# Suggested Answers

VCE Chemistry 2010 Year 12 Trial Exam Unit 4

Multiple Choice Answers – Section A

Q1 C A fundamental principle of equilibrium is that the rates of the forward and reverse reactions are equal.
The forward reaction $CH_3CH_2OH(l) + CH_3COOH(l) \rightarrow CH_3COOCH_2CH_3(l) + H_2O(l)$ – a condensation reaction, proceeds at the same rate as the reverse reaction $CH_3COOCH_2CH_3(l) + H_2O(l) \rightarrow CH_3CH_2OH(l) + CH_3COOH(l)$, an hydrolysis reaction.
There is insufficient data to determine the value of the concentration fraction at equilibrium (or K) or the relative concentrations of ethanol and ethyl ethanoate at equilibrium.
The pH of the mixture is most likely to be less than 7 because of the acid catalyst.

Q2 D The electrochemical cell is an electrolysis cell, since it has an external power supply.
The only species present are $Zn^{2+}(l)$ and $Cl^-(l)$
At the anode (+), i.e. electrode Q, the reductant, $Cl^-(l)$ is oxidised according to $2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$,
i.e. chloride ions lose electrons at electrode Q.
At the cathode (-), i.e. electrode P, the oxidant, $Zn^{2+}(l)$ is reduced according to $Zn^{2+}(l) + 2e^- \rightarrow Zn(s)$

Q3 A Since K is 92 at 250°C and 2.3×10^4 at 30°C, it increases as the temperature decreases. Since K is effectively the ratio [products] / [reactants], there are a greater proportion of products at the lower temperature indicating that the forward reaction is exothermic.
Hence $\Delta H < 0$ and the yield of Z increases as the temperature decreases, and decreases as the temperature increases.

Q4 D The concentration-time graph shows that the [OH⁻] decreases instantaneously as a result of the imposed change and then increases, but not back to original equilibrium concentration as the system returns to equilibrium.
A. the addition of NaOH will cause an immediate increase in the [OH⁻], which will then decrease as the system returns to equilibrium.
B. Since the self-ionisation of water $2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$ is endothermic, cooling pure water will favour the reverse reaction and so the [OH⁻] will decrease but NOT suddenly. It will be a gradual decrease as the temperature drops to the new temperature.
C. Since the self-ionisation of water $2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$ is endothermic, heating pure water will favour the forward reaction and so the [OH⁻] will increase.

D. $NH_3(aq)$ is in equilibrium according to

$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$
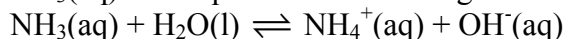
Doubling the volume by adding an equal volume of water will cause all concentrations to instantaneously halve. So the $[OH^-]$ instantaneously halves. Because H_2O is also a reactant, the reaction moves to the right as the system returns to equilibrium. Hence the $[OH^-]$ increases but is still lower than its concentration before the addition of water when equilibrium is again established.

Q5   C $n(NaOH)$ reacting $= 1.0 \times 25 \times 10^{-3} = 0.025$ mol
$n(HCl)$ reacting $= 1.0 \times 25 \times 10^{-3} = 0.025$ mol
So the energy released in to 50 mL solution from the reaction of 0.025 mol of HCl and 0.025 mol NaOH will cause the temperature to increase by 6°C.

A. 0.050 mol HCl and 0.050 mol NaOH would release twice as much energy, so the temperature of 50 mL of solution would increase by 12 °C.

B. 0.10 mol HCl and 0.10 mol NaOH, i.e. four times as much as the original reaction, will release four times as much energy, so the total solution volume of 100 mL would increase by 12 °C.

C. 0.050 mol HCl and 0.050 mol NaOH would release twice as much energy, so the total solution volume of 100 mL, (i.e. twice the original total volume of 50 mL) would increase by 6 °C.

D. 0.20 mol HCl and 0.20 mol NaOH, i.e. eight times as much as the original reaction, will release eight times as much energy, so the total solution volume of 200 (i.e. four times the original total volume of 50 mL) would increase by 12 °C.

