

VCE CHEMISTRY 2008

DETAILED SOLUTIONS

SAMPLE WRITTEN EXAMINATION 2



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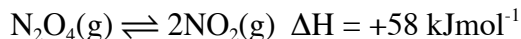
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WEB LINK

Web Links have been added to the answers for some questions so that this document can be used as a teaching tool. Internet links do not last forever. Please let us know if any links are "dead".

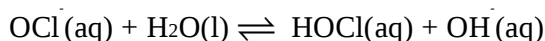
SECTION A

Question 1 ANS D

[WEB LINK](#)

In the balanced equation 1 mol of N_2O_4 produces 2 mol of NO_2 . Therefore, according to Le Chatelier's Principle, an increase in pressure will favour the production of the smaller number of molecules. The equilibrium will shift to the left. The forward reaction is endothermic. Hence, an increase in temperature will force the reaction to the right. More NO_2 will be produced. The equilibrium yield of NO_2 will go **down** with **increasing pressure** and **up** with **increasing temperature**. Only graph **D** shows this.

Question 2 ANS B



When two drops of 5.0 M NaOH are added, this will not significantly affect the volume of the equilibrium mixture. When two drops of 5.0 M NaOH are added to an equilibrium mixture of OCl^- in water at constant temperature, only some (but not all) of the added OH^- ions are used up as the equilibrium shifts to the left. Hence, there is an increase in the $[\text{OH}^-]$. The pH of the mixture will **increase** because of the extra OH^- . **A** is false. The concentration of HOCl will **decrease** as the equilibrium shifts to the left. **C** is false. The ratio given in D (the equilibrium constant without the water) will not change since the temperature is constant. **D** is false.

Question 3 ANS A

Increasing the proportion of a reactant in the reaction mixture will always results in an equilibrium shift to favour products. **I** is true. Increasing the temperature will only increase the amount of product present at equilibrium if the forward reaction is endothermic. **II** is false. Decreasing the pressure (by increasing the volume) will only increase the amount of product present at equilibrium if the forward reaction produces a larger number of mole of gas. **III** is false. Adding a catalyst changes the rate at which equilibrium is reached. A catalyst has no effect on the amount of product present at equilibrium. **IV** is false.

Question 4 ANS A

Since the pH is the same in both beakers, the concentration of $\text{H}^+(\text{aq})$ in beaker 1 is equal to the concentration of $\text{H}^+(\text{aq})$ in beaker 2. **III** is false. The concentration of $\text{HY}(\text{aq})$ is much less (0.0001M) than the concentration of $\text{HX}(\text{aq})$ (0.10M) but both acids produce the same hydrogen ion concentration in solution. Therefore, $\text{HY}(\text{aq})$ is more ionised than $\text{HX}(\text{aq})$. Therefore, $\text{HY}(\text{aq})$ is the stronger acid. Hence, $\text{HY}(\text{aq})$ has the higher K_a value. **II** is false. Only **I** is correct.

Question 5 ANS C**WEB LINK**

On page 11 of the Data Book, the K_a value for benzoic acid is given as 6.4×10^{-5} .

The equilibrium is $\text{C}_6\text{H}_5\text{COOH}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{C}_6\text{H}_5\text{COO}^-(\text{aq})$

$$\text{Hence, } K_a = \frac{[\text{C}_6\text{H}_5\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{C}_6\text{H}_5\text{COOH}]} = 6.4 \times 10^{-5}$$

$$n(\text{C}_6\text{H}_5\text{COOH}) = \frac{m}{M_r} = \frac{0.500}{122} = 4.1 \times 10^{-3}$$

$$c(\text{C}_6\text{H}_5\text{COOH}) = \frac{n}{V} = \frac{4.1 \times 10^{-3}}{0.200} = 0.0205 \text{ M}$$

$$K_a = \frac{[\text{C}_6\text{H}_5\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{C}_6\text{H}_5\text{COOH}]} = 6.4 \times 10^{-5}$$

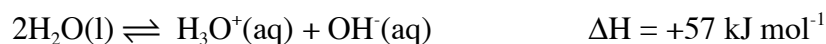
Since $[\text{C}_6\text{H}_5\text{COO}^-] = [\text{H}_3\text{O}^+]$ and $[\text{C}_6\text{H}_5\text{COOH}] \approx 0.0205 \text{ M}$

$$[\text{H}_3\text{O}^+]^2 = 6.4 \times 10^{-5} \times 0.0205 = 1.3 \times 10^{-6}$$

$$[\text{H}_3\text{O}^+] = 1.145 \times 10^{-3} \text{ M}$$

$$\text{pH} = -\log_{10} [\text{H}_3\text{O}^+] = -\log_{10} (1.145 \times 10^{-3}) = -2.94$$

