Insight Publications

# *INSIGHT*

***Trial Exam Paper***

# 2008
# CHEMISTRY
# Written examination 2

## *Solutions book*

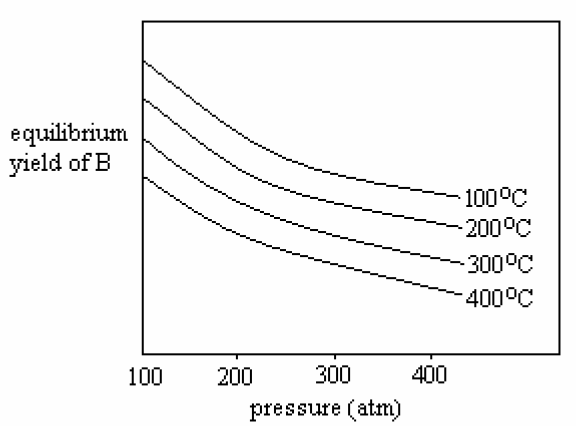
**This book presents:**

- correct solutions
- explanatory notes
- mark allocations
- tips and guidelines

## SECTION A – Multiple-choice questions

### Question 1

Reactant A exists in equilibrium with product B. No other chemical species is present. The graph below shows the effect of increasing pressure and temperature on the equilibrium yield of product B.




Which one of the following could be a correct equation and $\Delta H$ value for this reaction?

**A.** **$A(g) \rightleftharpoons 2B(g)$** **$\Delta H = -64$ kJ mol$^{-1}$**

**B.** $2A(g) \rightleftharpoons B(g)$ $\Delta H = +64$ kJ mol$^{-1}$

**C.** $A(g) \rightleftharpoons 2B(g)$ $\Delta H = +64$ kJ mol$^{-1}$

**D.** $A(aq) \rightleftharpoons 2B(aq)$ $\Delta H = -64$ kJ mol$^{-1}$

***Answer is A.***

**Explanatory notes**

- A is correct. Le Chatelier's principle states that if a reaction system is subject to change it will adjust itself to partially oppose the effects of that change. According to the graph, increasing pressure at constant temperature decreases the yield of product B. This response indicates that there are more particles on the product side of the equation because increasing the pressure will cause a shift in the direction of fewer particles, in this case backwards.

  Also according to the graph, increasing temperature at constant pressure also decreases the yield of product B. This response indicates the forward direction of the reaction must be exothermic because increasing the temperature always causes a shift in the endothermic direction, in this case backwards. The $\Delta H$ value will be negative to indicate that the forward reaction is exothermic.
- B is incorrect because the equation shows more particles on the reactant side and the $\Delta H$ value is positive, indicating the reaction is endothermic. This reaction would increase the yield of product B when both pressure and temperature were increased.
- C is incorrect because the reaction is exothermic, not endothermic.
- D is incorrect because the states of reactant A and product B are aqueous, indicating they are in solution. Only gases will respond to changes in pressure. Solutions will respond to changes in concentration.

**Tips**

- *Remember that a negative $\Delta H$ value indicates that a reaction is exothermic and a positive $\Delta H$ value indicates that a reaction is endothermic. A reaction proceeding in the exothermic direction produces heat and increases the temperature of its surroundings. A reaction proceeding in the endothermic direction absorbs heat and lowers the temperature of its surroundings.*

**Question 2**

Which of these Lowry-Bronsted acids would have the **strongest** conjugate base?

**A.** **$H_3BO_3$**
**B.** HF
**C.** HCl
**D.** HOCl

***Answer is A.***

**Explanatory notes**

- A is correct because $H_3BO_3$ has the lowest acidity constant, $K_a$, of $5.8 \times 10^{-10}$. A conjugate acid-base pair differ by a single proton, $H^+$. The acidity constant is a measure of the strength of an acid. The lower the $K_a$ then the weaker the acid; the stronger the conjugate base then the greater the extent of the backwards reaction. Therefore, the conjugate base of $H_3BO_3$ is $H_2BO_3^-$.
- B is incorrect because the $K_a$ value of HF, which is $7.6 \times 10^{-4}$, is not the smallest $K_a$ value. $F^-$ is not the strongest conjugate base.
- C is incorrect because HCl is a strong acid. It completely ionises in water. Its conjugate base, $Cl^-$, is a very weak base.
- D is incorrect because the $K_a$ value of HOCl, which is $2.9 \times 10^{-8}$, is not the smallest $K_a$ value. $OCl^-$ is not the strongest conjugate base.

**Tips**

- *The $K_a$ values of these and some other weak acids are located in the data booklet. The larger the value of the equilibrium constant, the greater the extent of the forward reaction. So, with respect to the ionisation of acids, the larger the $K_a$, the greater the percentage ionisation and the stronger the acid.*

**Question 3**

Which of the following changes will **always** ensure an increase in the rate of a chemical reaction?

I Adding a catalyst

II Increasing temperature

III Increasing concentration

IV Increasing the surface area of the reactant

**A.** I only

**B.** **I, II and III only**

**C.** I, II and IV only

**D.** I, II, III and IV

***Answer is B.***

**Explanatory notes**

- B is correct. Adding a reaction-specific catalyst always increases the reaction rate of a chemical reaction by providing an alternative pathway for the reaction with a lower activation energy. Increasing the temperature always increases the reaction rate because the reactant particles, on average, have more energy. They collide more often and, more importantly, a greater proportion of collisions are fruitful because it is easier for the particles to overcome the activation energy barrier and react. Increasing concentration always increases the reaction rate because more particles in the same volume will collide more frequently. Adding additional reactant to the mixture will not always increase the rate of a chemical reaction. If the addition does not also increase the concentration, the particles will not collide any more frequently.
- A is incorrect because II and III also increase the reaction rate.
- C is incorrect because III increases the reaction rate and IV does not ensure it is increased.
- D is incorrect because IV does not ensure an increase in the reaction rate.

**Tips**

- *Read these types of questions very carefully to check whether the question is referring to the rate or yield of a chemical reaction. Rate refers to how quickly a reaction will reach equilibrium, whereas yield refers to the proportion of product in the reaction mixture at equilibrium.*

*Questions 4 and 5 refer to the following information.*

Hydrocyanic acid, HCN, is a weak acid. When added to water it reacts according to the equation

$$HCN(aq) + H_2O(l) \rightleftharpoons CN^-(aq) + H_3O^+(aq)$$

**Question 4**

When three drops of 3.0 M NaOH are added to an equilibrium mixture of HCN in water at constant temperature

**A.** the pH of the solution decreases.

**B.** the ratio $\frac{[CN^-][H_3O^+]}{[HCN]}$ decreases.

**C.** the concentration of HCN increases.

**D.** **the concentration of $CN^-$ increases.**

***Answer is D.***

**Explanatory notes**

- D is correct. Le Chatelier's principle states that if a reaction system is subject to change it will adjust itself to partially oppose the effects of that change. NaOH is a strong base. The $OH^-$ ions will react with $H_3O^+(aq)$ ions, effectively removing them from the solution. The system will respond to this change by shifting to the right to produce more $H_3O^+$ ions. The concentration of the other product, $CN^-$, will increase.
- A is incorrect. Addition of a strong base will cause the solution to become less acidic and its pH will increase.
- B is incorrect because the addition or removal of product and changing concentrations will not affect the *K* value (equilibrium constant), which is calculated using this ratio. Changing the temperature is the only change that will alter the *K* value.
- C is incorrect because the system responds to the addition of NaOH by shifting to the right. The concentration of HCN decreases.

**Question 5**

The amount, in mol, of HCN that needs to be dissolved completely in water to form a 1.0 L solution of hydrocyanic acid with a pH of 4.9 is closest to

**A.** $3.4 \times 10^{-4}$

**B.** **0.25**

**C.** 6.8

**D.** 3.7

***Answer is B.***

**Explanatory notes**

- B is correct according to these steps.
  Step 1: Calculate the concentration of $H_3O^+$ ions using the pH.
  $[H_3O^+] = 10^{-pH}$
  $= 10^{-4.9}$
  $= 1.3 \times 10^{-5}$ M

  Step 2: Calculate the concentration of HCN using the $K_a$ value and the assumption that $[H_3O^+] = [CN^-]$.

  $$K_a = \frac{[H_3O^+][CN^-]}{[HCN]}$$

  $$K_a = \frac{[H_3O^+]^2}{[HCN]}$$

  $$6.3 \times 10^{-10} = \frac{(1.3 \times 10^{-5})^2}{[HCN]}$$

  $$[HCN] = \frac{(1.3\times10^{-5})^2}{6.3\times10^{-10}}$$

  $$= \frac{1.6\times10^{-10}}{6.3\times10^{-10}}$$

  $= 0.25$ M

  Step 3: Calculate the amount, in mol, of HCN in the 1.0 L solution.

  $n(HCN) = cV$
  $= 0.25 \times 1.0$
  $= 0.25$ mol
- A is incorrect because HCN is a weak acid and does not completely ionise in water. [HCN] is not equal to $[H_3O^+]$.
- C is incorrect because 6.8 is the mass of HCN required in *grams* and not the amount in *mol*, as asked for.
- D is incorrect because $[HCN] = \frac{[H_3O^+]^2}{K_a}$ , not $[HCN] = \frac{K_a}{[H_3O^+]^2}$ .

