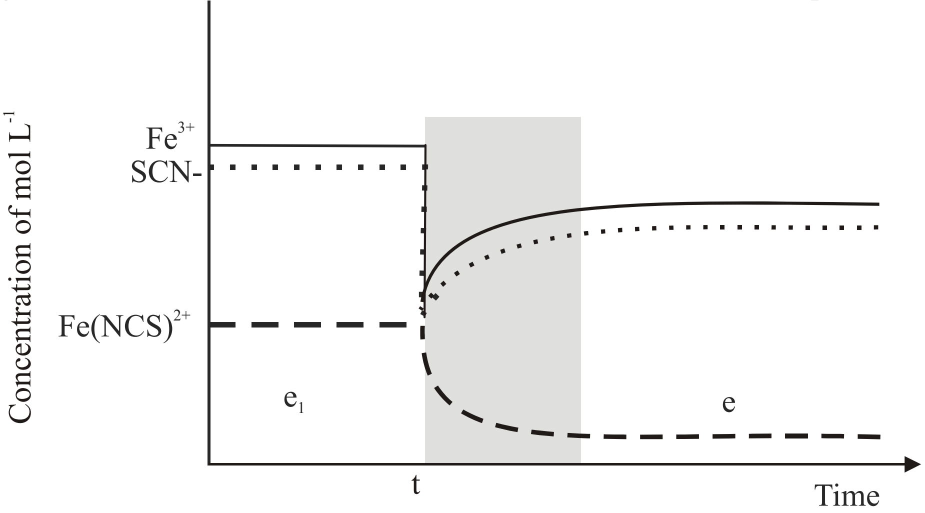
Suggested Answers

VCE Chemistry 2011 Year 12 Trial Exam Unit 4

SECTION A – Multiple Choice Answers

Q1 **A** In the Data Book the specific heat capacity of water is given as 4.18 J g^{-1} $°C^{-1}$. Since the accepted density of water at 25°C is 1 g mL^{-1}, then 100 mL of water is 100 g of water.
The amount of energy required to raise the 100 mL of water in the calorimeter by 1°C (or 1 K) is 418 J.
However **the components in calorimeter, such as the steel container, the electric heater, the thermometer and the stirrer also absorb energy during calibration**. However since their heat capacities are lower than water's, and there is less mass to heat, not a large quantity of extra energy will be required, and so 560 J $°C^{-1}$ or **560 J K^{-1} would be a reasonable value of the calorimeter constant for a well insulated calorimeter.**
Since the calorimeter constant is calculated as Electrical Energy / ΔT, a higher temperature change would lead to a significantly smaller calorimeter constant.
If the calorimeter contained only 80 mL of water, the temperature change would be larger and the calculated calibration constant would be smaller.
Temperature changes are exactly the same in K as in °C, e.g. 25°C (298 K) to 27°C (300 K) is a two degree change in both units.

Q2 **C** Adding water to the equilibrium $Fe^{3+}(aq) + SCN^{-}(aq) \rightleftharpoons Fe(NCS)^{2+}(aq)$ causes all concentrations to decrease and the reaction quotient (concentration fraction) $[Fe(NCS)^{2+}] / \{[Fe^{3+}][SCN^{-}]\}$ to become > K.
So to get back to equilibrium, the position of equilibrium moves to the left, producing more $Fe^{3+}(aq)$ and $SCN^{-}(aq)$. However, as the concentration-time graph below shows, the increase in amounts of $Fe^{3+}(aq)$ and $SCN^{-}(aq)$ is not enough to get the concentrations back to their concentrations at the first equilibrium.



This is also consistent with Le Chatelier's principle, with the system moving to produce more particles in the larger volume caused by the addition of water.
At the new equilibrium $[Fe(NCS)^{2+}]$ is lower, $n(Fe(NCS)^{2+})$ is lower; $[SCN^{-}]$ is lower, $n(SCN^{-})$ is higher; $[Fe^{3+}]$ is lower, $n(Fe^{3+})$ is higher.

Q3 B Original equilibrium equation

$NO(g) + \frac{1}{2}O_2(g) \rightleftharpoons NO_2(g)$; $\Delta H = -57.0$ kJ mol^{-1}; $K = 843$

To get to the required equilibrium equation, the original equilibrium equation needs to be doubled and then reversed

1. Original equation doubled.
   ΔH value doubled, K value raised to power 2
   $2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$; $\Delta H = -57.0\text{x}2$ kJ mol^{-1}; ($K = 843^2$)
   $\Delta H = -114$ kJ mol^{-1}; $K = 7.11\text{x}10^5$
2. Equation from 1. is reversed
   Sign of previous ΔH value is swapped, reciprocal of previous K value taken.
   $2NO_2(g) \rightleftharpoons 2NO(g) + O_2(g)$; $\Delta H = +114$ kJ mol^{-1}; $K = 1 / (7.11\text{x}10^5)$
   $\Delta H = +114$ kJ mol^{-1}; $K = 1.41\text{x}10^{-6}$

Q4 C According to the equation, 5 mol $C_6H_{12}O_6$ releases 11925 kJ of energy.

