

Suggested Answers

VCE Chemistry 2009 Year 12 Trial Exam Unit 4

Multiple Choice Answers - Section A

Question 1

- C. The graph shows that the quantity on the vertical axis increases with time, but the rate of increase – as represented by the gradient of the curve – decreases with time. Since HCl(aq) is a strong acid, the reaction may be represented by the ionic equation $\text{Zn(s)} + 2\text{H}^+(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{H}_2(\text{g}); \Delta H = -154 \text{ kJ mol}^{-1}$. Consider the alternatives:
- A. As the reaction proceeds, the number of collisions per second between the reactant particles will decrease as the $[\text{H}^+]$ decreases.
 - B. The rate of reaction decreases with time as the $[\text{H}^+]$ decreases.
 - C. Since the reaction is **exothermic** the **temperature of the solution will increase**, rapidly initially and then more slowly as the rate of reaction decreases.
 - D. Number of ions decreases with time as H^+ ions are converted into H_2 molecules.

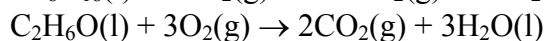
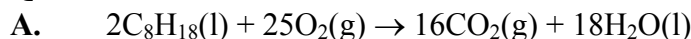
Question 2

- D. According to oxidation numbers, **Zn is oxidised** (from 0 in Zn to +2 in Zn(OH)_2) and **O is reduced** (from 0 in O_2 to -2 in Zn(OH)_2). Since oxidation occurs at the negative electrode when a galvanic cell is converting chemical energy to electrical energy, then **Zn must be the negative electrode**. The half-equations can be deduced from the electrochemical series
- Anode (-)** $\text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$ oxidation
- Cathode (+)** $\text{O}_2(\text{g}) + 2\text{H}_2\text{O(l)} + 4\text{e}^- \rightarrow 4\text{OH}^-(\text{aq})$ reduction
- Multiply the oxidation half-equation by 2 and add the reduction half-equation to get the overall equation $2\text{Zn(s)} + \text{O}_2(\text{g}) + 2\text{H}_2\text{O(l)} \rightarrow 2\text{Zn}^{2+}(\text{aq}) + 4\text{OH}^-(\text{aq}) \rightarrow 2\text{Zn(OH)}_2(\text{s})$

Question 3

- D. The rate of a chemical reaction is determined by the **proportion of reactant molecule collisions that have energy greater than the activation energy for the reaction**. This proportion is reflected, on a kinetic energy distribution curve, as the **area under the curve to the right of the activation energy**. This area is **greatest for reaction 4**, hence this reaction should occur at the fastest rate.

Question 4



Energy released = 2000 kJ

According to the Data Book

$\Delta H_c(\text{C}_8\text{H}_{18}) = -5464 \text{ kJ mol}^{-1}$; $\Delta H_c(\text{C}_2\text{H}_6\text{O}) = -1364 \text{ kJ mol}^{-1}$

$n(\text{C}_8\text{H}_{18}) \text{ reacting} = 2000 \text{ kJ} / 5464 \text{ kJ mol}^{-1}$
 $= 0.3660 \text{ mol}$

$n(\text{CO}_2) \text{ produced from C}_8\text{H}_{18} = (16/2) \times n(\text{C}_8\text{H}_{18})$
 $= 8 \times 0.3660$
 $= \mathbf{2.928 \text{ mol}}$

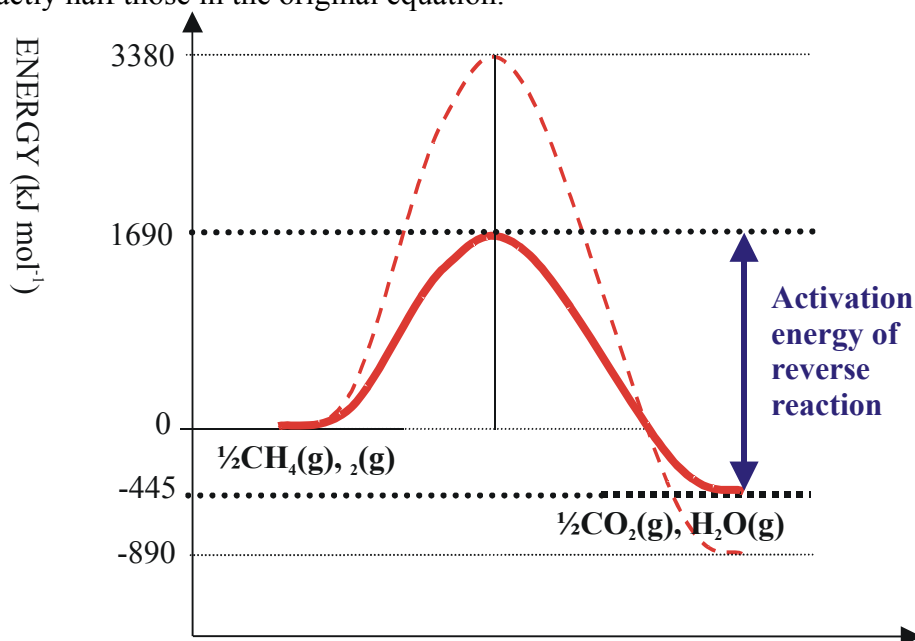
$n(\text{C}_2\text{H}_6\text{O}) \text{ reacting} = 2000 \text{ kJ} / 1364 \text{ kJ mol}^{-1}$
 $= 1.466 \text{ mol}$

$n(\text{CO}_2) \text{ produced from C}_2\text{H}_6\text{O} = 2 \times n(\text{C}_2\text{H}_6\text{O})$
 $= 2 \times 1.466$
 $= \mathbf{2.933 \text{ mol}}$

Ratio $n(\text{CO}_2) \text{ produced from C}_8\text{H}_{18} : n(\text{CO}_2) \text{ produced from C}_2\text{H}_6\text{O} = 2.928 : 2.933$
 $= 2.928/2.933 : 1$
 $= \mathbf{1 : 1}$

Question 5

- B. The energy profile for $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$ can be adjusted to give the energy profile for $\frac{1}{2}\text{CH}_4(\text{g}) + \text{O}_2(\text{g}) \rightarrow \frac{1}{2}\text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g})$, by halving all the values on the original energy profile, and the ΔH value, since all the mole ratios are exactly half those in the original equation.