Q6   B   The self-ionisation constant of water $= 1.0 \times 10^{-14}$ at 25 °C.
In pure water, $[H_3O^+]$ is always equal to $[OH^-]$ but they are only both 10^{-7} M at 25°C.
The higher K_w at 45°C indicates that the equilibrium

$2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$

lies further to the right at 45 °C than at 25°C,
i.e. $[H_3O^+] > 10^{-7}$ M and $[OH^-] > 10^{-7}$ M
Consequently at 15°C the equilibrium lies further to the left than at 25°C and so the $[H_3O^+] < 10^{-7}$ M and $[OH^-] < 10^{-7}$ M
Since $[H_3O^+] < 10^{-7}$ M at 25°C, the pH of pure water is > 7.
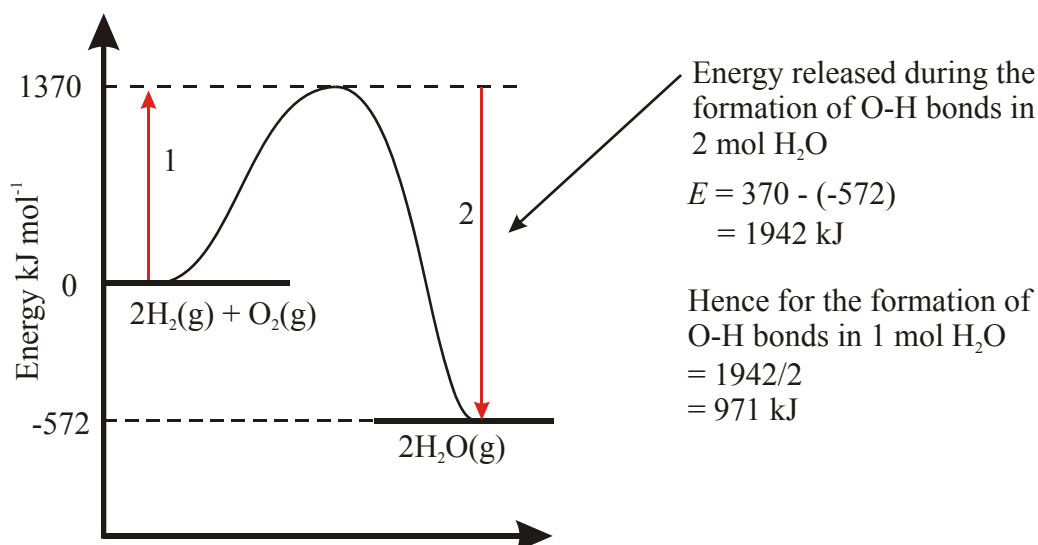
Q7   B   According to the electrochemical series, in 1 M $H_2SO_4(aq)$ the strongest oxidant present is $H^+(aq)$ and this is reduced to $H_2(g)$ at the cathode (-) during electrolysis; $2H^+(aq) + 2e^- \rightarrow H_2(g)$.
The strongest reductant present is $H_2O(l)$, which is oxidised at the anode (+) according to
$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$
$n(H_2)$ produced $= 2.0$ g / 2.0 g mol^{-1}
            $= 1.0$ mol
Since according to the half-equations
$n(H_2) = \frac{1}{2}$ x $n(e^-)$, and $n(O_2) = \frac{1}{4}$ x $n(e^-)$, then
$n(O_2) = \frac{1}{2}$ x $n(H_2)$
This is also consistent with the overall redox equation $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$
$n(O_2) = \frac{1}{2}$ x $1.0 = 0.50$ mol
$m(O_2) = 0.50$ mol x 32.0 g mol^{-1}
            $= 16.0$ g

Q8   B $E(cell) = E^0(oxidant\ half\text{-}cell) - E^0(reductant\ half\text{-}cell)$
So if $X^{2+}(aq)$ / $X(s)$ contains the oxidant
$0.94 = E^0(X^{2+}/X) - E^0(Ag^+/Ag)$
        $= E^0(X^{2+}/X) - 0.80$
$E^0(X^{2+}/X) = 1.74$ V
So if $X^{2+}(aq)$ / $X(s)$ contains the reductant
$0.94 = E^0(Ag^+/Ag) - E^0(X^{2+}/X)$
        $= 0.80 - E^0 (X^{2+}/X)$
$E^0(X^{2+}/X) = -0.14$ V
There is no half-equation on the electrochemical series with $E^0 = 1.74$, but at $E^0 = -0.14$ is the redox half-equation $Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$.
Hence X is most likely to be tin.

Q9   A   According to the equilibrium $2X_2(g) + 3Y_2(g) \rightleftharpoons 2X_2Y_3$,
2 mol X_2 reacts with 3 mol Y_2 to produce 2 mol X_2Y_3, so
1 mol X_2 could react with 1.5 mol Y_2 to produce 1 mol X_2Y_3.
However if all 1.0 mol X_2 initially present reacted, the system would not be at equilibrium – all reactants and products must be present if the system is at equilibrium.
So less than 1.0 mol X_2 will have reacted when equilibrium has been achieved and less than 1.0 mol X_2Y_3 will be present.