So **1 mol $C_6H_{12}O_6$** → 11 925 / 5 = **2385 kJ**

According to information in Table 13 of the Data Book,

the **combustion of 1 mol $C_6H_{12}O_6$ releases 2816 kJ**

The maximum amount of energy available from 1 mol $C_6H_{12}O_6$ in the bacteria driven reaction is **less** than that available from the combustion of 1 mol $C_6H_{12}O_6$.

Q5 D
Q6 C The rate-time graphs show:

1. The rates of the forward and reverse reactions are initially equal – the system is at equilibrium.
2. At time **t** the rates of both the forward and reverse reactions increase suddenly, with the rate of the forward reaction increasing more than the rate of the reverse reaction.
3. The rate of the forward reaction decreases and the rate of reverse reaction increases as the system returns to equilibrium.

Consider the alternatives for **Question 5.**

A. Adding more reactant would cause a sudden increase in the rate of the forward reaction only.
B. A temperature decrease would cause a decrease in reaction rates.
C. Removing some product would cause a sudden decrease in the rate of the reverse reaction only.
D. A **volume decrease would immediately increase the concentrations of both reactants and products** and, since reaction rates increase with increased concentration, **increase the rates of both the forward and reverse reactions**.

Consider **Question 6.**

The greater increase in the rate of the forward reaction at time **t** indicates there are **more particles on the reactant side** of the associated equilibrium. Also for a **decrease in volume** on an equilibrium system, the **system will respond by shifting in the direction of the side with fewer particles.**

Since the rate of the forward reaction decreases and reverse increases as the system responds to the imposed change, the **forward reaction is favoured**. Hence there will be **fewer particles on the product side of the equilibrium.**

Only alternative C shows fewer particles on the product side.

Q7 B According to the relevant equations from the electrochemical series

$Br_2(l) + 2e^- \rightleftharpoons 2Br^-(aq)$ 1.09 V

$Fe^{3+}(aq) + e^- \rightleftharpoons \mathbf{Fe^{2+}}(aq)$ 0.77 V

$\mathbf{Fe^{2+}}(aq) + 2e^- \rightleftharpoons Fe(s)$ -.0.44 V

$Zn^{2+}(aq) + 2e^- \rightleftharpoons \mathbf{Zn}(s)$ -0.76 V

These half-equations show that, depending on the species with which it is reacting, **Fe^{2+}(aq) can act as an oxidant** (oxidising Zn to Zn^{2+}) **and a reductant** (reducing Br_2 to Br^-). In these roles Fe^{2+}(aq) will be converted to Fe(s) and Fe^{3+}(aq) respectively.

Q8 A When the cell is discharging electrons flow from the (-) electrode to the (+) electrode. Since electrons always move from the site of oxidation to the site of reduction, the oxidation number will decrease at the (+) electrode.

$$\overset{+1}{Ag_2O(s)} + \overset{0}{Zn(s)} + H_2O(l) \rightarrow \overset{0}{2Ag(s)} + \overset{+2}{Zn(OH)_2(s)}$$

Change in oxidation number = final value – initial value.

Oxidation number of Zn increases from 0 to +2 at the (-) electrode, i.e. a change of +2

Oxidation number of Ag decreases from +1 to 0 at the **(+) electrode**, i.e. a **change of -1**

Q9 C Since electrolysis is occurring, electrons are being forced from the (+) electrode to the (-) electrode.

Hence reduction is occurring at the (-) electrode.

In each solution there are **two possible oxidants – the metal cation and H_2O – which may be reduced.**

If water is the stronger oxidant, i.e. above the cation on the electrochemical series, it will be reduced according to

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

and **H_2(g) will be evolved at the negative electrode**.

If the metal cation is the stronger oxidant, i.e. above H_2O on the electrochemical series, **it will be reduced to the solid metal.**

Consider the alternatives.

A. Oxidants Al^{3+}(aq), K^+(aq), Li^+(aq) and H_2O.
   H_2O is the stronger oxidant in all three solutions, so H_2 gas is produced at the (-) electrode in all three.

B. Oxidants Co^{2+}(aq), Cu^{2+}(aq), Ca^{2+}(aq) and H_2O.
   Both Cu^{2+}(aq) and Co^{2+}(aq) are stronger oxidants than H_2O, so Cu(s) and Co(s) will be produced with H_2(g) evolved in only one cell.

