PEARSON EDEXCEL A-LEVEL CHEMISTRY · 9CH0

CHEMISTRY



EXAM QUESTION PACK

TOPIC 11: EQUILIBRIUM II

SCORE	PERCENTAGE
/211	%

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Q1.

This question is about an experiment to determine the equilibrium constant, K_c , for an esterification reaction producing propyl ethanoate. The equation for the reaction is

$$\underset{\text{ethanoic acid}}{CH_3COOH(l)} + \underset{\text{propan-1-ol}}{CH_3CH_2CH_2OH(l)} \rightleftharpoons \underset{\text{propyl ethanoate}}{CH_3COOCH_2CH_2CH_3(l)} + H_2O(l)$$

In an experiment to determine the equilibrium constant, K_c , the following steps were carried out.

- 6.0 cm^3 of ethanoic acid (0.105 mol), 6.0 cm^3 of propan-1-ol (0.080 mol) and 2.0 cm^3 of dilute hydrochloric acid were mixed together in a sealed boiling tube.
  In this pre-equilibrium mixture, there is 0.111 mol of water
- The mixture was left for one week, at room temperature and pressure, to reach equilibrium
- The equilibrium mixture and washings were transferred to a volumetric flask and the solution made up to exactly 250.0 cm^3 using distilled water
- 25.0 cm^3 samples of the **diluted** equilibrium mixture were titrated with a solution of sodium hydroxide, concentration 0.200 mol dm^{-3}, using phenolphthalein as the indicator
- The mean titre was 23.60 cm^3 of 0.200 mol dm^{-3} sodium hydroxide solution.

(a) State the role of the hydrochloric acid in the esterification reaction.

(1)

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(b) (i) Calculate the total amount, in moles, of acid present in the **volumetric flask** in the equilibrium mixture.

(2)

(ii) The 2.0 cm^3 of dilute hydrochloric acid contained 0.00400 mol of $H^+(aq)$ ions.

Use this and your answer to part (b)(i) to calculate the amount, in moles, of ethanoic acid present in the equilibrium mixture.

(1)

(c) (i) The initial mixture in the boiling tube contained 0.105 mol of ethanoic acid.

Use your answer to (b)(ii) to calculate the amount, in moles, of ethanoic acid that reacted to form the ester in the equilibrium mixture.

(1)

(ii) Use information given in the method, and your answer to (c)(i), to calculate the amounts, in moles, of propan-1-ol, propyl ethanoate and water that are present in the equilibrium mixture.

(3)

Moles of propan-1-ol at equilibrium

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Moles of propyl ethanoate at equilibrium

.....

Moles of water at equilibrium

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(d) (i) Write the expression for the equilibrium constant, K_c , for this reaction.

$$CH_3COOH(l) + CH_3CH_2CH_2OH(l) \rightleftharpoons CH_3COOCH_2CH_2CH_3(l) + H_2O(l)$$

(1)

(ii) Explain why it is possible, in this case, to calculate K_c using equilibrium amounts in moles, rather than equilibrium concentrations.

(2)

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(iii) Calculate the value of K_c.

Give your answer to an appropriate number of significant figures.

(2)

(e) The pink colour of the phenolphthalein fades after the end-point of the titration has been reached.

Give a possible explanation for this observation.

(2)

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(f) Explain what you could do to confirm that one week is sufficient time for the mixture to reach equilibrium.

(2)

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(g) A student repeated the experiment, but left the mixture in a water bath at 40 °C until equilibrium was reached.

$$CH_3COOH(l) + CH_3CH_2CH_2OH(l) \rightleftharpoons CH_3COOCH_2CH_2CH_3(l) + H_2O(l) \qquad \Delta_r H^{\ominus} = +21.4\,kJ\,mol^{-1}$$

Deduce the effect, if any, on this student's value for K_c compared with that obtained in part (d)(iii).

(2)

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(Total for question = 19 marks)

Q2.

An experiment to determine the equilibrium constant for the reaction between sodium fluoride solution and magnesium hydroxide is carried out.

100 cm^3 of sodium fluoride solution, of concentration 0.250 mol dm^{-3}, is added to a conical flask. Excess solid magnesium hydroxide is added and the mixture is left for 24 hours to reach equilibrium.