**Question 6**

Water self-ionises according to the equation

$$2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq) \qquad \Delta H = +57 \text{ kJ mol}^{-1}$$

Pure water will have an ionisation constant, $K_w$, of $7.29 \times 10^{-14}$ $M^2$ when the pH and temperature are

| | pH | Temperature |
|---|---|---|
| **A.** | 7.00 | 25°C |
| **B.** | 6.57 | 5°C |
| **C.** | 6.26 | 55°C |
| **D.** | **6.57** | **55°C** |

***Answer is D.***

**Explanatory notes**

- D is correct. The pH can be calculated according to the steps below.

  Step 1: Determine the concentration of $[H_3O^+]$ ions according to the $K_w$ value.

  $[H_3O^+][OH^-] = K_w$

  $[H_3O^+]^2 = K_w$ because $H_3O^+$ ions and $OH^-$ ions are produced in the same ratio.

  $[H_3O^+] = \sqrt{7.29 \times 10^{-14}}$

  $= 2.70 \times 10^{-7}$ M

  Step 2: Calculate the pH.

  $pH = -\log_{10}[H_3O^+]$

  $= -\log_{10}(2.70 \times 10^{-7})$

  $= 6.57$

  The pH of pure water at 25°C is 7.00. This decreased pH is due to an increased concentration of $H_3O^+$ ions, meaning the equilibrium has shifted to the right from where it was at 25°C. The reaction is endothermic, so the temperature must have increased to cause this right shift in the endothermic direction.
- A is incorrect because the ionisation constant of pure water at 25°C is $1.00 \times 10^{-14}$. The $K_w$ in this question is different, so the water must be a different temperature.
- B is incorrect because the self-ionisation of water is endothermic. The shift to the right that increased $[H_3O^+]$ must have been caused by an increase in temperature and not a decrease.
- C is incorrect since the ratio of $H_3O^+$ and $OH^-$ in equation is 1 : 1 and not 2 : 1.

**Tips**

- *The pH of pure water can vary with changing temperatures, however, pure water is still considered neutral because $[H_3O^+] = [OH^-]$.*

**Question 7**

A student has two solutions of hydrochloric acid, HCl. Solution A has a pH of 2 and solution B has a pH of 3. It can be deduced that

**A.** solution A is twice as concentrated as solution B.

**B.** solution A has a volume ten times greater than solution B.

**C.** HCl has ionised more completely in solution A than in solution B.

**D.** **solution A is ten times more concentrated than solution B.**

***Answer is D.***

**Explanatory notes**

- D is correct. $[H_3O^+] = 10^{-pH}$, so solution A with a pH of 2 is 10 times more concentrated than solution B.
- A is incorrect because solution A is ten times more concentrated than solution B. The pH scale is logarithmic, so a difference of +1 in pH equates to a decrease of ×10 in $[H_3O^+]$.
- B is incorrect because the volume of a solution does not affect its pH given the concentration of the solute remains constant. pH is a measure of the concentration of $H_3O^+$ ions.
- C is incorrect because HCl is a strong acid and always ionises completely in solution.

**Question 8**

Iron(III) nitrate is added to tin(II) chloride and a reaction occurs. The reactant at the cathode and the polarity of the cathode are

| | Cathode | Polarity of cathode |
|---|---|---|
| **A.** | **$Fe^{3+}$** | **positive** |
| **B.** | $Sn^{2+}$ | positive |
| **C.** | $Sn^{4+}$ | negative |
| **D.** | $Fe^{3+}$ | negative |

***Answer is A.***

**Explanatory notes**

- A is correct. A spontaneous reaction occurs so the polarity of the cathode is positive. $Fe^{3+}$ is the reactant because on the left side of the electrochemical series, it is higher than $Sn^{2+}$, making it the strongest oxidant present. $Fe^{3+}$ is reduced and reduction always occurs at the cathode.
- B is incorrect because $Sn^{2+}$ is a weaker oxidant than $Fe^{3+}$ so is not reduced at the cathode. $Sn^{2+}$ will be oxidised at the anode.
- C is incorrect because $Sn^{4+}$ is not present in the cell so cannot be reduced at the cathode. It is also a weaker oxidant than $Fe^{3+}$ because it is below it on the left side of the electrochemical series. Also, the polarity of the cathode is positive if a spontaneous reaction occurs.
- D is incorrect because the polarity of the cathode is positive if a spontaneous reaction occurs.

**Tips**

- *Reduction is the gain of electrons. It always occurs at the cathode and the species that undergoes reduction is referred to as the oxidant. Oxidation is the loss of electrons and always occurs at the anode. The species that undergoes oxidation is referred to as the reductant. AN OIL RIG CAT is a useful way to remember that at the ANode Oxidation Is Loss and Reduction Is Gain at the CAThode.*

**SECTION A** – continued

**Question 9**

Three half-cells were set up under standard conditions.

| Half-cell | Electrode | Electrolyte |
|---|---|---|
| I | metal A | $A^{2+}(aq)$ |
| II | metal B | $B^{3+}(aq)$ |
| III | metal C | $C^{+}(aq)$ |

When a galvanic cell is constructed from half-cell I and half-cell II, the electrode in half-cell I is positive.

When a galvanic cell is constructed from half-cell I and half-cell III, metal C is deposited at one electrode.

The **increasing** order of reactivity of the metals from least reactive to most reactive is

**A.** A, B, C

**B.** B, A, C

**C.** **C, A, B**

**D.** C, B, A

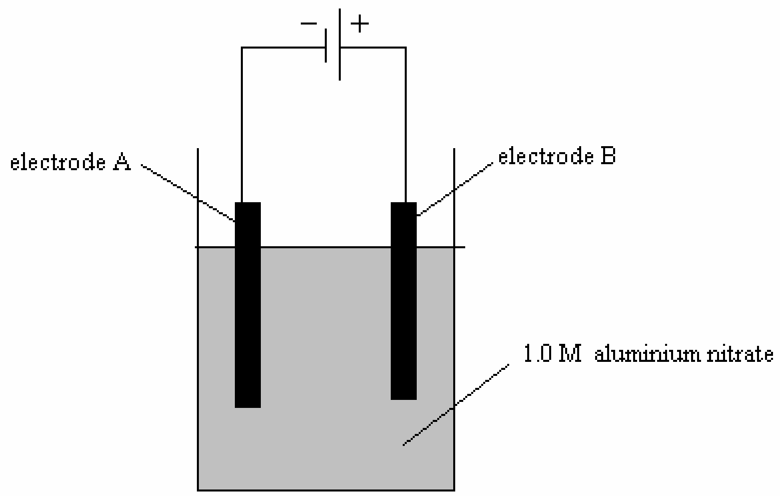
***Answer is C.***

**Explanatory notes**

- C is correct. In a galvanic cell the positive electrode is the cathode at which reduction occurs. In the galvanic cell constructed from half-cell I and half-cell II, metal A is the cathode at which reduction is occurring. The reduction reaction must be $A^{2+}(aq) + 2e^{-} \rightleftharpoons A(s)$. Oxidation must be occurring in half-cell II according to the reaction $B(s) \rightleftharpoons B^{3+}(aq) + 3e^{-}$. It can be concluded that B is a more reactive metal than A.
  In the galvanic cell constructed from half-cell I and half-cell III, metal C is deposited, so $C^{+}$ ions are being reduced according to the equation $C^{+}(aq) + e^{-} \rightleftharpoons C(s)$. Oxidation must be occurring in half-cell I according to the reaction $A(s) \rightleftharpoons A^{2+}(aq) + 2e^{-}$. It can be concluded that A is a more reactive metal than C. The increasing order of reactivity of the metals must be C, A, B.
- A is incorrect because in the second constructed galvanic cell, metal A reacted preferentially to metal C; hence, metal A must be more reactive than metal C.
- B is incorrect because it places the metals in the reverse order of reactivity. Metal B is more reactive than metal A because it reacts preferentially to metal A in the first constructed cell, and metal A is more reactive than metal C because it reacts preferentially to metal C in the second constructed cell.
- D is incorrect because in the first constructed galvanic cell, metal B reacted preferentially to metal A; hence, metal B must be more reactive than metal A.

**Question 10**

Consider the electrolytic cell shown below, which operates at 25°C. Several drops of indicator solution are added.




A student observing the cell would see

**A.** bubbles of gas on electrode A and an overall increase in the pH of the solution.

**B.** **bubbles of gas on electrodes A and B.**

**C.** a metallic coating on electrode A and bubbles of gas on electrode B.

**D.** bubbles of gas on electrode B and an overall decrease in the pH of the solution.

***Answer is B.***

**Explanatory notes**

- B is correct. Electrode A is the cathode because it is connected to the negative terminal of the electrical source. Electrons are sent to electrode A and reduction of the strongest oxidant occurs. The species present in the solution are $Al^{3+}(aq)$, $NO_3^-(aq)$ and $H_2O(l)$. The strongest oxidant, according to the electrochemical series, is $H_2O(l)$. It will react according to the equation

  $2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$. Bubbles of $H_2$ gas will appear at electrode A.

