INSIGHT

Trial Exam Paper

2011

CHEMISTRY

Written examination 2

Solutions book

This book presents:

- correct solutions with full working
- explanatory notes
- mark allocations
- tips and guidelines

SECTION A – Multiple-choice questions

Question 1

The energy density of methane gas in kJ g^{-1} is

A. 0.0180

B. 27.8

C. **55.6**

D. 889

Answer is C.

Explanatory notes

- C is correct because the molar enthalpy of methane is –889 kJ mol^{-1}. The molar mass of methane is 16.0 g mol^{-1}. Therefore the energy density is 889 kJ per 16.0 g, which is equivalent to $\frac{889}{16.0} = 55.6$ kJ g^{-1}.
- A is incorrect because the conversion from kJ mol^{-1} to kJ g^{-1} is $\frac{m}{M}$, not $\frac{M}{m}$.
- B is incorrect because the molar mass of methane is 16.0. It does not need to be multiplied by 2.
- D is incorrect because this value is the molar enthalpy in kJ mol^{-1}.

Tip

- *The molar enthalpy values of the combustion of many common fuels are listed in the Data Book.*

Question 2

Hydrogen gas is produced industrially by the following reaction

$$CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3H_2(g) \quad \Delta H = +207 \text{ kJ mol}^{-1}$$

The rate at which hydrogen gas is produced would be optimised by

A. low pressure, low temperature.

B. low pressure, high temperature.

C. high pressure, low temperature.

D. **high pressure, high temperature.**

Answer is D.

Explanatory notes

- D is correct because high pressure and high temperature will always cause an increase in the rate of the reaction, regardless of whether the reaction is exothermic or endothermic or whether there are more reactant or product particles.
- A is incorrect because low pressure and low temperature will result in a slower rate of reaction.
- B is incorrect because low pressure will result in a slower rate of reaction than high pressure will.
- C is incorrect because low temperature will result in a slower rate of reaction than high temperature will.

Tip

- *Read the question carefully to determine whether it is asking about rate (which is not affected by equilibrium considerations) or yield (which is affected by equilibrium considerations).*

Question 3

Propane and butane can both be burned to produce heat energy. The volume of propane gas, in L, at SLC that produces the same amount of energy as 1.50 L of butane gas at SLC is

A. 0.794

B. 1.16

C. 1.50

D. **1.94**

Answer is D.

Explanatory notes

- D is correct according to the following. Butane has a molar enthalpy value of –2874 kJ mol^{-1} and propane has a molar enthalpy value of –2217 kJ mol^{-1}. Both of these values are from the Data Book.

  Step 1: Determine the amount, in mol, of butane gas in 1.50 L at SLC.

  $n(C_4H_{10}) = \frac{V}{V_M}$

  $= \frac{1.5}{24.5}$

  $= 0.0612$ mol

Step 2: Determine the amount of energy released by this amount of butane.

$E = 0.061 \times 2874$

$= 176$ kJ

Step 3: Determine the amount, in mol, of propane required to produce 176 kJ of energy.

$n(C_3H_8) = \frac{176}{2217}$

$= 0.0794$ mol

Step 4: Calculate the volume of propane at SLC.

$V = n \times V_M$

$= 0.0794 \times 24.5$

$= 1.94$ L

- A is incorrect because 0.0794 is the amount, in mol, of propane required to produce the same amount of energy as 1.5 L of butane. It needs to be converted to volume at SLC.
- B is incorrect because butane and propane have different molar enthalpy values. 1.16 is the volume of butane required to produce 2217 kJ of energy. It does not answer the question.
- C is incorrect because the volume of propane will be different because the number of mole of propane will be different as propane and butane have different molar enthalpy values.

Question 4

Consider the following equilibrium system

$$4HCl(g) + O_2(g) \rightleftharpoons 2H_2O(g) + 2Cl_2(g) \qquad \Delta H = -116 \text{ kJ mol}^{-1}$$

Which of the following reaction conditions is most likely to result in the most improved rate **and** yield of reaction?

A. excess HCl, low temperature

B. excess O_2, high temperature

C. addition of a catalyst, low pressure

D. **addition of a catalyst, low temperature**

Answer is D.

Explanatory notes

- D is correct because the addition of a catalyst will increase the rate of the reaction while the low temperature will result in a higher equilibrium yield because this is an exothermic reaction. Le Chatelier's Principle states that when an equilibrium system is subject to change (in this case, lowering the temperature), it will adjust itself to partially oppose the effects of the change (in this case, moving to increase the temperature by shifting in the exothermic direction).
- A is incorrect because although the excess HCl and low temperature will improve the yield of the reaction, the low temperature will reduce the rate.
- B is incorrect because although the excess O_2 will improve yield and the high temperature will improve rate, the high temperature will also act to reduce yield as this is an exothermic reaction.
- C is incorrect because although addition of a catalyst will improve rate, the low pressure will reduce yield as it favours the side with the most particles, which in this case is the reactant side.

Question 5

A 2.00 g sample of benzoic acid, C_6H_5COOH (relative mass = 122), was used to calibrate a bomb calorimeter.

1 mol of benzoic acid releases 3227 kJ of heat energy. Complete combustion of the sample of benzoic acid resulted in the temperature of the water bath increasing from 23.10°C to 26.45°C. The calibration factor, in kJ °C^{-1}, for the calorimeter is

A. 2.00

B. **15.8**

C. 52.9

D. 1.93×10^3

Answer is B.

Explanatory notes

- B is correct according to the following steps:

  Step 1: Calculate the amount, in mol, of benzoic acid.

  $$n(C_6H_5COOH) = \frac{m}{M}$$
  $$= \frac{2.00}{122}$$
  $$= 0.0164 \text{ mol}$$

  Step 2: Calculate the amount of energy released by the benzoic acid.

  $$E = 0.0164 \times 3227$$
  $$= 52.9 \text{ kJ}$$

  Step 3: Calculate the calibration factor.

  $$\text{Calibration factor} = \frac{E}{\Delta T}$$
  $$= \frac{52.9}{3.35}$$
  $$= 15.8 \text{ kJ °C}^{-1}$$

- A is incorrect because the temperature change is 26.45 – 23.10 = 3.35°C, not 26.45°C.
- C is incorrect because this is the total amount of energy released by the benzoic acid that increased the temperature by 3.35°C. It needs to be changed to kJ $°C^{-1}$.
- D is incorrect because 1 mol of benzoic acid released 3227 kJ of heat, not 1.0 g. The 2.00 g needs to be converted to number of mole of benzoic acid to determine the amount of energy released.

Question 6

Which of the following best explains the increase in reaction rate observed when a catalyst is added to a reaction mixture?

A. Adding a catalyst increases the activation energy of the reaction, making it easier for the reaction to occur.

B. Adding a catalyst causes the reactant particles to collide with each other more often.

C. **Adding a catalyst provides a surface to which the reactant particles can form bonds and so react with each other more easily.**

D. Adding a catalyst lowers the activation energy of the reaction, meaning the particles collide with each other more frequently.

Answer is C.