The mixture is filtered and 25.0 cm^3 of the solution is removed and titrated with hydrochloric acid.

11.70 cm^3 of 0.100 mol dm^{-3} HCl(aq) is required to neutralise the NaOH formed.

(a) Draw a dot-and-cross diagram of the ions in $Mg(OH)_2$.
Show outer electrons only.

(2)

(b) (i) Use the data to calculate the equilibrium constant, K_c for the reaction between NaF and $Mg(OH)_2$.
Give your answer to an appropriate number of significant figures and give the units, if any.

$$2NaF(aq) + Mg(OH)_2(s) \rightleftharpoons 2NaOH(aq) + MgF_2(s)$$

(5)

(b) (ii) The experiment was repeated on different days.
Justify whether this would improve the reliability of the data collected.

(1)

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(Total for question = 8 marks)

Q3.

The gas phase reaction between hydrogen and iodine is reversible.

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

(a) (i) Write the expression for the equilibrium constant, K_c, for this reaction.

(1)

(ii) If the starting concentration of both hydrogen and iodine was **a** mol dm^{-3} and it was found that 2**y** mol dm^{-3} of hydrogen iodide had formed once equilibrium had been established, write the K_c expression in terms of **a** and **y**.

(2)

(b) The expression for the equilibrium constant in (a)(ii) can be rearranged as shown.

$$\mathbf{y} = \frac{\mathbf{a}\sqrt{K_c}}{2+\sqrt{K_c}}$$

In an experiment, air was removed from a 1 dm^3 flask and amounts of hydrogen and iodine gases were mixed together such that their initial concentrations were both **a** mol dm^{-3}. This mixture was allowed to reach equilibrium at 760 K. The equilibrium concentration of iodine was then measured.

The experiment was repeated for various initial concentrations, **a** mol dm^{-3}, and the results recorded in the table.

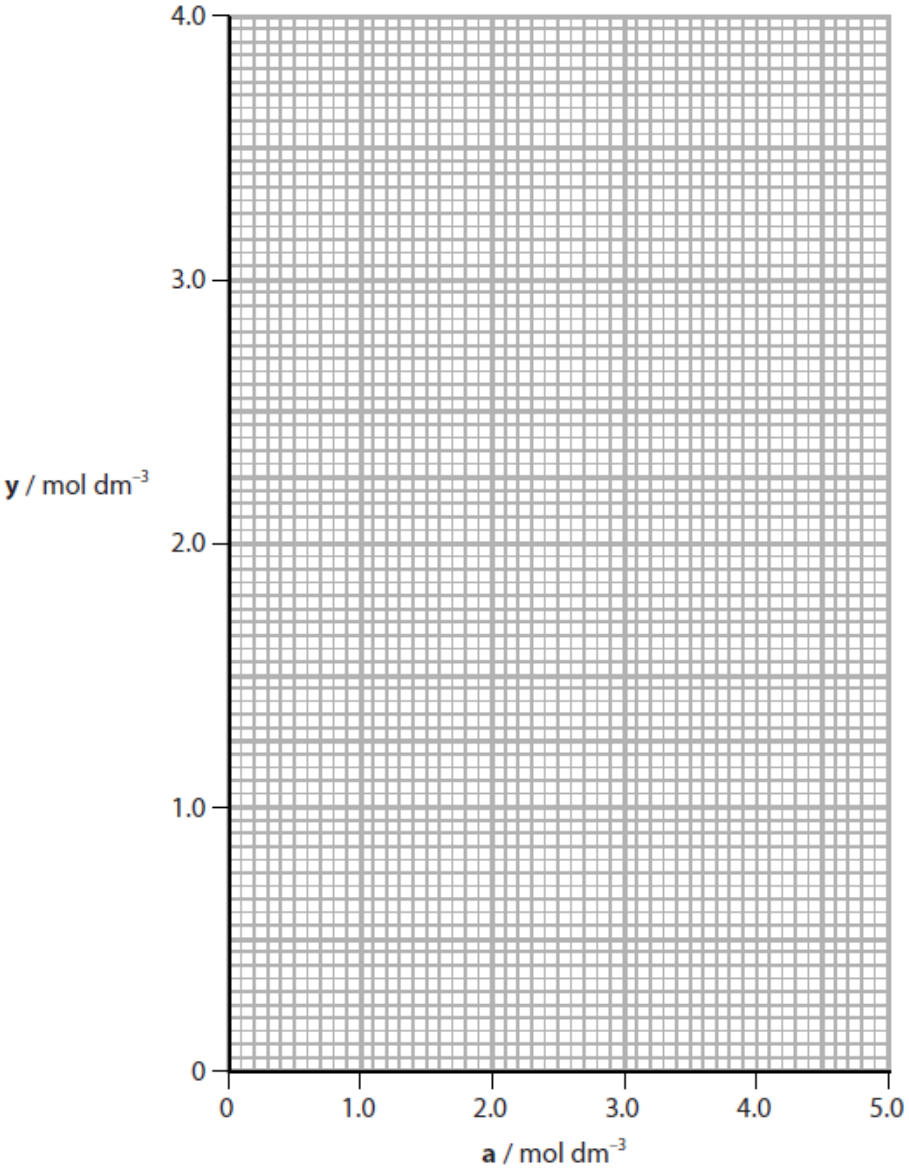
(i) Complete the table to give the two remaining values of **y** mol dm^{-3}, to **two** decimal places.