From the energy profile for $\frac{1}{2}\text{CH}_4(\text{g}) + \text{O}_2(\text{g}) \rightarrow \frac{1}{2}\text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g})$, the activation energy for $\frac{1}{2}\text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightarrow \frac{1}{2}\text{CH}_4(\text{g}) + \text{O}_2(\text{g})$, i.e. the reverse reaction is $1690 - (-445) = \mathbf{2135 \text{ kJ}}$

Alternatively:

From the energy profile for $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$
the activation energy for the reaction $\text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g}) \rightarrow \text{CH}_4(\text{g}) + 2\text{O}_2(\text{g})$ is
 $3380 - (-890) = 4270 \text{ kJ mol}^{-1}$.

Hence the activation energy for the reaction $\frac{1}{2}\text{CH}_4(\text{g}) + \text{O}_2(\text{g}) \rightarrow \frac{1}{2}\text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g})$
is $4270 / 2 = \mathbf{2135 \text{ kJ mol}^{-1}}$.

Question 6

- C. In a galvanic cell electrons move spontaneously from the anode (-) to the cathode (+).
So the electrode signs indicate that oxidation is occurring in the Fe^{2+}/Fe half-cell i.e.
 $\text{Fe}(\text{s}) \rightarrow \text{Fe}^{2+}(\text{aq}) + \text{e}^-$.

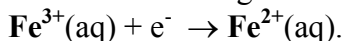
Since electrons leave the site of oxidation and go to the site of reduction, the left-hand half-cell must contain an oxidant which is 1.21 V higher up on the electrochemical series than the reductant $\text{Fe}(\text{s})$.

Scanning the electrochemical series and looking for a difference of 1.21 V indicates that the oxidant must be $\text{Fe}^{3+}(\text{aq})$, according to



$$E(\text{cell}) = 0.77 - (-0.44) \\ = 1.21 \text{ V}$$

Reaction occurring in the half-cell 1 is



Since all the species in the half-equation are ions, a platinum electrode is used in the half-cell.

Alternatively:

Since $E(\text{cell}) = E^\circ(\text{oxidant half-cell}) - E^\circ(\text{reductant half-cell})$

$$1.21 \text{ V} = E^\circ(\text{oxidant}) - E^\circ(\text{Fe}^{2+}(\text{aq})/\text{Fe}(\text{s}))$$

$$= E^\circ(\text{oxidant}) - (-0.44 \text{ V})$$

$$= E^\circ(\text{oxidant}) + 0.44 \text{ V}$$

$$\text{Hence } E^\circ(\text{oxidant}) = 1.21 - 0.44 \\ = 0.77 \text{ V}$$

From the electrochemical series $\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Fe}^{2+}(\text{aq}); E^\circ = +0.77 \text{ V}$

Question 7

- C. To decide whether the system is or is not at equilibrium, it is necessary to calculate the value of the reaction quotient (Q) and compare it with the equilibrium constant.

$$Q = [\text{NO}]^2 / \{[\text{N}_2] \times [\text{O}_2]\} \\ = (4.2 \times 10^{-3})^2 / \{0.50 \times 0.25\} \\ = 1.4 \times 10^{-4}$$

Q (1.4×10^{-4}) is less than K_c (1.7×10^{-3}) so, for the system to get to equilibrium, Q must increase until its value is the same as the equilibrium constant.

Since Q is effectively the ratio of product concentrations to reactant concentrations, the reaction must proceed in the forward direction to give more products and less reactants. In order for the reaction to move to the right, the **rate of the forward reaction must be greater than the rate of the reverse reaction.**

Irrespective of the reaction direction, or if the system is at equilibrium, the pressure in the container will remain the same because the total number of particles present does not change.

Question 8

D. Calibration Factor = Energy added during heating / ΔT_c
= $VIt / \Delta T_c$
= $6.00 \times 1.40 \times 54.5 / 0.390$
= $457.8 / 0.390$
= $1174 \text{ J } ^\circ\text{C}^{-1}$

Energy released by hydrocarbon = $CF \times \Delta T_r$
= $1174 \times (22.17 - 21.45)$
= $1174 \text{ J } ^\circ\text{C}^{-1} \times 0.72 ^\circ\text{C}$
= 845 J

Energy per mol hydrocarbon = $845 \text{ J} / 9.50 \times 10^{-4} \text{ mol}$
= $8.90 \times 10^5 \text{ J}$
= $8.90 \times 10^2 \text{ kJ}$
= **890 kJ**

So the hydrocarbon has a molar enthalpy (ΔH_c) of combustion of **-890 kJ mol⁻¹**.
According to Table 13 in the Data Book. $\Delta H_c(\text{methane}) = \text{-889 kJ mol}^{-1}$

Question 9

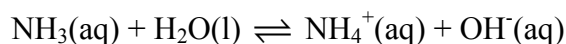
- A. The supplied data show clearly that the ‘**yield of product**’ **increases**
– as the **temperature decreases**
– as the **pressure increases**
If the yield is increasing as the temperature is decreasing, the **forward reaction** is favoured by low temperature, i.e. is **exothermic**.
If the yield is increasing as the pressure increases, then since the equilibrium system responds to a pressure increase, by striving to counteract the pressure increase by moving to decrease the pressure, the position of equilibrium is pushed to the **side with fewer particles**, so the **product** must be on that side.