  Electrode B is the anode because it is connected to the positive terminal of the electrical source. Oxidation of the strongest reductant occurs here. The species present in the solution are $Al^{3+}(aq)$, $NO_3^-(aq)$ and $H_2O(l)$. The strongest reductant, according to the electrochemical series, is $H_2O(l)$. It will react according to the equation $2H_2O(l) \rightleftharpoons O_2(g) + 4H^+(aq) + 4e^-$. Bubbles of $O_2$ gas will appear at electrode B. The overall reaction will be $2H_2O(l) \rightleftharpoons 2H_2(g) + O_2(g)$. There will be no overall change in the pH of the solution.
- A is incorrect because there is no overall change in the pH of the solution as the overall reaction is $2H_2O(l) \rightleftharpoons 2H_2(g) + O_2(g)$.
- C is incorrect because $H_2O(l)$ is a stronger oxidant than $Al^{3+}(aq)$ and so will react preferentially.
- D is incorrect because there is no overall change in the pH of the solution as the overall reaction is $2H_2O(l) \rightleftharpoons 2H_2(g) + O_2(g)$.

**Tips**

- *Whenever a solution is electrolysed, water must be considered as a possible reactant. The only way to remove water from consideration is to have the reactant in molten form rather than aqueous. Al, Mg, Na, Ca, K or Li cannot be produced from solutions of their salts since $H_2O$ is a stronger oxidant and so will react preferentially.*
- *Remember that the electrochemical series is applicable only under certain conditions, which are 1 atm, 25°C and 1.0 M solutions. If any of these are varied, predictions based on the series may not be valid.*

**Question 11**

Which of the following conditions are required for the reliability of the electrochemical series?

| | |
|---|---|
| I | 25°C |
| II | pH of 7.0 |
| III | 1 M |
| IV | 1 atm |

**A.** I, II and III only

**B.** **I, III and IV only**

**C.** I, II, and IV only

**D.** I and III only

***Answer is B.***

**Explanatory notes**

- B is correct because one of the limitations of the electrochemical series is that it can predict reactions reliably only at conditions of 25°C, 1 M solutions and 1 atm pressure.
- A is incorrect because a pH of 7.0 is not a condition required for the reliability of the electrochemical series.
- C is incorrect because a pH of 7.0 is not a condition required for the reliability of the electrochemical series.
- D is incorrect because condition IV, 1 atm pressure, is also a condition that is needed for the electrochemical series to be valid.

**Question 12**

Which one of the following solutions will have the highest pH if they all have equal concentration?

**A.** 1.0 M $HNO_3$

**B.** 1.0 M $H_2SO_4$

**C.** **1.0 M $NH_3$**

**D.** 1.0 M HF

***Answer is C.***

**Explanatory notes**

- C is correct because $NH_3$ is the molecular formula of ammonia, a strong base. Bases have a pH greater than 7. The order of the solutions from lowest pH to highest pH is $H_2SO_4$, $HNO_3$, HF, $NH_3$.
- A is incorrect because $HNO_3$ is nitric acid, a strong acid. Acids have a pH below 7 and the stronger the acid, the lower the pH.
- B is incorrect because $H_2SO_4$ is a strong acid and is also diprotic, meaning that each acid molecule can donate two protons. It will have the lowest pH.
- D is incorrect because HF is a weak acid (and is listed in the table of weak acids in the data booklet). It will have a pH below 7 because it is an acid, but is not as low as a strong acid at the same concentration.

**Question 13**

A solution calorimeter was calibrated by passing 1.30 A through the electric heater for 2.00 minutes with a potential difference of 7.50 V. A temperature rise of 0.453°C was recorded. The calibration factor of the calorimeter, in J °C$^{-1}$, is

**A.** $3.87 \times 10^{-4}$

**B.** 43.0

**C.** 530

**D.** $\mathbf{2.58 \times 10^{3}}$

***Answer is D.***

**Explanatory notes**

- D is correct according to the following:

  Calibration factor $= \dfrac{VIt}{\Delta T}$

  $= \dfrac{7.50 \times 1.30 \times 2 \times 60}{0.453}$

  $= \dfrac{1170}{0.453}$

  $= 2.58 \times 10^{3}$ J °C$^{-1}$

- A is incorrect because the calibration factor $= \dfrac{VIt}{\Delta T}$, not $\dfrac{\Delta T}{VIt}$.
- B is incorrect because the time must be in seconds, not minutes.
- C is incorrect because the calibration factor $= \dfrac{VIt}{\Delta T}$, not $VIt \times \Delta T$.

**Question 14**

Which one of the following statements about a secondary cell when it is recharging is correct?

**A.** Oxidation is occurring at the negatively charged anode.

**B.** Electrons are flowing in the direction of cathode to anode.

**C.** **Electrical energy is being converted to chemical energy.**

**D.** Electrons are being lost by a chemical species at the negatively charged cathode.

***Answer is C.***

**Explanatory notes**

- C is correct. In the recharging process an electrical source is attached to the cell with the negative terminal attached to the intended cathode and the positive terminal attached to the intended anode. When switched on, electrical energy is converted to chemical energy.
- A is incorrect because the reaction in a recharging secondary cell is not spontaneous. Oxidation does occur at the anode during electrolysis but the polarity of the anode is positive.
- B is incorrect because electrons always flow in the direction of anode to cathode. Oxidation, the loss of electrons, occurs at the anode and the electrons flow towards the cathode, where they are gained in the reduction reaction.
- D is incorrect because although the cathode does carry a negative charge when the cell is recharging, reduction always occurs at the cathode and reduction is the gain of electrons.

**Question 15**

Which of the following is an advantage of nuclear fission over brown coal as an energy source?

I cost

II energy output per unit mass of fuel

III renewable

**A.** **II only**

**B.** II and III only

**C.** I and II only

**D.** I, II and III

***Answer is A.***

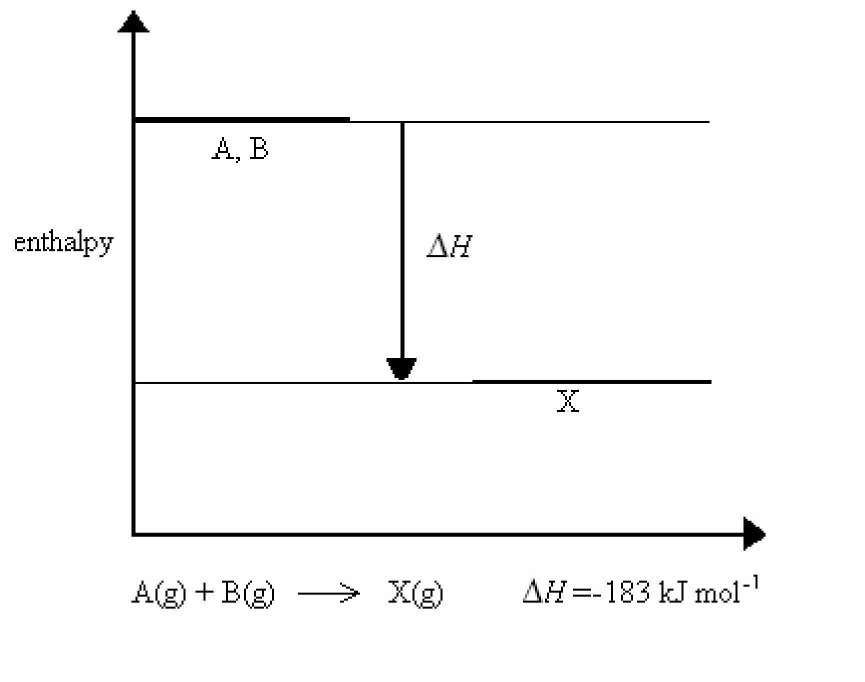
**Explanatory notes**

- A is correct because the energy output per unit mass of fuel of nuclear fission is much higher than for coal; however, it is more expensive. Both nuclear fission and coal are non-renewable sources of energy.
- B is incorrect because both nuclear fission and coal are non-renewable sources of energy.
- C is incorrect because nuclear fission is more expensive than coal and both nuclear fission and coal are non-renewable sources of energy.
- D is incorrect because nuclear fission is more expensive than coal and both nuclear fission and coal are non-renewable sources of energy.

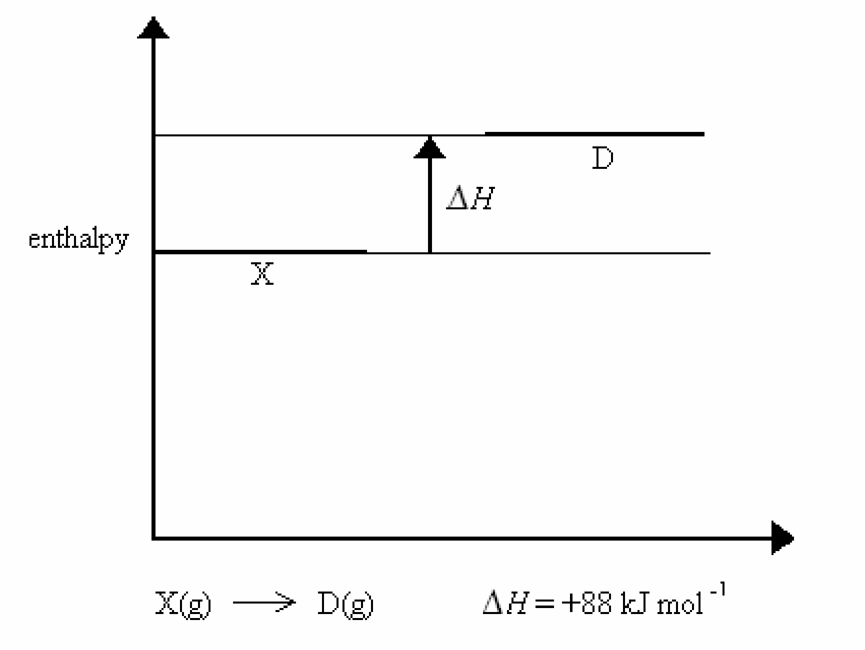
**Question 16**

Consider the following energy profiles for reactions between A and B.