(1)

a / mol dm^{-3}	$[I_2]_{eq}$ / mol dm^{-3}	y / mol dm^{-3}
0.20	0.02	0.18
0.80	0.25	0.55
1.50	0.37	
2.10	0.57	1.53
2.80	0.65	2.15
3.80	0.87	
4.90	1.15	3.75

(ii) Plot a graph to show how **y** mol dm^{-3} varies with the initial concentrations of hydrogen and iodine, **a** mol dm^{-3}.

(2)



(iii) Determine the gradient of your graph.

Show your working on the graph.

(2)

(iv) Use your answer to (b)(iii) and the expression $\mathbf{y} = \frac{\mathbf{a}\sqrt{K_c}}{2+\sqrt{K_c}}$ to calculate the value of K_c.

(2)

(c) Identify a safety issue associated with this experiment.

(1)

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(d) One of the experiments in part (b) was repeated using the same molar quantities of hydrogen and iodine but in a 500 cm^3 flask instead of the 1 dm^3 flask.

Deduce the effect, if any, that this would have on the rate of reaction and on the value of K_c calculated.

(2)

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(e) The equation for the reaction between hydrogen and iodine is

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g) \qquad \Delta H = -9.6 \text{ kJ mol}^{-1}$$

(i) Explain the effect, if any, on the value of K_c when the temperature is increased.

(2)

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(ii) On your graph in (b)(ii), draw and label the line you would expect if the experiment was carried out at 1000 K instead of 760 K.

(1)

(Total for question = 16 marks)

Q4.

Ammonia is manufactured by the Haber Process.

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

$$K_p = \frac{p(NH_3)^2}{p(N_2)p(H_2)^3}$$

A mixture of 1.0 mol of nitrogen and 3.0 mol of hydrogen is left to reach equilibrium at 700 K.

Calculate the total pressure, in atmospheres, needed to produce a yield of 0.30 mol of ammonia at 700 K.
Give your answer to an appropriate number of significant figures.

You must show your working.

[$K_p = 7.76 \times 10^{-5}$ atm^{-2} at 700 K]

(5)

(Total for question = 5 marks)

Q5.

Ammonia is produced industrially by reacting nitrogen and hydrogen.

Nitrogen and hydrogen were reacted together at 673 K and 200 atm pressure in a closed vessel.

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

The reaction mixture was allowed to reach equilibrium. The number of moles of each gas at equilibrium were found to be:

$N_2(g)$ 2.88 $H_2(g)$ 8.64 $NH_3(g)$ 6.48

Calculate the value of K_p at 673 K, giving appropriate units.

(5)

Q6.

The following reversible reaction is used in industry to make methanol, CH_3OH:

$$CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g) \quad \Delta H = -91 \text{ kJ mol}^{-1}$$

(a) Which change would affect both the value of the equilibrium constant, K_c, and the proportion of methanol present in an equilibrium mixture of the three gases?