This is closest to 3.0

Question 6 ANS A

From the balanced equation above, it can be seen that equal numbers of moles of hydronium ions and hydroxide ions are produced when water dissociates. $[\text{H}_3\text{O}^+] = [\text{OH}^-]$. Pure water is neutral **at any temperature** because the concentration of hydronium ions is always equal to the concentration of hydroxide ions.. At 25°C , for pure water, $[\text{H}_3\text{O}^+] = 10^{-7} \text{ M}$ and the pH is 7. However, as the forward reaction shown above is endothermic, an increase in temperature to 90°C will favour the forward reaction and produce a higher concentration of hydronium ions, that is, greater than 10^{-7} M . Therefore, the pH will decrease to less than 7.

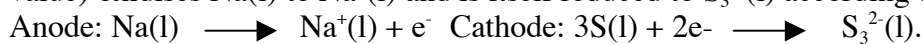
Question 7 ANS D

On this diagram, the hydrogen ion concentration is **decreasing** from left to right. **B** is false. The pH scale is a logarithmic scale. A difference of 1 pH unit indicates a hydrogen ion concentration difference 10 times greater or 10 times less. **A** and **C** are false. As pH decreases, hydrogen ion concentration increases. Hence, distilled water with a pH of 7 compared to soap with a pH of 10, has a hydrogen ion concentration 1000 times greater. This can be shown mathematically as

$$\frac{[\text{H}_3\text{O}^+] \text{ at pH } 7}{[\text{H}_3\text{O}^+] \text{ at pH } 10} = \frac{10^{-7}}{10^{-10}} = 10^3 = 1000$$

Question 8 ANS A[WEB LINK](#)

In a galvanic cell, spontaneous chemical reactions produce a flow of electrons. Oxidation occurs at the anode (electrons are lost), so the anode is negative. Reduction occurs at the cathode (electrons are gained), so the cathode is positive. In this galvanic cell, the stronger oxidant S(l) (higher E° value) oxidises Na(l) to Na⁺(l) and is itself reduced to S₃²⁻(l) according to the equations:



Sodium metal in liquid form is the reactant at the anode which is negative.

Question 9 ANS D

This is a galvanic cell in which lithium metal is the anode. A spontaneous oxidation occurs here and **A** is the half-reaction for this discharge. CORRECT. **B** is the reduction half-reaction for the discharge as Co changes its oxidation number from +4 to +3. CORRECT. **C** is the overall reaction for the discharge as Li loses one electron and Co gains one electron. CORRECT. **D** is the overall cell reaction for **recharge** (the opposite of **C**). Li metal is the anode and it is being oxidized. **D** is the only reaction where Li is being reduced and this reaction occurs at the cathode. CANNOT OCCUR during discharging.

Question 10 ANS C

To predict the products at the electrodes of an electrolytic cell (electrical energy is converted to chemical energy) under standard conditions (1M concentrations of aqueous solutions, 25°C temperature, 1 atm pressure for gases used), the electrochemical series on page 4 of the Data Book must be used. *This question is defective in that it does not use standard conditions and does not specify the materials in the electrodes.* We will assume that the electrodes are inert and that the E° values apply to this 0.5 M solution. The Cu²⁺(aq) ions and the positive ends of the water molecules will be attracted to the negative electrode (the cathode in an electrolytic cell). Cu²⁺(aq) (+0.34V) is a stronger oxidant than H₂O(l) (-0.83V) and so Cu²⁺(aq) ions will accept electrons in preference to H₂O(l) molecules. Cu(s) will be deposited on the cathode. The Br⁻(aq) ions and the negative ends of the water molecules will be attracted to the positive electrode (the anode in an electrolytic cell). Br⁻(aq) (+1.09V) is a stronger reductant than H₂O(l) (+1.23V) and so Br⁻(aq) ions will give up electrons in preference to H₂O(l) molecules. Br₂(l) will be produced at the anode. It is predicted that no gases will be produced in this electrolytic cell. In a real experiment over an extended period of time . . . what do you think?