**Energy profile A**




**Energy profile B**




The value of $\Delta H$, in kJ $mol^{-1}$, for the reaction A(g) + B(g) → D(g) is

**A.** **–95**

**B.** +95

**C.** –366

**D.** +366

***Answer is A.***

**Explanatory notes**

- A is correct. $A(g) + B(g) \rightarrow X(g)$ is exothermic and $X(g) \rightarrow D(g)$ is endothermic so $\Delta H(\text{for } A(g) + B(g) \rightarrow D(g)) = \Delta H(\text{for } A(g) + B(g) \rightarrow X(g)) + \Delta H(\text{for } X(g) \rightarrow D(g))$

  $\Delta H(\text{for } A(g) + B(g) \rightarrow D(g)) = -183 + 88$

  $= -95 \text{ kJ mol}^{-1}$

  Another way to visualise this might be:

  $$\begin{array}{lll} A(g) + B(g) \rightarrow X(g) & \Delta H = & -183 \\ \text{added} \quad X(g) \rightarrow D(g) & \Delta H = & 88 \\ \hline A(g) + B(g) \rightarrow D(g) & = & -95 \text{ kJ mol}^{-1} \end{array}$$

- B is incorrect because $\Delta H$ for $A(g) + B(g) \rightarrow D(g) = -183 + 88$. The overall reaction is exothermic and so has a negative $\Delta H$.
- C is incorrect because $\Delta H$ for $A(g) + B(g) \rightarrow D(g) = -183 + 88$, not $-183 - 88$.
- D is incorrect because $\Delta H$ for $A(g) + B(g) \rightarrow D(g) = -183 + 88$. Also, the overall reaction is exothermic and so has a negative $\Delta H$.

**Question 17**

An electroplating cell was constructed to deposit a zinc coating onto an iron nail. A voltage of 5.40 V and a current of 0.850 A are passed through a solution of 1 M zinc nitrate, $Zn(NO_3)_2$, for 300 seconds. The mass of zinc produced, in grams, is closest to

**A.** $1.4 \times 10^{-3}$
**B.** **0.0864**
**C.** 0.173
**D.** 0.467

***Answer is B.***

**Explanatory notes**

- B is correct according to the steps below.

  Step 1: Calculate the charge, in coulombs, applied to the cell.

  $Q = It$

  $= 0.850 \times 300$

  $= 255$ C

  Step 2: Determine the amount, in mol, of electrons.

  $n(e^-) = \frac{Q}{F}$

  $= \frac{255}{96\,500}$

  $= 0.00264$ mol

  Step 3: Write a half-equation for the reduction of zinc ions.

  $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$

  Step 4: Calculate the amount, in mol, of zinc produced.

  $n(e^-) : n(Zn)$

  2 : 1

  So, $n(Zn) = \frac{1}{2} \times n(e^-)$

  $= \frac{1}{2} \times 0.00264$

  $= 0.00132$ mol

  Step 5: Calculate the mass, in grams, of zinc deposited.

  $m = nM$

  $= 0.00132 \times 65.4$

  $= 0.0864$ g

- A is incorrect because the time must be expressed in seconds, not changed to 5.00 minutes.
- C is incorrect because the ratio of $n(Zn) : n(e^-)$ is 1 : 2, not 1 : 1.
- D is incorrect because $Q = It$, not $Q = VIt$. The voltage is not required for this calculation.

**Tips**

- *Be very clear on your formulas. $Q = It$ and $E = VIt$ are similar formulas and can be confused, especially when you are provided with the voltage in a question that does not require it.*

**Question 18**

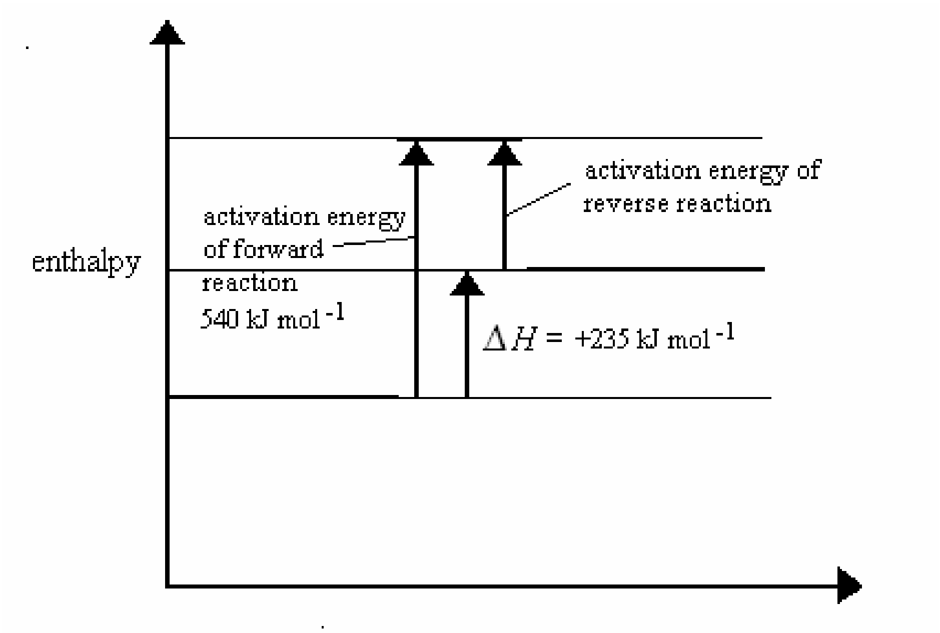
A chemical reaction has a $\Delta H$ of 235 kJ $mol^{-1}$. The activation energy for its forward reaction is 540 kJ $mol^{-1}$. The activation energy for its reverse reaction, in kJ $mol^{-1}$, is

**A.** **+305**
**B.** –305
**C.** +540
**D.** –540

***Answer is A.***

**Explanatory notes**

- A is correct because the backward reaction is exothermic. Its activation energy will be the activation energy of the forward (endothermic) reaction, therefore, the $\Delta H$ value = 540 – 235 = +305 kJ $mol^{-1}$. This is demonstrated in the energy profile below.




- B is incorrect because activation energy never has a negative charge, as it represents energy that must be absorbed to break the bonds in the reactants. It is always positive, even if the overall $\Delta H$ value is negative.
- C is incorrect because the activation energy of the forward reaction is always a different quantity to the reverse reaction.
- D is incorrect because the activation energy of the forward reaction is always a different quantity to the reverse reaction Also, activation energy never has a negative charge because it represents energy that must be absorbed to break the bonds in the reactants. It is always positive, even if the overall $\Delta H$ value is negative.

**Question 19**

Which of the following is a demonstration of Faraday's second law of electrolysis?

**A.** Current (amps) = time (seconds) × charge (coulombs)
**B.** The mass of lead deposited at the anode of an electrolytic cell is directly proportional to the quantity of energy passed through the cell.
**C.** The highest reaction in the electrochemical series that can occur in the forward direction is likely to occur at the cathode.

**D.** **The production of one mole of lead requires two moles of electrons to be consumed.**

***Answer is D.***

***Explanatory notes***

- D is correct because Faraday's second law of electrolysis shows that in order to produce one mole of a metal, one, two, three or another whole number of moles of electrons must be consumed.
- A is incorrect because this equation is incorrect, and it is not part of Faraday's second law.
- B is incorrect because the metal is always deposited at the cathode, where reduction occurs.
- C is incorrect because although this is a correct statement, it is not part of Faraday's second law.

## Question 20

Which of the following will cause the calculated enthalpy value for an exothermic reaction carried out in a bomb calorimeter to be lower than the actual value?

I poor insulation

II a thermometer that reads 3°C lower than the actual temperature

III a volume of water that was incorrectly measured and is more than it should be

IV a poorly fitting lid

**A.** I only

**B.** I, II and IV

**C.** **I, III and IV**

**D.** I, II, III and IV

***Answer is C.***

**Explanatory notes**

- C is correct because a thermometer that reads 3°C lower than it should will not lower the calculated enthalpy value. This is because it is the $\Delta T$ value that is used in the calculation of the calibration factor and enthalpy. A thermometer consistently reading 3°C lower than it should will not affect the $\Delta T$ value.
- A is incorrect because II will not result in a lower enthalpy value; however, I and IV will. Poor insulation and a poorly fitting lid will allow some heat to escape to the environment, making the measured $\Delta T$ of the water lower than it should, resulting in the calculated enthalpy value being lower.
- B is incorrect because II will not result in a lower enthalpy value; however, III will. If the volume of water is more than it should be then the heat released by the reaction will result in a lower $\Delta T$ of the water, so the calculated enthalpy value will be lower.
- D is incorrect because II will not result in a lower enthalpy value.

