

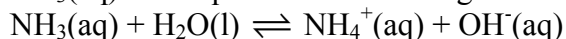
Suggested Answers

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 $\text{CH}_3\text{CH}_2\text{OH}(\text{l}) + \text{CH}_3\text{COOH}(\text{l}) \rightarrow \text{CH}_3\text{COOCH}_2\text{CH}_3(\text{l}) + \text{H}_2\text{O}(\text{l})$ – a condensation reaction, proceeds at the same rate as the reverse reaction $\text{CH}_3\text{COOCH}_2\text{CH}_3(\text{l}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{CH}_3\text{CH}_2\text{OH}(\text{l}) + \text{CH}_3\text{COOH}(\text{l})$, 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 $\text{Zn}^{2+}(\text{l})$ and $\text{Cl}^-(\text{l})$
At the anode (+), i.e. electrode Q, the reductant, $\text{Cl}^-(\text{l})$ is oxidised according to $2\text{Cl}^-(\text{l}) \rightarrow \text{Cl}_2(\text{g}) + 2\text{e}^-$,
i.e. chloride ions lose electrons at electrode Q.
At the cathode (-), i.e. electrode P, the oxidant, $\text{Zn}^{2+}(\text{l})$ is reduced according to $\text{Zn}^{2+}(\text{l}) + 2\text{e}^- \rightarrow \text{Zn}(\text{s})$
- 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 $[\text{products}] / [\text{reactants}]$, 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 $[\text{OH}^-]$ 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 $[\text{OH}^-]$, which will then decrease as the system returns to equilibrium.
B. Since the self-ionisation of water $2\text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{OH}^-(\text{aq})$ is endothermic, cooling pure water will favour the reverse reaction and so the $[\text{OH}^-]$ 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 $2\text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{OH}^-(\text{aq})$ is endothermic, heating pure water will favour the forward reaction and so the $[\text{OH}^-]$ will increase.

D. $\text{NH}_3(\text{aq})$ is in equilibrium according to



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

Q5 C $n(\text{NaOH}) \text{ reacting} = 1.0 \times 25 \times 10^{-3} = 0.025 \text{ mol}$

$n(\text{HCl}) \text{ reacting} = 1.0 \times 25 \times 10^{-3} = 0.025 \text{ mol}$

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 .

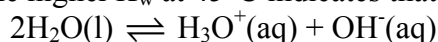
C. 0.050 mol HCl and 0.050 mol NaOH would release twice as much energy, so the total solution volume of 100 mL, (i.e. twice the original total volume of 50 mL) would increase by 6°C .

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 = 1.0×10^{-14} at 25°C .

In pure water, $[\text{H}_3\text{O}^+]$ is always equal to $[\text{OH}^-]$ but they are only both 10^{-7} M at 25°C .

The higher K_w at 45°C indicates that the equilibrium



lies further to the right at 45°C than at 25°C ,

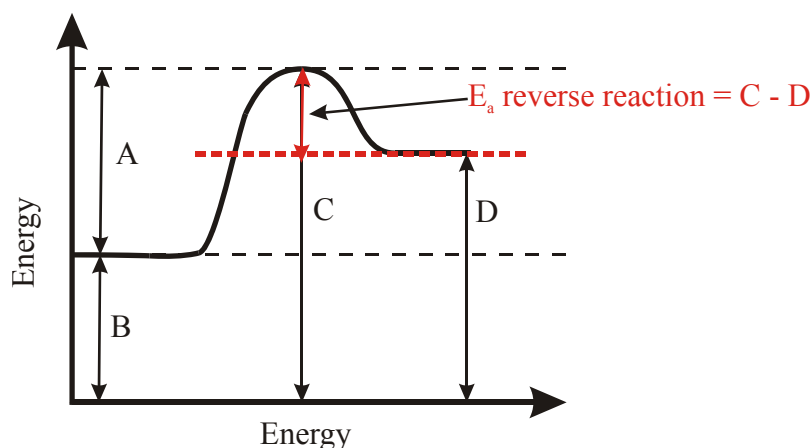
i.e. $[\text{H}_3\text{O}^+] > 10^{-7} \text{ M}$ and $[\text{OH}^-] > 10^{-7} \text{ M}$