Question 10

- D. Fluorine is produced by the electrolysis of salts containing the fluoride ions. However according to the electrochemical series, particularly
 $\text{F}_2(\text{g}) + 2\text{e}^- \rightarrow 2\text{F}^-(\text{aq}) \quad +2.87 \text{ V}$
 $\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}(\text{l}) \quad +1.23 \text{ V}$
aqueous solutions of fluoride salts cannot be used because **H₂O(l) is a stronger reductant than F⁻(aq) and would be preferentially oxidised**.
Fluorine is produced by the electrolysis of molten salts such as KF(l), for which the electrode half-equations are
(+) oxidation: $2\text{F}^-(\text{l}) \rightarrow \text{F}_2(\text{g}) + 2\text{e}^-$
(-) reduction: $\text{K}^+(\text{l}) + \text{e}^- \rightarrow \text{K}(\text{l})$

Question 11

- A. The addition of the 20 mL of 0.10 M $\text{NH}_3(\text{aq})$ to 50 mL of water pushes the equilibrium



to the right.

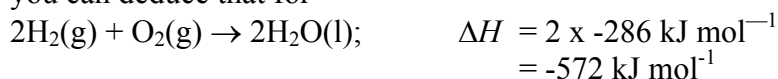
This is consistent with Le Chatelier's principle – the system moves to compensate for the extra reactant, H_2O , by favouring the forward reaction until equilibrium is again reached.

So the **number of NH_4^+ ions increases** and the number of OH^- ions (the link to pH) also increases.

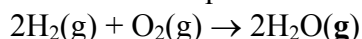
However since the total volume has increased, the **increase in $n(\text{OH}^-)$ will not be enough to compensate for the volume increase**. The $[\text{OH}^-]$ will be **lower** at the new equilibrium and consequently the **pH will be lower**.

Question 12

- C. Given $\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l}); \Delta H = -286 \text{ kJ mol}^{-1}$
you can deduce that for



However the equation for which the ΔH is required is

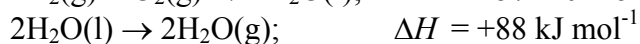
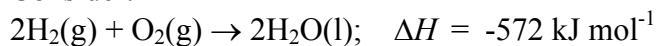


The **formation of $\text{H}_2\text{O}(\text{g})$ from $\text{H}_2(\text{g})$ and $\text{O}_2(\text{g})$ releases less energy per mole than the formation of $\text{H}_2\text{O}(\text{l})$** . This is consistent with the fact that we add energy to boil water, i.e. the reaction $\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{O}(\text{g})$ is endothermic, and hence $\text{H}_2\text{O}(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l})$ is exothermic.

So for $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g}); \Delta H = -484 \text{ kJ mol}^{-1}$ is more accurate.

It is also consistent with $\text{H}_2\text{O}(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l}); \Delta H = -44 \text{ kJ mol}^{-1}$

Consider:

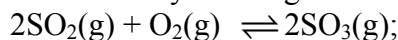


Hence

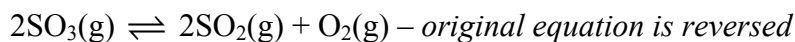


Question 13

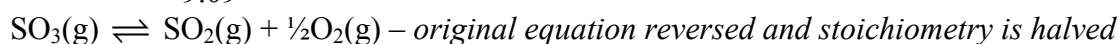
- B. Consider the variations that occur in the equilibrium law expression as the equation stoichiometry is changed.



$$K_1 = [\text{SO}_3]^2 / \{[\text{SO}_2]^2[\text{O}_2]\} \\ = 0.11$$



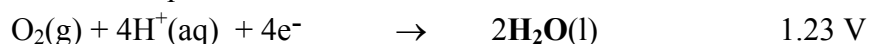
$$K_2 = \{[\text{SO}_2]^2[\text{O}_2]\} / [\text{SO}_3]^2 \\ = 1 / ([\text{SO}_3]^2 / \{[\text{SO}_2]^2[\text{O}_2]\}) \\ = 1 / K_1 \\ = 1 / 0.11 \\ = 9.09$$



$$K_3 = \{[\text{SO}_2][\text{O}_2]^{1/2}\} / [\text{SO}_3] \\ = (\{[\text{SO}_2]^2[\text{O}_2]\} / [\text{SO}_3]^2)^{1/2} \\ = K_2^{1/2} \\ = \sqrt{K_2} \\ = \sqrt{9.09} \\ = 3.0$$

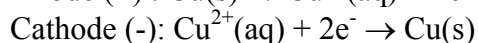
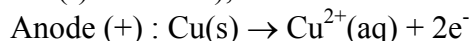
Question 14

- A. The relevant half-equations from the electrochemical series are



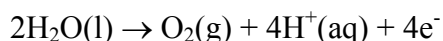
Because electroplating is electrolysis, oxidation occurs at the (+) electrode and reduction occurs at the (-) electrode.

On the basis of the reaction between the strongest oxidant and strongest reductant principle (and the clue of $\text{Cu}^{2+}(\text{aq})$ ions released at the (+) electrode and consumed at the (-) electrode), the reactions occurring in the normal copper-plating cell are



If the copper anode, i.e. the (+) electrode, is replaced with a carbon electrode, the only reductant present at the carbon anode will be H_2O .

So instead of Cu being oxidised at the (+) electrode, H_2O will be oxidised according to



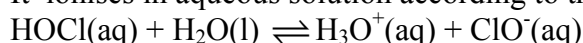
Thus bubbles of oxygen gas will be produced at the (+) electrode. Also, because of the increase in $[\text{H}^+]$ in the solution, the **pH** will **decrease**.

$\text{Cu}(\text{s})$ would still be produced at the (-) electrode until the $c(\text{Cu}^{2+})$ has dropped to an unworkable level.

Question 15

- D. Hypochlorous acid, HOCl(aq), is listed in Table 12 of the Data Book as a weak acid with $K_a = 2.9 \times 10^{-8}$.