Q10   C   Calculate the reaction quotient (concentration fraction), $Q = [NO_2]^2 / ([NO]^2[O^2])$
                                    $= 0.10^2 / (0.10^2$ x $0.010)$
                                    $= 100$
Since Q (100) $< K_c$ ($3x10^6$), the system is not at equilibrium. Since Q must increase to get to the same value as K_c at equilibrium, the forward reaction must dominate as the system moves towards equilibrium. So the rate of the forward reaction will be faster than the rate of the reverse reaction and, since there are fewer particles on the product side, the pressure will decrease.

Q11 D
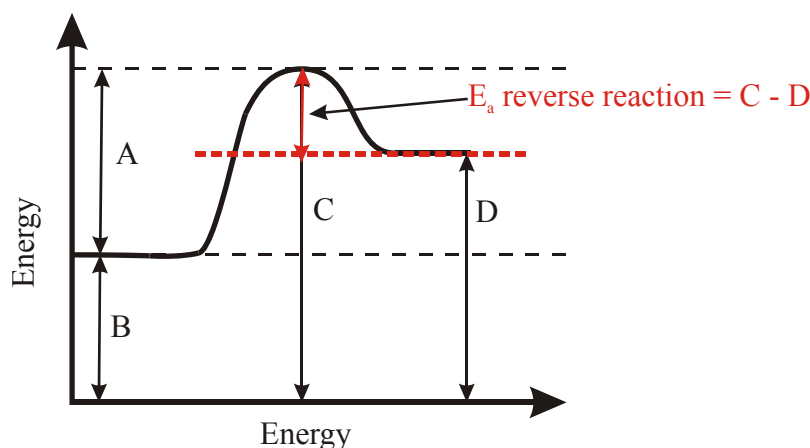
Q12 A   When 50 mL of water is added to 50 mL of the equilibrium mixture, all
concentrations instantaneously halve.
The concentration fraction $[Fe(NCS)^{2+}] / ([Fe^{3+}][SCN^-])$ doubles, i.e. becomes
greater than the equilibrium constant.
Since the concentration fraction must decrease for the system to get back to
equilibrium, the reverse reaction will be favoured. This will
A. increase the $m(Fe^{2+})$
B. increase the $[SCN^-]$ but the concentration at the new equilibrium will be
smaller than at the original equilibrium.
C. have no effect on the equilibrium constant since the temperature is the same at
both the initial and final equilibria.
D. increase the total number of particles, because there are more particles on the
reactant side of the equation.

Q13 C   Since all the changes in concentration as a response to the change in conditions
are gradual, rather than any of them being instantaneous, the change must have
been a change in temperature.
Since the concentrations of the reactants increase as a result of the temperature
change, the reverse reaction is favoured. According to the ΔH information, the
forward reaction is exothermic, hence the reverse reaction is endothermic and
favoured by an increase in temperature.

Q14 B   The fundamental difference between sea-water and tap water as far as this
equilibrium is concerned was the presence of $Cl^-(aq)$ ions in sea-water.
Thus the sea-water has higher $c(Cl^-)$. This has the effect of pushing the
equilibrium in sea-water further to the left, relative to the equilibrium in tap water.
So
A. the $[ClO^-]$ would be lower in sea-water
B. the sea-water would have a lower $[H_3O^+]$ hence a higher pH
C. since sea-water has a lower $[H_3O^+]$ and $[H_3O^+][OH^-]$ is constant, then the
$[OH^-]$ will be higher in sea-water.
D. it is as if $Cl^-(aq)$ was added to the tap water equilibrium; the $[Cl^-]$ increases
instantaneously and then decreases as the system moves back to equilibrium.
However it remains higher at the new equilibrium than at the original
equilibrium. So the $[Cl^-]$ will be higher in the sea-water equilibrium.

Q15 C The energy profile shows the energy needed to break the bonds (**1.**) in 2 mol H_2 and 1 mol O_2, i.e. the activation energy of the reaction, and the energy released (**2.**) when bonds are formed in 2 mol H_2O.
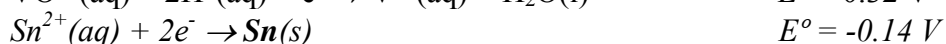

Energy released during the formation of O-H bonds in 2 mol H_2O

$E = 370 - (-572)$
    $= 1942$ kJ

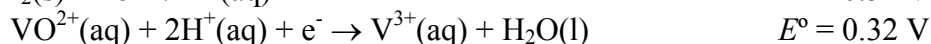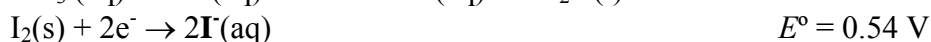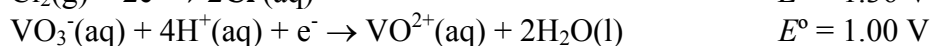Hence for the formation of O-H bonds in 1 mol H_2O
    $= 1942/2$
    $= 971$ kJ