Consequently at 15°C the equilibrium lies further to the left than at 25°C and so the $[\text{H}_3\text{O}^+] < 10^{-7} \text{ M}$ and $[\text{OH}^-] < 10^{-7} \text{ M}$

Since $[\text{H}_3\text{O}^+] < 10^{-7} \text{ M}$ at 25°C , the pH of pure water is > 7 .

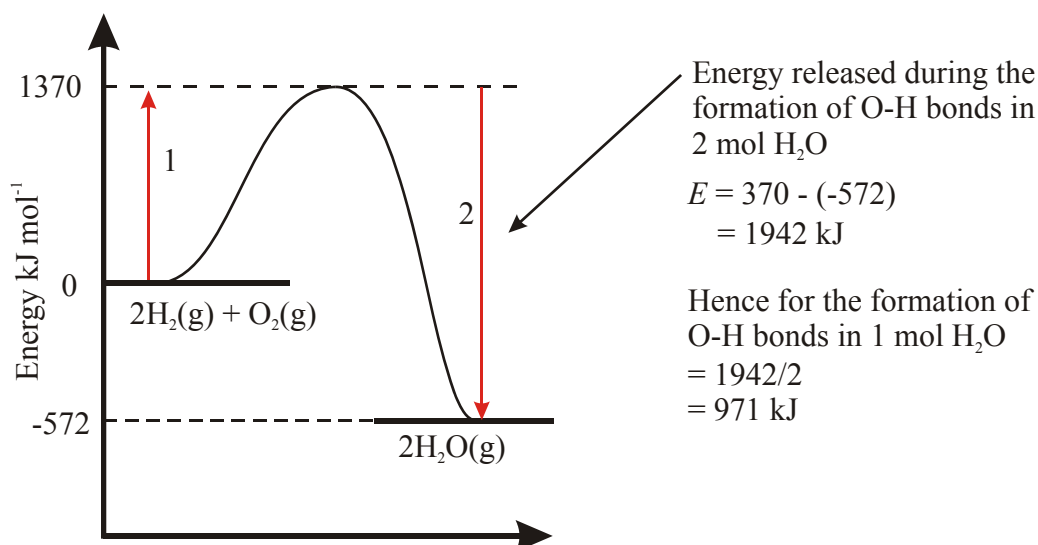
- Q7 B According to the electrochemical series, in 1 M $\text{H}_2\text{SO}_4(\text{aq})$ the strongest oxidant present is $\text{H}^+(\text{aq})$ and this is reduced to $\text{H}_2(\text{g})$ at the cathode (-) during electrolysis;
 $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$.
 The strongest reductant present is $\text{H}_2\text{O}(\text{l})$, which is oxidised at the anode (+) according to
 $2\text{H}_2\text{O}(\text{l}) \rightarrow \text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^-$
 $n(\text{H}_2) \text{ produced} = 2.0 \text{ g} / 2.0 \text{ g mol}^{-1}$
 $= 1.0 \text{ mol}$
 Since according to the half-equations
 $n(\text{H}_2) = \frac{1}{2} \times n(\text{e}^-)$, and $n(\text{O}_2) = \frac{1}{4} \times n(\text{e}^-)$, then
 $n(\text{O}_2) = \frac{1}{2} \times n(\text{H}_2)$
 This is also consistent with the overall redox equation $2\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{H}_2(\text{g}) + \text{O}_2(\text{g})$
 $n(\text{O}_2) = \frac{1}{2} \times 1.0 = 0.50 \text{ mol}$
 $m(\text{O}_2) = 0.50 \text{ mol} \times 32.0 \text{ g mol}^{-1}$
 $= 16.0 \text{ g}$
- Q8 B $E(\text{cell}) = E^0(\text{oxidant half-cell}) - E^0(\text{reductant half-cell})$
 So if $\text{X}^{2+}(\text{aq}) / \text{X}(\text{s})$ contains the oxidant
 $0.94 = E^0(\text{X}^{2+}/\text{X}) - E^0(\text{Ag}^+/\text{Ag})$
 $= E^0(\text{X}^{2+}/\text{X}) - 0.80$
 $E^0(\text{X}^{2+}/\text{X}) = 1.74 \text{ V}$
 So if $\text{X}^{2+}(\text{aq}) / \text{X}(\text{s})$ contains the reductant
 $0.94 = E^0(\text{Ag}^+/\text{Ag}) - E^0(\text{X}^{2+}/\text{X})$
 $= 0.80 - E^0(\text{X}^{2+}/\text{X})$
 $E^0(\text{X}^{2+}/\text{X}) = -0.14 \text{ V}$
 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 $\text{Sn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Sn}(\text{s})$.
 Hence X is most likely to be tin.
- Q9 A According to the equilibrium $2\text{X}_2(\text{g}) + 3\text{Y}_2(\text{g}) \rightleftharpoons 2\text{X}_2\text{Y}_3$,
 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), $Q = [\text{NO}_2]^2 / ([\text{NO}]^2[\text{O}_2])$
 $= 0.10^2 / (0.10^2 \times 0.010)$
 $= 100$
 Since $Q (100) < K_c (3 \times 10^6)$, 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 $[\text{Fe}(\text{NCS})^{2+}] / ([\text{Fe}^{3+}][\text{SCN}^-])$ 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
- increase the $m(\text{Fe}^{2+})$
 - increase the $[\text{SCN}^-]$ but the concentration at the new equilibrium will be smaller than at the original equilibrium.
 - have no effect on the equilibrium constant since the temperature is the same at both the initial and final equilibria.
 - 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 $\text{Cl}^-(\text{aq})$ ions in sea-water. Thus the sea-water has higher $c(\text{Cl}^-)$. This has the effect of pushing the equilibrium in sea-water further to the left, relative to the equilibrium in tap water. So
- the $[\text{ClO}^-]$ would be lower in sea-water
 - the sea-water would have a lower $[\text{H}_3\text{O}^+]$ hence a higher pH
 - since sea-water has a lower $[\text{H}_3\text{O}^+]$ and $[\text{H}_3\text{O}^+][\text{OH}^-]$ is constant, then the $[\text{OH}^-]$ will be higher in sea-water.
 - it is as if $\text{Cl}^-(\text{aq})$ was added to the tap water equilibrium; the $[\text{Cl}^-]$ 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 $[\text{Cl}^-]$ 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 H₂ and 1 mol O₂, i.e. the activation energy of the reaction, and the energy released (2.) when bonds are formed in 2 mol H₂O.