Explanatory notes

- C is correct because a catalyst provides an alternative pathway with a reduced activation energy, making it easier for reactants to have successful collisions and react. The reactants must come in contact with the surface of the catalyst for it to be effective.
- A is incorrect because adding a catalyst decreases the activation energy required for a reaction to occur. Increasing the activation energy will always make it more difficult for a reaction to occur.
- B is incorrect because the addition of a catalyst does not affect the number of reactant collisions. Rather it makes the collisions that are already occurring more successful.
- D is incorrect because the addition of a catalyst does not affect the number of reactant collisions. Rather it makes the collisions that are already occurring more successful.

Question 7

The pH of a standard hydrogen half-cell is

A. **0**

B. 1

C. 7

D. 13

Answer is A.

Explanatory notes

- A is correct because in a standard hydrogen half-cell, the concentration of H^+ ions is 1 M.

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

  $= -\log_{10}(1)$

  $= 0$
- B is incorrect because 1 is the concentration of H^+ ions, in mol L^{-1}, not the pH.
- C is incorrect because the solution is not neutral. The H^+ concentration is 1 M, so the pH can be calculated using $pH = -\log_{10}[H^+]$.
- D is incorrect because the concentration of H^+ ions is 1 M and pH is calculated using $pH = -\log_{10}[H^+]$.

Question 8

Hydrogen and iodine gas react to produce hydrogen iodide according to the equation

$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$

2.00 mol of H_2 and 1.00 mol of I_2 are added to a 2.0 L reaction vessel and allowed to reach equilibrium. If 0.500 mol of HI(g) is present at equilibrium, the equilibrium concentrations of $H_2(g)$ and $I_2(g)$ are

	[H_2]	**[I_2]**
A.	**0.875**	**0.375**
B.	1.75	0.750
C.	0.250	0.250
D.	1.00	0.500

Answer is A.

Explanatory notes

- A is correct according to the following steps:

  Step 1: Determine the amount, in mol, of H_2 and I_2 that have reacted.

  $n(H_2) : n(HI)$ is 1 : 2

  so $n(H_2)$ reacted = $\frac{1}{2} \times 0.500 = 0.250$ mol

  $n(I_2) : n(HI)$ is 1 : 2

  so $n(H_2)$ reacted = $\frac{1}{2} \times 0.500 = 0.250$ mol

  Step 2: Determine the amounts, in mol, of H_2 and I_2 at equilibrium.

  $n(H_2) = n(H_2)$ initially added – $n(H_2)$ reacted

  = 2.00 – 0.250

  = 1.75 mol

  $n(I_2) = n(I_2)$ initially added – $n(I_2)$ reacted

  = 1.00 – 0.250

  = 0.750 mol

  Step 3: Calculate the equilibrium concentrations of H_2 and I_2.

  $[H_2] = \frac{n}{V}$

  $= \frac{1.75}{2.00}$

  = 0.875 M

  $[I_2] = \frac{n}{V}$

  $= \frac{0.750}{2.00}$

  = 0.375

- B is incorrect because these are the equilibrium amounts, in mol, and need to be divided by volume to determine the equilibrium concentrations.
- C is incorrect because these are the amounts of H_2 and I_2 that reacted, not the equilibrium amounts, which must be determined by taking the amount that reacted from their initial amounts.
- D is incorrect because these are the concentrations of the initial amounts, not the equilibrium amounts.

Question 9

A mixture of N_2O_4(g) and NO_2(g) is at equilibrium according to the equation

$$N_2O_4(g) \rightleftharpoons 2NO_2(g) \qquad \Delta H = +57 \text{ kJ mol}^{-1}$$

Which of the following changes to the equilibrium mixture will result in an increase to the value of *K* for this reaction?

I Doubling the volume

II Doubling the temperature

III Addition of N_2O_4(g)

A. I and II only

B. **II only**

C. I, II and III

D. none of the above

Answer is B.

Explanatory notes

- B is correct because the only change that can be made to an equilibrium mixture that will alter *K* is a change in temperature. This reaction is endothermic so an increase in temperature will result in a shift to the right. The amount of products will increase, so *K* will also increase.
- A is incorrect because changing the volume will never alter *K* of an equilibrium mixture.
- C is incorrect because changing the volume and adding or removing a reactant never alters *K* of an equilibrium mixture.
- D is incorrect because changing the temperature will alter *K* of an equilibrium mixture. This reaction is endothermic so an increase in temperature results in a shift to the right meaning products and *K* will increase.

Tip

- *It is useful to remember that an increase in temperature of an equilibrium reaction always produces a shift in the endothermic direction. A decrease in temperature always results in a shift in the exothermic direction.*

Question 10

Which of the following acids will show the greatest percentage ionisation in solution?

A. boric acid

B. hypobromous acid

C. hydrocyanic acid

D. **lactic acid**

Answer is D.

Explanatory notes

- D is correct because lactic acid has the highest K_a of all the options, meaning it is the strongest of the weak acids so will undergo the greatest percentage ionisation.
- A is incorrect because boric acid's K_a is lower than those of the other options, so it is the weakest acid. The weaker the acid, the lower the percentage ionisation will be.
- B is incorrect because hypobromous acid has a lower K_a than lactic acid, meaning it is a weaker acid.
- C is incorrect because hydrocyanic acid has a lower K_a than lactic acid, meaning it is a weaker acid.

Tip

- *The K_a values of a number of weak acids are listed in the Data Book. Remember that the stonger the acid, the larger the K_a and the greater the percentage ionisation that occurs in solution.*

Question 11

The concentration, in mol L^{-1}, of a solution of barium hydroxide, $Ba(OH)_2$, which has a pH of 11.4 is

A. 1.99×10^{-12}

B. 3.98×10^{-12}

C. $\mathbf{1.26 \times 10^{-3}}$

D. 2.51×10^{-3}

Answer is C.

Explanatory notes

- C is correct according to the following calculations:

  Step 1: Calculate the concentration of $[H_3O^+]$ from the pH.

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

  $= 10^{-11.4}$ M

  $= 3.98 \times 10^{-12}$ M

  Step 2: Calculate the concentration of $[OH^-]$.

  $[OH^-] = \frac{10^{-14}}{[H_3O^+]}$

  $= \frac{10^{-14}}{3.98 \times 10^{-12}}$

  $= 2.51 \times 10^{-3}$ M

  Step 3: Calculate the concentration of $Ba(OH)_2$. There are two hydroxide ions in the formula.

  $c(Ba(OH)_2) = \frac{[OH^-]}{2}$

  $= \frac{2.51 \times 10^{-3}}{2}$

  $= 1.26 \times 10^{-3}$ M
- A is incorrect because barium hydroxide is a base, so $[OH^-]$ needs to be determined from $[H_3O^+]$ to enable the concentration of the base to be calculated.
- B is incorrect because barium hydroxide is a base, so $[OH^-]$ needs to be determined from $[H_3O^+]$ to enable the concentration of the base to be calculated.
- D is incorrect because barium hydroxide has two OH^- ions in each formula, so $[OH^-]$ needs to be divided by 2 to determine the concentration of the base.

Question 12

Magnesium reacts with oxygen according to the thermochemical equation

$2Mg(s) + O_2(g) \rightarrow 2MgO(s) \qquad \Delta H = -1200 \text{ kJ mol}^{-1}$

The mass, in g, of magnesium that would need to be burned to release 4000 kJ of energy is

A. 7.29

B. 14.6

C. 81.0

D. **162**

Answer is D.