**END OF SECTION A**

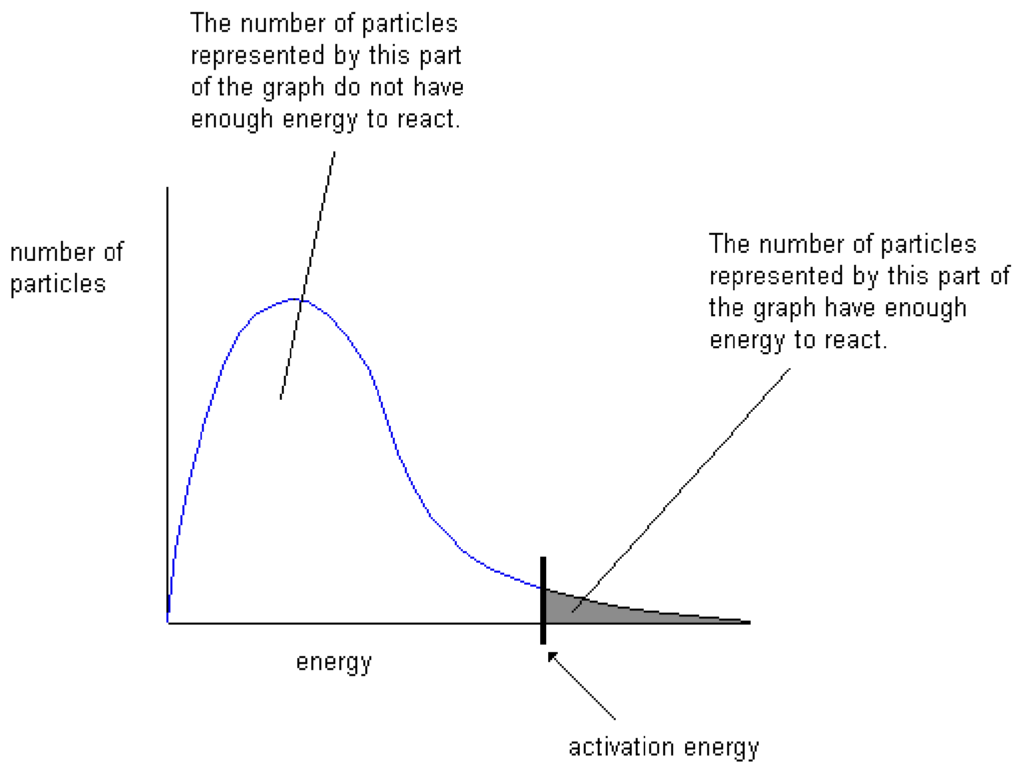
## SECTION B – Short-answer questions

### Question 1

Ethene and hydrogen chloride can react to produce chloroethane according to the equation

$$CH_2CH_2(g) + HCl(g) \rightarrow CH_3CH_2Cl(g)$$

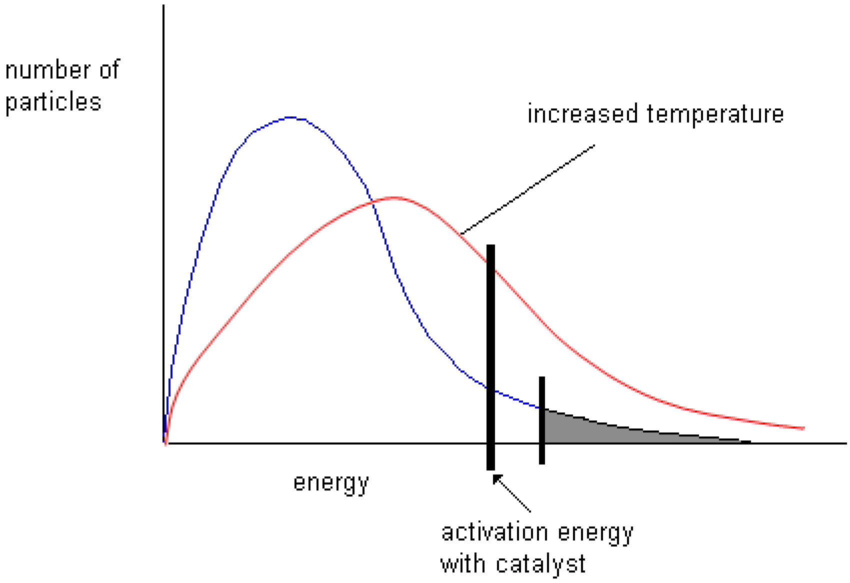
Some ethene and hydrogen chloride are placed in a sealed container and allowed to react at a fixed temperature. The graph below shows the range of energies possessed by the particles in the container and the number of particles with each particular energy.




**1a.** **i.** On the graph above, draw a line to show how the energies of the particles would differ if the temperature of the reaction vessel is increased. **Label this line.**

**ii.** On the graph above, mark with a line the new position of the activation energy if a catalyst was added to the vessel. **Label this line.**

**Solution**




1 + 1 = 2 marks

**Mark allocation**

- 1 mark for drawing an increased temperature curve that is more spread out along the *x*-axis but with the same area under the curve and shifted slightly to the right of the original curve.
- 1 mark for marking the activation energy with a catalyst as being lower than without.

**Explanatory notes**

Increasing the temperature of a reaction increases the average kinetic energy of all of the particles. This increases the rate of reaction because it is easier for more particles to overcome the activation energy barrier. There will be a greater proportion of fruitful collisions.

When a catalyst is used to increase the reaction rate it does so by providing an alternative reaction pathway with a lower activation energy. A greater proportion of the collisions will have enough energy to overcome the activation energy barrier and be successful.

**1b.** The collision theory model can be used to explain the rates of chemical reactions.

**i.** In terms of collision theory, explain what is needed for a chemical reaction to occur.

**ii.** Indicate what effects the following changes would have on a particular reaction mixture by placing ticks in appropriate boxes.

| **Change** | **Changes the activation energy** | **Increases the total number of collisions between reactant particles** | **Increases the fraction of collisions with energy above the activation energy** |
|---|---|---|---|
| Adding increased amounts of ethene and hydrogen chloride to the same volume container | | | |
| Increasing the temperature | | | |

**Solution**

**i.** For a chemical reaction to occur, particles must collide with enough energy (and correct collision geometry/orientation) to overcome the activation energy barrier.

**ii.**

| **Change** | **Changes the activation energy** | **Increases the total number of collisions between reactant particles** | **Increases the fraction of collisions with energy above the activation energy** |
|---|---|---|---|
| Adding increased amounts of ethene and hydrogen chloride to the same volume container | | ✓ | |
| Increasing the temperature | | ✓ | ✓ |

1 + 2 = 3 marks

**Mark allocation**

- 1 mark for part **i**.
- 1 mark for 'adding increased amounts of hydrogen' having the one correct tick.
- 1 mark for 'increasing the temperature' having the two correct ticks.

**Explanatory notes**

Increasing the amount of gaseous reactants in the same volume increases the pressure and will cause the reactant particles to collide with each other more often. This will increase the reaction rate because, although the fraction of collisions with energy above the activation energy is unchanged, the increased number of collisions results in more successful collisions. Hence, the activation energy remains unchanged.

Increasing the temperature increases the kinetic energy of the particles. The increase in particle speed increases the number of collisions and the increase in particle energy increases the fraction of collisions that are able to overcome the activation energy barrier. Hence, the activation energy remains unchanged.

**Tips**

- *It is useful to remember that the only method of increasing a reaction rate that changes the activation energy is using a suitable reaction-specific catalyst.*

**c.** As the reaction between ethene and hydrogen chloride proceeds, would you expect the reaction rate to increase or decrease? Give a reason for your answer.

**Solution**

Decrease because as the number of reactant particles decreases, the number of collisions will also decrease.

1 mark

**Explanatory notes**

For a reaction to occur the reactant particles must collide. As the reaction proceeds and the number of reactant particles decreases, they will collide less often. Hence, the reaction rate will decrease.

Total 2 + 3 + 1 = 6 marks

**Question 2**

The equilibrium constant for a chemical reaction occurring at a fixed temperature with gaseous reactants and products can be expressed as

$$K = \frac{[A]^2[B]}{[C][D]^3} = 7.46 \times 10^5 \text{ M}^{-1}$$

**a.** Write a balanced chemical equation for the reaction involving gases $A$, $B$, $C$ and $D$.

**Solution**

$C(g) + 3D(g) \rightleftharpoons 2A(g) + B(g)$

2 marks

**Mark allocation**

- 1 mark for correct species and coefficients.
- 1 mark for correct states.

**Explanatory notes**

The equilibrium law is an expression of $\frac{[\text{products}]}{[\text{reactants}]}$.

In addition, the index of each concentration in the expression of the equilibrium constant is the same as the coefficient for the substances in the chemical equation for the reaction.

$$K = \frac{[A]^2[B]}{[C][D]^3}$$

$$C(g) + 3D(g) \rightleftharpoons 2A(g) + B(g)$$

**Tips**

- *Always include states when writing any chemical or ionic equation.*

**b.** What does the value of this equilibrium constant indicate about the rate of this chemical reaction?

**Solution**

Nothing. It gives no information about rate.

1 mark

**Explanatory notes**

The equilibrium constant only gives an indication about the position of equilibrium for a chemical reaction. A very large value of $K$ (i.e. $> 10^4$) indicates that the equilibrium mixture is mostly products. A very small value of $K$ (i.e. $< 10^{-4}$) indicates that the equilibrium mixture is mostly reactants.

**c.** What does the value of this equilibrium constant indicate about the ratio of reactants to products when this reaction mixture is at equilibrium?

**Solution**

The ratio of reactants to products is very low; that is, there are mostly products in the equilibrium mixture.

1 mark

**Explanatory notes**

The $K$ value of $7.46 \times 10^5$ $M^{-1}$ is very large. A very large value of $K$ (i.e. $> 10^4$) indicates that the equilibrium mixture is mostly products.