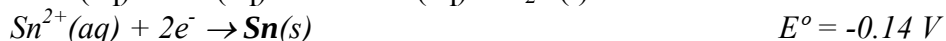
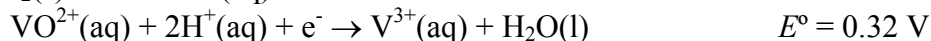
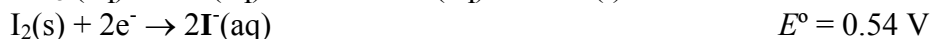
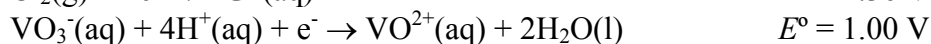
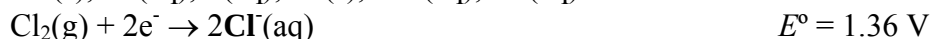
- Q16 B Access information for hypochlorous acid from the Data Book
 $\text{HOCl(aq)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{ClO}^-\text{(aq)} + \text{H}_3\text{O}^+\text{(aq)}; \quad K_a = 2.9 \times 10^{-8}$
 $[\text{ClO}^-][\text{H}_3\text{O}^+] / [\text{HOCl}] = 2.9 \times 10^{-5}$
 Weak acid assumptions:
 $[\text{HOCl}]_{\text{equilibrium}} = \text{initial } [\text{HOCl}] = 0.050 \text{ M}$
 $[\text{ClO}^-] = [\text{H}_3\text{O}^+]$
 So $[\text{H}_3\text{O}^+]^2 / 0.050 = 2.9 \times 10^{-8}$
 $[\text{H}_3\text{O}^+]^2 = 0.050 \times 2.9 \times 10^{-8}$
 $[\text{H}_3\text{O}^+] = \sqrt{(0.050 \times 2.9 \times 10^{-8})}$
 $= 3.8 \times 10^{-5} \text{ M}$
 At 25°C, $[\text{OH}^-] = 10^{-14} / (3.8 \times 10^{-5})$
 $= 2.6 \times 10^{-10} \text{ M}$