Explanatory notes

- D is correct according to the following steps:

  Step 1: Determine the amount, in mol, of magnesium that would need to be burned.

  2 mol releases 1200 kJ.

  x mol releases 4000 kJ.

  $\frac{4000}{1200} \times 2 = 6.667$ mol

  Step 2: Calculate the mass of magnesium.

  $m(\text{Mg}) = nM$

  $= 6.667 \times 24.3$

  $= 162$ g

- A is incorrect because the amount, in mol, of Mg required to be burned is $\frac{4000}{1200}$, not $\frac{1200}{4000}$. Also, the equation tells us that 2 mol of Mg releases 41200 kJ so $\frac{4000}{1200}$ must be multiplied by 2 to determine the amount, in mol of Mg.
- B is incorrect because the amount, in mol, of Mg that needs to be burned is $\frac{4000}{1200} \times 2$, not $\frac{1200}{4000} \times 2$.
- C is incorrect because the equation tells us that 2 mol of Mg releases 41200 kJ so $\frac{4000}{1200}$ must be multiplied by 2 to determine the amount, in mol of Mg.

Question 13

Consider the reaction system

$$2A(l) + 3B(g) \rightarrow 2C(g) + 4D(g) \qquad \Delta H = +345 \text{ kJ mol}^{-1}$$

It is correct to say that for this equation

I for each mol of A reacting, an enthalpy change of 345 kJ is produced.

II this is an endothermic reaction.

III the activation energy for this reaction is less than the ΔH value.

A. I and III only

B. **II only**

C. II and III only

D. I, II and III

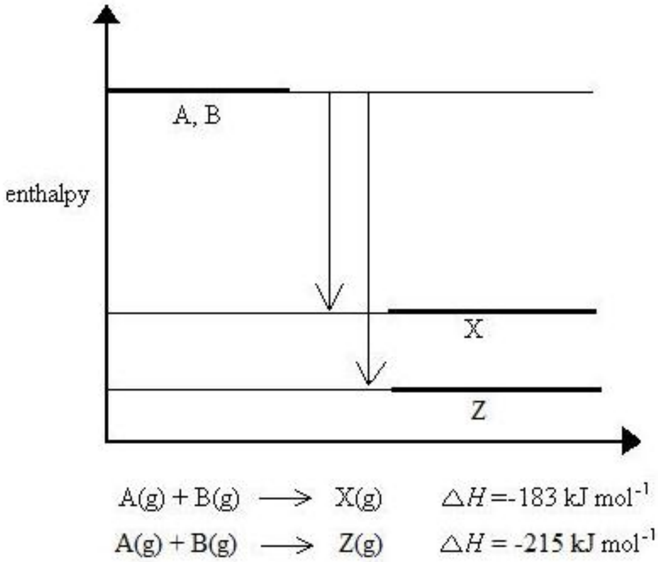
Answer is B.

Explanatory notes

- I is an incorrect statement because the thermochemical equation tells us that for every 2 mol of A that reacts, 345 kJ is absorbed, so for every 1 mol of A that reacts, 173 kJ is absorbed.
- II is a correct statement because the ΔH value is positive, indicating this is an endothermic reaction.
- III is an incorrect statement because in an endothermic reaction, the activation energy is always greater than the ΔH value.
- B is correct.
- A, C and D are incorrect because they do not identify the correct statements.

Question 14

The energy diagram below relates to two reactions.



The Δ*H* value for the reaction 2X(g) → 2Z(g) will be

A. +32 kJ mol^{-1}

B. –32 kJ mol^{-1}

C. +64 kJ mol^{-1}

D. **–64 kJ mol^{-1}**

Answer is D.

Explanatory notes

- D is correct. X has a higher enthalpy value than Z so the reaction X → Z is exothermic, meaning the Δ*H* value will be negative. The difference in enthalpy between X and Y is
  –215 – (–183) = –32 kJ mol^{-1}. In this reaction there are 2 mol of X producing 2 mol of Z so the energy change will be 2 × (–32) = –64 kJ mol^{-1}.
- A is incorrect because this is an exothermic reaction and there are 2 mol of X reacting to produce 2 mol of Y, not just 1 mol.
- B is incorrect because there are 2 mol of X reacting to produce 2 mol of Y, not just 1 mol.
- C is incorrect because this is an exothermic reaction as X has a higher enthalpy than Y, so energy will be released. The Δ*H* value will be negative.

Question 15

The strongest reductant from the following list Fe^{3+}, Sn^{2+}, H_2O, K^+, Br^-, Au, Pb^{2+} is

A. K^+

B. Fe^{3+}

C. $\mathbf{Sn^{2+}}$

D. Au

Answer is C.

Explanatory notes

- C is correct because Sn^{2+} is a reductant and is the lowest listed on the right side of the electrochemical series.
- B is incorrect because Fe^{3+} is an oxidant. It is only listed on the left side of the electrochemical series.
- A is incorrect because K^+ is an oxidant. It is only listed on the left side of the electrochemical series.
- D is incorrect because although Au is a reductant and listed on the right side of the electrochemical series, it is higher in the electrochemical series than Sn^{2+}, meaning it is a weaker reductant.

Tip

- *It can be helpful in the examination to write 'oxidants' above the left side of your electrochemical series in the Data Book and 'reductants' on the right so that you can remember which is which. Drawing an arrow going up the oxidants side and down the reductants side can also help you remember the orders of increasing strength.*

Question 16

An electric current is passed through a solution of 1.0 M KCl. The reaction that will occur at the cathode is

A. $K^+(aq) + e^- \rightarrow K(s)$

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

C. $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

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

Answer is B.

Explanatory notes

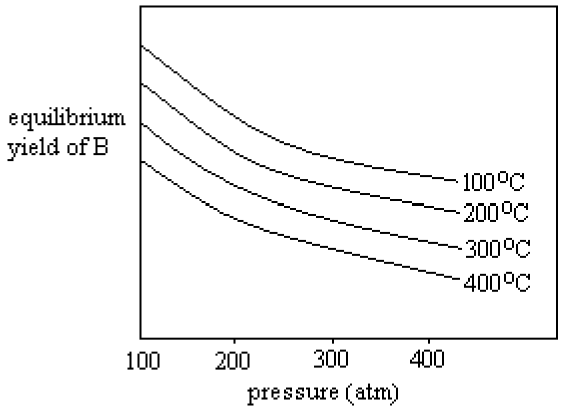
- B is correct because $H_2O(l)$ is a stronger oxidant than $K^+(aq)$, so will react preferentially and be reduced at the cathode.
- A is incorrect because this is a solution that contains water molecules and H_2O is a stronger oxidant than $K^+(aq)$, so will react preferentially and be reduced at the cathode.
- C is incorrect because this is an oxidation reaction and the cathode is always the site of reduction.
- D is incorrect because this is an oxidation reaction and the cathode is always the site of reduction.

Tips

- *AN OILRIG CAT can be a helpful way to remember that occurring at the ANode, Oxidation Is Loss while Reduction Is Gain at the CAThode.*

Question 17

Gaseous reactant A exists in equilibrium with gaseous product B. No other chemical species are present. The following graph shows the effect of increasing pressure and temperature on the equilibrium yield of product B.



Which of the following conclusions about the reaction could be correct?