Question 11 ANS C**WEB LINK**

Use the electrochemical series on page 4 of the Data Book. $\text{Fe}^{2+}(\text{aq})$ ions can act as either an oxidant or a reductant. As an oxidant, the E^0 value is -0.44V . Hence, $\text{Fe}^{2+}(\text{aq})$ can oxidise (be reduced by) $\text{Zn}(\text{s})$ which has the lower E^0 value of -0.76V . As a reductant, the E^0 of $\text{Fe}^{2+}(\text{aq})$ value of is $+0.77\text{V}$. Hence, $\text{Fe}^{2+}(\text{aq})$ can reduce (be oxidised by) $\text{Br}_2(\text{l})$ which has the higher E^0 value of $+1.09\text{V}$.

Question 12 ANS C

Use the molar enthalpy of combustion of ethanol in section 13 on page 11 of the Data Book. This shows that when 1 mole of ethanol reacts completely with oxygen, 1364 kJ of energy is released. The balanced chemical equation is $\text{C}_2\text{H}_5\text{OH}(\text{l}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{g})$. Note that in this case the **heat of combustion** is the same as the **enthalpy of combustion** when written as a balanced chemical equation. *This is not always the case.* Be clear about this point.

Heat released by 5.0 g of ethanol

$$= 1364 \times n$$

$$= 1364 \times \frac{5.0}{(24.0 + 6.0 + 16.0)}$$

$$= 1364 \times \frac{5.0}{46.0}$$

$$= 1364 \times 0.1087$$

$$= 148.26\text{ kJ}$$

Temperature change of water

$$= \frac{\text{energy released in kJ}}{\text{calorimeter calibration factor in kJ K}^{-1}}$$
$$= \frac{148.26}{3.34}$$

$$= 44.4\text{ K (}^\circ\text{C)}$$

Temperature change is closest to 44°C

Question 13 ANS B

Use the electrochemical series on page 4 of the Data Book to predict the reactions at each of the negative electrodes. $\text{H}_2\text{O}(\text{l})$ is a stronger oxidant than $\text{Ca}^{2+}(\text{aq})$ and, therefore, no $\text{Ca}(\text{s})$ will be produced in the first cell since hydrogen gas and hydroxide ions will be produced at the cathode in preference to $\text{Ca}(\text{s})$. **A** and **C** are incorrect. $\text{Ag}^+(\text{aq})$ and $\text{Ni}^{2+}(\text{aq})$ are both stronger oxidants than $\text{H}_2\text{O}(\text{l})$. Hence, both $\text{Ag}(\text{s})$ and $\text{Ni}(\text{s})$ will be produced at the negative electrodes in preference to hydrogen gas and hydroxide ions. When 2 mol of electrons flows through the circuit, 2 mol of $\text{Ag}(\text{s})$ will be produced but only 1 mol of $\text{Ni}(\text{s})$ because of the relationships from the balanced half-equations: $n(\text{Ag}(\text{s})) = n(\text{e}^-)$ and $n(\text{Ni}(\text{s})) = \frac{1}{2} n(\text{e}^-)$.

The correct ratio is $0\text{ mol Ca}(\text{s}) : 2\text{ mol Ag}(\text{s}) : 1\text{ mol Ni}(\text{s})$.

Question 14 ANS A

In every galvanic cell that is discharging, whether primary (non-rechargeable) or secondary (rechargeable), a spontaneous oxidation reaction occurs at the anode which is, therefore, negatively charged since electrons are being produced at this electrode. The electrons flow through the external circuit from the anode to the cathode and the negatively charged ions in the electrolyte – the anions – move in the same circular direction. That is, they flow towards the negative electrode – the anode. **B**, **C** and **D** are incorrect because primary and secondary galvanic cells operate on the same principles when discharging.

Question 15 ANS C

You can use the electrochemical series on page 4 of the Data Book to help with this question but you need to use fundamental reasoning to work out the answer since one of the metals (X) is unknown. From the Data Book, you can see that $\text{Fe}^{3+}(\text{aq})$ is a stronger oxidant than $\text{Co}^{2+}(\text{aq})$ which, in turn, is a stronger oxidant than $\text{Fe}^{2+}(\text{aq})$.

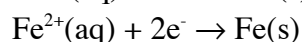
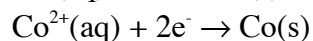
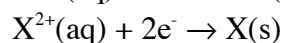
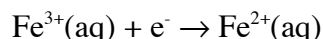
Half cells 1 and 3 are joined: Since the cobalt electrode is negative in this galvanic cell, this is the anode and, therefore, $\text{X}^{2+}(\text{aq})$ is a stronger oxidant (electron acceptor) than $\text{Co}^{2+}(\text{aq})$.