- Q17 A At the cathode (-) the strongest oxidant present, i.e. Zn²⁺(aq), will be reduced according to
 $\text{Zn}^{2+}\text{(aq)} + 2\text{e}^- \rightarrow \text{Zn(s)}$
 Charge passed $Q = It$
 $= 1.2 \times 5.0 \times 60$
 $= 360 \text{ C}$
 $n(\text{e}^-) = Q / F$
 $= 360 / 96500$
 $= 3.73 \times 10^{-3} \text{ mol}$
 $n(\text{Zn}) = \frac{1}{2} \times n(\text{e}^-)$
 $= \frac{1}{2} \times 3.73 \times 10^{-3}$
 $= 1.87 \times 10^{-3} \text{ mol}$
 $m(\text{Zn}) = 1.87 \times 10^{-3} \times 65.4$
 $= 0.12 \text{ g}$

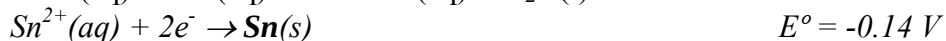
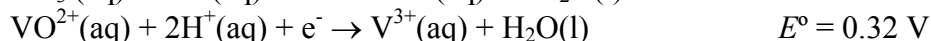
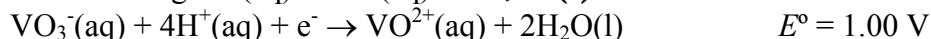
Q18 D Ammonium vanadate dissociates in aqueous solution $\text{NH}_4\text{VO}_3(\text{aq}) \rightarrow \text{NH}_4^+(\text{aq}) + \text{VO}_3^-(\text{aq})$

To get the green solution, the vanadate ion, $\text{VO}_3^-(\text{aq})$, has to be reduced through to $\text{V}^{3+}(\text{aq})$. So the reductant used must be capable of reducing $\text{VO}_3^-(\text{aq})$ to $\text{VO}^{2+}(\text{aq})$ and then to $\text{V}^{3+}(\text{aq})$. But it must not be strong enough to reduce $\text{V}^{3+}(\text{aq})$ to $\text{V}^{2+}(\text{aq})$.

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

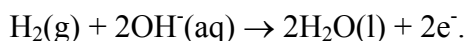


This shows that of the four reductants added, $\text{Cl}^-(\text{aq})$, $\text{I}^-(\text{aq})$, $\text{Sn}(\text{s})$ and $\text{Zn}(\text{s})$, the only one capable of reducing $\text{VO}_3^-(\text{aq})$ to $\text{VO}^{2+}(\text{aq})$ and then $\text{VO}^{2+}(\text{aq})$ to $\text{V}^{3+}(\text{aq})$ without reducing $\text{V}^{3+}(\text{aq})$ to $\text{V}^{2+}(\text{aq})$ is **tin, Sn(s)**.