It ionises in aqueous solution according to the equilibrium



For which $K_a = [\text{H}_3\text{O}^+]_e [\text{ClO}^-]_e / [\text{HOCl}]_e$

Since 20 mL of 0.050 M HOCl(aq) is added to 40 mL of water, we need to determine the initial concentration of the diluted solution

$$\begin{aligned} n(\text{HOCl}) \text{ in } 20 \text{ mL } 0.050 \text{ M} &= cV \\ &= 0.050 \times 20 \times 10^{-3} \\ &= 1.0 \times 10^{-3} \text{ mol} \\ c(\text{HOCl}) \text{ in } 60 \text{ mL solution} &= n(\text{HOCl}) / V \\ &= 1.0 \times 10^{-3} / 60 \times 10^{-3} \\ &= 0.0167 \text{ mol L}^{-1} \end{aligned}$$

Assumptions for a weak acid

1. Limited ionisation of HOCl, so that $[\text{HOCl}]_e = [\text{HOCl}]_{\text{initially}} = 0.0167 \text{ M}$

2. $[\text{H}_3\text{O}^+]_e = [\text{ClO}^-]_e$

Since $K_a = 2.9 \times 10^{-8}$

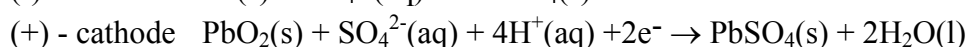
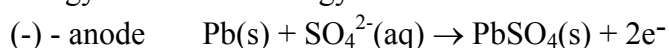
$$\begin{aligned} 2.9 \times 10^{-8} &= [\text{H}_3\text{O}^+]_e [\text{ClO}^-]_e / [\text{HOCl}]_e \\ &= [\text{H}_3\text{O}^+]^2 / 0.0167 \end{aligned}$$

$$\begin{aligned} 0.0167 \times 2.9 \times 10^{-8} &= [\text{H}_3\text{O}^+]^2 \\ [\text{H}_3\text{O}^+]^2 &= 4.84 \times 10^{-10} \\ [\text{H}_3\text{O}^+] &= \sqrt{4.84 \times 10^{-10}} \\ &= 2.20 \times 10^{-5} \end{aligned}$$

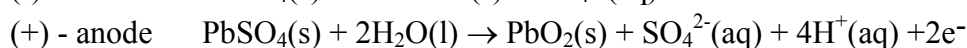
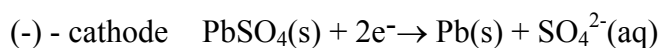
$$\begin{aligned} \text{pH} &= -\log_{10} [\text{H}_3\text{O}^+] \\ &= -\log_{10} (2.20 \times 10^{-5}) \\ &= -(-4.7) \\ &= 4.7 \end{aligned}$$

Question 16

- A. The electrode signs for the discharging reactions in each cell which convert chemical energy to electrical energy are



So the reactions occurring at these electrodes during the recharging reactions in each cell 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 numbers of Pb change from +2 to +4 at the anode, and +2 to 0 at the cathode

Question 17

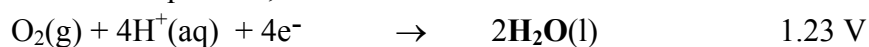
B. Changes that will increase the yield of SO_2Cl_2 by pushing the position of equilibrium to the right and are consistent with Le Chatelier's principle include:

- **decreasing the temperature**, since the forward reaction is exothermic
- **increasing the pressure / decreasing the volume**, since the system moves to produce fewer particles in the larger volume
- **increasing reactant concentration by adding a reactant**
- **decreasing product concentration by removing a product**

However whilst a **catalyst** speeds up the rate of both the forward and reverse reactions, by lowering their respective activation energies, it has **no effect on the position of equilibrium**, hence no effect on the yield of SO_2Cl_2 .

Question 18

C The relevant half-equations, from the electrochemical series are



$$\begin{aligned} n(\text{Cu}^{2+}) \text{ in cell before electrolysis} &= c \times V \\ &= 0.10 \times 200 \times 10^{-3} \\ &= 0.020 \text{ mol} \end{aligned}$$

$$\begin{aligned} n(\text{e}^-) \text{ passed through the cell} &= Q / F \\ &= 9650 / 96500 \\ &= 0.1000 \text{ mol} \end{aligned}$$

Strongest oxidant present in the cell is $\text{Cu}^{2+}(\text{aq})$, and the strongest reductant present is $\text{H}_2\text{O}(\text{l})$. So the reactions at the electrodes would be

Anode (+) reaction: $2\text{H}_2\text{O}(\text{l}) \rightarrow \text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^-$.

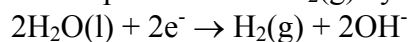
$$\begin{aligned} n(\text{O}_2) \text{ produced} &= n(\text{e}^-) / 4 \\ &= \mathbf{0.025 \text{ mol}} \end{aligned}$$

Cathode (-) reaction: $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s})$

$$\begin{aligned} n(\text{Cu}) \text{ produced} &= n(\text{Cu}^{2+}) \\ &= 0.020 \text{ mol} \end{aligned}$$

$$\begin{aligned} n(\text{e}^-) \text{ used to produce Cu} &= 2 \times n(\text{Cu}) \text{ produced} \\ &= 0.040 \text{ mol} \end{aligned}$$

However this leaves $0.1000 - 0.040 = 0.060 \text{ mol e}^-$ for another cathode reaction, i.e. for the production of $\text{H}_2(\text{g})$ by the reduction of H_2O .



$$\begin{aligned} n(\text{H}_2) \text{ produced} &= \frac{1}{2} \times n(\text{e}^-) \text{ used} \\ &= \frac{1}{2} \times 0.060 \\ &= \mathbf{0.030 \text{ mol}} \end{aligned}$$

So the maximum amounts of gaseous products that could be formed would be **0.025 mol O_2** (at the anode) and **0.030 mol H_2** (at the cathode).