$\text{X}^{2+}(\text{aq})$ oxidises $\text{Co}(\text{s})$ to $\text{Co}^{2+}(\text{aq})$ and is itself reduced to $\text{X}(\text{s})$.

Half cells 2 and 3 are joined: Since the metal X electrode is negative in this galvanic cell, this is the anode and, therefore, $\text{Fe}^{3+}(\text{aq})$ is a stronger oxidant (electron acceptor) than $\text{X}^{2+}(\text{aq})$.

$\text{Fe}^{3+}(\text{aq})$ oxidises $\text{X}(\text{s})$ to $\text{X}^{2+}(\text{aq})$ and is itself reduced to $\text{Fe}^{2+}(\text{aq})$.

The order on the electrochemical series is



The strongest oxidant is $\text{Fe}^{3+}(\text{aq})$.

Extension question: From the Data Book, identify the metals that could be metal X.

Question 16 ANS D[WEB LINK](#)

According to the balanced equation, 1 mol of phosphorus solid reacts exactly with 10 mol of chlorine gas to produce 4 mol of phosphorus pentachloride gas.

$$n(\text{PCl}_5) \text{ produced} = \frac{m}{M_r} = \frac{6.49}{31.0 + (5 \times 35.5)} = \frac{6.49}{208.5} = 0.0311 \text{ mol}$$

The ΔH for a chemical reaction is per mol of equation **as written**.

There are 4 mol of PCl_5 in the equation so the ΔH value is the energy released when 4 mol of PCl_5 is produced. 0.0311 mol produces 11.7 kJ, so 4 mol produces x kJ.

$$x = \Delta H = 11.7 \times \frac{4}{0.0311} = 1504 \text{ kJ mol}^{-1}$$

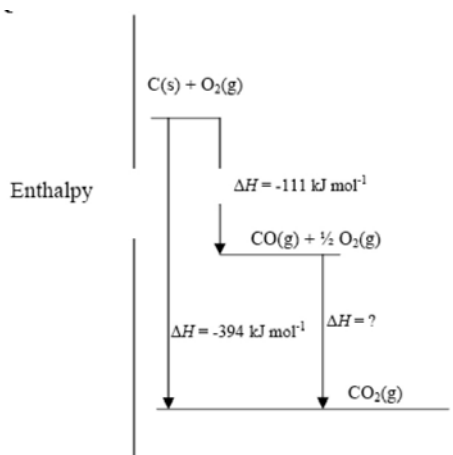
This is closest to 1500 kJ mol^{-1} .

Question 17 ANS A**WEB LINK**

Use the molar enthalpy of combustion for each of these fuels given in Section 13 on Page 11 of the Data Book. The molar enthalpy of combustion (ΔH_c) of a fuel is the energy released during the complete combustion of 1 mol of the fuel at 298 K and 101.3 kPa.

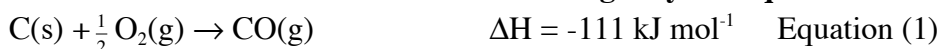
(It is **not** necessarily the same value as ΔH for the balanced equation.)

1 mol of methane (CH_4) produces 1 mol of CO_2 . Energy released = 889 kJ per mol of CO_2 .	1 mol of butane (C_4H_{10}) produces 4 mol of CO_2 . Energy released $= \frac{2874}{4} = 718.5 = 719$ kJ per mol of CO_2 .
1 mol of octane (C_8H_{18}) produces 8 mol of CO_2 . Energy released $= \frac{5464}{8} = 683$ kJ per mol of CO_2 .	1 mol of ethanol ($\text{C}_2\text{H}_5\text{OH}$) produces 2 mol of CO_2 . Energy released $= \frac{1364}{2} = 682$ kJ per mol of CO_2 .

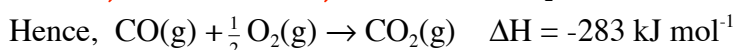
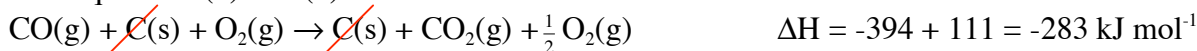
Question 18 ANS B

From the energy profile shown above,
the unknown $\Delta H = -394 - (-111) = -394 + 111 = -283 \text{ kJ mol}^{-1}$.