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 $\text{Ni}(\text{OH})_2$ and is itself oxidised. Since the electrolyte is alkaline, i.e. contains $\text{OH}^-(\text{aq})$, the oxidation half-equation must contain $\text{OH}^-(\text{aq})$ ions and may be obtained from the electrochemical series in the Data Book, i.e.



Q20 B $n(\text{KOH}) = 1.00 / 56.1$
 $= 0.0178 \text{ mol}$

Since the reaction is exothermic, energy is released when KOH dissolves.

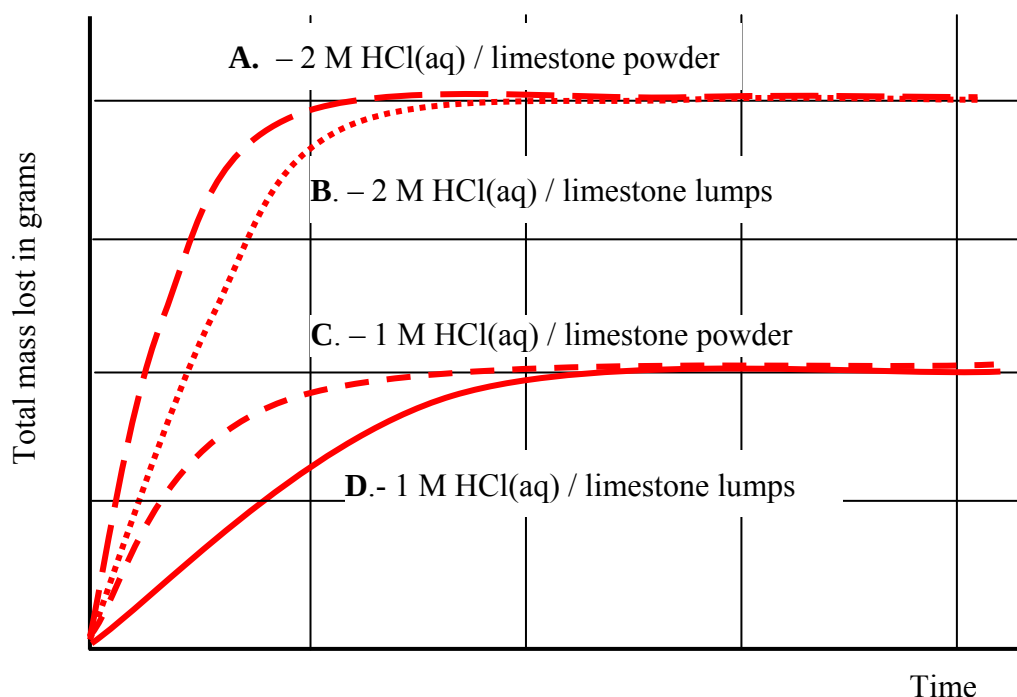
$$\begin{aligned} \text{Energy released} &= n(\text{KOH}) \times \text{Energy per mole} \\ &= 0.0178 \times 55.6 \\ &= 0.991 \text{ kJ} \end{aligned}$$

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. $\text{CaCO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{CaCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$ ❶ *or*
 $\text{CaCO}_3(\text{s}) + 2\text{H}^+(\text{aq}) \rightarrow \text{Ca}^{2+}(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$
- b. In the reaction $\text{CaCO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{CaCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$, 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.