Question 19

- B.** If the reaction $2\text{CH}_3\text{OH}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{CH}_2\text{O}(\text{g}) + 2\text{H}_2\text{O}(\text{g})$ proceeded to completion, 2 mol CH_3OH reacts with 1 mol O_2 to produce 2 mol CH_2O .
Hence if 3 mol CH_3OH was mixed with 3 mol O_2 and the reaction proceeded to completion, 3 mol CH_3OH would react with 1.5 mol O_2 to produce 3 mol CH_2O .
However this would result in no CH_3OH remaining, which cannot be the case if the system reaches equilibrium. Hence when equilibrium has been reached, less than 3 mol CH_3OH must have been converted to CH_2O , **and less than 3 mol CH_2O would be present at equilibrium.**
Although 1.5 mol is less than 3 mol, alternative A is not correct because there was not enough information supplied to determine exactly how much CH_2O is present at equilibrium.

Question 20

- A.** The fundamental difference in a nuclear power station is that the first energy change is the **conversion of nuclear energy (in uranium nuclei) to thermal energy.**
The energy change happens as **large uranium nuclei are split into smaller nuclei of different elements.**
In the conversion of **chemical energy to thermal energy**, the atoms are rearranged as **chemical bonds are broken and formed, but there is no change in the elements present.**
In a nuclear power station, the energy changes are
1. nuclear to thermal – in the **reactor**
 2. thermal to thermal – **production of steam** by boiling water
 3. thermal to mechanical – steam driving the **turbine**
 4. mechanical to electrical – in the **generator**

Short Answer (Answers) - Section B

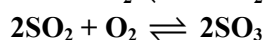
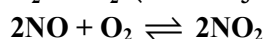
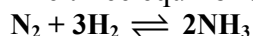
Question 1

- a. i. The CPOX reaction is **exothermic** ($\Delta H < 0$), whereas the SMR process is **endothermic** ($\Delta H > 0$). ❶
Hence the **heating costs should be less with the CPOX process**. Also the **heat released in the CPOX process can be used elsewhere in the process**, e.g. to heat incoming gases. ❶
- ii. Because the CPOX equilibrium is exothermic, the **forward reaction** and yield of Syngas is **favoured by low temperatures**. ❶ However since reaction rates are slower at lower temperatures then **the rate of production of Syngas may be too slow**. ❶
- iii. **Lower pressure**. ❶ Since both processes have more particles on the product side of the equilibrium, the position of equilibrium will shift to the right as the system adjusts to a pressure decrease.

iv. v.

❶ Ammonia – NH_3	Sulfuric Acid – H_2SO_4	Ethene – C_2H_4	Nitric Acid – HNO_3
❶ $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$	$2\text{SO}_2 + \text{O}_2 \rightleftharpoons 2\text{SO}_3$	$\text{C}_3\text{H}_8 \rightleftharpoons \text{C}_2\text{H}_4 + \text{CH}_4$	$2\text{NO} + \text{O}_2 \rightleftharpoons 2\text{NO}_2$
❶ Lower temperature	Lower temperature	Decrease pressure	Lower temperature

The three equilibria



like the CPOX equilibrium all have $\Delta H < 0$ and so the forward reaction is favoured by low temperature. However unlike the CPOX equilibrium, the forward reaction is favoured by high pressure.

The equilibrium $\text{C}_3\text{H}_8 \rightleftharpoons \text{C}_2\text{H}_4 + \text{CH}_4$ has $\Delta H > 0$ and so the forward reaction is favoured by high temperature. However, like the CPOX equilibrium it has more particles on the product side and so the forward reaction is favoured by low pressure.

- vi. $\text{CO(g)} + 2\text{H}_2\text{(g)} \rightleftharpoons \text{CH}_3\text{OH(g)}$ ❶
- b. i. $\text{CH}_3\text{OH(l)} + \text{H}_2\text{O(l)} \rightarrow \text{CO}_2\text{(g)} + 6\text{H}^+\text{(aq)} + 6\text{e}^-$ ❶
- ii. $\text{O}_2\text{(g)} + 2\text{H}^+\text{(aq)} + 2\text{e}^- \rightarrow 2\text{H}_2\text{O(l)}$ ❶
- iii. **Liquid methanol is easier to transport and store** than gaseous hydrogen. ❶
There is also a significant issue with the ‘explosiveness’ of hydrogen

Question 2

- a. From Data Book; $\Delta H_c(\text{ethanol}) = -1364 \text{ kJ mol}^{-1}$
 $n(\text{C}_2\text{H}_5\text{OH})$ used in calibration = $5.000 \text{ g} / 46.0 \text{ g mol}^{-1}$
 = 0.109 mol
 Energy released = $0.109 \text{ mol} \times 1364 \text{ kJ mol}^{-1}$
 = 148 kJ ❶
 Temperature change = $53.19 - 23.11$
 = 30.08°C ❶
 Calibration factor = $148 \text{ kJ} / 30.08^\circ\text{C}$
 = $4.93 \text{ kJ } ^\circ\text{C}^{-1}$ ❶

- b. Energy released = calorimeter constant x temperature change
 $= 4.93 \text{ kJ } ^\circ\text{C}^{-1} \times (41.6 - 23.1)^\circ\text{C}$ ❶
 $= 4.93 \times 18.5 \text{ kJ}$
 $= \mathbf{91.2 \text{ kJ}}$ ❶
- c. $5.865 \text{ g C}_3\text{H}_6\text{O}_3 \rightarrow 91.2 \text{ kJ}$
 To get ΔH you need to determine the energy released per mol $\text{C}_3\text{H}_6\text{O}_3$
 $n(\text{C}_3\text{H}_6\text{O}_3) = m(\text{C}_3\text{H}_6\text{O}_3) / M(\text{C}_3\text{H}_6\text{O}_3)$
 $= 5.865 \text{ g} / (3 \times 12.0 + 6 \times 1.00 + 3 \times 16.0) \text{ g mol}^{-1}$
 $= 5.865 / 90.0 \text{ mol}$
 $= 0.0652 \text{ mol}$ ❶
 Energy per mol $= 91.2 \text{ kJ} / 0.0652 \text{ mol}$
 $= 1.40 \times 10^3 \text{ kJ mol}^{-1}$ ❶
 So for $\text{C}_3\text{H}_6\text{O}_3(\text{aq}) + 3\text{O}_2(\text{g}) \rightarrow 3\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l})$, $\Delta H = -1.40 \times 10^3 \text{ kJ mol}^{-1}$ ❶

Question 3

- a. **20 minutes** ❶ - the concentrations stop changing at 20 minutes and the graphs are horizontal.

