

Suggested Answers VCE Chemistry Year 12 Trial Exam Unit 4

Multiple Choice Answers – Section A

Question 1

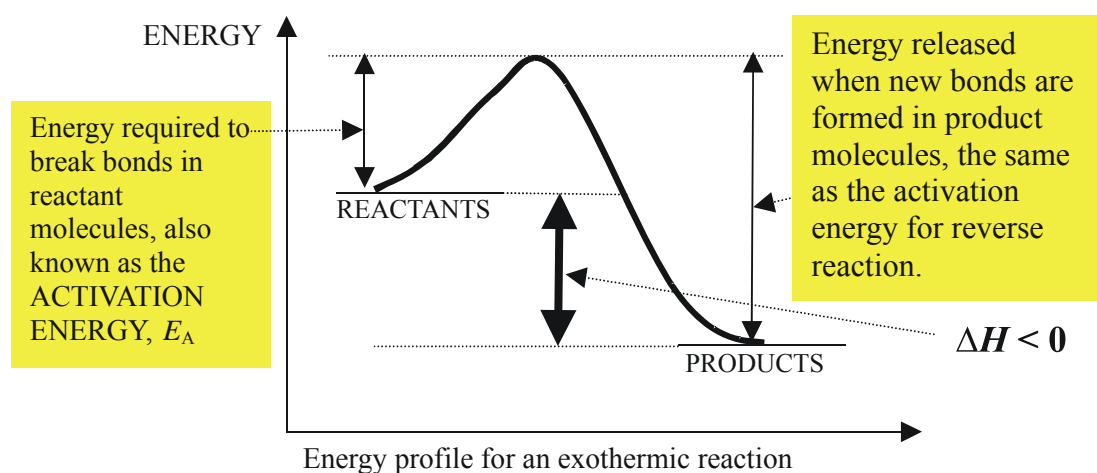
- D. With the volume of CO_2 levelling off with time, the rate of reaction is clearly decreasing. Since the surface area of the CaCO_3 particles must decrease (not increase) as it is consumed alternatives A and C are not viable. The rate of reaction decreases as the acid concentration decreases, i.e. becomes more dilute.

Question 2

- A. If X is favoured by an increase in pressure, then X must be on the side of the equilibrium with fewer particles. Hence, alternatives A or B must be correct. Equilibrium systems counteract an increase in pressure by moving the position of equilibrium to the side with fewer particles. If the amount of X increases with an increase in temperature, then the production of X from Y is an endothermic reaction and has $\Delta H > 0$, i.e. $3Y \rightleftharpoons 2X \quad \Delta H > 0$. This means that the production of Y from X is exothermic as is represented in the equation $2X \rightleftharpoons 3Y, \quad \Delta H < 0$. For this equilibrium an increase in pressure produces more X and an increase in temperature (favours endothermic reverse reaction) produces more X.

Question 3

- B. Exothermic reactions release energy and so the energy content of the product molecules is lower than the energy content of the reactant molecules. This is consistent with the **bonds (in total) within the product molecules** (2 H_2O molecules) being **stronger than the bonds (in total) within the reactant molecules** (2 H_2 molecules and 1 O_2 molecule). There is also evidence for this in that the activation energy for the reverse reaction is greater as indicated on the energy profile for an exothermic reaction.



Question 4

- B.** The concentration time graph shows that in getting to equilibrium
The substance represented by I has increased from 1 M to 4 M, i.e. by 3 M or 3 mol in 1 L.
The substance represented by II has increased from 2 M to 3 M, i.e. by 1 M or 1 mol in 1 L.
The substance represented by III has decreased from 4 M to 2M, i.e. by 2 M or 2 mol in 1 L.
This is consistent with 3 mol reacting with 1 mol to produce 2 mol,
i.e. $3Z + Y \rightarrow 2Z$
Y is best represented by **graph II**

Question 5

- A** $\text{CH}_4(\text{g}) + 2\text{H}_2\text{S}(\text{g}) \rightleftharpoons \text{CS}_2(\text{g}) + 4\text{H}_2(\text{g}); K = 3.59$