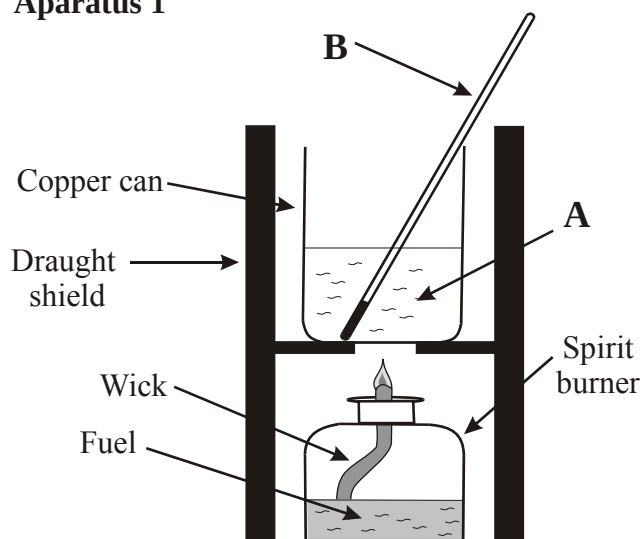


1. limestone lumps and 200 mL of 2.0 M $\text{HCl}(\text{aq})$:- B
 2. limestone lumps and 200 mL of 1.0 M $\text{HCl}(\text{aq})$:- D
 3. powdered limestone and 200 mL of 2.0 M $\text{HCl}(\text{aq})$:- A
 4. powdered limestone and 200 mL of 1.0 M $\text{HCl}(\text{aq})$:- C ❶ ❶ - 4 x ½
- e. Powdered limestone provides a greater surface area for reaction than do the limestone lumps, and the 2 M $\text{HCl}(\text{aq})$ provides more particles (H^+ ions) to collide with the $\text{CaCO}_3(\text{s})$ than does 1 M $\text{HCl}(\text{aq})$.
 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 ❶ and hence a faster reaction rate.

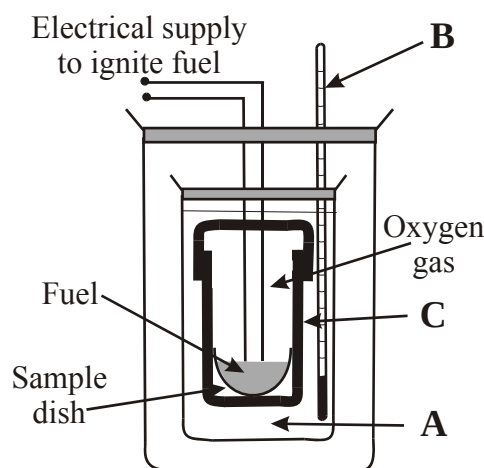
- f. Since the CaCO_3 is in excess the $m(\text{CO}_2)$ produced depends on the $n(\text{HCl})$. 200 mL of 2.0 M $\text{HCl}(\text{aq})$ has twice the $n(\text{HCl})$ as 200 mL of 1.0 M $\text{HCl}(\text{aq})$ and so should produce approximately twice as much CO_2 . ❶
- g. Either **measure the volume of CO_2 produced** by collecting it in a gas syringe *or* use a pH meter to **monitor the pH** of the solution (since the $[\text{H}^+]$ 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 = $1.52 \text{ kJ} / 2.78^\circ\text{C}$
 $= 0.547 \text{ kJ } ^\circ\text{C}^{-1}$ ❶

$$\text{Energy released by } 5.00 \times 10^{-3} \text{ mol fuel} = 0.547 \text{ kJ } ^\circ\text{C}^{-1} \times 37.5^\circ\text{C} = 20.5 \text{ kJ}$$