I There are more gaseous reactant particles than gaseous product particles in the equation.

II The reaction is exothermic.

A. I and II only

B. I only

C. **II only**

D. neither I nor II

Answer is C.

Explanatory notes

- I is an incorrect statement. It can be seen from the graph that as pressure increases, the yield of B decreases. Le Chatelier's Principle tells us that as pressure increases on a gaseous equilibrium, the system will move to the side with the least number of gaseous particles to partially oppose the increase in pressure. If yield is decreasing, the least number of particles must be the reactant side.
- II is a correct statement. It can be seen from the graph that as temperature increases, the yield of B decreases. Le Chatelier's Principle tells us that as temperature increases, the system will move in the endothermic direction to absorb some of the extra heat and partially oppose the increase in temperature. If yield is decreasing, the endothermic direction must be backwards, making this an exothermic direction.
- C is correct.
- A, B and D are incorrect.

Question 18

Which of the following best describes the source of energy in a nuclear power station?

A. **Nuclei are split into smaller particles and energy is released in the process.**

B. Nuclei are burnt in oxygen, releasing energy in the process.

C. Nuclei fuse with each other to make bigger particles, releasing energy in the process.

D. Nuclei release neutrons one at a time and release energy in the process.

Answer is A.

Explanatory notes

- A is correct because nuclear fission involves the splitting of nuclei to create smaller nuclei. This releases an enormous amount of energy in the process according to the equation $E = mc^2$.
- B is incorrect because nuclei do not burn in oxygen. They split after being knocked by smaller particles in a chain reaction process.
- C is incorrect because nuclear fission involves nuclei splitting apart, not fusing together.
- D is incorrect because when a nucleus splits, it may release several neutrons at once. These go on to split other nuclei. It does not release one neutron at a time.

Question 19

An electroplating cell is devised to place a gold coating on a brass ring. A voltage of 2.80 V and a current of 1.35 A are passed through a solution of 1.00 M Au^+ ions, for 280 seconds.

The mass of gold produced, in grams, is closest to

A. 0.0129

B. 0.386

C. **0.772**

D. 2.16

Answer is C.

Explanatory notes

- C is correct according to the following steps:

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

  $Q = It$

  $= 1.35 \times 280$

  $= 378$ C

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

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

  $= \frac{378}{96\ 500}$

  $= 0.003\ 92$ mol

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

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

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

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

  1 : 1

  So, $n(Au) = n(e^-)$

  $= 0.003\ 92$ mol

  Step 5: Calculate the mass, in g, of gold deposited.

  $m(Au) = nM$

  $= 0.003\ 92 \times 197.0 = 0.772$ g
- A is incorrect because t has to be in seconds for use in $Q = It$. It is not converted to minutes.
- B is incorrect because the gold ion is Au^+, not Au^{2+}.
- D is incorrect because the calculation of charge is $Q = It$, not VIt.

Tips

- *Be very clear on your formulae. Q = It and E = VIt are similar formulae and can be confused, especially when voltage is provided but the question does not require it to be used.*
- *Although you should be very familiar and confident with your ion valencies by now, the electrochemical series in your Data Book can be a useful reference. It includes the gold reduction half-equation, which tells us that the valency of a gold ion is +1.*

Question 20

Water self-ionises according to the equation

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

At 25°C the pH of pure water is 7.00. At 35°C the pH of pure water is

A. 7.00 because the concentration of H_3O^+ ions has not changed.

B. 7.00 because although the reaction has shifted to the right, it is still neutral.

C. less than 7.00 because the reaction has shifted to the right and the solution is now slightly acidic.

D. **less than 7.00 because the reaction has shifted to the right; however, the solution is still neutral**

Answer is D.

Explanatory notes

- D is correct because the increase in temperature causes a shift in the endothermic direction which in this case is to the right. $[H_3O^+]$ increases so pH decreases. However, the solution is still neutral because $[H_3O^+] = [OH^-]$
- A is incorrect because the H_3O^+ concentration and pH have changed. This is an endothermic reaction so will respond to a change in temperature according to Le Chatelier's Principle.
- B is incorrect because the reaction has shifted to the right and it is still neutral, however, any change in $[H_3O^+]$ must cause a change in pH as $pH = -\log_{10}[H_3O^+]$.
- C is incorrect because although the system has shifted to the right and the pH has changed, it is still a neutral solution because $[H_3O^+] = [OH^-]$.

Tip

- *A neutral solution is any solution in which the concentrations of H_3O^+ and OH^- are equal. This will always be the case in pure water, regardless of the pH because they are produced at the same rate by the self-ionisation of water. An acidic solution is one in which $[H_3O^+] > [OH^-]$.*

END OF SECTION A

SECTION B – Short-answer questions

Question 1

Nitrosyl bromide ($NOBr_2$) decomposes according to the reaction

$$2NOBr_2(g) \rightleftharpoons 2NO(g) + Br_2(g) \qquad \Delta H < 0$$

a. **i.** Write an equilibrium expression for this reaction.

1 mark

Solution

$K = \frac{[NO]^2[Br_2]}{[NOBr_2]^2}$

Explanatory notes

The equilibrium constant, K, is calculated as $K = \frac{[\text{products}]}{[\text{reactants}]}$. Coefficients for each species in the equation become indices in the equilibrium expression.

ii. 3.00 mol of $NOBr_2$ was added to a 4.00 L reaction vessel at 300 K. Once equilibrium was reached, 0.805 mol of Br_2 was present in the mixture. Calculate the equilibrium constant for the reaction at this temperature.

3 marks

Solution

	$2NOBr_2(g)$	**2NO(g)**	**$Br_2(g)$**
Initial (mol)	3.00	0	0
Change (mol)	–1.61	+1.61	+0.805
Equilibrium (mol)	1.39 mol	1.61 mol	0.805 mol
$c = \frac{n}{V}$	0.348 M	0.403 M	0.201 M

$K = \frac{[NO]^2[Br_2]}{[NOBr_2]^2}$

$= \frac{[0.403]^2[0.201]}{[0.348]^2}$

$= \frac{0.0326}{0.121}$

$= 0.269$ M

Mark allocation

- 1 mark for calculating the correct equilibrium amounts, in mol, or concentrations, in M, of each species.
- 1 mark for determining the correct equilibrium concentrations of each species.
- 1 mark for correctly calculating *K*.

Explanatory notes

You can earn consequential marks if, for example, your equilibrium amounts are incorrect but you use these values to find concentrations and *K* correctly.

iii. In a second reaction mixture, also in a 4.0 L reaction vessel and at 300 K, the concentrations of the gases were found to be

$[NOBr_2] = 0.507$ M $[Br_2] = 0.0405$ M

Calculate the concentration, in M, of NO in the reaction mixture.

2 marks

Solution

$K = \frac{[NO]^2[Br_2]}{[NOBr_2]^2}$

$0.269 = \frac{[NO]^2(0.0405)}{(0.507)^2}$

$[NO]^2 = 1.71$

$[NO] = 1.31$ M

Mark allocation

- 1 mark for using the equilibrium constant calculated in **part ii**, showing some working.
- 1 mark for calculating the correct concentration of [NO], using the equilibrium expression written in **part i**.

b. Would the equilibrium yield for this reaction be favoured by a low or high temperature? Use Le Chatelier's Principle to explain your answer.