C. Oxidants Mg^{2+}(aq), Ni^{2+}(aq), K^+(aq) and H_2O.
   H_2O is a stronger oxidant than Mg^{2+}(aq) and K^+(aq), so **H_2 gas is produced at the (-) electrode in two of the cells, Ni(s) will be produced at the (-) electrode in the other cell.**

D. Oxidants Ag^+(aq), Sn^{2+}(aq), Li^+(aq) and H_2O.
   Both Ag^+(aq) and Sn^{2+}(aq) are stronger oxidants than H_2O, so Ag(s) and Sn(s) will be produced and H_2(g) evolved in only one cell.

Q10 A $Ca(OH)_2(aq) \rightarrow Ca^{2+}(aq) + 2OH^-(aq)$

pH 10.0 → $[H_3O^+] = 10^{-10}$

$= 1x10^{-10}$ M

Since $[H_3O^+][OH^-] = 10^{-14}$ at 25°

$[OH^-] = 10^{-14} / 1x10^{-10}$

$= 1x10^{-4}$ M

$= 0.0001$ M

$[Ca(OH)_2] = ½ \text{ x } [OH^-]$

$= ½ \text{ x } 0.0001$

= **0.00005** M

Q11 D $HA(aq) \rightleftharpoons H^+(aq) + A^-(aq)$

To identify the weak acid, use the data to work out the value of K_a and refer to Table 12 in the Data Book.

$K_a = [H^+][A^-] / [HA]$

$= [H^+] \text{ x } ([A^-]/[HA])$

pH 4.5 → $\mathbf{[H^+]} = 10^{-4.5}$ M

$= \mathbf{3.16x10^{-5} \text{ M}}$

$[HA] / [A^-] = 1.32x10^4 \rightarrow \mathbf{([A^-]/[HA])} = 1 / (1.32x10^4)$

$= \mathbf{7.58x10^{-5}}$

$K_a = 3.16x10^{-5} \text{ x } 7.58x10^{-5}$

$= \mathbf{2.4x10^{-9}}$

Hypobromous acid

Q12 B Consider the alternatives:

A. The specific heat capacity of H_2O is 4.18 J g^{-1} $°C^{-1}$

Energy needed to increase 1 kg H_2O by 1 °C = 4.18 x 1000 x 1

= 4180 J

= 4.18 kJ

B. Electrical energy, as in calorimeter calibration, is determined from

$E = VIt$

So for a current of 1 A, at 1 V for 1000 s

$E = 1 \text{ x } 1 \text{ x } 1000$

= 1000 J

= **1 kJ**

C. $n(H_2) = 2 \text{ g} / 2.0 \text{ g mol}^{-1}$

= 1 mol

According to Table 13 of the Data Book, the combustion of 1 mol H_2 will release 286 kJ of energy.

D. Since the combustion of 1 mol H_2 to form H_2O releases 286 kJ of energy, then the conversion of H_2O to produce 1 mol H_2 will require the input of a similar amount of energy.

Q13 B $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$.

Adding 50 mL of water doubles the volume so all concentrations halve.
Addition of water pushes the equilibrium to the right, causing more NH_3 to ionise, i.e. **increasing the percentage ionisation**.
The amount of $OH^-(aq)$ present in the solution also increases.
Although there is a more **OH^-(aq)** it is in a **larger volume**, the amount of $OH^-(aq)$ produced in the equilibrium shift only partially compensates for the volume increase by the time equilibrium is regained.
Overall **$[OH^-]$** is **lower** → $[H_3O^+]$ is higher → **lower pH.**
So the **pH decreases and the percentage ionisation increases.**

Q14 D **Diagram X**
This shows the energy distribution at two different temperatures. Graph 2 represents a higher **temperature** situation.
In the case of the higher temperature, where the collision energies are higher, the greater shaded area under the curve, for energy greater than E_a, represents a **greater proportion of collisions, with energy greater than the activation energy**, hence **increasing temperature increases the rate of reaction**.
Diagram Y
This shows the impact of a **catalyst.** Activation energy $\boldsymbol{E_{A2}}$ is **less than** activation energy $\boldsymbol{E_{A1}}$. This **decrease in the activation energy** is caused by the use of a **catalyst** (the only possibility)**.** Hence **many more collisions will have energy greater than the activation energy at $\boldsymbol{E_{A2}}$** (as indicated by greater shaded area under graph above E_{A2}) and so the **rate of reaction increases in the presence of a catalyst.**

Q15 C $2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$; $\Delta H = -113 \text{ kJ mol}^{-1}$

Consider the alternatives.

A. Since the forward reaction is exothermic, decreasing the temperature favours the forward reaction, **increasing the amount of NO_2**.