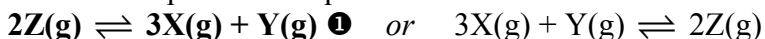
- b. i.

	Initial Concentration	Equilibrium Concentration	How the concentration changes as the system moves to equilibrium
X	1 mol L ⁻¹	4 mol L ⁻¹	Increases by 3 mol L⁻¹ ❶
Y	2 mol L ⁻¹	3 mol L ⁻¹	Increases by 1 mol L⁻¹ ❶
Z	4 mol L ⁻¹	2 mol L ⁻¹	Decreases by 2 mol L⁻¹ ❶

- ii. The concentrations of X and Y increase so more of both is being made as the system moves to equilibrium and Z is being consumed.

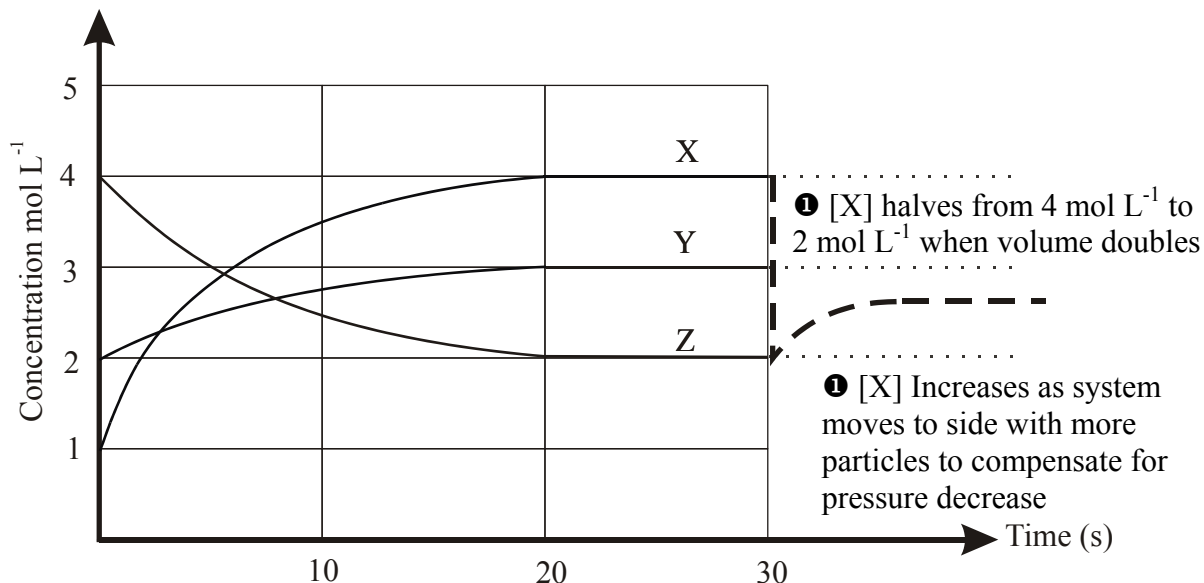
Since it is a 1 L vessel, **2 mol Z has reacted to produce 3 mol of X and 1 mol of Y.** ❶

Hence the equilibrium equation would be



- c. $K = [\text{X}]^3[\text{Y}] / [\text{Z}]^2$ ❶ or $K = [\text{Z}]^2 / [\text{X}]^3[\text{Y}]$
 $= 4^3 \times 3 / 2^2 = 2^2 / (4^3 \times 3)$
 $= \mathbf{5 \times 10^1 \text{ (M}^2\text{)}}$ ❶ $= \mathbf{2 \times 10^{-1} \text{ (M}^{-2}\text{)}}$

- d.