An alternative solution is shown below using only the equations and their ΔH values.



Add equations (2) and (3)



Question 19 ANS B

Carbon dioxide is an acidic oxide. It will react with the basic solution NaOH(aq) according to the equation: $\text{CO}_2(\text{g}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{Na}_2\text{CO}_3(\text{aq}) + \text{H}_2\text{O}(\text{l})$. This would trap the carbon dioxide in solution as the ion $\text{CO}_3^{2-}(\text{aq})$ and prevent the CO_2 from reaching the atmosphere. Carbon dioxide is soluble in water but this would not be as effective as reacting it with sodium hydroxide. NaCl(aq) and HCl(aq) would no more effective than pure water.

Question 20 ANS C[WEB LINK](#)

The activation energy is the energy required to break the bonds in compounds to allow other compounds to form. A catalyst lowers the activation energy by providing an alternative pathway for the reaction. The activation energy is lowered for **both** the forward and the reverse reactions.

SECTION B**Question 1**

- a. When the catalyst changes from a finely ground powder to large clumps, the total surface area decreases. Hence, there will be a smaller number of collisions per second occurring between reactants on the surface of the catalyst. Therefore, the rate of reaction will decrease.
- b. i. To obtain the overall equation, first balance the electrons by multiplying the first half equation by 2 and the second half equation by 5. Next, add the two half equations.
$$2\text{MnO}_4^-(\text{aq}) + 16\text{H}^+(\text{aq}) + 5\text{C}_2\text{O}_4^{2-}(\text{aq}) \rightarrow 2\text{Mn}^{2+}(\text{aq}) + 8\text{H}_2\text{O}(\text{l}) + 10\text{CO}_2(\text{g})$$
- ii. As a reaction proceeds, the concentration of reactants decreases and, therefore, the number of collisions per second decreases. Hence, the reaction rate decreases.
- iii. It is possible that one of the products in the reaction (for example, $\text{Mn}^{2+}(\text{aq})$) is a catalyst for the reaction so that as the reaction proceeds and more catalyst is formed, the reaction rate increases.

OR

It is also possible that the reaction is exothermic and that an increase in temperature causes a higher proportion of successful collisions between reactants and, therefore, an increase in the reaction rate.

Question 2 [WEB LINK](#)

- a. The equilibrium constant expression is $K_c = \frac{[\text{CH}_3\text{OH}]_e}{[\text{CO}]_e [\text{H}_2]_e^2}$.

Note the subscript *e* which is to remind us that the concentrations used in this expression must be equilibrium concentrations.

b.

- i. The elevated temperatures of 850°C in Step 1 and 300°C in Step 2 are used to increase the proportion of successful collisions and, thereby, increase the rate at which equilibrium is reached. In the case of Step 1, which is an endothermic reaction, the elevated temperature is also used to increase the size of the equilibrium constant and shift the equilibrium position to the right so that the equilibrium yield of products increases. In the case of Step 2, which is an exothermic reaction, the temperature used is a lower compromise temperature so that the best yield of product is obtained in a reasonable time.
- ii. High pressures than atmospheric are used in both steps because higher pressures increase the number of collisions between reactants and, thereby, increase the rates of reaction. In Step 1, a lower compromise pressure is used since Le Chatelier's Principle indicates that a higher pressure will favour the reverse reaction (4 mol of gas gives 2 mol of gas). The pressure is chosen to give the best yield of product in a reasonable time. In Step 2, a higher pressure will increase the equilibrium yield of product as well as the rate since 3 mol of gas produces 1 mol of gas. No compromise is needed here from a chemistry angle but high pressures are expensive and so not used excessively.
- iii. In both steps a catalyst increases the rates of the forward and reverse reactions so that equilibrium is reached more quickly thus eliminating the need for more extreme temperatures and pressures.
- c. To make the production of methanol more efficient, the heat released in the exothermic reaction (Step 2) could be collected and used in the endothermic reaction (Step 1) to maintain the required temperature.
- d. When methanol is completely oxidised in a fuel cell, carbon dioxide is produced. An acidic electrolyte indicates the presence of hydrogen ions. Use your rules for writing half equations as follows:

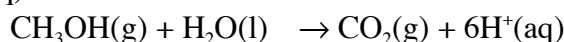
Balance the C atoms (already balanced).