B. Decreasing the volume increases the concentration of all species present. This makes the concentration fraction smaller than *K*. So the forward reaction is favoured as the system returns to equilibrium.
Also the pressure exerted by species in the equilibrium system increases, since they are occupying a smaller volume. In order to compensate for the pressure increase, the reaction which leads to an overall reduction in the number of particles present is favoured, i.e. the forward reaction, **increasing the amount of NO_2.**

C. N_2 gas is not part of the equilibrium system, does not react with any of the species in the equilibrium system, and does not change the position of equilibrium. It does not cause a change in the concentration fraction, so the system is not moved from equilibrium. Hence the **amount of NO_2** is **not affected.**
From the pressure perspective, whilst the total pressure in the container will increase, the **pressure exerted by the individual species in the equilibrium mixture does not change.**

D. Adding air increases the $[O_2]$ in the system which responds to this change by using up some of the added O_2. Hence the **forward reaction is favoured and the amount of NO_2 increases**.

Q16 A According to the equation

$NH_3(g) + 3Cl_2(g) \rightleftharpoons NCl_3(g) + 3HCl(g)$

if the reaction did proceed to completion, then with 3 mol NH_3 and 10 mol Cl_2 present at the start, 3 mol NH_3 would react with 9 mol Cl_2 to produce 3 mol NCl_3 and 9 mol HCl leaving 1 mol Cl_2 in excess.

However all 3 mol NH_3 cannot react if the mixture reaches equilibrium, since equilibrium requires the presence of all reactants and all products.

Hence less than 3 mol NH_3 will have reacted and **less than 3 mol NCl_3 will be present** when equilibrium is reached.

Q17 A The diagram shows the direction of movement of ions in the salt bridge. Anions, in this case NO_3^- ions, move towards the anode – the site of oxidation – which is (-) in a galvanic cell, and is in half-cell 2.

So **half-cell 2** contains the (-) electrode and is the **site of oxidation**.

Half-cell 1 contains the (+) electrode and is the site of reduction.

Q18 B Energy released = 200 MJ = 200×10^3 kJ

Energy available from one mole of C_8H_{18} = 5464 kJ (Data Book Table 13)

$n(C_8H_{18})$ = energy released / energy per mole C_8H_{18}

= 200×10^3 kJ / 5464 kJ mol^{-1}

= 36.6 mol

$\boldsymbol{m(C_8H_{18})}$ = 36.6 mol x 114 g mol^{-1}

= 4.17×10^3 g

= **4.17 kg**

Q19 D This question requires recognition that, in a fuel cell, electrons flow from the (-) electrode to the (+) electrode.

So oxidation occurs at the (-) electrode and **reduction at the (+) electrode**.

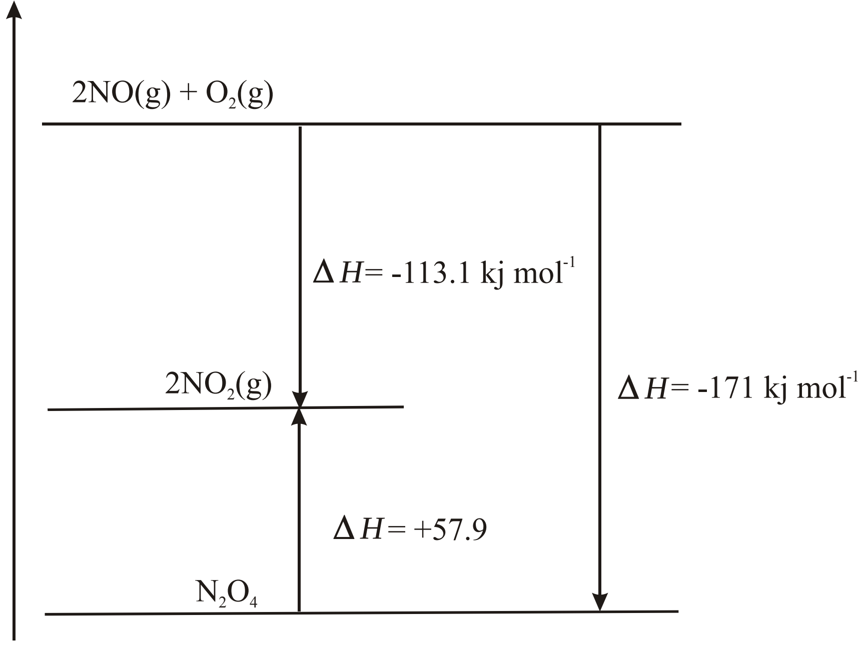
In a fuel cell, the fuel is oxidised at the (-) electrode and **O_2(g) is reduced at the (+) electrode**. In a H_2-O_2 v fuel cell, H_2 is the fuel.

An alkaline electrolyte contains **OH^-(aq) ions**.

Then it is a case of effective use of the electrochemical series – identifying the half-equation which includes both O_2(g) and OH^-(aq), i.e.