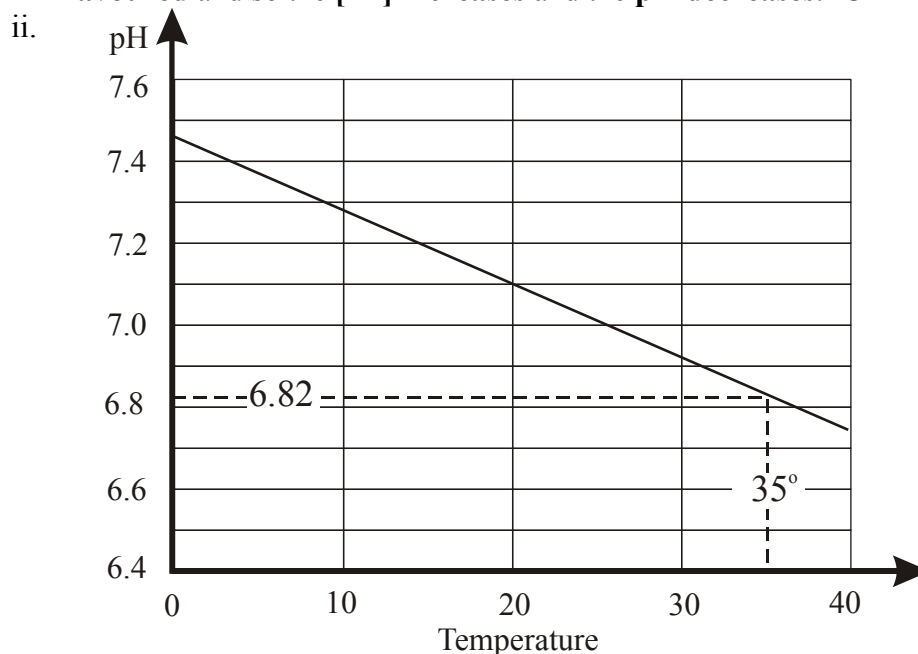


Question 4

- a. $Q = It = 2.5 \times 50 \times 60$
 $= 7.5 \times 10^3 \text{ C}$ ①
- b. $n(e^-) = Q / F = 7.5 \times 10^3 / 96500$
 $= 7.8 \times 10^{-2} \text{ mol}$ ①
- c. Whether or not a metal is deposited on the cathode during the electrolysis of an aqueous solution of a metal salt, depends on the relative strengths, as oxidants, of the metal cations and water.
According to the electrochemical series
 $2\text{H}_2\text{O(l)} + 2e^- \rightarrow \text{H}_2\text{(g)} + 2\text{OH}^-\text{(aq)} \quad -0.83 \text{ V}$
 $\text{K}^+\text{(aq)} + e^- \rightarrow \text{K(s)} \quad -2.93 \text{ V}$
 H_2O is a stronger oxidant than $\text{K}^+\text{(aq)}$ and will be preferentially reduced at the cathode, so no K will be deposited. ①
Since metal M is produced during the electrolysis of $\text{M(NO}_3)_2\text{(aq)}$, **$\text{M}^{2+}\text{(aq)}$ is a stronger oxidant than H_2O , and its reduction half-equation will be higher than that of H_2O on the electrochemical series.** ① i.e.
 $\text{M}^{2+}\text{(aq)} + 2e^- \rightarrow \text{M(s)} \quad > -0.83 \text{ V}$
 $2\text{H}_2\text{O(l)} + 2e^- \rightarrow \text{H}_2\text{(g)} + 2\text{OH}^-\text{(aq)} \quad -0.83 \text{ V}$
- d. Half-equations occurring in the KI(aq) cell are
Anode (+) $2\text{I}^-\text{(aq)} \rightarrow \text{I}_2\text{(s)} + 2e^-$
Cathode (-) $2\text{H}_2\text{O(l)} + 2e^- \rightarrow \text{H}_2\text{(g)} + 2\text{OH}^-\text{(aq)}$
 $n(\text{I}_2) \text{ formed} = \frac{1}{2} \times n(e^-)$
 $= \frac{1}{2} \times 7.8 \times 10^{-2}$
 $= 3.9 \times 10^{-2} \text{ mol}$ ①
 $m(\text{I}_2) \text{ formed} = 3.9 \times 10^{-2} \text{ mol} \times (2 \times 126.9)$
 $= 9.9 \text{ g}$ ①
- e. Reduction half-equation at cathode is $\text{M}^{2+}\text{(aq)} + 2e^- \rightarrow \text{M(s)}$
 $n(\text{M}) = \frac{1}{2} \times n(e^-)$
 $= \frac{1}{2} \times 7.8 \times 10^{-2}$
 $= 3.9 \times 10^{-2} \text{ mol}$ ①
 $\text{Molar mass of M} = m / n$
 $= 2.48 \text{ g} / 3.9 \times 10^{-2} \text{ mol}$
 $= 63.6 \text{ g mol}^{-1}$
Hence metal is **copper, Cu** ①

Question 5

- a. i. The graph shows that the pH of pure water decreases as the temperature increases.
This is consistent with the **exothermic** nature of the self-ionisation of water
 $2\text{H}_2\text{O(l)} \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{OH}^-(\text{aq})$ ①
As the temperature of pure water increases, the **endothermic reverse reaction is favoured** and so the **$[\text{H}^+]$ increases** and the **pH decreases**. ①



At 35°C, pH = 6.82
 $[\text{H}_3\text{O}^+] = 10^{-6.82}$
 $= 1.5 \times 10^{-7} \text{ M}$ ①

Since pure water is neutral, whatever the temperature $[\text{OH}^-] = [\text{H}_3\text{O}^+]$, so
 $[\text{OH}^-] = 1.5 \times 10^{-7} \text{ M}$ ①

- b. Propanoic acid, $\text{C}_2\text{H}_5\text{COOH}$, is a weak acid $K_a = 1.3 \times 10^{-5}$ (Data Book – Table 13)
 $\text{C}_2\text{H}_5\text{COOH(aq)} \rightleftharpoons \text{C}_2\text{H}_5\text{COO}^-(\text{aq}) + \text{H}^+(\text{aq})$
 $K_a = [\text{C}_2\text{H}_5\text{COO}^-][\text{H}^+] / [\text{C}_2\text{H}_5\text{COOH}]$
 Normal weak acids assumptions apply
 $[\text{C}_2\text{H}_5\text{COO}^-] = [\text{H}_3\text{O}^+]$ and $[\text{C}_2\text{H}_5\text{COOH}]_{\text{equilibrium}} = [\text{C}_2\text{H}_5\text{COOH}]_{\text{initially}} = 0.10 \text{ M}$
 So $1.3 \times 10^{-5} = [\text{H}_3\text{O}^+]^2 / 0.10$ ①
 $0.10 \times 1.3 \times 10^{-5} = [\text{H}_3\text{O}^+]^2$
 $[\text{H}_3\text{O}^+] = \sqrt{1.3 \times 10^{-6}}$
 $= 1.1 \times 10^{-3} \text{ M}$ ①
 $\text{pH} = -\log_{10} [\text{H}_3\text{O}^+]$
 $= -\log_{10} [1.1 \times 10^{-3}]$
 $= 2.9$ ①