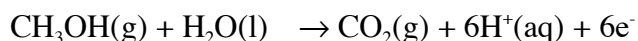
Balance the O atoms by adding H₂O(l).



Balance the H atoms by adding H⁺(aq).



Balance the charge by adding electrons.



Question 2 (continued)

e.

- i. As hydrogen is separated from the reaction mixture and removed, the equilibrium shifts to the right and more hydrogen is produced.
- ii. One disadvantage of this method of steam methane reforming is the cost of the palladium catalyst. Energy required and greenhouse considerations may also be disadvantages.
- iii. In the electrolysis of dilute potassium chloride, hydrogen gas is produced at the cathode and oxygen gas is produced at the anode according to the equations in the Data Book on Page 4 – The electrochemical series.
- $$2\text{H}_2\text{O(l)} \rightarrow \text{O}_2\text{(g)} + 4\text{H}^+\text{(aq)} + 4\text{e}^- \quad \text{Equation (1)}$$
- and $(2\text{H}_2\text{O(l)} + 2\text{e}^- \rightarrow \text{H}_2\text{(g)} + 2\text{OH}^-\text{(aq)}) \times 2 \quad \text{Equation (2)}$

Adding these equations together gives $2\text{H}_2\text{O(l)} \rightarrow \text{O}_2\text{(g)} + 2\text{H}_2\text{(g)}$

Question 3 [WEB LINK](#)

- a. i. To calculate the value of the equilibrium constant, the **equilibrium concentrations** are needed. The following table is a good way of doing this problem.
The balanced chemical equation is: $\text{COCl}_2\text{(g)} \rightleftharpoons \text{CO(g)} + \text{Cl}_2\text{(g)}$.

	$\text{CoCl}_2\text{(g)}$	CO(g)	$\text{Cl}_2\text{(g)}$
Initial number of mol	0.30	0.20	0.00
Number of mol reacting or produced	0.10 reacted	0.10 produced	0.10 produced
Number of mol at equilibrium	$0.30 - 0.10 = 0.20$	$0.20 + 0.10 = 0.30$	0.10
Concentration at equilibrium (mol L^{-1})	$\frac{0.20}{5.0} = 0.040$	$\frac{0.30}{5.0} = 0.060$	$\frac{0.10}{5.0} = 0.020$

To calculate the equilibrium constant for the forward reaction at 40°C :

$$K_1 = \frac{[\text{CO}][\text{Cl}_2]}{[\text{COCl}_2]}$$

$$K_1 = \frac{0.060 \times 0.020}{0.040} = 0.030 \text{ M}$$

- ii. The equilibrium constant for the reverse reaction at $40^\circ\text{C} = K_2$

$$K_2 = \frac{[\text{COCl}_2]}{[\text{CO}][\text{Cl}_2]} = \frac{1}{K_1}$$

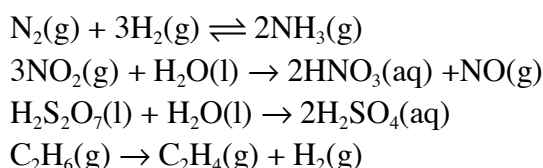
$$K_2 = \frac{0.040}{0.060 \times 0.020} = 33.3 \text{ M}^{-1}$$

Question 3 (continued)

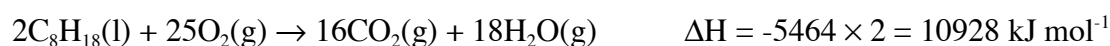
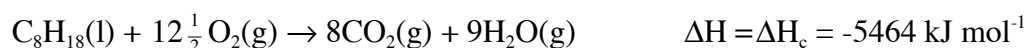
- b.** When the pressure is constant in an equilibrium reaction in which there is a change in the number of mol of gas from reactants to products, the rate of the forward reaction must be equal to the rate of the reverse reaction. Hence, the system is in dynamic equilibrium.
- c.i.**
1. An increase in volume at constant temperature causes a decrease in pressure. According to Le Chatelier's Principle, the equilibrium will shift in favour of the side of the equation containing the larger number of mole of gas. More $\text{Cl}_2(\text{g})$ will be produced and, therefore, the final colour of the mixture will be a **deeper green** in comparison with that in the initial equilibrium.
 2. The addition of more $\text{CO}(\text{g})$ at constant volume and temperature will cause the equilibrium to shift to the left and, therefore, some of the $\text{Cl}_2(\text{g})$ will be used up. The final colour of the mixture will be a **lighter green** in comparison with that in the initial equilibrium.
 3. The addition of a non-reacting gas at constant volume and temperature changes nothing in the equilibrium and, therefore, there will be no shift in the equilibrium position and the final colour of the mixture will be **unchanged** in comparison with that in the initial equilibrium.
- ii.** The value of the equilibrium constant depends only on the temperature and the equation written for the equilibrium. None of the changes in **i.** will change the value of the equilibrium constant.
- d.** To determine experimentally whether the forward reaction is exothermic or endothermic, increase the temperature of the equilibrium mixture and observe the change in colour. If the colour is darker green, then the forward reaction has been favoured and, therefore, the forward reaction is endothermic ($\Delta H = +\text{ive}$). Another method would be to react some $\text{CO}(\text{g})$ and $\text{Cl}_2(\text{g})$ and measure the temperature change.