$\mathbf{O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)}$

Q20 C Consider the thermochemical information via a simplified energy profile.



The profile enables you to deduce

$2NO_2(g) \rightarrow N_2O_4(g)$; ΔH = -57.9 kJ mol^{-1}

So the combination of

$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$; ΔH = -113.1 kJ mol^{-1}, and

$2NO_2(g) \rightarrow N_2O_4(g)$; ΔH = -57.9 kJ mol^{-1}

produces

$2NO(g) + O_2(g) \rightarrow N_2O_4(g)$; **$\Delta H$ = -171 kJ mol^{-1}**

SECTION B – Short Answer (Answers)

Question 1

a. NH_3 *or* C_2H_4 *or* H_2SO_4 *or* HNO_3 ❶

b. i. $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ *or*
$C_3H_8(g) \rightleftharpoons C_2H_4(g) + CH_4(g)$ / other cracking equation with C_2H_4 as a product *or*
$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ *or*
$2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$ / $2NO_2(g) \rightleftharpoons N_2O_4(g)$ ❶

ii. For reaction to commence the **reactant particles must have energy equal to or greater than the activation energy**. ❶

iii. When the reactants, e.g. SO_2 and O_2, are mixed together and start reacting, **the rate of the forward reaction** (e.g. $2SO_2 + O_2 \rightarrow 2SO_3$) is at its **maximum**. ❶
As the **forward reaction** proceeds, its **rate decreases**, due to the **decreasing concentration** of **reactants**. ❶
Simultaneously the **concentration of the products increases**, and so the **rate of the reverse reaction** (e.g. $2SO_3 \rightarrow 2SO_2 + O_2$) **increases**. ❶
Eventually the **rates of the forward and reverse reactions become equal** and the reacting system has reached **chemical equilibrium**. ❶

iv. When the pressure on an equilibrium system is increased, the system responds by favouring the side with fewer particles.
When the temperature in an equilibrium system is decreased, the system responds by favouring the direction of the exothermic reaction.
NH_3: Increasing pressure **<u>right</u>** ❶; Decreasing temperature **<u>right</u>** ❶
$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g); \Delta H < 0$
C_2H_4: Increasing pressure **<u>left</u>**; Decreasing temperature **<u>left</u>**
$C_3H_8(g) \rightleftharpoons C_2H_4(g) + CH_4(g)$; $\Delta H > 0$
H_2SO_4: Increasing pressure **<u>right</u>**; Decreasing temperature **<u>right</u>**
$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g); \Delta H < 0$
HNO_3: Increasing pressure **<u>right</u>**; Decreasing temperature **<u>right</u>**
$2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g); \Delta H < 0$ / $2NO_2(g) \rightleftharpoons N_2O_4(g); \Delta H < 0$

c. $NH_3(g) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$ ❶
$C_2H_4(g) + H_2O(g) \rightleftharpoons CH_3CH_2OH(l)$
$H_2SO_4(l) + H_2O(l) \rightarrow HSO_4^-(aq) + H_3O^+(aq)$
$HNO_3(l) + H_2O(l) \rightarrow NO_3^-(aq) + H_2O(l)$

Question 2

a. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g \text{ or } l)$ ❶

b. Chemical → Electrical ❶

c. In a **fuel cell the conversion from chemical energy to electrical energy is direct** and hence more efficient than the **multistep conversion (chemical → thermal → mechanical → electrical) of a conventional fuel (gas) fired power station**. ❶

d. i. (-) electrode (anode) $CH_4(g) + 4O^{2-}(l) \rightarrow CO_2(g) + 2H_2O(g) + 8e^-$ ❶

ii. (+) electrode (cathode) $O_2(g) + 4e^- \rightarrow 2O^{2-}(l)$ ❶

The description of solid-oxide fuel cells indicated that O_2 was reduced to O^{2-} ions at the cathode - *'bring in oxygen from the air to be reduced at the cathode, and then pass the oxide ions through a solid-oxide membrane to the anode'* – therefore both the CO_2 and H_2O, characteristic products of the use of methane as a fuel, had to be produced at the anode, where the only source of oxygen for the oxidation was $O^{2-}(l)$ ions.