Q16 B Access information for hypochlorous acid form the Data Book
        $HOCl(aq) + H_2O(l) \rightleftharpoons ClO^-(aq) + H_3O^+(aq)$; $K_a = 2.9 \times 10^{-8}$
        $[ClO^-][H_3O^+] / [HOCl] = 2.9 \times 10^{-5}$
        Weak acid assumptions:
        $[HOCl]_{equilibrium} =$ initial $[HOCl] = 0.050$ M
        $[ClO^-] = [H_3O^+]$
        So $[H_3O^+]^2 / 0.050 = 2.9 \times 10^{-8}$
            $[H_3O^+]^2 = 0.050 \times 2.9 \times 10^{-8}$
            $[H_3O^+] = \sqrt{(0.050 \times 2.9 \times 10^{-8})}$
                    $= 3.8 \times 10^{-5}$ M
        At 25°C, $[OH^-] = 10^{-14} / (3.8 \times 10^{-5})$
                    $= 2.6 \times 10^{-10}$ M

Q17 A At the cathode (-) the strongest oxidant present, i.e. $Zn^{2+}(aq)$, will be reduced according to
        $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$
        Charge passed $Q = It$
                        $= 1.2 \times 5.0 \times 60$
                        $= 360$ C
        $n(e^-) = Q / F$
            $= 360 / 96500$
            $= 3.73 \times 10^{-3}$ mol
        $n(Zn) = ½ \times n(e^-)$
            $= ½ \times 3.73 \times 10^{-3}$
            $= 1.87 \times 10^{-3}$ mol
        $m(Zn) = 1.87 \times 10^{-3} \times 65.4$
            $= 0.12$ g

Q18 D Ammonium vanadate dissociates in aqueous solution $NH_4VO_3(aq) \rightarrow NH_4^+(aq) + VO_3^-(aq)$

To get the green solution, the vanadate ion, $VO_3^-(aq)$, has to be reduced through to $V^{3+}(aq)$. So the reductant used must be capable of reducing $VO_3^-(aq)$ to $VO^{2+}(aq)$ and then to $V^{3+}(aq)$. But it must not be strong enough the reduce $V^{3+}(aq)$ to $V^{2+}(aq)$.

Including all the species added to the separate solutions of ammonium vanadate, i.e. Zn(s), $K^+(aq)$, $I^-(aq)$, Sn(s), $Na^+(aq)$, $Cl^-(aq)$ in the electrochemical series data.

$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$	$E^o = 1.36$ V
$VO_3^-(aq) + 4H^+(aq) + e^- \rightarrow VO^{2+}(aq) + 2H_2O(l)$	$E^o = 1.00$ V
$I_2(s) + 2e^- \rightarrow 2I^-(aq)$	$E^o = 0.54$ V
$VO^{2+}(aq) + 2H^+(aq) + e^- \rightarrow V^{3+}(aq) + H_2O(l)$	$E^o = 0.32$ V
$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$	*$E^o = -0.14$ V*
$V^{3+}(aq) + e^- \rightarrow V^{2+}(aq)$	$E^o = -0.26$ V
$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$	$E^o = -0.76$ V
$Na^+(aq) + e^- \rightarrow Na(s)$	$E^o = -2.71$ V
$K^+(aq) + e^- \rightarrow K(s)$	$E^o = -2.93$ V
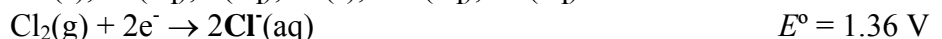
This shows that of the four reductants added, $Cl^-(aq)$, $I^-(aq)$, Sn(s) and Zn(s), the only one capable of reducing $VO_3^-(aq)$ to $VO^{2+}(aq)$ and then $VO^{2+}(aq)$ to $V^{3+}(aq)$ without reducing $V^{3+}(aq)$ to $V^{2+}(aq)$ is **tin, Sn(s)**.