Question 4 [WEB LINK](#)

- a.**
- i.** The selected chemical is one of NH_3 , HNO_3 , H_2SO_4 , C_2H_4 .
 - ii.** The chemical equations for the productions of each of the above are:



- b.** **Two** balanced equation for the complete combustion of octane are shown below

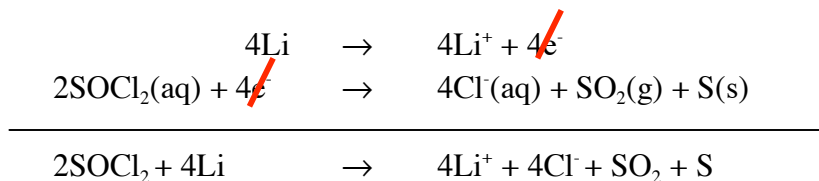


Note that heat of combustion (ΔH_c) is **per mol of the fuel** (that is, for 1 mol)

Note that the enthalpy of a reaction (ΔH) is **per mol of equation as written**.

Question 4 (continued)

- c. i. Lithium is the reductant according to the half-equation: $\text{Li} \rightarrow \text{Li}^+ + \text{e}^-$
- ii. The overall equation is obtained by balancing the electrons and then adding the half-equations as shown below



- iii. If the SO_2 produced in the cell did not dissolve in the electrolyte, dangerous pressure would build up and the cell could explode.

Question 5 [WEB LINK](#)

- a. These physical constants are supplied on page 5 of the Data Book.

- Specific heat capacity of water = $c = 4.18 \text{ J g}^{-1} \text{ K}^{-1}$.
- Density (d) of water at $25^\circ\text{C} = 1.00 \text{ g mL}^{-1}$.

Therefore, mass of water = $m = d \times V = 200 \text{ mL} \times 1.00 \text{ g mL}^{-1} = 200 \text{ g}$.

Amount of energy in Joules added during heating is given by

$$\begin{aligned}
 E &= m \times c \times \Delta T \\
 &= 200 \times 4.18 \times (85.6 - 22.7) \\
 &= 200 \times 4.18 \times 62.9 \\
 &= 52584 \text{ J} \\
 &= 52.6 \text{ kJ}
 \end{aligned}$$

- b. Mass of 1-propanol ($\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$) used = $125.62 - 122.89 = 2.73 \text{ g}$
(Note that the formula for 1-propanol is given on page 11 of the Data Book)

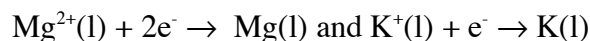
$$\text{Number of mol of 1-propanol used} = \frac{2.73}{36 + 8 + 16} = \frac{2.73}{60} = 0.0455 \text{ mol.}$$

$$\text{Enthalpy of combustion} = \frac{52.6}{0.0455} = 1156 \text{ kJ mol}^{-1}$$

- c. The percentage of chemical energy of the 1-propanol transferred to heat energy in the water
- $$= \frac{1156}{2016} \times 100 = 57.3\%$$
- d. The numerical value of the heat of combustion of 1-propanol obtained using a bomb calorimeter would be larger than that obtained using the spirit burner because
- (1) less heat would be lost to the surrounding since the container is insulated.
 - (2) pure oxygen would give more complete combustion.
 - (3) all of the heat goes into the water in the bomb calorimeter but some would go into the air in the spirit burner.