**d.** Some additional gas $C$ is added to the equilibrium mixture. State the effect this will have on the value of $K$.

**Solution**

no effect

1 mark

**Explanatory notes**

The value of $K$ for a particular chemical reaction is dependent only on temperature. It is not affected by any other changes that can be made, such as addition or removal of reactants or products, concentration or pressure changes or use of a catalyst.

**e.** This reaction is exothermic. What effect will an increase in temperature have on the value of $K$? Explain your answer.

**Solution**

The value of $K$ will decrease. An increase in temperature will cause the equilibrium to shift left in the endothermic direction. The relative concentration of products will decrease, so the value of $K$ will decrease.

2 marks

**Mark allocation**

- 1 mark for correctly stating that *K* will decrease.
- 1 mark for correctly stating that concentration of products will decrease.

**f.** Under a different set of temperature conditions, a 2.0 L vessel containing 0.5 mol of gas *C* and 0.750 mol of gas *D* are allowed to reach equilibrium. The amount, in mol, of gas *B* present at equilibrium is found to be 0.15 mol. Calculate the value of *K* at this temperature.

**Solution**

Step 1: Calculate the amounts of all gases present. This can be done using the initial amounts of gases *C* and *D*, the equilibrium amount of gas *B* and the mole ratio from the coefficients in the balanced chemical equation.

| | *A* | *B* | *C* | *D* |
|---|---|---|---|---|
| Initial amount (mol) | 0.0 | 0.0 | 0.50 | 0.750 |
| Given amount of a gas at equilibrium (mol) | | 0.15 | | |
| Amount reacted of given gas (mol) | | +0.15 | | |
| Mole ratio of each gas to *B* according to balanced chemical equation | *A* : *B*<br>2 : 1 | | *C* : *B*<br>1 : 1 | *D* : *B*<br>3 : 1 |
| Amount reacted of each species using ratio and *n*(*B*) reacted (mol) | $\frac{2}{1} \times B = 0.30$ | | $\frac{1}{1} \times B = 0.15$ | $\frac{3}{1} \times B = 0.45$ |
| Equilibrium amount | 0.0 + 0.30<br>= 0.30 mol | 0.15 mol | 0.50 – 0.15<br>= 0.35 mol | 0.750 – 0.45<br>= 0.30 mol |

Step 2: Calculate the concentrations of each gas at the new equilibrium.

$[A] = \frac{n}{V}$

$= \frac{0.30}{2.0}$

$= 0.15$ M

$[B] = \frac{n}{V}$

$= \frac{0.15}{2.0}$

$= 0.075$ M

$[C] = \frac{n}{V}$

$= \frac{0.35}{2.0}$

$= 0.18$ M

$[D] = \frac{n}{V}$

$= \frac{0.30}{2.0}$

$= 0.15$ M

Step 3: Calculate the value of $K$ at the new temperature, using the given expression of the equilibrium constant.

$$K = \frac{[A]^2[B]}{[C][D]^3}$$

$$= \frac{(0.15)^2(0.075)}{(0.18)(0.15)^3}$$

$$= \frac{1.7 \times 10^{-3}}{6.1 \times 10^{-4}}$$

$$= 2.8 \text{ M}^{-1}$$

4 marks

**Mark allocation**

- 1 mark for correct calculation of amounts, in mol, of each species.
- 1 mark for calculating equilibrium concentrations of each species.
- 1 mark for substituting equilibrium concentrations into equilibrium expression.
- 1 mark for correct answer of 2.8 $\text{M}^{-1}$ (*NB*: equilibrium constant units are not required).

**Tips**

- *As a general rule, errors should be penalised only once; that is, consequential marks should be awarded if calculations are all correct but the chemical equation in part* ***a*** *is not balanced correctly.*

Total 2 + 1 + 1 + 1 + 2 + 4 = 11 marks

**Question 3**

Hydrogen gas and chlorine gas are introduced to a reaction vessel according to the equation

$H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g)$

Once equilibrium is reached, the reaction vessel is heated.

**a.** State the effect of increasing temperature on the concentration of HCl.

**Solution**

HCl concentration will decrease.

1 mark

**Explanatory notes**

According to Le Chatelier's principle, if an equilibrium system is subject to change it will adjust itself to partially oppose the effects of the change. In this case, heating the reaction vessel adds energy to the system. It responds by shifting in the endothermic direction, which in this case is backwards. The amount and concentration of HCl will decrease.

**Tips**

- *The introductions to questions can contain very important information. In this case, the statement that the 'temperature inside the vessel increases' as the reaction is progressing in a net forward direction indicates the forward direction is exothermic.*

**b.** State the effect of increasing temperature on the pressure within the vessel.

**Solution**

There is the same number of particles on the reactant and product sides of the chemical equation. Neither the forward nor reverse reactions produce more or less particles, so pressure will remain unchanged as the reaction proceeds.

1 mark

In solution, the following equilibrium is established between $Fe^{3+}$, which is a yellow colour; $SCN^-$, which is colourless; and $FeSCN^{2+}(aq)$, which is a dark red colour.

$Fe^{3+}(aq) + SCN^-(aq) \rightleftharpoons FeSCN^{3+}(aq)$
yellow dark red

**c.** Complete the table by predicting the effect of each of the following changes on the colour of the solution, once a new equilibrium is established, by placing a tick in the appropriate boxes. The temperature is kept constant.

| Change | No change | Colourless | More yellow | Darker red |
|---|---|---|---|---|
| Addition of $FeCl_3(aq)$ | | | | |
| Addition of KSCN(aq) | | | | |
| Dilution with water | | | | |

**Solution**

| Change | No change | Colourless | More yellow | Darker red |
|---|---|---|---|---|
| Addition of $FeCl_3(aq)$ | | | | ✓ |
| Addition of KSCN(aq) | | | | ✓ |
| Dilution with water | | | ✓ | |

3 marks

**Mark allocation**

- 1 mark for each correct tick.

**Explanatory notes**

- Le Chatelier's principle states that if an equilibrium system is subject to change it will adjust itself to partially oppose the effects of the change.
- The addition of $FeCl_3(aq)$ is addition of a reactant in the form of $Fe^{3+}(aq)$ ions, which causes a net forward reaction to remove some of the added ions. More $FeSCN^{3+}(aq)$ is produced so the solution becomes a darker red.
- The addition of KSCN(aq) is also addition of a reactant in the form of $SCN^{-}(aq)$ ions, which also causes a net forward reaction producing more $FeSCN^{3+}(aq)$ and a darker red colour.
- Diluting the solution reduces the concentration of all of the ions. The system responds by moving in the direction of more particles to increase their overall concentration. There are two particles on the reactant side to one on the product side; hence, a net backward reaction results and the amount of $Fe^{3+}$ increases. The solution becomes more yellow.

An equilibrium mixture contains the gases $Cl_2$, $F_2$ and $ClF_3$.

$Cl_2(g) + 3F_2(g) \rightleftharpoons 2ClF_3(g)$

**d.** A sample of argon is added to the mixture. State the effect of adding argon on the amount of $ClF_3$ present.

**Solution**

no effect

1 mark

**Explanatory notes**

Argon is a noble gas and is chemically inert. It will increase the overall pressure in the container but will not affect the partial pressures of any of the gases present, and so will not cause any shift in the position of equilibrium.

**e.** The gas mixture is transferred to a new container that is half the volume of the original container. State the effect of this change in volume on the amount of $ClF_3$ present. Explain your answer.

**Solution**

increase

Halving the volume increases the concentration of all of the gas particles. The system responds by moving in the direction of fewest particles to decrease their overall concentration. There are four particles on the reactant side of the equation ($Cl_2 + 3F_2$) to two on the product side ($2ClF_3$). Hence, a net forward reaction results and the amount of $ClF_3$ will increase.

2 marks

**Marking allocation**

- 1 mark for correctly stating that there is an increase.
- 1 mark for explaining that movement occurs in the direction of fewest particles.

Benzoic acid is a weak acid used as a food preservative. In solution, the following equilibrium is established.

$$C_6H_5COOH(aq) \rightleftharpoons C_6H_5COO^-(aq) + H^+(aq)$$

**f.** Calculate the initial concentration, in mol $L^{-1}$, of benzoic acid that is required to produce an aqueous solution of benzoic acid that has a pH of 2.54.

**Solution**

Step 1: Calculate $[H^+]$ from the pH.

$pH = -\log_{10}[H^+]$

So, $[H^+] = 10^{-pH}$

$= 10^{-2.54}$

$= 2.88 \times 10^{-3}$ M

Step 2: Write an expression for the acidity constant.

$$K_a = \frac{[C_6H_5COO^-][H^+]}{[C_6H_5COOH]}$$

Step 3: Use the expression of the acidity constant to calculate the concentration of benzoic acid.

We can assume $[C_6H_5COO^-] = [H^+]$, as they are produced in a 1 : 1 ratio.

From the data booklet, the $K_a$ of benzoic acid is $6.4 \times 10^{-5}$ M.

$$6.4 \times 10^{-5} = \frac{(2.88 \times 10^{-3})^2}{[C_6H_5COOH]}$$

$$[C_6H_5COOH] = \frac{(2.88 \times 10^{-3})^2}{6.4 \times 10^{-5}}$$

$$= \frac{8.31 \times 10^{-6}}{6.4 \times 10^{-5}}$$

$= 0.130$ M

2 marks

**Mark allocation**

- 1 mark for the assumption that $[C_6H_5COO^-] = [H^+]$.
- 1 mark for correctly calculating 0.130 M.