$VO_3^-(aq) + 4H^+(aq) + e^- \rightarrow VO^{2+}(aq) + 2H_2O(l)$	$E^o = 1.00$ V
$VO^{2+}(aq) + 2H^+(aq) + e^- \rightarrow V^{3+}(aq) + H_2O(l)$	$E^o = 0.32$ V
$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$	*$E^o = -0.14$ V*
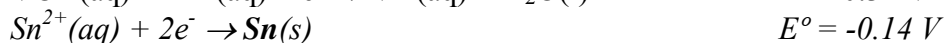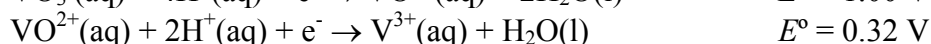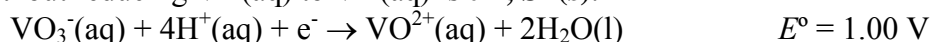
Q19 D Since NiOOH is reduced, it must be the oxidant. This suggests that H_2 must be the reductant.

Since H_2 is the reductant, it causes the reduction of NiOOH to $Ni(OH)_2$ and is itself oxidised. Since the electrolyte is alkaline, i.e. contains $OH^-(aq)$, the oxidation half-equation must contain $OH^-(aq)$ ions and may be obtained from the electrochemical series in the Data Book, i.e.

$H_2(g) + 2OH^-(aq) \rightarrow 2H_2O(l) + 2e^-$.
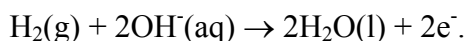
Q20 B $n(KOH) = 1.00 / 56.1$
$= 0.0178$ mol

Since the reaction is exothermic, energy is released when KOH dissolves.
Energy released $= n(KOH)$ x Energy per mole
$= 0.0178$ x 55.6
$= 0.991$ kJ

Because 991 kJ of energy is released in the reaction, the temperature inside the calorimeter rises. So to return the calorimeter to its temperature prior to the reaction, 0.991 kJ, i.e. 991 J of energy must be removed from the calorimeter.

Short Answer (Answers) – Section B

Question 1

a. $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$ ❶ *or*
    $CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$

b. In the reaction $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$, the law of conservation of mass holds because the total of the masses of $CaCl_2$, H_2O and CO_2 produced, i.e. the products is equal to the total of the masses of $CaCO_3$ and HCl reacting, i.e. the reactants.
    However because CO_2 is produced as a gas, if the reaction is carried out in an open flask the CO_2 will escape, i.e. be 'lost' to the atmosphere.❶ Hence monitoring the total decrease in mass of the reaction flask as the reaction proceeds will be equal to the total mass of CO_2 produced.❶

c. Since the $CaCO_3$ is in excess, the total mass loss will become constant when all of the hydrochloric acid has reacted. ❶

d.
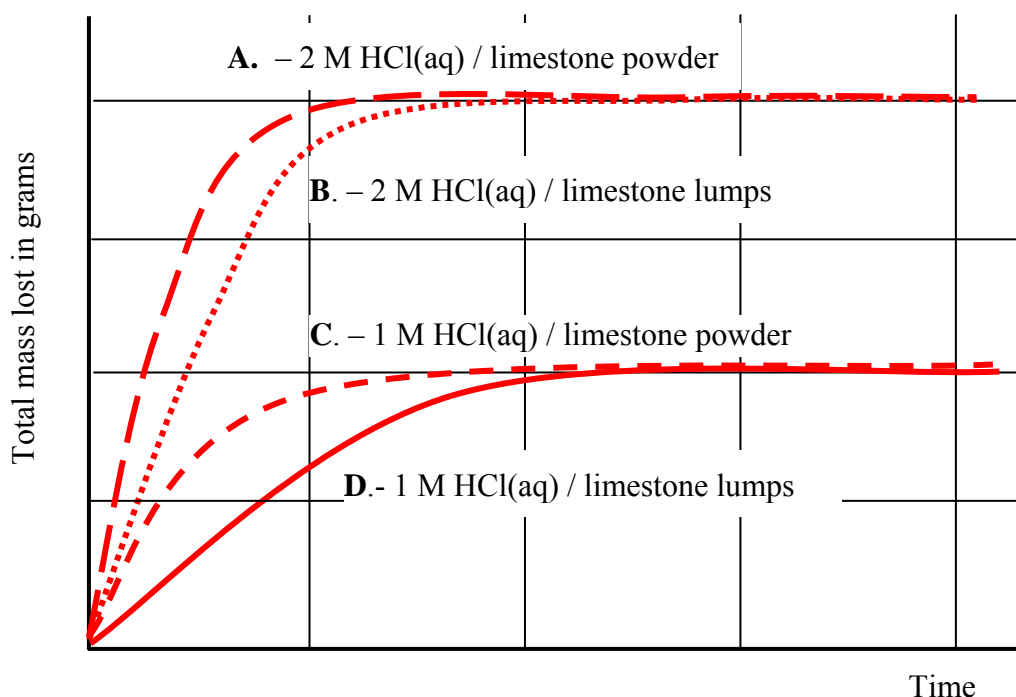
1. limestone lumps and 200 mL of 2.0 M HCl(aq) :- ***B***