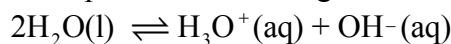
Comparison of the concentration fraction (*CF*), [or reaction quotient (*Q*)] based on the supplied data with the value of the equilibrium constant provides insight into the 'position' of the reaction.

$$\begin{aligned} CF &= [\text{CS}_2][\text{H}_2]^4 / [\text{CH}_4][\text{H}_2\text{S}]^2 \\ &= (0.755 \times 1.08^4) / (0.575 \times 0.600^2) \\ &= 1.027 / 0.207 \\ &= 4.96 \end{aligned}$$

Since the *CF* (4.96) is not equal to *K* (3.59), the reaction is not at equilibrium. For equilibrium to be established the value of the concentration fraction must decrease. This requires a decrease in the product concentrations relative to the reactant concentrations, so the **reaction must continue to the left** for equilibrium to be established.

Question 6

- C.** The self-ionisation of water proceeds according to



So in pure water $[\text{H}_3\text{O}^+] = [\text{OH}^-]$

$$\begin{aligned} \text{Since } [\text{H}_3\text{O}^+][\text{OH}^-] &= K_w \\ [\text{H}_3\text{O}^+]^2 &= 1.20 \times 10^{-15} \\ [\text{H}_3\text{O}^+] &= \sqrt{1.20 \times 10^{-15}} \\ &= 3.46 \times 10^{-8} \text{ M} \end{aligned}$$

$$\begin{aligned} \text{pH} &= -\log_{10}[\text{H}_3\text{O}^+] \\ &= -\log_{10}(3.46 \times 10^{-8}) \\ &= -(-7.46) \\ &= \mathbf{7.46} \end{aligned}$$

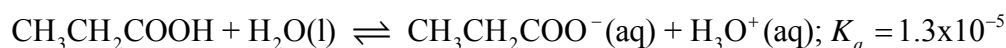
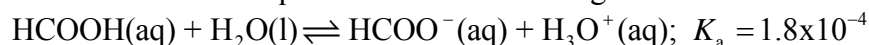
Question 7

- A.** According to the data book, the acidity constants for the two acids are

Methanoic acid – 1.8×10^{-4}

Propanoic acid – 1.3×10^{-5}

The acids ionise in aqueous solution according to



Since **HCOOH**(aq) has the higher acidity constant, it will ionise slightly more than **CH₃CH₂OH**(aq), thus producing a **higher *c*(H₃O⁺)** and a **lower pH**.

Question 8

- C. Consider the variations that occur in the equilibrium law expression as the equation stoichiometry is changed.

$$2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g});$$
$$K_1 = \frac{[\text{NO}_2]^2}{\{[\text{NO}]^2[\text{O}_2]\}}$$
$$= 7.10 \times 10^5$$

$$2\text{NO}_2(\text{g}) \rightleftharpoons 2\text{NO}(\text{g}) + \text{O}_2(\text{g}) - \textit{original equation is reversed}$$
$$K_2 = \frac{\{[\text{NO}]^2[\text{O}_2]\}}{[\text{NO}_2]^2}$$
$$= 1 / ([\text{NO}_2]^2 / \{[\text{NO}]^2[\text{O}_2]\})$$
$$= 1 / K_1$$
$$= 1 / 7.10 \times 10^5$$
$$= 1.41 \times 10^{-6}$$

$$\text{NO}_2(\text{g}) \rightleftharpoons \text{NO}(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) - \textit{original equation reversed and stoichiometry is halved.}$$
$$K_3 = \frac{\{[\text{NO}][\text{O}_2]^{1/2}\}}{[\text{NO}_2]}$$
$$= (\{[\text{NO}]^2[\text{O}_2]\} / [\text{NO}_2]^2)^{1/2}$$
$$= K_2^{1/2}$$
$$= \sqrt{K_2}$$
$$= \sqrt{(1.41 \times 10^{-6})}$$
$$= \mathbf{1.19 \times 10^{-3}}$$

Question 9

- B. Maximising energy efficiency is linked to ‘sustainability’ of energy resources and minimising waste is clearly a vital aspect of ‘waste management’. However all five strategies listed are amongst the principles of ‘green chemistry’.