Total 1 + 1 + 3 + 1 + 2 + 2 = 10 marks

**Question 4**

In VCE Chemistry Unit 4, you were required to investigate the industrial production of one chemical selected from ammonia, ethene, sulfuric acid or nitric acid.

**a.** State the name **and** describe a major use of the chemical you studied.

**Solution**

One of the following:

Ammonia – manufacture of fertilisers

Ethene – manufacture of polyethene OR plastics

Sulfuric acid – manufacture of fertilisers

Nitric acid – manufacture of the fertiliser ammonium nitrate

1 mark

**Explanatory notes**

- Students need to discuss only their chosen chemical. The options for answers given in parts **a**–**d** may not contain all possible answers for each chemical, so student answers that differ should be considered carefully.

**Tips**

- *Simply writing 'fertilisers' for ammonia, nitric acid and sulfuric acid does not earn the mark. It must be specified that they are used in the manufacture of fertilisers.*

**b.** Write a balanced chemical equation for the final reaction that occurs during the production of this chemical; that is, where your chosen chemical will be a product.

**Solution**

One of the following:

Ammonia – $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

Ethene – Several cracking reactions will be appropriate; e.g. $C_2H_6(g) \rightleftharpoons C_2H_4(g) + H_2(g)$ or $C_3H_8(g) \rightleftharpoons C_2H_4(g) + CH_4(g)$

Sulfuric acid – $H_2S_2O_7(l) + H_2O(l) \rightarrow 2H_2SO_4(l)$

Nitric acid – $3NO_2(g) + H_2O(l) \rightarrow 2HNO_3(aq) + NO(g)$

1 mark

**c.** Briefly describe one way in which waste generated by the production of this chemical is managed.

**Solution**

One of the following:

Ammonia – carbon dioxide waste can be liquefied and sold OR aqueous solutions of ammonia can be used in the manufacture of urea.

Ethene – unreacted feedstock is recycled back into the converter OR ethyne is converted to ethene OR $CO_2$ and $H_2S$ are removed by treatment with NaOH(aq) OR hydrogen and methane are used as fuels for furnaces OR excess heat produced in the quenching process is used for the cracking process.

Sulfuric acid – $SO_2$ waste is recycled back to the converter for additional passes over the catalyst OR spent catalyst has vanadium removed before disposal in landfill OR excess energy produced is used in production of other chemicals or the cogeneration of electricity.

Nitric acid – NO and $NO_2$ wastes can be heated with a fuel to reduce them to $N_2$.

1 mark

**d.** Describe a safety risk associated with the production of this chemical and describe what steps are taken as a precaution against this risk.

**Solution**

One of the following:

Ammonia – Ammonia is a toxic gas so areas are well ventilated and workers wear protective clothing, including gloves, face masks, rubber boots and rubber aprons OR the hydrogen used as a reactant is explosive and can cause fires so the plant design and operation is planned carefully.

Ethene – Ethene can form explosive mixtures with air so fire-fighting prevention and fire-fighting strategies are used by industry OR high and low temperatures used can cause burns so protective clothing is used.

Sulfuric acid – Sulfuric acid is highly corrosive and can burn skin and eyes severely so protective clothing is worn OR inhaled sulfuric acid mist is harmful so areas are well ventilated and breathing apparatus must be used.

Nitric acid – Nitric acid is corrosive so stainless steel pipes and reaction vessels are sometimes used OR nitric acid can cause severe burns to the skin and eyes so workers wear protective clothing OR the fumes are harmful if inhaled so breathing apparatus must be used.

2 marks

**Mark allocation**

- 1 mark for explaining the safety risk.
- 1 mark for describing the precautions taken.

**Explanatory notes**

- The precaution given must match the safety risk (i.e. wearing safety gloves or safety glasses does not match 'the fumes are harmful if inhaled'), so the second mark would not be earned if these were given together.

Ethanol can be produced industrially by reacting ethene with steam according to the equation

$$CH_2CH_2(g) + H_2O(g) \rightleftharpoons CH_3CH_2OH(g) \qquad \Delta H = -45 \text{ kJ mol}^{-1}$$

The reaction conditions are:

$T$ = 300°C

$P$ = 60–70 atm

Catalyst = phosphoric(V) acid

**e.** A conflict is involved in choosing the best temperature for this reaction. Suggest a reason for this conflict and explain how it will be resolved.

**Solution**

A high temperature is required for a fast reaction rate and a low temperature is required for optimum equilibrium yield. To resolve the conflict, a moderate temperature is used with a catalyst.

2 marks

**Mark allocation**

- 1 mark for explaining that a fast rate requires high temperature and that a high yield requires low temperature.
- 1 mark for explaining that moderate temperature and a catalyst are to be used.

**Explanatory notes**

- The reaction rate of a chemical reaction will always be increased with high temperature, regardless of whether the reaction is exothermic or endothermic. However, this reaction is exothermic so produces heat in the forward direction. Yield will be maximised by lower temperatures.

**f.** If a higher pressure was used it would result in a greater equilibrium yield of ethanol. Suggest why a higher pressure is not used.

**Solution**

too expensive OR too dangerous

1 mark

**Explanatory notes**

- A higher equilibrium yield would be obtained with high pressure because there is one product particle and two reactant particles. However, a good enough yield is obtained without going to the expense of high pressure.

Total 1 + 1 + 1 + 2 + 2 + 1 = 8 marks

**Question 5**

An instant gas hot water service is listed on the online auction site eBay. The water is heated by the combustion of liquid petroleum gas (LPG).

**a.** Calculate the quantity of energy, in kJ, required to heat the water required for a 4-minute shower to 40°C above room temperature. An average shower uses 15.0 L of water per minute.

**Solution**

The specific heat capacity of water is 4.184 J $g^{-1}$ °$C^{-1}$.

A 4-minute shower will use 15.0 × 4 = 60.0 L of water.

The density of water is 1.00 g $mL^{-1}$, so 60.0 L of water has a mass of $60.0 \times 10^3$ g.

Energy (J) = specific heat capacity × mass (g) × temperature rise (°C)

$= 4.184 \times 60.0 \times 10^3 \times 40$

$= 1.0 \times 10^7$ J

$$\text{Energy (kJ)} = \frac{4.0 \times 10^6}{1000}$$

$= 1.0 \times 10^4$ kJ

2 marks

**Mark allocation**

- 1 mark for calculation of energy in joules = $4.184 \times 60.0 \times 10^3 \times 40$.
- 1 mark for conversion to kilojoules = $1.0 \times 10^4$ kJ.

**Tips**

- *The specific heat capacity and density of water are both listed in the data booklet.*

LPG is a mixture of hydrocarbon gases, primarily propane, $C_3H_8$, and butane, $C_4H_{10}$. A particular brand of LPG contains 60% by mass of propane and 40% by mass of butane.

**b**. Write thermochemical equations for the complete combustion of propane and butane. Include the value and sign of Δ*H* for each equation.

**i.** Propane

**Solution**

$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g) \quad \Delta H = -2217 \text{ kJ mol}^{-1}$

2 marks

**Mark allocation**

- 1 mark for correctly balanced chemical equation with states.
- 1 mark for $\Delta H = -2217$ kJ $mol^{-1}$.

**Explanatory notes**

- Combustion is a reaction with oxygen. The only products of the complete combustion of a hydrocarbon are carbon dioxide and water. The water is in a gaseous state due to the heat produced by the combustion. All combustion reactions are exothermic and so have a negative Δ*H*.

**Tips**

- *The molar enthalpy value for the combustion of some common fuels, including propane and butane, is in the data booklet.*

**ii.** Butane

**Solution**

$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$ $\Delta H = -5748$ kJ mol$^{-1}$

2 marks

**Mark allocation**

- 1 mark for correctly balanced chemical equation with states.
- 1 mark for $\Delta H = -5748$ kJ mol$^{-1}$.

**Explanatory notes**

- The molar enthalpy value for the combustion of one mole of butane is –2874 kJ mol$^{-1}$. In the balanced chemical equation, two moles of butane are being consumed, so the $\Delta H$ value for the equation is $-2874 \times 2 = -5748$ kJ mol$^{-1}$.

**c.** Calculate the mass of LPG required to heat the water for a 4-minute shower, as described in part **a**.

**Solution**

Step 1: Calculate the energy released by 100 g of LPG.

LPG is 60% propane, so $m(C_3H_8) = 60.0$ g

$$n(C_3H_8) = \frac{m}{M}$$

$$= \frac{60.0}{3 \times 12.0 + 8 \times 1.0}$$

$$= \frac{60.0}{44.0}$$

$$= 1.36 \text{ mol}$$

Energy released by propane = 1.36 × 2217

= 3023 kJ

LPG is 40% butane, so $m(C_4H_{10}) = 40.0$ g

$$n(C_4H_{10}) = \frac{m}{M}$$

$$= \frac{40.0}{4 \times 12.0 + 10 \times 1.0}$$

$$= \frac{40.0}{58.0}$$

$$= 0.689 \text{ mol}$$

Energy released by butane $= \frac{0.689}{2} \times 5748$

= 3964 kJ

*E* from 100 g LPG = *E* released by 60 g propane + *E* released by 40 g butane

= 3023 + 3964

$= 6.99 \times 10^3$ kJ

Step 2: Calculate the proportion of *E* required for the shower of the *E* released by 100 g LPG.

$$\text{Ratio} = \frac{4.0\times10^{3}}{6.99\times10^{3}}$$

$= 0.57$

Step 3: Calculate the mass of LPG required using the ratio determined in step 2.