3 marks

Solution

A low temperature.

The reaction is exothermic, so the yield will be favoured by low temperatures.

Le Chatelier's Principle tells us that when an equilibrium system is subject to change (decreased temperature), it will adjust itself to partially oppose the effects of the change. In this case that means shifting in the exothermic direction to increase the temperature. More products are produced.

Mark allocation

- 1 mark for low temperature.
- 1 mark for indicating that this is an exothermic reaction.
- 1 mark for a valid explanation that makes reference to Le Chatelier's Principle.

Explanatory notes

It is not enough to simply say that in exothermic reactions, yield is favoured by low temperatures. You are asked specifically to make reference to Le Chatelier's Principle.

Total 9 marks

Tip

- *Although you were not given an actual ΔH value, you are told it is less than 0, so it must be negative, which indicates that this is an exothermic reaction.*

Question 2

A student obtains four 0.10 M solutions of nitric acid, ethanoic acid, benzoic acid and sulfuric acid.

a. Calculate the pH of the

i. nitric acid solution

1 mark

Solution

$[H_3O^+]$ = $[HNO_3]$

= 0.10 M

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

= $-\log_{10}(0.100)$

= 1.0

Explanatory notes

Nitric acid is a strong acid, so completely ionises in water and the assumption that $[H_3O^+] = [HNO_3]$ is valid.

ii. ethanoic acid solution

3 marks

Solution

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

$1.7 \times 10^{-5} = \frac{[H_3O^+]^2}{[0.100]}$

$[H_3O^+]^2 = 1.7 \times 10^{-6}$

$[H_3O^+] = 0.0013$ M

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

= $-\log_{10}(0.0013)$

= 2.9

Mark allocation

- 1 mark for correct use of the expression for the acidity constant.
- 1 mark for $[H_3O^+] = 0.0013$ M.
- 1 mark for correctly using $[H_3O^+]$ to calculate pH.

Explanatory notes

Nitric acid is a weak acid so does not completely ionise in water. Hence, the assumption that $[H_3O^+] = [CH_3COOH]$ is not valid. Instead the K_a value must be used, as well as the assumptions that $[CH_3COOH]$ does not change since so little ionises and that $[H_2O]$ stays relatively constant. We also know that $[H_3O^+] = [CH_3COO^-]$ as they are produced at the same rate. The K_a value is obtained from the Data Book.

b. Will the molar concentration of H_3O^+ ions be the same, higher or lower in the benzoic acid solution compared to the ethanoic acid solution? Explain your answer.

2 marks

Solution

Higher.

Benzoic acid has a higher acidity constant than ethanoic acid, meaning it is a slightly stronger acid and will ionise slightly more in water.

Mark allocation

- 1 mark for higher.
- 1 mark for a valid explanation that refers to benzoic acid being a slightly stronger acid.

c. List the four 0.100 M solutions in order from lowest pH to highest pH.

2 marks

Solution

Sulfuric acid, nitric acid, benzoic acid, ethanoic acid

Mark allocation

- 0 marks if more than one is out of order.
- 1 mark if one is out of order.
- 2 marks for correct order.

Explanatory notes

Sulfuric acid has the lowest pH because it is diprotic. H_2SO_4 is a strong acid, so completely ionises. Although HSO_4^- is a weaker acid, it will ionise to some extent, producing more H_3O^+ ions. Nitric acid is a strong monoprotic acid, so will completely ionise. Ethanoic and benzoic acids are weak acids. Benzoic acid is slightly stronger as it has a higher K_a, so it ionises slightly more and will have a lower pH than ethanoic acid.

d. 100 mL of distilled water is added, separately, to the nitric acid solution and to the ethanoic acid solution.
Indicate the effect on the pH and percentage ionisation of each acid by ticking the appropriate box in the table below.

4 marks

Solution

	Decrease	**Increase**	**Not change**
Effect on pH of nitric acid		√	
Effect on percentage ionisation of nitric acid			√
Effect on pH of ethanoic acid		√	
Effect on percentage ionisation of ethanoic acid		√	

Mark allocation

- 1 mark for each correct tick.

SECTION B – continued
TURN OVER

Explanatory notes

A dilution will make the nitric acid solution less acidic; hence, the pH will increase. But dilution will not alter the percentage ionisation of HNO_3 molecules as it is a strong acid that was fully ionised anyway.

A dilution will also make the ethanoic acid solution less acidic; hence pH will increase. However, according to Le Chatelier's Principle, adding a reactant (H_2O) will cause a shift to the right of the reaction and the percentage ionisation will increase. There will be more H_3O^+ molecules in the solution; however, their concentration is still decreased due to the increase in volume.

Total 12 marks

Question 3

A bomb calorimeter can be used to determine the enthalpy changes in a chemical reaction. In one reaction, pentane is combusted in the bomb of a calorimeter that is surrounded by a 300 mL water bath.

a. Write a balanced thermochemical equation for the complete combustion of pentane.

2 marks

Solution

$C_5H_{12}(g) + 8O_2(g) \rightarrow 5CO_2(g) + 6H_2O(g)$ $\Delta H = -3509 \text{ kJ mol}^{-1}$

Mark allocation

- 1 mark for correct species and a balanced equation.
- 1 mark for the ΔH value

Explanatory notes

Complete combustion of organic compounds involves a reaction with oxygen gas and produces carbon dioxide and water.

Tip

- *To balance a combustion reaction, start by balancing the C atoms, then the H atoms and finally the O atoms.*

b. **i.** Calculate the mass of pentane that would need to undergo combustion in a calorimeter to produce a temperature change in the water of 1.65°C. Be sure to give your answer with the correct number of significant figures.

4 marks

Solution

E required to heat water = specific heat capacity of water × m × ΔT

$= 4.18 \times 300 \times 1.65$

$= 2.07 \times 10^3$ J

$= 2.07$ kJ

$n(C_5H_{12}) = \frac{E}{\text{Energy released by 1 mole}}$

$= \frac{2.07}{3509}$

$= 5.90 \times 10^{-3}$ mol

$m(C_5H_{12}) = 5.90 \times 10^{-3} \times 72.0$

$= 0.425$ g

Mark allocation

- 1 mark for correctly calculating the energy required to heat 300 mL of water.
- 1 mark for correctly determining the number of mole of pentane required to produce that energy.
- 1 mark for correctly calculating the mass of pentane from its number of mole.
- 1 mark for giving the answer to 3 significant figures.

ii. State one assumption you made when calculating your answer to **part i**.

1 mark

Solution

No heat was lost to the environment, i.e. all the energy from the combustion reaction was used to heat the water in the water bath.

Total 7 marks

Question 4

During this semester, you looked at one of the following reactions as part of your study of the industrial production of an important chemical. Place a tick in the box next to the reaction that is part of the production of the chemical you studied.

	$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$
	$C_3H_8(g) \rightleftharpoons C_2H_4(g) + CH_4(g)$
	$2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$
	$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

a. **i.** What conditions of temperature and pressure favour optimum yield for your chosen reaction?