e. The **operating temperature of SOFCs (500-1000°C) is well beyond the reference temperature (25°C) of the electrochemical series.** ❶

f. i. $H_2(g) \rightarrow 2H^+(aq) + 2e^-$ ❶

ii. $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$ ❶

g. $2H_2(g) + O_2(l) \rightarrow 2H_2O(l)$ ❶

Question 3

a. $NaOH(aq) + HCOOH(aq) \rightarrow HCOONa(aq) + H_2O(l)$

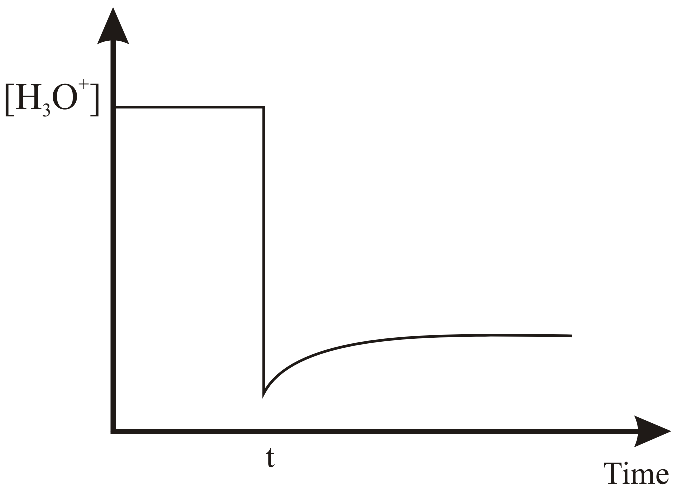
$n(HCOOH)$ reacting $= n(NaOH)$ required
$= 0.155 \text{ mol L}^{-1} \times 17.25\times10^{-3} \text{ L}$
$= 2.67\times10^{-3} \text{ mol}$ ❶

b. i. Since the equation for the ionisation of $HCOOH(aq)$ is
$HCOOH(aq) + H_2O(l) \rightleftharpoons HCOO^-(aq) + H_3O^+(aq)$
the equilibrium law is
$K_a = [HCOO^-][H_3O^+] / [HCOOH]$ ❶

ii. $K_a(HCOOH) = 1.8\times10^{-4}$ (from Data Book)
Since 18.5 mL of the methanoic acid solution contained 2.67×10^{-3} mol HCOOH
$c(HCOOH) = 2.67\times10^{-3} \text{ mol} / 18.5\times10^{-3} \text{ L}$
$= 0.144 \text{ M}$ ❶
Using the standard weak acid assumptions: $[HCOOH]_{equilibrium} = 0.144$ M
$[HCOO^-] = [H_3O^+]$, i.e. negligible H_3O^+ from the self-ionisation of water.

$1.8\times10^{-4} = [H_3O^+]^2 / 0.144$ ❶
$[H_3O^+] = \sqrt{(0.144 \times 1.8\times10^{-4})}$
$= 5.10\times10^{-3}$ M ❶
$pH = -\log_{10}(5.10\times10^{-3})$
$= 2.3$ ❶

c. i.



When 10 mL methanoic acid is diluted to 100 mL with water, the **[H_3O^+] immediately decreases by a factor of 10**. However because the **dilution decreases all concentrations, the reaction quotient** (concentration fraction), i.e. [$HCOO^-$][H_3O^+] / [HCOOH], decreases and **become less than K_a**. ❶For the system to return to equilibrium, the **position of equilibrium moves to the right and the [H_3O^+] increases**, ❶ i.e. there are more H_3O^+ ions present, but **the concentration is still well below its value prior to dilution when equilibrium is regained.**

ii. Since the [H_3O^+] decreases then increases as the equilibrium moves to the right, the **pH will increase and then decrease**. ❶ However the **pH after dilution will be higher than before dilution** ❶
[pH of 0.0144 M HCOOH(aq) = 2.8].

iii. The **dilution of a strong acid**, e.g. HCl(aq). ❶ When HCl(aq) ionises, effectively all the HCl molecules have donated H^+(aq) and so there is no increase in the amount of H_3O^+ present as a result of the dilution.

Question 4

a. According to the supplied data, the calorimeter was calibrated by the combustion of ethanol in the reaction bomb.

n(ethanol) used = 0.853 g / 46.0 g mol^{-1}
= 1.85×10^{-2} mol
ΔH_c(ethanol) = -1364 kJ mol^{-1} (Data Book)
Energy released = 1.85×10^{-2} mol x 1364 kJ mol^{-1}
= 25.3 kJ ❶
ΔT calibration = 24.10 - 21.30
= 2.80 °C
Calibration Constant = 25.3 kJ / 2.80°C
= 9.03 kJ $°C^{-1}$
= **9.03 kJ K^{-1}** ❶

*A temperature **change** is the same magnitude in both °C and K*

b. ΔT combustion of olive oil $= 22.391 - 22.150$
$= 0.241\ °C$
Energy released $= 9.03\ kJ\ K^{-1} \times 0.241\ K$
$= \mathbf{2.18\ kJ}$ ❶

c. $M(C_{57}H_{104}O_6) = 57 \times 12.0 + 104 \times 1.0 + 6 \times 16.0$
$= 884.0\ g\ mol^{-1}$
$n(C_{57}H_{104}O_6) = 1.00\ g / 884.0\ g\ mol^{-1}$
$= 1.13 \times 10^{-3}\ mol$ ❶
Heat of combustion olive oil $= 2.18\ kJ / 1.13 \times 10^{-3}\ mol$
$= \mathbf{1.93 \times 10^3\ kJ\ mol^{-1}}$ ❶

d. This is simply a matter of effective combustion equation balancing.