(1)

☒ **A** adding a catalyst

☒ **B** changing the temperature

☒ **C** increasing the concentration of carbon monoxide

☒ **D** increasing the pressure

(b) The expression for the equilibrium constant, K_c, for this reaction is

$$K_c = \frac{[CH_3OH(g)]}{[CO(g)][H_2(g)]^2}$$

0.200 mol of CO(g) and 0.400 mol of $H_2(g)$ are mixed in a sealed container of volume 1.2 dm^3 at a temperature of 500 K and a pressure of 100 atmospheres and allowed to reach equilibrium.

The equilibrium mixture is found to contain 0.086 mol of $CH_3OH(g)$.

(i) Calculate K_c for this reaction. Give your answer to an appropriate number of significant figures and state the units.

(5)

(ii) The equilibrium mixture of CO(g), $H_2(g)$ and $CH_3OH(g)$ is heated in the same sealed container to a temperature higher than 500 K. Since the gas volume remains the same, the increased temperature results in an increase in pressure.

Explain why it is difficult to predict the effect on the yield of CH_3OH.

(3)

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Q7.

Hydrogen is produced on a large scale by several different processes.

Carbon monoxide reacts with steam.

$$CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$$

At 1100 K, $K_c = 1.00$

In an experiment, 1 mol of carbon monoxide was mixed with 1 mol of steam, 2 mol of carbon dioxide and 2 mol of hydrogen.

Deduce, with reasons, the direction in which the reaction will shift to reach equilibrium.

(3)

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(Total for question = 3 marks)

Q8.

What are the units of the equilibrium constant (K_c) for the hypothetical reaction below?

$$2A(aq) + B(aq) \rightleftharpoons 4C(aq) + D(aq)$$

☒ **A** $mol^2\ dm^{-9}$

☒ **B** $mol^{-2}\ dm^9$

☒ **C** $mol^2\ dm^{-6}$

☒ **D** $mol^{-2}\ dm^6$

(Total for question = 1 mark)

Q9.

This question is about an experiment to determine the equilibrium constant, K_c, for the reaction between ethanoic acid and ethanol to form ethyl ethanoate and water.

Two sealed test tubes were prepared.

The first test tube contained 0.0400 mol ethanoic acid, 0.0400 mol of ethanol and 0.20 cm^3 of concentrated hydrochloric acid.

The second test tube contained 0.0400 mol ethyl ethanoate, 0.0400 mol of water and 0.20 cm^3 of concentrated hydrochloric acid.

After standing at 25°C for two weeks, to ensure equilibrium is reached, the contents of each test tube were separately titrated with 0.200 mol dm^{-3} sodium hydroxide solution.

0.20 cm^3 of concentrated hydrochloric acid was also titrated with the same sodium hydroxide solution.

(a) (i) Using data from the Data Booklet, calculate the volume, in cm^3, of 0.0400 mol of ethanoic acid.

(2)

(ii) What would be the best piece of apparatus to measure out the volumes of the liquids for the sealed test tubes?

(1)

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(iii) Suggest a reason why the test tubes were sealed.

(1)

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(iv) Suggest a suitable indicator for the titration of the equilibrium mixture in either test tube, with the expected colour change. Justify your suggestion.

(3)

Indicator

.....

Colour change from ..... to .....

Justification

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(b) In this experiment, the following titres were obtained.

Titration	Volume of 0.200 mol dm^{-3} sodium hydroxide solution/cm^3
Contents of first test tube	77.10
Contents of second test tube	77.05
0.20 cm^3 concentrated hydrochloric acid	11.70

(i) Write the equation for the reaction between ethanoic acid and ethanol to form ethyl ethanoate and water, using structural formulae. State symbols are not required.

(1)

(ii) Calculate the number of moles of ethanoic acid present at equilibrium in the first test tube.

(2)

(iii) Deduce the number of moles of ethanol present at equilibrium in the first test tube.

(1)

(iv) Calculate the number of moles of ethyl ethanoate formed at equilibrium in the first test tube.

(1)

(v) Write an expression for the equilibrium constant, K_c, for the reaction.
Assuming the number of moles of water and ethyl ethanoate present at equilibrium are the same, calculate the equilibrium constant, K_c.

(2)

(vi) Explain why the equilibrium constant for this reaction has no units.

(1)

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(vii) Why, in fact, is the number of moles of water present in the equilibrium mixture greater than the number of moles of ethyl ethanoate?

