2]^3 / ([CH_4]\ [NH_3]) = K_c$

$(0.163 \times (0.489)^3) / (0.249 \times 0.662) = 0.116\ M^{-2}$

d) Calculate the theoretical yield of HCN in mol.

Methane is clearly the limiting reactant

=> mol of HCN that should be produced = 1.00

e) Calculate the actual yield of HCN

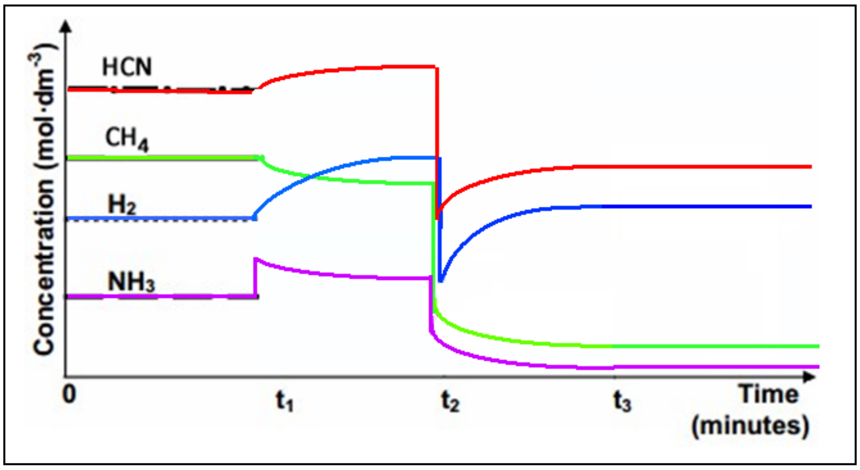
0.326 mol

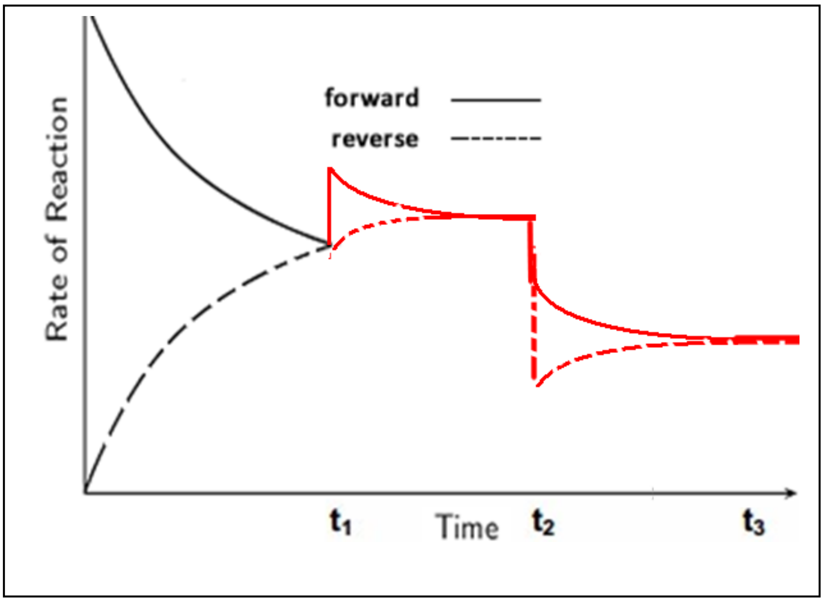
f) Calculate the percentage yield for this reaction at equilibrium

(0.326 / 1.00) X 100 = 32.6%

g) The temperature of the equilibrium established in d) above was changed. The graph below shows the concentrations for the different species at equilibrium. Sketch how the concentration of each species as well as the rate of the forward and reverse reactions change over time when at:

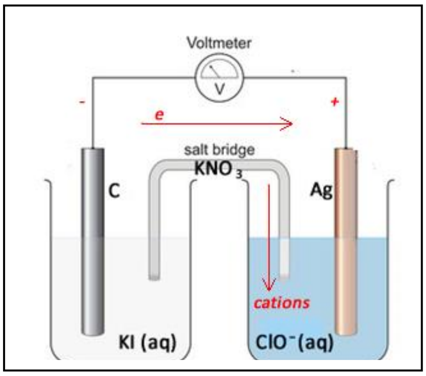
- T_1 extra NH_3 is added to the system and equilibrium is established before T_2
- T_2 the volume of the vessel is doubled and equilibrium is established before T_3
- T_3 neon gas is added to the system and the pressure inside the vessel is doubled