1. Reactants and products
   $C_{57}H_{104}O_6(l) + \_\_\_O_2(g) \rightarrow \_\_\_CO_2(g) + \_\_\_H_2O(g)$
2. Balance for ‘C’
   $C_{57}H_{104}O_6(l) + \_\_\_O_2(g) \rightarrow \mathbf{57}CO_2(g) + \_\_\_H_2O(g)$
3. Balance for ‘H’
   $C_{57}H_{104}O_6(l) + \_\_\_O_2(g) \rightarrow \mathbf{57}CO_2(g) + \mathbf{52}H_2O(g)$
4. Balance for ‘O’
   On RHS – 2x57 + 52 = 166 ‘O’
   On LHS – 6 ‘O’
   So 160 ‘O’, i.e. 80 O_2 is required
   $C_{57}H_{104}O_6(l) + \mathbf{80}O_2(g) \rightarrow \mathbf{57}CO_2(g) + \mathbf{52}H_2O(g)$
   The thermochemical equation, which includes ΔH, is
   $\mathbf{C_{57}H_{104}O_6(l) + 80O_2(g) \rightarrow 57CO_2(g) + 52H_2O(g);\ \Delta \mathit{H} = -1.93 \times 10^3\ kJ\ mol^{-1}}$
   ❶ *for correct chemical formulae and states*
   ❶ *for correct balancing*
   ❶ *for ΔH consistent with the equation.*

e. Specific heat capacity of water = $4.18\ J\ g^{-1}\ K^{-1}$ (Data Book)
Energy into water $= 4.18\ J\ g^{-1}\ K^{-1} \times m(H_2O) \times \Delta T$
$2.18 \times 10^3\ J = 4.18\ J\ g^{-1}\ K^{-1} \times m(H_2O) \times 1.49\ K$ ❶
$m(H_2O) = 2.18 \times 10^3 / (4.18 \times 1.49)$
$= \mathbf{350\ g}$ ❶

Question 5

a. Using $pV = nRT \rightarrow \boldsymbol{p = nRT / V}$
Since you are given the concentrations of all gases and $\boldsymbol{c = n/V}$
Then $\boldsymbol{p = cRT}$ and the **total pressure depends on the total concentration.**
Total concentration $= 0.00250 + 0.00350 + 0.0250 + 0.0400$
$= 0.0710\ mol\ L^{-1}$ ❶
$p = 0.0710 \times 8.31 \times (460 + 273)$
$= 0.0710 \times 8.31 \times 733$
$= 433\ kPa$ ❶
According to the data book, 101.3 kPa = 1 atm
$p = 433 / 101.3\ atm$
$= \mathbf{4.27\ atm}$ ❶

b. i. To show the reaction mixture is at equilibrium, calculate the reaction quotient, *Q*, (concentration fraction) and compare it with the stated equilibrium constant.

$Q = [NO][SO_3] / [SO_2][NO_2]$

$= [0.0250 \times 0.0400 / 0.00250 \times 0.00350]$

$=$ **114** ❶

Since ***Q* > *K*$_c$, the system is not at equilibrium**.

To get to equilibrium the **value of *Q* must decrease and so the reverse (backward) reaction will dominate.** ❶

ii. When the system is at equilibrium, the **pressure will remain constant.** ❶ Forward and reverse reactions are occurring at the same rate.

c. Since the reverse reaction is favoured as the system shifts to equilibrium, and there is a 1:1 mole ratio through the equation, each reactant concentration must each increase by 0.00050 mol L^{-1} and each product concentration decrease by 0.0050 mol L^{-1}.

Hence **the total concentration, and number of moles, in the reaction vessel does not change** and so the pressure at equilibrium will be the same as before equilibrium, i.e. **432 kPa** ❶

This is also consistent with the equilibrium having the same number of particles on both sides so that the pressure does not change as reaction proceeds in either direction.

d. i. Le Chatelier's principle states that 'if a system at equilibrium is disturbed by a change in conditions, the reaction moves in the direction that will counteract the effect of the change and, if possible, return the system to equilibrium'. ❶

ii. The equilibrium system $H_2CO_3(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + H_3O^+(aq)$ will respond to an increase in pH, i.e. a decrease in $[H_3O^+]$, by moving to increase the $[H_3O^+]$ – and decrease the pH, by **favouring the forward reaction and moving the position of equilibrium to the right**. ❶

This occurs because **when the $[H_3O^+]$ decreases, the rate of the reverse (backward) reaction drops relative to the rate of the forward reaction.** This **drives the reaction forward** until the rates of forward and reverse reactions are again equal and equilibrium is regained. ❶

Question 6

a. i. The cell voltage, 1.02 V, is the difference between the *E*° values of the half-cell containing the oxidant and the half-cell containing the reductant.