(1)

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(c) (i) What is the type of reaction that took place in each test tube?

(2)

First test tube

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Second test tube

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*(ii) Comment on the value of the titre for the equilibrium mixture in the second test tube compared to the first test tube.

What characteristic feature of equilibrium reactions is demonstrated by the values of these

titres?

(2)

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(iii) State the role of the concentrated hydrochloric acid in the equilibrium reaction.

(1)

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(Total for question = 21 marks)

Q10.

Methane hydrate is found on continental shelves deep in oceans. It forms methane in an endothermic equilibrium reaction, which may be represented as

$$CH_4.6H_2O(s) \rightleftharpoons CH_4(g) + 6H_2O(l)$$

(a) Which of the following changes would **increase** the equilibrium yield of methane?

(1)

☒ **A** Increasing the temperature and decreasing the pressure.

☒ **B** Decreasing both the temperature and the pressure.

☒ **C** Increasing both the temperature and the pressure.

☒ **D** Decreasing the temperature and increasing the pressure.

(b) Which of the following would **decrease** the value of the equilibrium constant, K_p, for the above equilibrium?

(1)

☒ **A** Decreasing the pressure

☒ **B** Increasing the pressure

☒ **C** Decreasing the temperature

☒ **D** Increasing the temperature

(Total for question = 2 marks)

Q11.

Methane hydrate is found on continental shelves deep in oceans. It forms methane in an endothermic equilibrium reaction, which may be represented as

$$CH_4.6H_2O(s) \rightleftharpoons CH_4(g) + 6H_2O(l)$$

(a) Which of the following changes would **decrease** the equilibrium yield of methane?

(1)

☒ **A** Decreasing the temperature and decreasing the pressure.

☒ **B** Increasing the temperature and decreasing the pressure.

☒ **C** Decreasing the temperature and increasing the pressure.

☒ **D** Increasing the temperature and increasing the pressure.

(b) Which of the following would **increase** the value of the equilibrium constant, K_p, for the above equilibrium?

(1)

☒ **A** Decreasing the pressure

☒ **B** Increasing the pressure

☒ **C** Decreasing the temperature

☒ **D** Increasing the temperature

(Total for question = 2 marks)

Q12. (a) For the equilibrium reaction between hydrogen and iodine

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

increasing the pressure of the system

(1)

☒ **A** has no effect on the rate or the position of equilibrium.

☒ **B** increases the rate but does not affect the position of equilibrium.

☒ **C** increases the rate and shifts the equilibrium to the right.

☒ **D** increases the rate and shifts the equilibrium to the left.

(b) The equation for the equilibrium reaction between hydrogen and iodine may also be written as

$$\tfrac{1}{2}H_2(g) + \tfrac{1}{2}I_2(g) \rightleftharpoons HI(g)$$

This change to the equation, compared to that in part (a),

(1)

☒ **A** has no effect on the value of the equilibrium constant.

☒ **B** halves the value of the equilibrium constant.

☒ **C** doubles the value of the equilibrium constant.

☒ **D** square roots the value of the equilibrium constant.

(Total for Question = 2 marks)

Q13. Ammonium chloride decomposes on heating:

$$NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$$

The equilibrium constant, K_p, for this reaction equals

☒ **A** $P_{NH_3} \times P_{HCl}$

☒ **B** $\dfrac{1}{P_{NH_3} \times P_{HCl}}$

☒ **C** $\dfrac{P_{NH_3} \times P_{HCl}}{P_{NH_4Cl}}$

☒ **D** $\dfrac{P_{NH_4Cl}}{P_{NH_3} \times P_{HCl}}$

(Total for Question = 1 mark)

Q14. This question is about the equilibrium reaction below.

$$Fe^{2+}(aq) + Ag^{+}(aq) \quad \Delta S^{\ominus}_{total} = R \ln K \quad Fe^{3+}(aq) + Ag(s)$$

The equilibrium is reached slowly.

*(a) Describe the changes you would see if aqueous solutions of iron(II) sulfate and silver nitrate were mixed and allowed to stand for a few hours.

(2)

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(b) The concentration of silver ions in the equilibrium mixture can be found by titration with potassium thiocyanate. Silver thiocyanate precipitates.

$$Ag^{+}(aq) + CNS^{-}(aq) \quad \Delta S^{\ominus}_{total} = R \ln K \quad AgCNS(s)$$

When all the silver ions have reacted, a deep red complex ion of iron(III) thiocyanate forms.