Question 10

- A. For the equilibrium

$\text{CH}_4(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightleftharpoons \text{CO}(\text{g}) + 3\text{H}_2(\text{g}); \Delta H = +206.1 \text{ kJ mol}^{-1}$, the concentration of hydrogen in the equilibrium mixture (yield of H_2) will be increased by factors that move the position of equilibrium to the right.

- A. An increase in pressure will cause the position of equilibrium to move to the side with fewer particles, in this case to the left. Hence this **decreases the yield of H_2** .
- B. An increase in temperature will cause the position of equilibrium to move in the direction of the endothermic reaction, in this case to the right.
- C. Removing hydrogen will cause the position of equilibrium to move to the right to compensate for the loss of hydrogen.
- D. Increasing the amount of CH_4 in the reaction mixture will cause the position of equilibrium to move the right to compensate for the increase in $[\text{CH}_4]$.

Question 11

- B. According to the equation, the reaction is endothermic ($\Delta H > 0$) and 7.2 kJ of energy is absorbed from the solution when one mol $\text{Na}_2\text{S}_2\text{O}_3$ dissolves.

$$n(\text{Na}_2\text{S}_2\text{O}_3) \text{ dissolving} = m / M$$
$$= 15.8 \text{ g} / 158.2 \text{ g mol}^{-1}$$
$$= 0.0999 \text{ mol}$$

$$\text{Energy absorbed} = n(\text{Na}_2\text{S}_2\text{O}_3) \times \text{energy absorbed by one mole}$$
$$= 0.0999 \text{ mol} \times 7.2 \text{ kJ mol}^{-1}$$
$$= \mathbf{0.72 \text{ kJ}}$$

Question 12

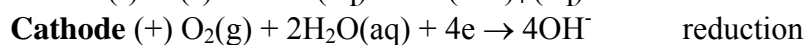
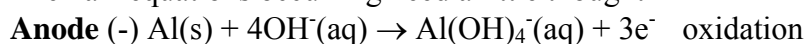
- D.** The key to this question was the direction of movement of the ions in the salt-bridge.
Anions – NO_3^- – always move towards the anode, which is the site of oxidation, i.e. Y
Cations – K^+ – always move towards the cathode, which is the site of reduction, i.e. X
Hence, reduction is occurring at electrode X, and **oxidation is occurring at electrode Y.**

In galvanic cells the electrons move spontaneously from the (-) electrode to the (+) electrode.

Question 13

- B.** According to the oxidation number **Al is oxidised** (from 0 in Al to +3 in $\text{Al}(\text{OH})_4^-$) and **O is reduced** (from 0 in O_2 to -2 in $\text{Al}(\text{OH})_4^-$)
Since oxidation occurs at the negative electrode (spontaneous reaction) then **Al** must be the **negative electrode.**

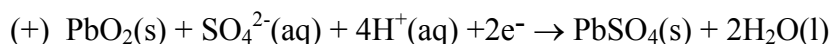
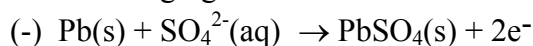
The half-equations occurring need a little thought



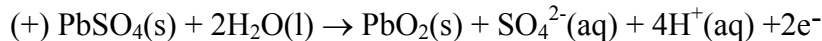
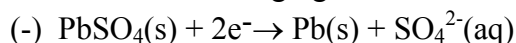
Multiply the oxidation half-equation by 4 and the reduction half-equation by 3 to get the overall redox equation $4\text{Al}(\text{s}) + 3\text{O}_2(\text{g}) + 6\text{H}_2\text{O}(\text{l}) + 4\text{OH}^-(\text{aq}) \rightarrow 4\text{Al}(\text{OH})_4^-(\text{aq}).$

Question 14

- A.** The discharging reactions with electrode signs are



Therefore the recharging reactions are



Checking the various alternatives

- A. Pb is produced at the (-) electrode
B. pH will decrease as H^+ ions are produced at (+) electrode
C. PbSO_4 is consumed
D. Oxidation number changes of Pb are from +2 to 0 and +2 to +4

Question 15

- C.** Useful facts about electrochemical cells.