1 mark

Solution

Ammonia – low temperature, high pressure

Ethene – high temperature, low pressure

Nitric acid – low temperature, high pressure

Sulfuric acid – low temperature, high pressure

Explanatory notes

The reactions that are part of the production process for ammonia, nitric acid and sulfuric acid are exothermic and so are favoured by low temperatures. The reaction for ethene production is endothermic and so is favoured by high temperatures. The reactions for ammonia, nitric acid and sulfuric acid have less reactant particles than product particles and so are favoured by high pressure, whereas the reaction for ethene production has more product particles and so is favoured by high pressure.

ii. Are these conditions used on an industrial scale? Give a reason for your answer.

2 marks

Solution

Ammonia – No, a moderate temperature is used as well as a catalyst to keep the rate reasonable.

Ethene – Yes, the pressure is kept below 1 atm and the temperature is 750°C to 900°C.

Nitric acid – Yes, the low temperature still gives a fast reaction rate.

Sulfuric acid – No, a moderate temperature is used as well as a catalyst so the reaction is not too slow. Also, acceptable yields are obtained at 1 atm so high pressure is not used as it would be expensive.

Mark allocation

- 1 mark for ‘No’ or ‘Yes’ as appropriate.
- 1 mark for reasonable explanation.

b. **i.** Identify one specific risk to human health associated with the production of your selected chemical.

1 mark

Solution

Ammonia – Irritation to skin, eyes and respiratory system; explosive so can cause physical injury and burns

Ethene – Dizziness, damage to nervous system; explosive so can cause physical injury and burns

Nitric acid – Irritation to skin and eyes.

Sulfuric acid –Blindness, respiratory diseases.

ii. Give one precaution that is taken in response to the risk you identified in **part i.**

1 mark

Solution

Ammonia – Safety clothing, safety glasses, breathing apparatus, gloves.

Ethene – Safety clothing, breathing apparatus, signage.

Nitric acid – Safety glasses, breathing apparatus, stainless steel equipment.

Sulfuric acid – Safety glasses, air ventilation systems, face mask.

c. **i.** Name one useful commercial product formed from the chemical you chose.

1 mark

Solution

Ammonia – fertiliser

Ethene – polyethene

Nitric acid – fertiliser

Sulfuric acid – fertiliser

ii. Write a chemical equation to show the formation of the commercial product named in **part i**.

1 mark

Solution

Ammonia: $NH_3 + HNO_3 \rightarrow NH_4NO_3$

Ethene: $nC_2H_4 \rightarrow -(CH_2CH_2)_n-$

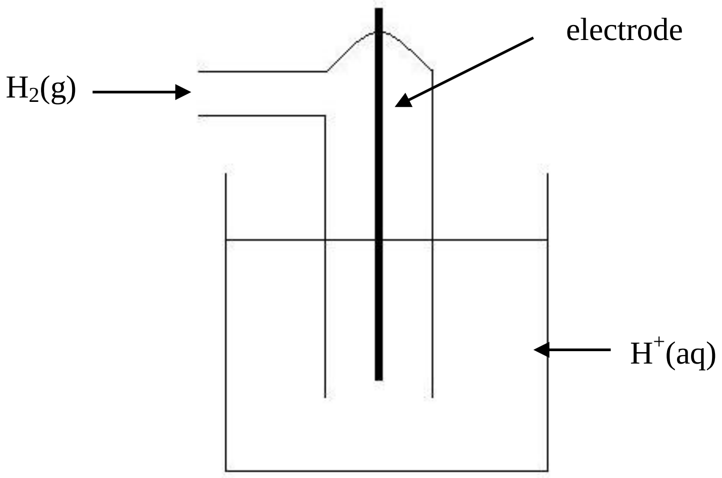
Nitric acid: $HNO_3 + NH_3 \rightarrow NH_4NO_3$

Sulfuric acid: $H_2SO_4 + 2NH_3 \rightarrow (NH_4)_2SO_4$

Total 7 marks

Question 5

The $H^+(aq)/H_2(g)$ half-cell is the standard used to obtain the *E*° values listed in the electrochemical series. A diagram of the cell is shown below.



a. State the temperature and pressure required for this cell to be used as a half-cell.

1 mark

Solution

25°C and 1 atm or 298 K and 101.3 kPa

Explanatory notes

Both conditions must be given to receive the mark. A galvanic cell is set up from the following two half-cells:

Half-cell 1: the $H^+(aq)/H_2(g)$ standard half cell.

Half-cell 2: a platinum electrode in a solution containing $Fe^{3+}(aq)$ and $Fe^{2+}(aq)$ ions.

b. **i.** After several minutes, will the pH in the $H^+(aq)/H_2(g)$ half-cell have increased, decreased or be unchanged?

1 mark

Solution

Decreased

ii. Explain your answer to **part i**.

2 marks

Solution

The Fe^{3+}(aq)/Fe^{2+}(aq) half-equation is higher in the electrochemical series than the H^+(aq)/H_2(g) half-equation. H_2(g) is a stronger reductant than Fe^{2+}(aq) so will be oxidised at the anode. The concentration of H^+ ions will increase, so pH will decrease.

Mark allocation

- 1 mark for reference to H_2(g) being a stronger reductant than Fe^{2+}(aq) so undergoing oxidation.
- 1 mark for reference to $[H^+]$ increasing.

c. i. Write an overall ionic equation for a galvanic cell in which one half-cell is the Fe^{3+}(aq)/Fe^{2+}(aq) cell and which has a potential difference of 1.05 V.

1 mark

Solution

$2Fe^{3+}$(aq) + Co(s) → $2Fe^{2+}$(aq) + Co^{2+}(aq)

Mark allocation

- 1 mark for a correct equation including all correct state symbols.

Explanatory notes

The Fe^{3+}(aq)/Fe^{2+}(aq) half-equation has an *E*° of +0.77. The potential difference of a galvanic cell is equal to the *E*° of the higher half-equation minus the *E*° of the lower half-equation. Fe^{3+}(aq)/Fe^{2+}(aq) must be combined with a half-cell that has an *E*° of +1.82 or –0.28. According to the electrochemical series in the Data Book, the Co^{2+}(aq)/Co(s) has an *E*° of –0.28.

ii. Identify the strongest reductant in this reaction.

1 mark

Solution

Co

Explanatory notes

The strongest reductant is the species that undergoes oxidation. Co(s) is losing electrons to form Co^{2+}(aq) so is undergoing oxidation and must be the strongest reductant.

Total 7 marks

Question 6

Primary, secondary and fuel cells are all types of galvanic cells used for the production of electricity.

a. What is the main difference between a primary cell and a fuel cell?

1 mark

Solution

A primary cell has a set amount of reactants in it. A fuel cell has a continuous supply of reactants when the cell is running.

b. The nickel–cadmium cell is a secondary cell used to power small appliances. When the cell is being used the overall reaction is

$$NiO_2(s) + 2H_2O(l) + Cd(s) \rightarrow Ni(OH)_2(s) + Cd(OH)_2(s)$$

i. What feature of this cell enables it to be recharged?

1 mark

Solution

The products of the discharge reaction are in the solid state, so they remain in contact with the electrodes in a reversible form.

ii. When the cell is recharging, which species is reacting at the negative electrode?