In an experiment, 25.0 cm^3 of 0.100 mol dm^{-3} silver nitrate solution was added to 25.0 cm^3 of 0.100 mol dm^{-3} of iron(II) sulfate solution, mixed thoroughly, and allowed to stand overnight in an air-tight container.

10.0 cm^3 samples of the reaction mixture were then titrated with 0.0200 mol dm^{-3} potassium thiocyanate solution. The average titre was 5.60 cm^3.

(i) The initial concentrations of silver ions and iron(II) ions **in the reaction mixture** are the same.

Calculate this initial concentration in mol dm^{-3}.

(1)

(ii) Calculate the number of moles of silver ions in the 10.0 cm^3 sample at equilibrium and hence calculate the equilibrium concentration of silver ions in the mixture.

(2)

(iii) Deduce the equilibrium concentration of iron(II) ions.

(1)

(iv) Hence calculate the equilibrium concentration of iron(III) ions.

(1)

(v) Write the expression for the equilibrium constant, K_c, for the reaction

$$Fe^{2+}(aq) + Ag^{+}(aq) \quad \Delta S^{\ominus}_{total} = R \ln K \quad Fe^{3+}(aq) + Ag(s)$$

Calculate its value and give your answer, with appropriate units, to **three** significant figures.

(4)

(c) (i) The relationship between the total entropy change for a reaction and the equilibrium constant is

$$\Delta S^{\ominus}_{total} = R \ln K$$

Calculate the total entropy change for this reaction, giving a sign and appropriate units.

[R = 8.31 J mol^{-1} K^{-1}]

(2)

(ii) $Fe^{2+}(aq) + Ag^{+}(aq) \quad \Delta S^{\ominus}_{total} = R \ln K \quad Fe^{3+}(aq) + Ag(s)$

$\Delta S^{\ominus}_{system}$ for this reaction is −208.3 J mol^{-1} K^{-1}

Use ideas about entropy to explain why this value is negative.

(2)

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(iii) Calculate the entropy change of the surroundings, $\Delta S^{\ominus}_{surroundings}$.

(1)

(iv) Use your answer to (c)(iii) to calculate the standard enthalpy change, $\Delta H^{\ominus}$, for this reaction at 298 K.

Hence state and explain the effect of increasing temperature on the value of $^{\ominus}$

(3)

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*(d) After the samples from the original mixture for the titration are taken, the remainder was filtered and then allowed to stand overnight, in an air-tight container at the same temperature. Another 10.0 cm^3 sample was taken and titrated. How, if at all, would you expect the titre to change?

Justify your answer. A calculation is not required.

(2)

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(Total for Question = 21 marks)

Q15.

Ammonia is manufactured by the Haber Process.

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

$$K_p = \frac{p(NH_3)^2}{p(N_2)p(H_2)^3}$$

The equilibrium constants for K_p and K_c are related by the equation

$$K_p = \frac{K_c}{(RT)^{\Delta n}}$$

where Δn is the number of moles of reactants minus the number of moles of products.

Calculate the value of K_c at 500 K when the value of $K_p = 3.55 \times 10^{-2}$ atm^{-2}.
Include the units for K_c .

[Use the value of $R = 0.0821$ dm^3 atm K^{-1} mol^{-1}]

(4)

(Total for question = 4 marks)

Q16.

Ammonia is manufactured by the Haber Process.

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

$$K_p = \frac{p(NH_3)^2}{p(N_2)p(H_2)^3}$$

The value of the equilibrium constant, K_p, varies with temperature.
The equation relating the values of the equilibrium constant at two temperatures is

$$\ln\left[\frac{K_2}{K_1}\right] = \frac{\Delta H}{R}\left[\frac{1}{T_1} - \frac{1}{T_2}\right]$$

The equilibrium constant, K_1, for the formation of ammonia is 6.76×10^5 atm^{-2} when the temperature $T_1 = 298$ K.
The enthalpy change $\Delta H = -92\,400$ J mol^{-1}.

Calculate the value of the equilibrium constant for this reaction at 310 K.