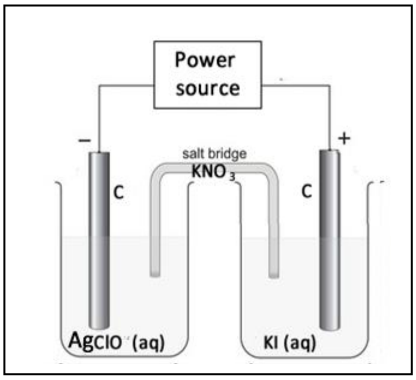


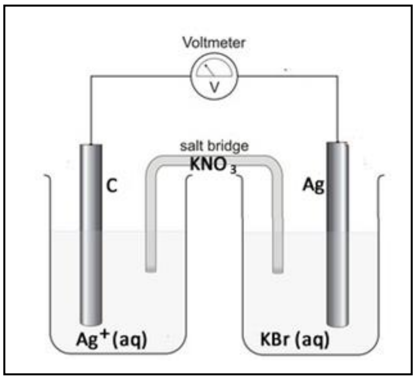


3) Consider the **Standard Reduction Potentials at 25°C (298K) shown below**

$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$	+1.33
$O_2 + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$	+1.23
$ClO_4^-(aq) + 2H^+(aq) + 2e^- \rightarrow ClO_3^-(aq) + H_2O(l)$	+1.23
$MnO_2(s) + 4H^+(aq) + 2e^- \rightarrow Mn^{2+}(aq) + 2H_2O(l)$	+1.23
$IO_3^-(aq) + 6H^+(aq) + 5e^- \rightarrow I_2(s) + 3H_2O(l)$	+1.20
$Br_2(l) + 2e^- \rightarrow 2Br^-(aq)$	+1.09
$HNO_2(aq) + H^+(aq) + e^- \rightarrow NO(g) + H_2O(l)$	+1.00
$NO_3^-(aq) + 4H^+(aq) + 3e^- \rightarrow NO(g) + 2H_2O(l)$	+0.96
$ClO^-(aq) + H_2O(l) + 2e^- \rightarrow Cl^-(aq) + 2OH^-(aq)$	+0.89
$ClO^-(aq) + 2H^+(aq) + e^- \rightarrow NO_2(g) + H_2O(l)$	+0.80
$Ag^+(aq) + e^- \rightarrow Ag(s)$	+0.80
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	+0.77
$BrO^-(aq) + H_2O(l) + 2e^- \rightarrow Br^-(aq) + 2OH^-(aq)$	+0.76
$MnO_4^{2-}(aq) + 2H_2O(l) + 2e^- \rightarrow MnO_2(s) + 4OH^-(aq)$	+0.60
$H_3AsO_4(aq) + 2H^+(aq) + 2e^- \rightarrow H_3AsO_3(aq) + H_2O(l)$	+0.56
$MnO_4^{2-}(aq) + e^- \rightarrow MnO_4^{2-}(aq)$	+0.56
$H_3AsO_4(aq) + 2H^+(aq) + 2e^- \rightarrow H_3AsO_3(aq) + H_2O(l)$	+0.56
$I_2(s) + 2e^- \rightarrow 2I^-(aq)$	+0.54
$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$	-0.83