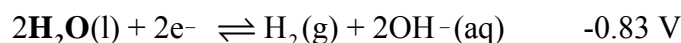
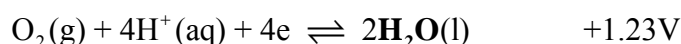
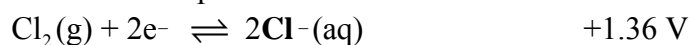
In all electrochemical cells electrons move from the site of oxidation (anode) to the site of reduction (cathode).

In galvanic cells electrons move spontaneously from the negative electrode to the positive electrode, whilst in electrolytic cells, electrons are forced to move from the negative electrode to the positive electrode.

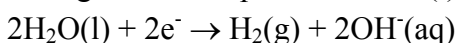
Reduction always occurs at the cathode, which is (+) in galvanic cells but (-) in electrolysis cells. Oxidation always occurs at the anode, which is (-) in galvanic cells but (+) in electrolysis cells.

Question 16

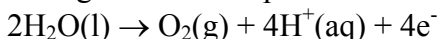
- B.** During electrolysis of dilute KCl(aq), the strongest oxidant present will be reduced at the cathode (-) and the strongest reductant present will be oxidised at the anode (+). The main species present prior to electrolysis are $K^+(aq)$, $Cl^-(aq)$ and $H_2O(l)$. Consider their positions on the relevant sections of the electrochemical series



Strongest oxidant present is $H_2O(l)$, so the reaction at the cathode (-) will be



Strongest reductant present is $H_2O(l)$, so the reaction at the anode (+) will be

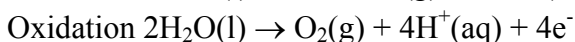
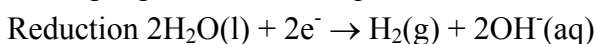


Therefore, alternatives B, C and D are not correct.

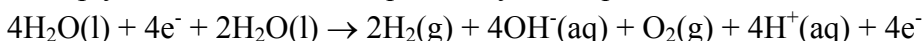
The **production of $OH^-(aq)$ ions** at the negative electrode will turn the **phenolphthalein indicator pink**.

Question 17

- D.** Adding together the half-equations identified in *Question 16*



Multiply the reduction half-equation by 2 to equalise the electrons



This simplifies to

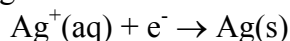


Which with $4H^+(aq) + 4OH^-(aq) \rightarrow 4H_2O$, further simplifies to



Question 18

- C.** Silver would be plated according to the reduction



$$\begin{aligned} n(\text{Ag}) \text{ plated} &= m / M \\ &= 0.300 \text{ mol} / 107.9 \text{ g mol}^{-1} \\ &= 2.78 \times 10^{-3} \text{ mol} \end{aligned}$$

$$n(e^-) \text{ required} = n(\text{Ag}) = 2.78 \times 10^{-3} \text{ mol}$$

$$\text{Since } Q = It \text{ and } Q = n(e^-) \times F$$

$$\text{Then } It = n(e^-) \times F$$

$$\begin{aligned} \text{and } t &= n(e^-) \times F / I \\ &= 2.78 \times 10^{-3} \times 96500 / 6.00 \\ &= \mathbf{44.7 \text{ seconds}} \end{aligned}$$

Question 19

- D. Using the specific heat of water, i.e. $4.18 \text{ J g}^{-1} \text{ }^\circ\text{C}^{-1}$ from data booklet
- $$\begin{aligned} \text{Energy released from 0.0100 mol of the alkanol} &= 4.18 \text{ J g}^{-1} \text{ }^\circ\text{C}^{-1} \times 200 \text{ g} \times 24.0^\circ\text{C} \\ &= 2.01 \times 10^4 \text{ J} \\ \text{Energy released per mol of the alkanol} &= 2.01 \times 10^4 \text{ J} / 0.0100 \text{ mol} \\ &= 2.01 \times 10^6 \text{ J mol}^{-1} \\ &= \mathbf{2.01 \times 10^3 \text{ kJ mol}^{-1}} \end{aligned}$$

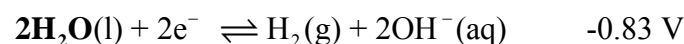
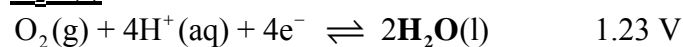
Hence the molar enthalpy of combustion (Table 13 – Data Book) must be close to 2010 kJ mol^{-1} .