[Use the value of $R = 8.31$ J mol^{-1} K^{-1}]

(4)

(Total for question = 4 marks)

Q17.

This question is about equilibrium systems.

Sulfur dioxide and oxygen form an equilibrium with sulfur trioxide.

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

The composition of an equilibrium mixture at 698 K and a total pressure of 2.40 atm is shown in the table.

Substance	$SO_2(g)$	$O_2(g)$	$SO_3(g)$
Number of moles /mol	0.0160	0.0120	0.772

(i) Calculate the value of K_p at this temperature.

Include units, if appropriate.

(5)

(ii) Calculate the number of sulfur dioxide molecules present in this equilibrium mixture.

(1)

(iii) Deduce, by referring to K_p, how the number of sulfur dioxide molecules will change if more oxygen is added to the equilibrium mixture.

(2)

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(Total for question = 8 marks)

Q18.

Ethanoic acid and ethanol react together to form the ester ethyl ethanoate, $CH_3COOC_2H_5$, and water.

$$CH_3COOH(l) + CH_3CH_2OH(l) \rightleftharpoons CH_3COOCH_2CH_3(l) + H_2O(l)$$

(a) (i) Give the expression for K_c.

(1)

(ii) An equilibrium was reached when the amounts of substances shown in the table below were used.

Complete the table to show the amounts of each substance present at equilibrium.

(2)

Component	$CH_3COOH(l)$	$CH_3CH_2OH(l)$	$CH_3COOCH_2CH_3(l)$	$H_2O(l)$
Initial amount / mol	0.40	0.30	0.00	0.15
Equilibrium amount / mol	0.20			

(iii) Explain why K_c for this reaction has no units.

(1)

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(iv) Calculate the numerical value of K_c.

(1)

(b) The esterification reaction above was carried out in the presence of hydrochloric acid as the catalyst.

State the effect on the equilibrium position and the rate of attainment of equilibrium if the concentration of the acid catalyst were to be increased.

(2)

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(c) (i) Identify which bonds are broken and which bonds are made in the esterification reaction.

(2)

Bonds broken:

Bonds made:

(ii) Explain why ΔH for this reaction is not **exactly** zero.
(A calculation is not required.)

(1)

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(d) (i) State the relationship between ΔS_{total} and the equilibrium constant, K, of a reaction.

(1)

*(ii) Use entropy considerations and your answer to (d)(i) to predict any effect of an increase in temperature on the value of the equilibrium constant of a reaction for which ΔH is zero. Assume

that ΔS_{system} does not change with temperature.

(3)

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(e) An alternative method for preparing ethyl ethanoate is to react ethanoyl chloride with ethanol.

(i) Give the equation for the reaction.

(1)

(ii) Draw the **skeletal** formula of ethyl ethanoate.

(1)

(iii) Ethanoyl chloride also reacts with concentrated ammonia. Draw the **displayed** formula of the organic product of this reaction.

(1)

(f) (i) Complete the equation below for the alkaline hydrolysis of ethyl ethanoate using sodium hydroxide. State symbols are **not** required.

(1)

$CH_3COOCH_2CH_3 + NaOH \rightarrow$

(ii) Explain why the reaction in (f)(i) gives a better yield of the alcohol compared with acid hydrolysis of the ethyl ethanoate.

(1)

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(Total for question = 19 marks)

Q19.

The decomposition of hydrogen iodide to form iodine and hydrogen is an equilibrium reaction.

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

The equilibrium was investigated by taking sealed tubes containing the same mass of hydrogen iodide and heating them at 700 K for some time. At this temperature, the equilibrium takes about two days to be established.

The tubes were rapidly cooled to room temperature, which maintained the equilibrium concentrations, because at this temperature the reaction is extremely slow.

Each tube was opened under an aqueous solution of potassium iodide, which dissolved the hydrogen iodide and the iodine. The amount of iodine was found by titration and the composition of the equilibrium mixture calculated. From the number of moles of each substance at equilibrium, and the volume of the tubes, the equilibrium concentrations were calculated.

(a) (i) How would the appearance of the contents of a tube change as it was cooled to room temperature?

(2)

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(ii) How could you show that equilibrium had been established?

(2)

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(b) The equilibrium concentrations of one experiment are shown in the table below.