2. limestone lumps and 200 mL of 1.0 M HCl(aq) :- ***D***

3. powdered limestone and 200 mL of 2.0 M HCl(aq) :- ***A***

4. powdered limestone and 200 mL of 1.0 M HCl(aq) :- ***C***   ❶❶ - 4 x ½

e. Powdered limestone provides a greater surface area for reaction than do the limestone lumps, and the 2 M HCl(aq) provides more particles (H^+ ions) to collide with the $CaCO_3(s)$ than does 1 M HCl(aq).
    So in reaction A there is a both a greater surface area of $CaCO_3$ ❶ with which H^+ ions can collide and more H^+ ions ❶ available to collide with the $CaCO_3$. This allows more collisions per second between H^+ ions and $CaCO_3$ collisions, more successful collisions (collisions with energy greater than the activation energy) per second ❶and hence a faster reaction rate.

f. Since the $CaCO_3$ is in excess the $m(CO_2)$ produced depends on the $n(HCl)$. 200 mL of 2.0 M HCl(aq) has twice the $n(HCl)$ as 200 mL of 1.0 M HCl(aq) and so should produce approximately twice as much CO_2. ❶

g. Either **measure the volume of CO₂ produced** by collecting it in a gas syringe *or* use a pH meter to **monitor the pH** of the solution (since the $[H^+]$ decreases the pH should increase as the reaction proceeds). ❶

Question 2

Aparatus 1
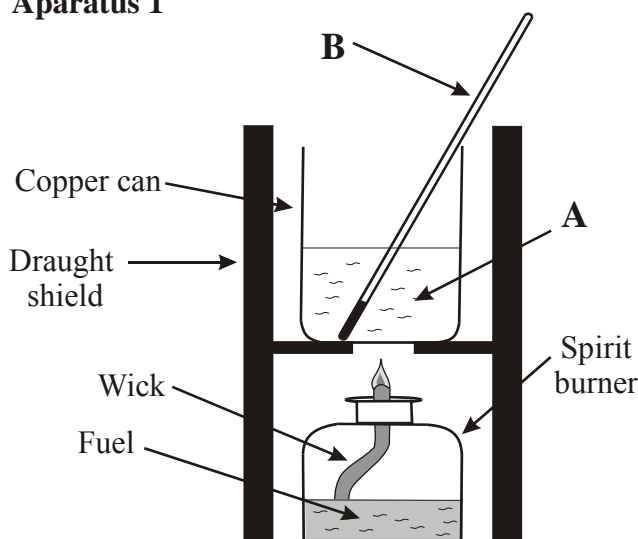
Aparatus 2
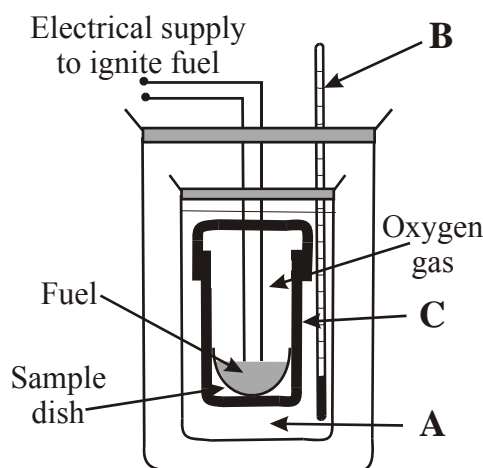
a. A:water / H_2O ❶; B: thermometer ❶; C: reaction bomb. ❶

b. Calibration factor = 1.52 kJ / 2.78°C
= 0.547 kJ °C⁻¹ ❶

Energy released by 5.00×10^{-3} mol fuel = 0.547 kJ °C⁻¹ x 37.5°C
= 20.5 kJ

Energy released by one mole of fuel = 20.5 / 5.00×10^{-3}
= 4.10×10^3 kJ ❶

ΔH_c(fuel) = -4.10×10^3 kJ mol⁻¹
According to the Data Book, the fuel is most likely to be hexane, C_6H_{14} ❶

c. i. Energy absorbed by water = SHC (H_2O) x $m(H_2O)$ x ΔT
Since the density is 1.00 g mL⁻¹, $m(H_2O)$ = 200 g, ΔT = 12.35 °C = 12.35 K
Energy absorbed by water = 4.18 J g⁻¹ K⁻¹ x 200 g x 15.35 K
= 1.28×10^4 J
= 12.8 kJ ❶

ii. $m(C_6H_{14})$ reacting = 86.450 – 86.025
= 0.425 g
$n(C_6H_{14})$ consumed = 0.425 g / 86 g mol⁻¹
= 4.94×10^{-3} mol ❶

iii. Energy released by 1 mol C_6H_{14} = 12.8 kJ / 4.94×10^{-3} mol
= 2.59×10^3 kJ mol⁻¹