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

$$\Delta H_c(\text{fuel}) = -4.10 \times 10^3 \text{ kJ mol}^{-1}$$

According to the Data Book, the fuel is most likely to be hexane, C_6H_{14} ❶

- c. i. Energy absorbed by water = $\text{SHC}(\text{H}_2\text{O}) \times m(\text{H}_2\text{O}) \times \Delta T$
 Since the density is 1.00 g mL^{-1} , $m(\text{H}_2\text{O}) = 200 \text{ g}$, $\Delta T = 12.35^\circ\text{C} = 12.35 \text{ K}$
 Energy absorbed by water = $4.18 \text{ J g}^{-1} \text{ K}^{-1} \times 200 \text{ g} \times 12.35 \text{ K}$
 $= 1.04 \times 10^4 \text{ J}$
 $= 10.4 \text{ kJ}$ ❶
- ii. $m(\text{C}_6\text{H}_{14})$ reacting = $86.450 - 86.025$
 $= 0.425 \text{ g}$
 $n(\text{C}_6\text{H}_{14})$ consumed = $0.425 \text{ g} / 86 \text{ g mol}^{-1}$
 $= 4.94 \times 10^{-3} \text{ mol}$ ❶
- iii. Energy released by 1 mol C_6H_{14} = $10.4 \text{ kJ} / 4.94 \times 10^{-3} \text{ mol}$
 $= 2.11 \times 10^3 \text{ kJ mol}^{-1}$
 $\Delta H_c(\text{C}_6\text{H}_{14}) = -2.11 \times 10^3 \text{ kJ mol}^{-1}$ 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. $\text{HCOOH}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{HCOO}^-(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$ ❶
- b. $K = [\text{HCOO}^-][\text{H}_3\text{O}^+]/[\text{HCOOH}]$ ❶ *or*
 $[\text{HCOO}^-][\text{H}_3\text{O}^+]/[\text{HCOOH}] = 1.8 \times 10^{-4}$
- c. $n(\text{HCOOH}) = 4.0 \times 10^{-3} \text{ g} / 46.0 \text{ g mol}^{-1}$
 $= 8.7 \times 10^{-5} \text{ mol}$
 $c(\text{HCOOH}) = 8.7 \times 10^{-5} \text{ mol} / 2.5 \times 10^{-3} \text{ L}$
 $= 0.0348 \text{ M}$ ❶

Weak acid assumptions:

$$[\text{HCOOH}]_{\text{equilibrium}} = [\text{HCOOH}]_{\text{initially}} = 0.0348 \text{ M}$$

$$[\text{HCOO}^-] = [\text{H}_3\text{O}^+]$$

$$[\text{H}_3\text{O}^+]^2 / [\text{HCOOH}] = 1.8 \times 10^{-4}$$

$$[\text{H}_3\text{O}^+]^2 / 0.0348 = 1.8 \times 10^{-4}$$

$$[\text{H}_3\text{O}^+] = \sqrt{(0.0348 \times 1.8 \times 10^{-4})}$$

$$= 2.50 \times 10^{-3} \text{ M}$$
 ❶

$$\text{pH} = -\log_{10}(2.50 \times 10^{-3})$$

$$= 2.60$$
 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 $\text{CO}(\text{g}) + 2\text{H}_2(\text{g}) \rightleftharpoons \text{CH}_3\text{OH}(\text{g})$, 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. $2\text{CH}_3\text{OH}(\text{l}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\text{g})$ ❶ *or*
 $\text{CH}_3\text{OH}(\text{l}) + 1.5\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$
- e. $\text{CH}_3\text{OH}(\text{l})$ is oxidised to $\text{CO}_2(\text{g})$ at the (-) electrode (anode) in a fuel cell
 $\text{CH}_3\text{OH}(\text{l}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{CO}_2(\text{g}) + 6\text{H}^+(\text{aq}) + 6\text{e}^-$ ❶

Question 5

- a. i. The positive (electrode) **1**. In electrolysis or cell recharging the (+) terminal of the power supply is always connected to the (+) electrode.
- ii. $\text{Cu(s)} \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{e}^-$ **1**
- 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 $\text{Ni}^{2+}(\text{aq})$ because it is a stronger reductant than Cu. However because $\text{Cu}^{2+}(\text{aq})$ is a stronger oxidant, $\text{Ni}^{2+}(\text{aq})$ will not be reduced at the cathode, (-) electrode, and will remain in the electrolyte solution. **1** 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. **1**
- b. According to the electrochemical series, $\text{F}_2(\text{g})$ will react with $\text{H}_2\text{O}(\text{l})$ since the oxidant, F_2 , is higher than the reductant, H_2O . **1**
 The overall equation for the reaction can be established from the half-equations
 Reduction: $\text{F}_2(\text{g}) + 2\text{e}^- \rightarrow 2\text{F}^-(\text{aq})$
 Oxidation: $2\text{H}_2\text{O}(\text{l}) \rightarrow \text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^-$
 Overall equation: $2\text{F}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow 4\text{F}^-(\text{aq}) + 4\text{H}^+(\text{aq}) + \text{O}_2(\text{g})$ **1**
 The pH of the water decreases because of the increase in concentration of H^+ during the reaction. **1**