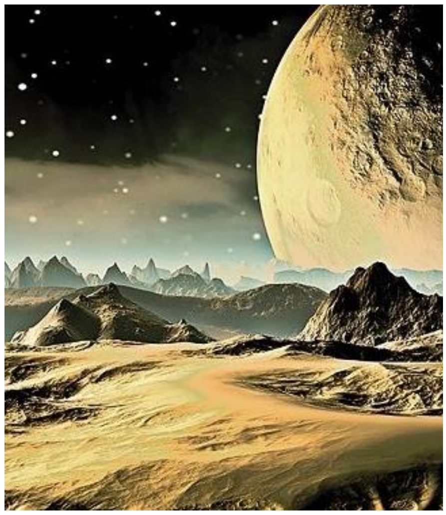






Indicate on the diagram the: - direction of electron flow - direction of cation movement Identify the Oxidant – *$ClO^-(aq)$* Reductant – *$I^-(aq)$* Write the - half equation at the negative electrode *$2I^-(aq) \rightarrow I_2(s) + 2e$* - half equation at the positive electrode *$ClO^-(aq) + 2H_2O + 2e \rightarrow Cl^-(aq) + 2OH^-(aq)$* Calculate the theoretical voltage. 0.35 V	Indicate on the diagram the: - direction of electron flow Identify the Oxidant – *Ag^+* Reductant – *I^-* Write the - oxidation half equation *$2I^-(aq) \rightarrow I_2(s) + 2e$ at the positive electrode* - reduction half equation *$Ag^+(aq) + e \rightarrow Ag(s)$* *$ClO^-$ is a stronger oxidant but an anion would be repelled by the negative electrode* At which electrode is each half equation taking place.	Will a reaction occur? *NO* Identify the Oxidant - Reductant - Write the - half equation at the negative electrode - half equation at the positive electrode Calculate the theoretical voltage.

4) On an alien planet a Human colony is established.



The planet is near a star and receives as much solar energy as Earth, however, unlike Earth has no atmosphere but has huge deposits of ice found in the shaded areas of craters. The planet rotates on its axis much like Earth, but has a gravitational pull one tenth that of Earth and has long periods of night and day. The colony has access to a variety of crops including corn and a range of micro-organisms such as yeast.

Using your chemical knowledge suggest plausible solutions for the following. Provide chemical equations to justify all your suggestions and diagrams of any electrochemical cells that are needed.

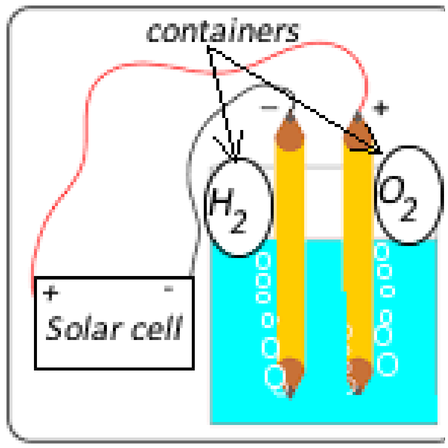
a) The need for:

- water for drinking and for the herbarium

Melting of ice via solar radiation

$H_2O(s) \rightarrow H_2O(l)$

- a reliable inorganic source of oxygen gas (draw a possible device labelling the terminals and their polarity as well as the chemical reactions taking place at each terminal)



Energy from the star can be used to power solar cells producing electricity to electrolyse water into hydrogen and oxygen gases.

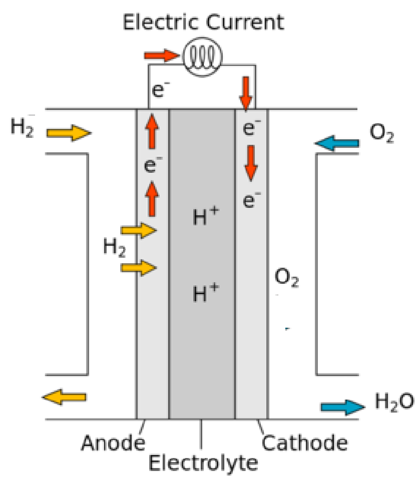
at the positive anode ---

$2H_2O(1) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$

at the negative cathode ---

$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$

- a reliable electrical supply in the night provided by fuel cells, with acidic electrolytes, powered by recyclable gases.



Since we can regenerate hydrogen and oxygen from the electrolysis of water we can use the hydrogen fuel cell.

- anode $H_2(g) \rightarrow 2H^+(aq) + 2e$

- cathode $O_2 + 4H^+(aq) + 2e \rightarrow 2\ H_2O(aq)$

- a low weight fuel to launch vehicles into space that has the highest energy density of all the available fuels produced by the colony. Provide evidence to justify your choice of fuel.

Ethanol and hydrogen are the two fuels that can be produced via fermentation and electrolysis of water respectively.

Hydrogen has the highest energy density when compared to ethanol at 141 kJ/g as compared to ethanol at 29.6 kJ/g.

The combustion reaction of hydrogen is $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

- a renewable organic fuel that is carbon neutral to power electric vehicles when travelling in long periods of darkness. Suggest why this fuel is renewable.

Fermentation of glucose, derived from corn starch, to produce ethanol.

$C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2CO_2(g)$

Ethanol can react in a combustion reaction with oxygen derived from the electrolysis of water or through photosynthesis.

$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$

Plants can photosynthesise to produce glucose at a rate that never sees glucose used up.

$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6\ O_2(g)$