$\Delta H_c(C_6H_{14})$ = -2.59×10^3 kJ mol⁻¹ ❶

d. Less energy per mole of fuel was calculated for Apparatus 1. Possible reasons include
   1. loss of heat released by the burning fuel directly to the atmosphere; a draught shield is not as effective as the insulation in Apparatus 2.
   2. loss of heat transferred from the water to the atmosphere through the open top of the copper can.
   3. heat released from the burning fuel transferred to components of the system, i.e. stand, copper can, thermometer
   4. evaporation of some fuel through the wick of the spirit burner, i.e. not all of the measured amount of fuel may have actually undergone combustion.
   5. incomplete combustion of the fuel
      ❶❶ - two marks for any two valid reasons.

Question 3

a. $HCOOH(aq) + H_2O(l) \rightleftharpoons HCOO^-(aq) + H_3O^+(aq)$ ❶

b. $K = [HCOO^-][H_3O^+][HCOOH]$ ❶ or
   $[HCOO^-][H_3O^+][HCOOH] = 1.8 \times 10^{-4}$

c. $n(HCOOH) = 4.0 \times 10^{-3}$ g $/ 46.0$ g mol^{-1}
   $= 8.7 \times 10^{-5}$ mol
   $c(HCOOH) = 8.7 \times 10^{-5}$ mol $/ 2.5 \times 10^{-3}$ L
   $= 0.0348$ M ❶

   Weak acid assumptions:
   $[HCOOH]_{equilibrium} = [HCOOH]$ initially $= 0.0348$ M
   $[HCOO^-] = [H_3O^+]$
   $[H_3O^+]^2 / [HCOOH] = 1.8 \times 10^{-4}$
   $[H_3O^+]^2 / 0.0348 = 1.8 \times 10^{-4}$
   $[H_3O^+] = \sqrt{(0.0348 \times 1.8 \times 10^{-4})}$
   $= 2.50 \times 10^{-3}$ M ❶
   pH $= -\log_{10}(2.50 \times 10^{-3})$
   $= 2.60$ ❶

Question 4

a. Since ΔH (-90 kJ mol^{-1}) is < 0, the forward reaction is exothermic and would be favoured at lower temperatures. Hence for this equilibrium, the yield would be greater at lower temperatures and smaller at higher temperatures. ❶

b. Because the forward reaction is exothermic, the yield of methanol is better at lower temperatures. However lower temperature means a slower rate of reaction. So the use of moderately low temperature (400°C) and a catalyst provides a compromise to this conflict between rate and yield and allows for both a good yield and a high rate of reaction. ❶

c. The pressure in a gaseous equilibrium system depends on the total number of particles present. According to the equation $CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$, the production of methanol reduces the total number of particles and so reduces the pressure. ❶
   According to Le Chatelier's principle, the system will move to decrease the pressure, and increase the yield of CH_3OH, if it is subjected to a pressure increase. ❶

d. $2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(g)$ ❶ or
   $CH_3OH(l) + 1.5O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

e. $CH_3OH(l)$ is oxidised to $CO_2(g)$ at the (-) electrode (anode) in a fuel cell
   $CH_3OH(l) + H_2O(l) \rightarrow CO_2(g) + 6H^+(aq) + 6e^-$ ❶

Question 5

a. i. The positive (electrode) ❶. In electrolysis or cell recharging the (+) terminal of the power supply is always connected to the (+) electrode.

ii. $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$ ❶

iii. The locations of Cu, Ni and Au on the electrochemical series provide the clues. Since Cu must be oxidised at the anode, (+) electrode, Ni(s) will also be oxidised, to $Ni^{2+}(aq)$ because it is a stronger reductant than Cu.
However because $Cu^{2+}(aq)$ is a stronger oxidant, $Ni^{2+}(aq)$ will not be reduced at the cathode, (-) electrode, and will remain in the electrolyte solution. ❶
Since Cu is a much stronger reductant than Au, the voltage applied to ensure oxidation of Cu in the impure copper will not be enough to oxidise the Au. Hence over time the Au atoms will fall to the bottom of the cell and collect under the impure copper electrode. ❶

b. According to the electrochemical series, $F_2(g)$ will react with $H_2O(l)$ since the oxidant, F_2, is higher than the reductant, H_2O. ❶
The overall equation for the reaction can be established from the half-equations
Reduction: $F_2(g) + 2e^- \rightarrow 2F^-(aq)$
Oxidation: $2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$
Overall equation: $2F_2(g) + 2H_2O(l) \rightarrow 4F^-(aq) + 4H^+(aq) + O_2(g)$ ❶
The pH of the water decreases because of the increase in concentration of H^+ during the reaction. ❶

Question 6

a. i. Since the amount of HI present is lower at equilibrium, the forward reaction dominates as the reaction moves to equilibrium.