Question 6

- a. i. Since the amount of HI present is lower at equilibrium, the forward reaction dominates as the reaction moves to equilibrium.
- | | | | | |
|-------------|---|---------------|----------------------|--------------|
| | $2\text{HI}(\text{g}) \rightleftharpoons \text{H}_2(\text{g}) + \text{I}_2(\text{g})$ | | | |
| Initially | 0.070 | - | - | mol |
| Reacting | 0.050 | $\rightarrow$ | 0.025 0.025 | mol |
| Equilibrium | 0.020 | | 0.025 0.025 | mol 1 |
- 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
- $$K = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2}$$
- $$= 0.025 \times 0.025 / 0.020^2$$
- $$= 1.6 \quad \textbf{1}$$
- ii. Since the amount of I_2 present is lower at equilibrium, the reverse reaction dominates as the reaction moves to equilibrium.
- | | | | | |
|-------------|---|--------------|----------------------|--------------|
| | $2\text{HI}(\text{g}) \rightleftharpoons \text{H}_2(\text{g}) + \text{I}_2(\text{g})$ | | | |
| Initially | - | | 0.040 0.025 | mol |
| Reacting | 0.040 | $\leftarrow$ | 0.020 0.020 | mol |
| Equilibrium | 0.040 | | 0.020 0.005 | mol 1 |
- 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
- $$K = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2}$$
- $$= 0.020 \times 0.005 / 0.040^2$$
- $$= 0.06 \quad \textbf{1}$$

- 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_2H_4	H_2SO_4
ii. ❶ ❶	$\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$ Exothermic K decreases	$2\text{NO} + \text{O}_2 \rightleftharpoons 2\text{NO}_2$ Exothermic K decreases	$\text{C}_3\text{H}_8 \rightleftharpoons \text{C}_2\text{H}_4 + \text{CH}_4$ Endothermic K increases	$2\text{SO}_2 + \text{O}_2 \rightleftharpoons 2\text{SO}_3$ 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. $Q = It$
 $= 2.50 \times 35.0 \times 60$
 $= 5.25 \times 10^3 \text{ C}$
 $n(e^-) = Q / F$
 $= 5.25 \times 10^3 / 96500$
 $= 0.0544 \text{ mol}$ ❶
Reduction half-equation: $\text{M}^{2+}(\text{aq}) + 2e^- \rightarrow \text{M}(\text{s})$
 $n(\text{M}) = \frac{1}{2} \times n(e^-)$
 $= \frac{1}{2} \times 0.0544$
 $= 0.0272 \text{ mol}$ ❶
Molar mass $M = m(\text{M}) / n(\text{M})$
 $= 3.06 / 0.0272$
 $= 112.5 \text{ g mol}^{-1}$

Element is cadmium, Cd ❶

b. i. $E(\text{cell}) = E^0(\text{Cu}^{2+}/\text{Cu}) - E^0(\text{Cd}^{2+}/\text{Cd})$
 $= 0.34 - (-0.40)$
 $= 0.74 \text{ V}$ ❶

If the solutions in the half-cells are less than 1 M, e.g. 0.10 M $\text{Cu}^{2+}(\text{aq})$ and 0.10 M $\text{Cd}^{2+}(\text{aq})$, the generated voltage would be less than 0.74 V. ❶

- ii. KNO_3
 $\text{NO}_3^-(\text{aq})$ ions flow from the salt bridge to counteract the build up of positive charge in the solution due to the oxidation half-equation ($\text{Cd}(\text{s}) \rightarrow \text{Cd}^{2+}(\text{aq}) + 2e^-$), and so maintain the half-cell solution charge neutrality essential for effective operation of the galvanic cell. ❶

End of Suggested Answers