The alkanol is **1-propanol** (2016 kJ mol^{-1})

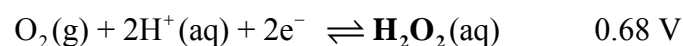
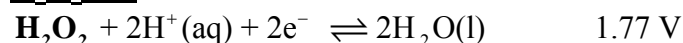
Question 20

- C. For a species to be able to act as either an oxidant or an oxidant it should appear on both sides of the electrochemical series.

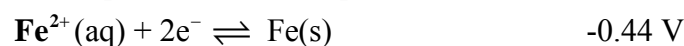
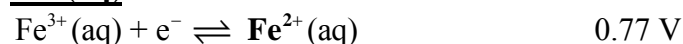
H₂O(l)



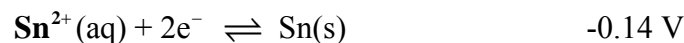
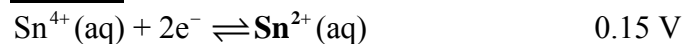
H₂O₂(aq)



Fe²⁺(aq)



Sn²⁺(aq)



Short Answer (Answers) - Section B

Question 1

NB In many parts of this question – particularly (e), (f) and (g) a variety of answers is possible.
Check your textbook and notes material for **your** 'selected chemical' for further options

	1.	2.	3.	4.
	$\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$	$2\text{NO} + \text{O}_2 \rightleftharpoons 2\text{NO}_2$	$\text{C}_3\text{H}_8 \rightleftharpoons \text{C}_2\text{H}_4 + \text{CH}_4$	$2\text{SO}_2 + \text{O}_2 \rightleftharpoons 2\text{SO}_3$
a. ①	ammonia	nitric acid	ethane	sulfuric acid
b. ①	exothermic	exothermic	endothermic	exothermic
c. ①①	low temperatures rate decreases at low temperatures 350-500°C and a porous iron catalyst are used.	low temperatures rate increases at low temperatures, so yield and rate both favoured. 30°C used.	high temperatures rate increases at high temperatures, so yield and rate both favoured. 750-900 °C used.	high temperatures rate decreases at low temperatures 400-500°C and a V ₂ O ₅ catalyst are used.
d. ①①	high pressures – system moves to side with fewer particles increasing yield of NH ₃	high pressures – system moves to side with fewer particles increasing yield of NO ₂ .	low pressures – system moves to side with more particles increasing yield of C ₂ H ₄	high pressures – system moves to side with fewer particles increasing yield of SO ₃
e. ①	waste heat used to preheat reaction mixture	maximising conversion of NO ₂ to nitric acid and minimising loss of NO / NO ₂ to the atmosphere	methane extracted and used as fuel for heating	SO ₃ emissions are minimised by using the double absorption process.
f. ①	NH ₃ is a severe, toxic irritant – workers must be protected against exposure	Concentrated HNO ₃ is corrosive and causes severe burns – workers must be protected against spills	ethene can form an explosive mixture with air – clear fire prevention and fighting strategies are used.	SO ₂ and SO ₃ are severe respiratory irritants. Work areas are well ventilated and protective clothing is used
g. ①	production of fertilisers	production of fertilisers	production of polymers	production of superphosphate.

Question 2

- a. The **energy released in the formation of the bonds in the products** [3380 - (-890) = 4270 kJ] is **greater than** the **energy used to break the bonds in the reactants** [3380 kJ]. ①① Hence the reaction is **exothermic**. ①

The profile also shows that the energy content of the products, i.e. 1 mol CO₂ and 2 mol H₂O, is lower than the energy content of the reactants, i.e. 1 mol CH₄ and 2 mol O₂, hence the reaction is exothermic.

- b. Activation energy for the reaction $\text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow \text{CH}_4(\text{g}) + 2\text{O}_2(\text{g})$ will be the difference between the 'top of the hill' and relative energy content of CO₂ and H₂O as given on the supplied profile.