$m(\text{LPG}) = 0.57 \times 100$

$= 57$ g

2 marks

**Mark allocation**

- 1 mark for correct use of percentages to calculate the energy released by a mass of LPG. It may not necessarily be 100 g.
- 1 mark for correctly calculating 57 g.

**Explanatory notes**

- 100 g was chosen as the mass of LPG from which to work out the energy released because it is an easy number to work with. Any mass could be chosen and, as long as the percentages and proportions were used correctly, the correct mass would be obtained.

Total 2 + 2 + 2 + 2 = 8 marks

**Question 6**

Give concise explanations for each of the following.

**a.** Lithium exists only in nature as part of a compound and not as an element.

**Solution**

Lithium is extremely reactive. It is the lowest reductant listed on the right side of the electrochemical series in the data booklet, and so is the most reactive metal listed.

1 mark

**Explanatory notes**

- Lithium is a highly reactive alkali metal. It corrodes rapidly in moist air to form a black tarnish.

**b.** A reaction will occur when nickel is placed in a 1.0 M solution of iron(III) nitrate, $Fe(NO_3)_3$(aq) but no reaction occurs when nickel is placed in a solution of iron(II) nitrate, $Fe(NO_3)_2$(aq).

**Solution**

Nickel is a reductant. It is lower in the series than the oxidant $Fe^{3+}$(aq) and so reduces these ions in solution. $Fe^{2+}$, as an oxidant, is below nickel in the series and so will not react. $Fe^{3+}$ is a stronger oxidant than $Fe^{2+}$.

1 mark

**Explanatory notes**

- Reference must be made to both $Fe^{3+}$(aq) reacting and $Fe^{2+}$(aq) not reacting to receive the mark.

**c.** Hydrogen and oxygen gas are the only products when a solution of potassium chloride undergoes electrolysis.

**Solution**

Water is a stronger oxidant than $K^+(aq)$ and a stronger reductant than $Cl^-(aq)$, so reacts preferentially to both.

1 mark

**Explanatory notes**

- The half-equations are $2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$

  and $2H_2O(l) \rightleftharpoons O_2(g) + 4H^+(aq) + 4e^-$

  giving an overall equation of $2H_2O(l) \rightleftharpoons 2H_2(g) + O_2(g)$

  Although $OH^-$ is also produced at the cathode and $H^+$ produced at the anode, the overall equation does not indicate this because species common to both reactant and product are cancelled out.

**Tips**

- *Successful electrolysis of potassium chloride requires it to be molten and in a non-aqueous environment.*

**d.** Less time is required to produce 10.0 g of gold by electrolysis using a current of 4.5 A than to produce 10.0 g of silver by electrolysis using the same current.

**Solution**

Gold has a molar mass of 197.0 g $mol^{-1}$ and silver has a molar mass of 107.9 g $mol^{-1}$. As 10.0 g of gold is a smaller amount, in mol, than 10.0 g of silver it will require less electricity to be produced.

1 mark

**Explanatory notes**

- The mass of metal produced during electrolysis will be determined by the current ($Q$), the time ($t$), the ratio of electrons to metal atoms in the reduction half-equation and the molar mass of the metal. In this question, the current and ratio are the same. The only difference is the molar masses. Calculations for calculating the time taken to deposit 10.0 g of each metal are shown below.

$$n(\text{Au}) = \frac{m}{M}$$

$$= \frac{10.0}{197.0}$$

$$= 0.0508 \text{ mol}$$

$$\text{Au}^+(aq) + e^- \rightarrow \text{Au}(s)$$

$$\text{So, } n(e^-) = \frac{1}{1} \times n(\text{Au})$$

$$= 0.0508 \text{ mol}$$

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

$= 0.0508 \times 96\ 500$

$= 4.90 \times 10^3$ C

$t = \frac{Q}{I}$

$= \frac{4.90 \times 10^3}{4.5}$

$= 1.09 \times 10^3$ s

$n(Ag) = \frac{m}{M}$

$= \frac{10.0}{107.9}$

$= 0.0927$ mol

$Ag^+(aq) + e^- \rightarrow Ag(s)$

So, $n(e^-) = \frac{1}{1} \times n(Ag)$

$= 0.0927$ mol

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

$= 0.0927 \times 96\ 500$

$= 8.94 \times 10^3$ C

$t = \frac{Q}{I}$

$= \frac{8.94 \times 10^3}{4.5}$

$= 1.99 \times 10^3$ s

Total 1 + 1 + 1 + 1 = 4 marks

**Question 7**

To determine the concentration of alcohol in a driver's breath, a fuel cell is used as one type of 'breathalyser'. Ethanol, $C_2H_5OH$, is oxidised to ethanoic acid, $CH_3COOH$, at one electrode and oxygen from air is converted to water at the other. An acidic electrolyte is used.

**a.** Write the equations for the half-reactions at the anode and the cathode.

**Solution**

Anode: $C_2H_5OH(aq) + H_2O(l) \rightarrow CH_3COOH(aq) + 4H^+(aq) + 4e^-$

Cathode: $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$

2 marks

**Explanatory notes**

- The question states that ethanol is oxidised in this cell to ethanoic acid and oxidation always occurs at the anode. The equations for half-reactions occurring in acidic media can be written by the following steps, which are summarised by K O H E S.

  K: Balance key elements (i.e. all those other than O and H).

  $C_2H_5OH \rightarrow CH_3COOH$

  O: Balance oxygen by adding water molecules to the appropriate side.

  $C_2H_5OH + H_2O \rightarrow CH_3COOH$

  H: Balance hydrogen by adding $H^+$ ions to the appropriate side.

  $C_2H_5OH + H_2O \rightarrow CH_3COOH + 4H^+$

  E: Add electrons to balance the charge.

  $C_2H_5OH + H_2O \rightarrow CH_3COOH + 4H^+ + 4e^-$

  S: Include states.

  $C_2H_5OH(aq) + H_2O(l) \rightarrow CH_3COOH + 4H^+(aq) + 4e^-$

  At the cathode, oxygen from air is reduced to water. The same steps can be followed to determine the half-equation:

  K: $O_2 \rightarrow H_2O$

  O: $O_2 \rightarrow 2H_2O$

  H: $O_2 + 4H^+ \rightarrow 2H_2O$

  E: $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$

  S: $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$

**b**. Write an overall equation for this reaction.

**Solution**

$C_2H_5OH(aq) + O_2(g) \rightarrow CH_3COOH(aq) + H_2O(l)$

1 mark

**Explanatory notes**

- An equation for an overall redox reaction is written by balancing the number of electrons lost in the oxidation and gained in the reduction. In this case, each half-reaction involves $4e^-$, so the oxidation and reduction half-reactions occur at a ratio of 1 : 1. Once added together, the electrons, hydrogen ions and one molecule of water cancel each other out.

$$C_2H_5OH(aq) + O_2(g) + 4H^+(aq) + \cancel{4e^-} \rightarrow CH_3COOH(aq) + 4H^+(aq) + \cancel{4e^-} + H_2O(l)$$

**c.** State the polarity of the anode in this cell.

**Solution**

negative

1 mark

**Explanatory notes**

- A fuel cell is a type of galvanic cell. The reaction is spontaneous and generates electricity. The anode in a galvanic cell is negative.

**d.** State the formula of the reductant in this cell.

**Solution**

$C_2H_5OH$

1 mark

**Explanatory notes**

- The reductant causes another species to be reduced and is itself oxidised.

**Tips**

- *Ensure a formula is written if that is what the question asks for.*

**e.** **i.** A car driver who has consumed alcohol blows into the breathalyser. If the breath entering the cell provides alcohol at a rate of $2.7 \times 10^{-5}$ g per second, calculate the amount, in mol, of ethanol blown into the fuel cell per second and, therefore, the amount of electrons available for reaction.

**Solution**

Step1: Calculate the amount, in mol, of ethanol entering the fuel cell per second.

$$n(C_2H_5OH) = \frac{m}{M}$$

$$= \frac{2.7\times10^{-5}}{2\times12.0+6\times1.0+16.0}$$

$$= \frac{2.7\times10^{-5}}{46.0}$$

$$= 5.9 \times 10^{-7} \text{ mol}$$

Step 2: Determine the amount, in mol, of electrons that would be produced.

According to the equation for the half-reaction

$n(e^-) : n(C_2H_5OH)$

4 : 1

So, $n(e^-) = \frac{4}{1} \times n(C_2H_5OH)$

$$= \frac{4}{1} \times 5.9 \times 10^{-7}$$

$$= 2.3 \times 10^{-6} \text{ mol}$$

2 marks

**ii.** Use your answer to part **e i.** to calculate the charge carried by this amount of electrons and, hence, calculate the current produced per second.

**Solution**

Step 1: Calculate the charge generated by this amount of electrons.

$$Q = n(e^-) \times F$$
$$= 2.3 \times 10^{-6} \times 96\,500$$
$$= 0.23 \text{ C}$$

Step 2: Calculate the current produced in one second.

$$I = \frac{Q}{t}$$
$$= \frac{0.23}{1}$$
$$= 0.23 \text{ A}$$

2 marks

Total 2 + 1 + 1 + 1 + 2 + 2 = 9 marks

**END OF SOLUTIONS**