	$2HI(g)$	$\rightleftharpoons$	$H_2(g)$	$+ I_2(g)$	
Initially	0.070		-	-	mol
Reacting	*0.050*	$\rightarrow$	*0.025*	*0.025*	*mol*
Equilibrium	0.020		**0.025**	**0.025**	mol ❶

Since there are equal numbers of particles on both sides of the equation, the volume of the container is not needed to work out the value of K. You can assume a volume of one litre
$$K = [H_2][I_2] / [HI]^2$$
$$= 0.025 \times 0.025 / 0.020^2$$
$$= 1.6 \text{ ❶}$$

ii. Since the amount of I_2 present is lower at equilibrium, the reverse reaction dominates as the reaction moves to equilibrium.

	$2HI(g)$	$\rightleftharpoons$	$H_2(g)$	$+ I_2(g)$	
Initially	-		0.040	0.025	mol
Reacting	*0.040*	$\leftarrow$	*0.020*	*0.020*	*mol*
Equilibrium	**0.040**		**0.020**	0.005	mol ❶

Since there are equal numbers of particles on both sides of the equation, the volume of the container is not needed to work out the value of K. You can assume a volume of one litre
$$K = [H_2][I_2] / [HI]^2$$
$$= 0.020 \times 0.005 / 0.040^2$$
$$= 0.06 \text{ ❶}$$

iii. The forward reaction is endothermic, and so is favoured by higher temperature. So the yield of products, and the value of K will increase as the temperature increases. Hence Experiment 1, with the higher K value, was at the higher temperature at equilibrium. ❶
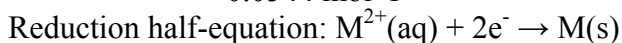
b.

i. ❶	NH$_3$	HNO$_3$	C$_2$H$_4$	H$_2$SO$_4$
ii. ❶ ❶	$N_2 + 3H_2 \rightleftharpoons 2NH_3$ Exothermic **K decreases**	$2NO + O_2 \rightleftharpoons 2NO_2$ Exothermic **K decreases**	$C_3H_8 \rightleftharpoons C_2H_4 + CH_4$ Endothermic **K increases**	$2SO_2 + O_2 \rightleftharpoons 2SO_3$ Exothermic **K decreases**
iii. ❶ ❶	e.g. Reductant Nitric acid production e.g. Base Fertiliser production	e.g. Acid Fertiliser production e.g. Oxidant Production of NO$_2$	e.g. Unsaturated hydrocarbon Addition polymer production.	e.g. Dehydrating agent Drying gases. e.g. Acid Superphosphate production e.g. Oxidant Reactions with metals

Question 7

a. Q $= It$
              $= 2.50 \times 35.0 \times 60$
              $= 5.25 \times 10^3$ C
    $n(e^-)$ $= Q / F$
              $= 5.25 \times 10^3 / 96500$
              $= 0.0544$ mol ❶
Reduction half-equation: $M^{2+}(aq) + 2e^- \rightarrow M(s)$
    $n(M)$ $= \frac{1}{2} \times n(e^-)$
              $= \frac{1}{2} \times 0.0544$
              $= 0.0272$ mol ❶
Molar mass M $= m(M) / n(M)$
              $= 3.06 / 0.0272$
              $= 112.5$ g mol^{-1}
Element is cadmium, Cd ❶

b.   i. $E(cell) = E^0(Cu^{2+}/Cu) - E^0(Cd^{2+}/Cd)$
              $= 0.34 - (-0.40)$
              $= 0.74$ V ❶
If the solutions in the half-cells are less than 1 M, e.g. 0.10 M $Cu^{2+}(aq)$ and 0.10 M $Cd^{2+}(aq)$, the generated voltage would be less than 0.74 V. ❶

    ii. KNO_3
$NO_3^-(aq)$ ions flow from the salt bridge to counteract the build up of positive charge in the solution due to the oxidation half-equation ($Cd(s) \rightarrow Cd^{2+}(aq) + 2e^-$), and so maintain the half-cell solution charge neutrality essential for effective operation of the galvanic cell. ❶

End of Suggested Answers