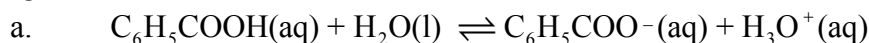
$$E_a = 3380 - (-890) \\ = 4270 \text{ kJ } \textcircled{1}$$

- c. Reaction between CH_4 and O_2 will only occur if the **collisions** between the molecules are fruitful / successful / effective, i.e. they collide with **energy greater than the activation energy**. ①
- d. A catalyst increases the reaction rate by providing an **alternative reaction pathway** which **lowers the activation energy**. Thus **increases the proportion of fruitful collisions**. ①
- However, the relative energy contents of the reactants and products does not change because **'difference' between the energy released in bond formation and the energy used in bond breaking does not change**. ①

Question 3

- a. Energy added during calibration, $E = VIt$
 $= 6.50 \times 2.90 \times 180$
 $= 3393 \text{ J}$ ①
- Temperature change during calibration, $\Delta T_c = 24.378 - 24.164$
 $= 0.224^\circ\text{C}$ ①
- Calibration factor = $E / \Delta T_c$
 $= 3393 \text{ J} / 0.224^\circ\text{C}$
 $= 1.51 \times 10^4 \text{ J } ^\circ\text{C}^{-1}$ ①
- b. Energy from combustion of C_5H_{12} = Calibration factor \times Temperature change during reaction
 $= 1.51 \times 10^4 \text{ J } ^\circ\text{C}^{-1} \times (24.164 - 21.250) ^\circ\text{C}$
 $= 1.51 \times 10^4 \times 2.914$ ①
 $= 4.41 \times 10^4 \text{ J}$
- $n(\text{C}_5\text{H}_{12}) = 1.44 / 72.0$
 $= 0.0200 \text{ mol}$ ①
- Energy per mol $\text{C}_5\text{H}_{12} = 4.41 \times 10^4 \text{ J} / 0.0200 \text{ mol}$
 $= 2.21 \times 10^6 \text{ J mol}^{-1}$
 $= 2.21 \times 10^3 \text{ kJ mol}^{-1}$
- Hence $\Delta H_c(\text{C}_5\text{H}_{12}) = -2.21 \times 10^3 \text{ kJ mol}^{-1}$ ①
- c. According to Data Booklet $\Delta H_c(\text{C}_5\text{H}_{12}) = -3509 \text{ kJ mol}^{-1}$
- The most likely reason for the lower calculated value would be **incomplete reaction of pentane** ① in the bomb calorimeter. This would lead to a lower temperature change during the reaction a lower calculated energy release from the combustion of pentane. The calculated value also depends on the accuracy of recorded readings of temperature, voltage, current and time, e.g. a higher than true recorded temperature during calibration would lead to a lower calibration factor and lower calculated energy release from the combustion of pentane.
- Other factors could include evaporation of some pentane during transfer to the reaction bomb.

Question 4



$$K_a = 6.4 \times 10^{-5} \text{ – from data booklet}$$

$$\text{pH} = 2.45 \rightarrow [\text{H}_3\text{O}^+] = 10^{-2.45} \\ = 3.58 \times 10^{-3} \quad \text{①}$$

Since $\text{C}_6\text{H}_5\text{COOH}(\text{aq})$ is a weak acid we may assume that

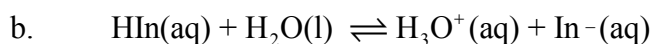
$$[\text{C}_6\text{H}_5\text{COO}^-] = [\text{H}_3\text{O}^+] \quad \text{①}$$

$$[\text{C}_6\text{H}_5\text{COOH}]_e = [\text{C}_6\text{H}_5\text{COOH}]_{\text{initially}}$$