- c. i. **HCl is a strong acid which ionises completely** according to
 $\text{HCl(aq)} + \text{H}_2\text{O(l)} \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{Cl}^-(\text{aq})$
0.10 M HCl(aq) $\rightarrow [\text{H}_3\text{O}^+] = 0.10 \text{ M} = 10^{-1} \text{ M}$ hence pH = 1 ①
 Diluting 10 mL of 0.10 M HCl(aq) to 100 mL with water gives
0.010 M HCl(aq) $\rightarrow [\text{H}_3\text{O}^+] = 0.010 \text{ M} = 10^{-2} \text{ M}$ hence pH = 2 ①
- ii. **Propanoic acid is a weak acid.**
 In part (b) it was shown that in **0.10 M $\text{C}_2\text{H}_5\text{COOH(aq)}$, $[\text{H}_3\text{O}^+] = 1.1 \times 10^{-3} \text{ M}$**
 Diluting 10 mL of 0.10 M $\text{C}_2\text{H}_5\text{COOH(aq)}$ to 100 mL with water gives
0.010 M $\text{C}_2\text{H}_5\text{COOH}$
 $K_a = [\text{C}_2\text{H}_5\text{COO}^-][\text{H}^+] / [\text{C}_2\text{H}_5\text{COOH}]$
 Normal weak acids assumptions apply
 $[\text{C}_2\text{H}_5\text{COO}^-] = [\text{H}_3\text{O}^+]$ and
 $[\text{C}_2\text{H}_5\text{COOH}]_{\text{equilibrium}} = [\text{C}_2\text{H}_5\text{COOH}]_{\text{initially}} = 0.010 \text{ M}$
 So $1.3 \times 10^{-5} = [\text{H}_3\text{O}^+]^2 / 0.010$ ①
 $0.010 \times 1.3 \times 10^{-5} = [\text{H}_3\text{O}^+]^2$
 $[\text{H}_3\text{O}^+] = \sqrt{(1.3 \times 10^{-7})}$
 $= 3.6 \times 10^{-4} \text{ M}$ ①
pH = $-\log_{10}[\text{H}_3\text{O}^+]$
 $= -\log_{10}(3.6 \times 10^{-4})$
 $= 3.4$ ①
- iii. As indicated in (i) **HCl(aq) is a strong acid** and so the **change in $[\text{H}_3\text{O}^+]$ is due only to the dilution factor**. Hence **increasing the volume by a factor of 10, decreases the $[\text{H}_3\text{O}^+]$ by a factor of 10 and thus increases the pH by 1**. ①
Propanoic acid, $\text{C}_2\text{H}_5\text{COOH}$, is a weak acid, ionising according to
 $\text{C}_2\text{H}_5\text{COOH(aq)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{C}_2\text{H}_5\text{COO}^-(\text{aq}) + \text{H}^+(\text{aq})$
 When the volume is increased by a factor of 10, the $[\text{H}_3\text{O}^+]$ immediately decreases by a factor of 10. **However, addition of more of the reactant $\text{H}_2\text{O(l)}$ pushes the position of equilibrium to the right, so increasing the $[\text{H}_3\text{O}^+]$** .
The combined effect of the dilution and the movement of the position of equilibrium is that $[\text{H}_3\text{O}^+]$ decreases but by a factor less than 10. ①
 Consequently the **change in pH** – it increases by less than 1 – is **smaller than that for HCl(aq)**. ①

Question 6

- a. **1 M $\text{H}^+(\text{aq})$ – e.g. 1 M HCl(aq), Pt electrode** ①
 $\text{H}_2(\text{g})$ at 101.3 kPa, 25°C ①
 A standard hydrogen half-cell, also known as the standard hydrogen electrode, contains a platinum electrode with hydrogen gas at 101.3 kPa, bubbled through a solution which is 1 M with respect to $\text{H}^+(\text{aq})$ at 25°C.
- b. According to the electrochemical series
 $\text{Cl}_2(\text{g}) + 2\text{e}^- \rightarrow 2\text{Cl}^-(\text{aq}) \quad 1.36 \text{ V}$
 $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g}) \quad 0.0 \text{ V}$
 The strongest oxidant Cl_2 reacts with strongest reductant H^+ as described by the overall equation
 $\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{H}^+(\text{aq}) + 2\text{Cl}^-(\text{aq})$ ①

- c. The decrease in pH indicates that the $[\text{H}^+]$ **increases** in the standard hydrogen half-cell.
 Half-reaction occurring must be $\text{H}_2(\text{g}) \rightarrow 2\text{H}^+(\text{aq}) + 2\text{e}^-$ ❶
 Since **oxidation** is occurring in the half-cell, and it is part of a galvanic cell, the **electrode must be negative (-)** ❶
- d. In the electrochemical series based on the standard hydrogen half-cell, the relevant half-equations are
 $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s}) \quad 0.34 \text{ V}$
 $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g}) \quad 0.00 \text{ V}$
 $\text{Ni}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Ni}(\text{s}) \quad -0.23 \text{ V}$
 In the alternative electrochemical series, the **differences between the standard electrode potentials will be same** – the relative strengths of oxidants and reductants remains the same – but the **standard electrode potential values change** because the **Cu^{2+}/Cu half-cell is set at 0.0.** ❶
 On the alternative electrochemical series, the half-equations would appear as
 $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s}) \quad 0.00 \text{ V}$
 $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g}) \quad -0.34 \text{ V}$
 $\text{Ni}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Ni}(\text{s}) \quad \underline{-0.57 \text{ V}}$ ❶
- e. According to the electrochemical series
 $2\text{H}_2\text{O}(\text{l}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g}) + 2\text{OH}^-(\text{aq}) \quad -0.83 \text{ V}$
 $\text{Al}^{3+}(\text{aq}) + 3\text{e}^- \rightarrow \text{Al}(\text{s}) \quad -1.67 \text{ V}$
 H_2O is a stronger oxidant ❶ than $\text{Al}^{3+}(\text{aq})$ and will be preferentially reduced ❶ during electrolysis

End of Suggested Answers