If the oxidant was in the Cu^{2+}(aq)/Cu(s) half-cell and the reductant in the 'unknown' half-cell then

1.02 V = *E*°(Cu^{2+}/Cu) – *E*°(unknown half-cell)

1.02 V = 0.34 - *E*°(unknown half-cell)

E°(unknown half-cell) = 0.34 – 1.02

= -0.68 V

There is no half-cell in the electrochemical series in the Data Book with *E*°= - 0.68 V

If the reductant was in the Cu^{2+}(aq)/Cu(s) half-cell and the oxidant in the 'unknown' half-cell then

1.02 V = *E*°(unknown half-cell) – *E*°(Cu^{2+}/Cu)

1.02 V = *E*°(unknown half-cell) – 0.34

***E*°(unknown half-cell)** = 1.02 + 0.34

= **1.36 V**

This corresponds to the ***E***° value for the Cl_2(g)/Cl^-(aq) half-cell, and the species are

A Cl_2(g) ❶
B 1 M Cl^-(aq) ❶
C Pt(s) *or* C(s) ❶

When the half-cell redox pair does not include a conducting solid, an inert electrode, Pt or C, is used.

ii. A salt-bridge is essential **to allow electrolytic conduction between the electrodes** and so **complete the circuit.** ❶

In the Cl_2(g)/Cl^-(aq)//Cu^{2+}(aq)/Cu(s) cell the cathode is in the Cl_2(g)/Cl^-(aq) half-cell.

The reduction half-equation $Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$ shows that the **release of Cl^-(aq) ions would increase the (-) charge in the solution around the cathode**. ❶ To counteract that situation and **keep the solution electrically neutral, cations flow through the salt bridge into the half-cell.** ❶

iii. **Ag^+(aq) ions would react with Cl^-(aq) to form insoluble AgCl.** ❶ This would interfere with the effective operation of the cell.

b. i. In the half-equation $O_3(g) + H_2O(l) + 2e^- \rightarrow O_2(g) + 2OH^-(aq)$ there is **no change in oxidation number** of any of the elements. ❶

The fundamental definition of reduction as the 'gain of electrons' applies.

ii. According to the electrochemical series, the relevant species $Y^{3+}(aq)$,$Cl^-(aq)$ and $H_2O(l)$ appear in the electrochemical series data in the half-equations.

$Cl_2(g) + 2e^- \rightleftharpoons 2\mathbf{Cl^-}(aq)$ $E° = 1.36$ V

$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2\mathbf{H_2O}(l)$ $E° = 1.23$ V

$2\mathbf{H_2O}(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$ $E° = -0.83$V

$\mathbf{Y^{3+}}(aq) + 2e^- \rightleftharpoons Y(s)$ $E° = -2.37$ V

The strongest oxidant, $H_2O(l)$, is reduced at the cathode, **the (-) electrode** , according to

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

The strongest reductant, $H_2O(l)$, is oxidised at the anode, **the (+) electrode** , according to

$\mathbf{2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-}$ ❶

iii. According to the electrochemical series data, only $H_2O(aq)$ and $In^{3+}(aq)$ can be reduced at the (-) electrode (cathode).

$\mathbf{In^{3+}}(aq) + 3e^- \rightleftharpoons In(s)$ $E° = -0.34$ V

$2\mathbf{H_2O}(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$ $E° = -0.83$V

Since $In^{3+}(aq)$ is the stronger oxidant, the reductant half-equation will be

$In^{3+}(aq) + 3e^- \rightarrow In(s)$

$n(In)$ produced $= m(In) / M(In)$

$= 0.448 \text{ g} / 114.8 \text{ g mol}^{-1}$

$= 0.00390$ mol ❶

$n(e^-)$ required $= 3 \times n(In)$

$= 3 \times 0.00390$

$= 0.0117$ mol ❶

$Q = n(e^-) \times F$

$= 0.0117 \text{ mol} \times 96500 \text{ C mol}^{-1}$

$= 1.13\times10^3$ C ❶

However the conversion of electrical energy to chemical energy is only 75 % efficient, so

$1.13\times10^3 \text{ C} = 0.75 \times Q_{TOTAL}$

$Q_{TOTAL} = 1.13\times10^3 / 0.75$

$= 1.51\times10^3$ C ❶

Since $Q = It$

$t = 1.51\times10^3 \text{ C} / 2.51 \text{ A}$

$= 600$ s

= **10.0 minutes** ❶

End of Suggested Answers