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

$$6.4 \times 10^{-5} = \frac{(3.58 \times 10^{-3})^2}{[\text{C}_6\text{H}_5\text{COOH}]_e} \quad \text{①}$$

$$[\text{C}_6\text{H}_5\text{COOH}]_e = \frac{(3.58 \times 10^{-3})^2}{6.4 \times 10^{-5}} \\ = \frac{1.28 \times 10^{-5}}{6.4 \times 10^{-5}} \\ = \mathbf{0.20 \text{ M}} \quad \text{①}$$



$$K_a = 6 \times 10^{-5} \text{ – from data booklet}$$

$$K_a = \frac{[\text{H}_3\text{O}^+]_e \times [\text{In}^-]_e}{[\text{HIn}]_e}$$

$$6 \times 10^{-5} = \frac{[\text{H}_3\text{O}^+]_e \times 1 / 100}{[\text{HIn}]_e} \quad \text{①}$$

$$[\text{H}_3\text{O}^+]_e = 6 \times 10^{-3} \text{ M} \quad \text{①}$$

$$\text{pH} = -\log_{10}(6 \times 10^{-3})$$

$$= \mathbf{2.2} \quad \text{①}$$

- c. i. The increase in concentrations of both reactants and products at time t_1 suggests that the **volume** of the system was **decreased**. ①
- ii. The decrease in volume caused an overall pressure increase. According to Le Chatelier, the system **moves to the side with fewer particles** to counteract the pressure increase. The **position of equilibrium moves to the left**. ①
- iii. **less** ① the position of equilibrium has shifted left thus decreasing the amount of NO_2 present. The concentration time graph shows that the $[\text{NO}_2]$ at the new equilibrium is still greater than at the original equilibrium.

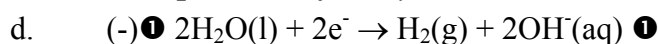
Question 5

a. The **use of crops that would normally be used in food production**. This may be driving up food prices. ①

b. **Production of CO_2** – contribution to **global warming** ①

SO_2 can also be released into the environment but in much smaller quantities than CO_2 .

c. photovoltaic cells which convert **solar energy** into electrical energy, or generators driven by **hydroelectricity**, **wind power**, **geothermal energy**, **nuclear energy** or **tidal power**. ① for any two.

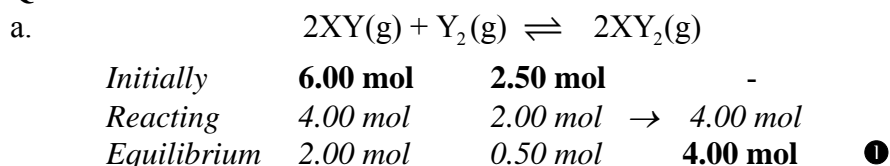


- e. In **pure water** the concentration of ions, $\text{H}_3\text{O}^+(\text{aq})$ and $\text{OH}^-(\text{aq})$ is very low and so the electrolytic conduction in the electrolysis cell is very low and so the **rate of reaction** will be **low**. ❶
 In dilute $\text{NaCl}(\text{aq})$ the **$\text{Na}^+(\text{aq})$ and $\text{Cl}^-(\text{aq})$ enhance the electrolytic conduction between the electrodes** and so the rate of production of hydrogen increases.
- f. $\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^- \rightarrow 4\text{OH}^-(\text{aq})$ ❶ *from electrochemical series.*
- g. The electrodes (1) must conduct, (2) be porous to allow contact between the gases and the electrolyte, (3) must be unreactive, (4) are usually impregnated with catalyst to speed up the rate of reaction. ❶❶ *for three properties, ❶ for two properties*
- h. $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l})$ ❶
- i. Water is the only product of its use. ❶