1 mark

Solution

$Cd(OH)_2$

Explanatory notes

When the cell is recharging, it is acting as an electrolytic cell so the negative electrode is the cathode. Reduction will be occurring. The above reaction will be going in the backwards direction. The oxidation number of Cd decreases from +2 in $Cd(OH)_2$ to 0 in Cd, indicating it is undergoing reduction.

c. The overall equation for a fuel cell used in 'breathalyser' instruments is

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

i. Write the equation for the half-reaction that occurs at the anode.

1 mark

Solution

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

Explanatory notes

Oxidation always occurs at the anode. The oxidation number of C increases from –2 in C_2H_5OH to 0 in CH_3COOH, so this is oxidation.

Tips

- *A redox half-equation in acidic media can be balanced using the KOHES steps:*

  K – Balance the key elements.

  O – Balance the oxygen atoms by adding H_2O molecules.

  H – Balance the hydrogen atoms by adding H^+ ions.

  E – Balance the charge by adding electrons (e^-).

  S – Add state symbols to all species except electrons.

ii. A motorist blows into a fuel cell for 5.0 seconds and provides 1.4×10^{-4} g ethanol at the rate of 2.8×10^{-5} g per second. Calculate the maximum current, in A, that the cell would produce.

4 marks

Solution

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

$= \frac{1.4\times10^{-4}}{46.0}$

$= 3.0 \times 10^{-6}$ mol

$n(e^-) = 4 \times n(C_2H_5OH)$

$= 4 \times 3.0 \times 10^{-6}$

$= 1.2 \times 10^{-5}$ mol

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

$= 1.2 \times 10^{-5} \times 96\,500$

$= 1.2$ C $I = \frac{Q}{t}$ $= \frac{1.2}{5.0}$ $= 0.23$ A

Mark allocation

- 1 mark for correct calculation of $n(C_2H_5OH)$.
- 1 mark for correct calculation of $n(e^-)$.
- 1 mark for correct calculation of Q.
- 1 mark for correct calculation of I.

iii. What is the energy, in kJ, that would be produced in **part ii** if the voltage is measured at 4.0 V?

1 mark

Solution

$E = VIt$ so $E = VQ$

$E = 4.0 \times 1.2$

$= 4.8$ J

$= 4.8 \times 10^{-3}$ kJ

Total 9 marks

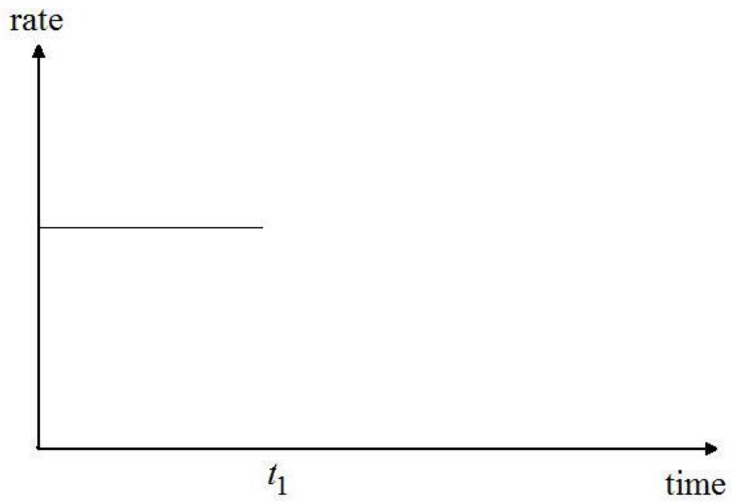
Question 7

The same amounts of reactants X and Z are placed in two sealed containers and allowed to react according to the equation

$$X(g) + Z(g) \rightleftharpoons 2Y(g) \quad \Delta H = -345 \text{ kJ mol}^{-1}$$

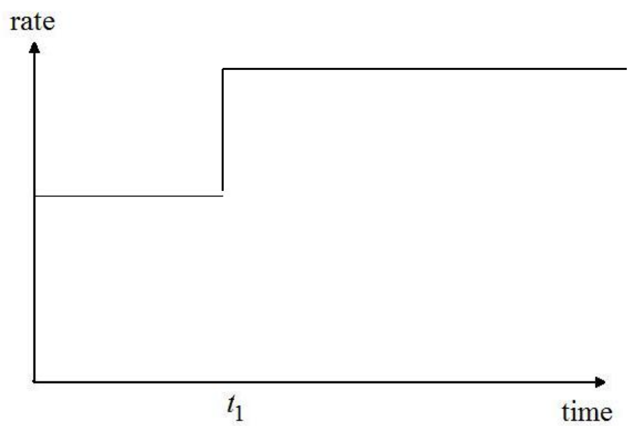
a. After the reaction reaches equilibrium in container 1, a small amount of catalyst is added to the container at time t_1.

i. On the rate–time graph below, continue the line to show what will happen to the rate of the forward reaction after time t_1 in container 1.



1 mark

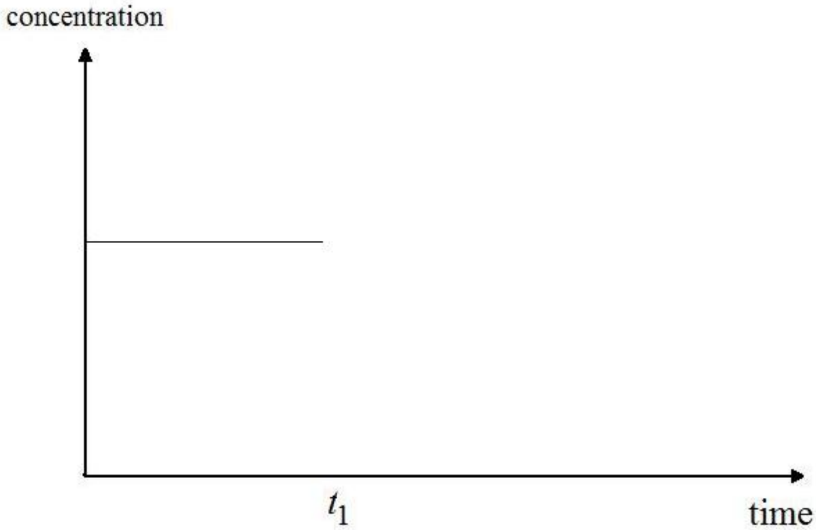
Solution



Explanatory notes

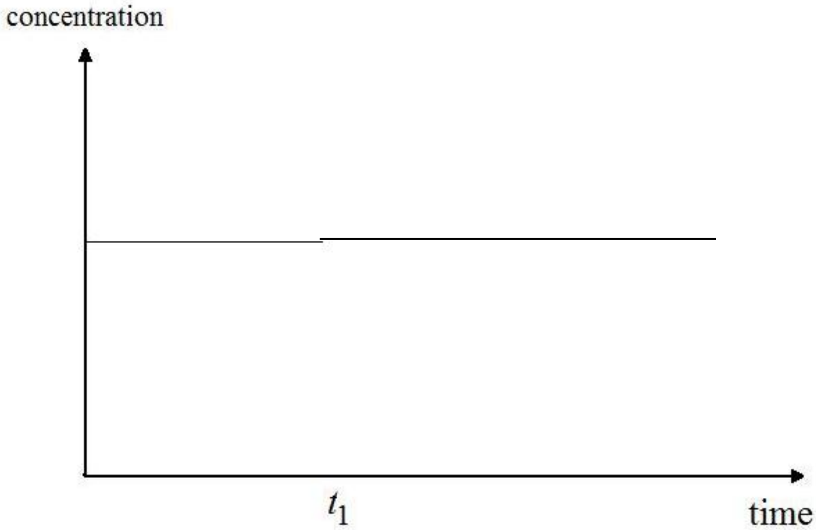
Adding a catalyst will increase the rate of the forward reaction.

ii. On the concentration–time graph below, continue the line to show what will happen to the concentration of reactant X after time t_1 in container 1.