Question 6

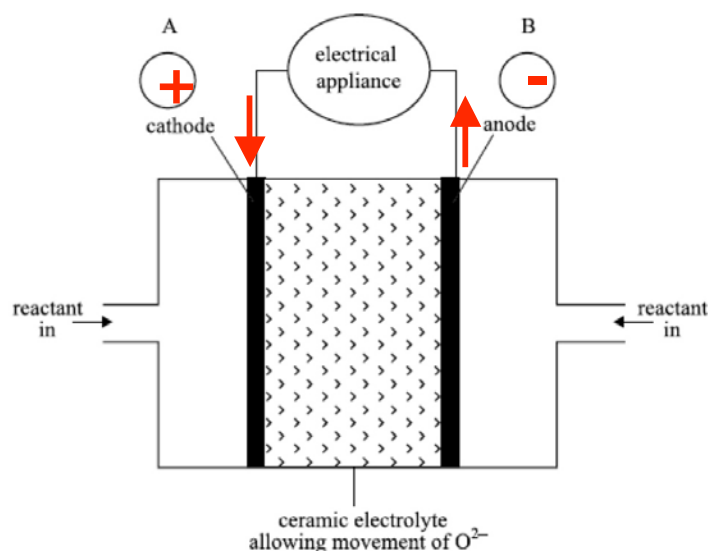
- a. The standard electrode potentials given in the electrochemical series predict the position of equilibrium for the redox reaction but they give no indication of the rate at which this equilibrium is achieved. In this case, the rate may be too slow to observe any reaction.
- b. When methane reacts with oxygen in a gas power station, a large amount of energy is lost as heat when the chemical energy is transformed through a series of steps to electrical energy. In the fuel cell the chemical energy of the methane is changed directly into electrical energy. The process is more efficient and less energy is lost as heat.
- c. From page 4 of the Data Book, it can be seen that water (-0.83 V) is a stronger oxidant than magnesium ions (-2.34 V) and, therefore, would be reduced in preference to the magnesium ions. It is important to note that all metals below water in the electrochemical series must be obtained from the electrolysis of **non-aqueous** electrolytes.
- d. The reactions occurring at the cathodes of these electrolytic cells are:



From the balanced equations, it can be seen that 2 mol of electrons ($2 \times 96500 \text{ C}$) is required to deposit 1 mol of Mg but only 1 mol of electrons ($1 \times 96500 \text{ C}$) is required to deposit 1 mol of K.

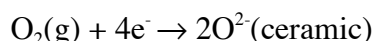
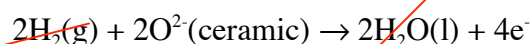
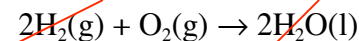
Question 7 [WEB LINK](#)

- a. In a galvanic cell such as a dry cell (primary) or lead-acid battery (secondary), the reactants are stored in the space occupied by the galvanic cell. The reactants are already present. In a fuel cell the reactants are supplied continuously from outside the fuel cell.
- b. Oxidation occurs at the anode and reduction occurs at the cathode. Electrons are produced at the anode and flow through the electrical appliance to the cathode. Hence, the anode is negative and the cathode is positive. This is shown on the diagram below.



Question 7 (continued)

- c. i. Hydrogen gas and oxygen gas are the reactants and water is the only product. Therefore, the overall cell reaction is $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l})$
- ii. The cathode reaction can be obtained by **subtracting** the anode reaction from the overall reaction as shown below.



- d. i. Electrical energy = $E = \text{voltage} \times \text{current} \times \text{time}$
 $E = VIt$
 $E = 0.600 \times 0.500 \times 10.0 \times 60$
 $E = 180 \text{ J}$
- ii. The charge produced by the cell = $Q = \text{current} \times \text{time}$
 $Q = It$
 $Q = 0.500 \times 10.0 \times 60$
 $Q = 300 \text{ C}$

iii.

$$n(\text{e}^-) \text{ produced} = \frac{Q}{F} = \frac{300}{96500} = 3.109 \times 10^{-3}$$

$$n(\text{H}_2) \text{ required} = \frac{1}{2} \times n(\text{e}^-)$$

$$= \frac{1}{2} \times 3.109 \times 10^{-3} \text{ mol}$$

$$= 1.554 \times 10^{-3} \text{ mol}$$

The cell is only 60.0% efficient. Therefore, MORE hydrogen will be needed than the calculated amount above.

$$n(\text{H}_2) \text{ consumed} = \frac{1.554 \times 10^{-3}}{0.600} = 2.59 \times 10^{-3} \text{ mol}$$

End of VCE Chemistry 2008 Sample Written Examination 2 Suggested Solutions

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