Question 6

- a. The position of fluorine on the electrochemical series shows that it is the **strongest oxidant**. Since electrodes must conduct electricity, and **metals are reductants**, the choice of electrode material is restricted to carbon (graphite) because fluorine will react explosively with most metals. ❶
- b. (+) $2\text{F}^-(\text{l}) \rightarrow \text{F}_2(\text{g}) + 2\text{e}^-$ ❶
 (-) $\text{K}^+(\text{l}) + \text{e}^- \rightarrow \text{K}(\text{l})$ ❶
In electrolysis electrons move from the (+) electrode to the (-) electrode.
- c. $2\text{F}^-(\text{l}) \rightarrow \text{F}_2(\text{g}) + 2\text{e}^-$
 $n(\text{F}_2) = 5.00 \text{ g} / 38.0 \text{ g mol}^{-1}$
 $= 0.1316 \text{ mol}$
 $n(\text{e}^-) = 2 \times 0.1316$
 $= 0.263 \text{ mol}$ ❶
 Since $Q = It$ and $Q = n(\text{e}^-)F$ then
 $It = n(\text{e}^-) \times F$
 $t = n(\text{e}^-) \times F / I$
 $= 0.263 \times 96500 / 5.00$ ❶
 $= 5076 \text{ seconds}$
 $= 84.6 \text{ minutes}$ ❶
- d. In $\text{KF}(\text{aq})$ there are two reductants – $\text{F}^-(\text{aq})$ and $\text{H}_2\text{O}(\text{l})$
 According to the electrochemical series
 $\text{F}_2(\text{g}) + 2\text{e}^- \rightleftharpoons 2\text{F}^-(\text{aq}) \quad 2.87 \text{ V}$
 $\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \rightleftharpoons 2\text{H}_2\text{O}(\text{l}) \quad 1.23 \text{ V}$
 $\text{H}_2\text{O}(\text{l})$ is a stronger reductant than $\text{F}^-(\text{aq})$ and so $\text{H}_2\text{O}(\text{l})$ will be oxidised ❶, according to $2\text{H}_2\text{O}(\text{l}) \rightarrow \text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^-$ ❶..... at the anode (+)
- e. $\text{Cu}^{2+}(\text{aq})$, $\text{Li}^+(\text{aq})$ and $\text{H}_2\text{O}(\text{l})$ are all oxidants.
 According to the electrochemical series
 $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Cu}(\text{s}) \quad 0.34 \text{ V}$
 $2\text{H}_2\text{O}(\text{l}) + 2\text{e}^- \rightleftharpoons \text{H}_2(\text{g}) + 2\text{OH}^-(\text{aq}) \quad -0.83 \text{ V}$
 $\text{Li}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Li}(\text{s}) \quad -3.02 \text{ V}$
 $\text{Cu}^{2+}(\text{aq})$ is a stronger oxidant than $\text{H}_2\text{O}(\text{l})$, so $\text{Cu}^{2+}(\text{aq})$ will be reduced to Cu ❶, but $\text{H}_2\text{O}(\text{l})$ is a stronger oxidant than Li^{2+} , so H_2O will be reduced and no Li will be produced ❶

Question 7



At equilibrium, $[\text{XY}] = 2.00 / 2 = 1.00 \text{ M}$

$[\text{Y}_2] = 0.50 / 2 = 0.25 \text{ M}$

$[\text{XY}_2] = 4.00 / 2 = 2.00 \text{ M}$ ①

$$\begin{aligned} K &= [\text{XY}_2]^2 / ([\text{XY}]^2 \times [\text{Y}_2]) \\ &= (2.00)^2 / ((1.00)^2 \times 0.25) \\ &= 16 \text{ ① (M}^{-1}\text{)} \end{aligned}$$

b. $2\text{XY}(\text{g}) + \text{Y}_2(\text{g}) \rightarrow 2\text{XY}_2(\text{g})$ is **exothermic**. ① Since less XY_2 is present at the higher temperature, the **forward reaction is not favoured by high temperatures**.

c. If the pressure exerted by the components of the equilibrium system increases, then according to Le Chatelier's principle the position of equilibrium should shift to the right to produce fewer particles. ① However since Neon is an inert gas and not part of the equilibrium system, it does not effect the pressure exerted by the equilibrium system, and so the position of equilibrium does not shift. ①

Alternatively the concentration fraction (reaction quotient) $(\text{XY}_2)^2 / ([\text{XY}]^2 \times [\text{Y}_2])$ is not affected by the addition of the non-reacting Neon and so the system is not pushed out of equilibrium.

- d.
- Position of equilibrium would shift to the right ① to counteract the loss of XY_2 .
 - The decrease in $[\text{XY}_2]$ due to its removal reduces the rate of the reverse reaction relative to the forward reaction. ① Hence the forward reaction dominates, and position of equilibrium shifts to the right, until the rates of forward and reverse reactions are again equal at the new equilibrium position. ①

End of Trial Exam