1 mark

Solution

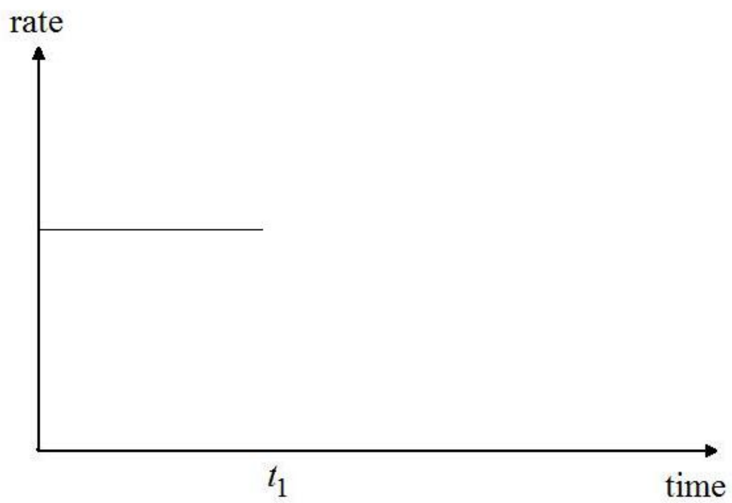


Explanatory notes

The reaction mixture was already at equilibrium before the catalyst was added. Adding a catalyst will not change the position of equilibrium as it increases the forward and backward reaction to the same extent, so the concentration of X does not change.

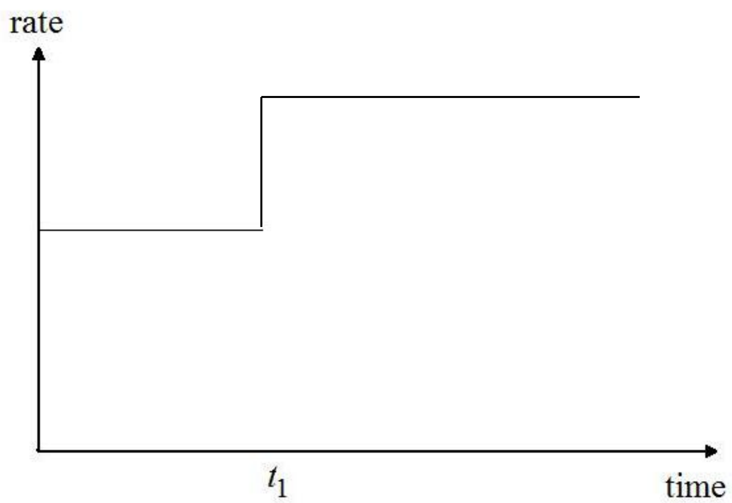
b. After the reaction reaches equilibrium in container 2, the volume of the reaction vessel is halved at time t_1.

i. On the rate–time graph below, continue the line to show what will happen to the rate of the forward reaction after time t_1 in container 2.



1 mark

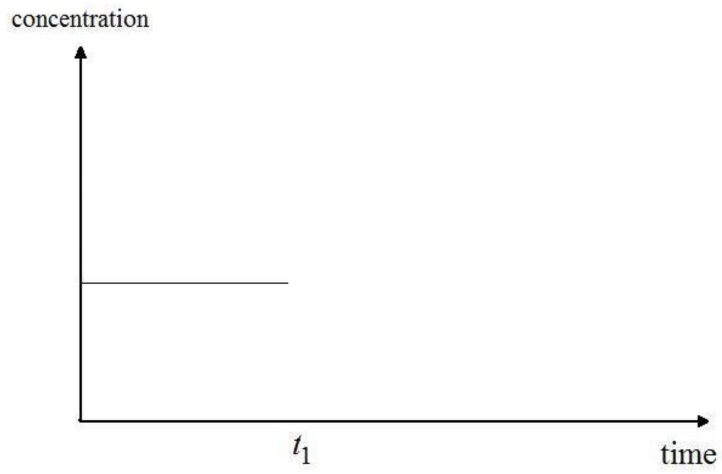
Solution



Explanatory notes

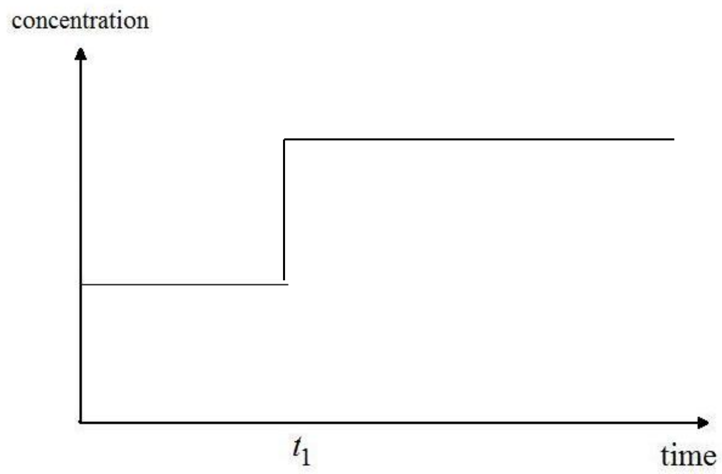
Halving the volume of the vessel doubles the pressure. The particles will collide much more frequently increasing the rate of the forward reaction.

ii. On the concentration–time graph below, continue the line to show what will happen to the concentration of reactant X at time t_1 in container 2.



2 marks

Solution



Mark allocation

- 1 mark for the concentration doubling.
- 1 mark for the concentration then remaining constant.

iii. Explain the graph you drew in **part i**.

1 mark

Solution

The position of equilibrium did not shift because there are two reactant particles and two product particles. Therefore, the concentration of X remains constant.

Total 6 marks

Question 8

A panel of experts is making recommendations to a government group regarding future choices for energy sources. The experts have been comparing the energy sources brown coal, natural gas, nuclear fission and ethanol (a biochemical fuel).

a. Which energy source would you recommend on the basis of sustainability? Give a reason for your answer.

1 mark

Solution

Ethanol. It is the only completely sustainable resource as it is derived from plants, which can be regrown.

b. Which energy source would you recommend on the basis of energy density; that is, which fuel has the highest kJ g^{-1}?

1 mark

Solution

Nuclear fission. (A very small mass releases an enormous amount of energy.)

c. Which **two** energy sources would you recommend as producing the least amount of greenhouse gas emissions per gram? Explain your answer.

2 marks

Solution

Ethanol as it is ‘carbon neutral’. Although it releases carbon dioxide on combustion, it was derived from glucose that was produced photosynthetically from carbon dioxide in the air. Nuclear fission as it does not release any carbon dioxide when it produces energy.

Mark allocation

- 1 mark for stating both ethanol and nuclear fission.
- 1 mark for both explanations.

Total 4 marks

END OF SOLUTIONS BOOK