[HI] /mol dm^{-3}	[H_2] /mol dm^{-3}	[I_2] /mol dm^{-3}
0.00353	0.00048	0.00048

*(i) The volume of the tube in this experiment was 30 cm^3. Calculate the initial mass of hydrogen iodide. Show your working.

(5)

(ii) Write an expression for the equilibrium constant, K_c, at 700 K.

(1)

(iii) Calculate the value for this equilibrium constant.

(1)

(iv) Does this equilibrium constant have units? Explain your answer.

(1)

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(c) The equation for the reaction at 700 K can also be written

$$HI(g) \rightleftharpoons \tfrac{1}{2}H_2(g) + \tfrac{1}{2}I_2(g)$$

(i) Write the equilibrium constant, K_c', for this reaction.

(1)

(ii) Using the same equilibrium concentrations as below, calculate the equilibrium constant, K_c'.

Deduce the relationship between this value and the value calculated in part (b)(iii).

(2)

[HI] /mol dm^{-3}	[H_2] /mol dm^{-3}	[I_2] /mol dm^{-3}
0.00353	0.00048	0.00048

(d) Consider the following equilibrium reaction.

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

For this reaction $K_p = \dfrac{P_{H_2}}{P_{HI}^2}$

Use the expression for K_p to explain the effect of an increase in total pressure on the position of the equilibrium.

(3)

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(Total for question = 18 marks)

Q20.

Sulfur dioxide and oxygen react together to form sulfur trioxide in a reversible reaction that can reach a position of equilibrium. The equation for the reaction is:

$$SO_2(g) + \tfrac{1}{2}O_2(g) \rightleftharpoons SO_3(g) \qquad \Delta_r H^{\ominus} = -98.0 \text{ kJ mol}^{-1}$$

The numerical value of the equilibrium constant, K_p, for this reaction at 298 K is 2.00×10^{12}.

The expression for K_p for this reaction is

(1)

☒ **A** $\dfrac{p^2_{SO_3}}{p^2_{SO_2} \cdot p_{O_2}}$

☒ **B** $\dfrac{p_{SO_3}}{p_{SO_2} \cdot p^{\frac{1}{2}}_{O_2}}$

☒ **C** $\dfrac{p^2_{SO_2} \cdot p_{O_2}}{p^2_{SO_3}}$

☒ **D** $\dfrac{p_{SO_2} \cdot p^{\frac{1}{2}}_{O_2}}{p_{SO_3}}$

(b) State what the value of K_p suggests about the relative amounts of the components in the equilibrium mixture at 298 K.

(1)

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(c) Explain, in terms of K_p, how an increase in temperature will affect the yield of sulfur trioxide in the equilibrium mixture.

(2)

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(d) Calculate the value for $\Delta_r G^{-}$ at 298 K using the equation $\Delta_r G^{\ominus} = -RT \ln K_p$.
Hence confirm the thermodynamic feasibility of the reaction at this temperature.

(2)

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(e) Give a reason why, in industry, the reaction is carried out at a temperature of 687 K.

(1)

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(Total for question = 7 marks)

Q21.

This question is about the solubility of metal hydroxides.

When excess magnesium hydroxide is added to water and shaken, a saturated solution is formed and the mixture reaches equilibrium.

$$Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2OH^-(aq)$$

The equilibrium constant, K_c , for this reaction is

$$K_c = [Mg^{2+}(aq)][OH^-(aq)]^2$$

(i) Give a reason why the magnesium hydroxide is not included in the expression for K_c.

(1)

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(ii) Give the units for K_c.

(1)

(iii) Calculate the enthalpy change of solution of magnesium hydroxide, using the following data.

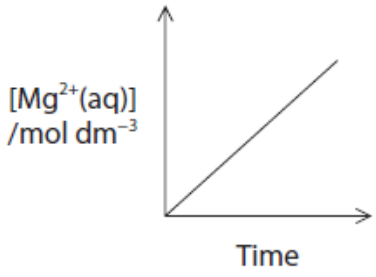
Energy or enthalpy change	Value / kJ mol^{-1}
Lattice energy of $Mg(OH)_2(s)$	−2842
$\Delta_{hyd}H$ ($Mg^{2+}(aq)$)	−1920
$\Delta_{hyd}H$ ($OH^-(aq)$)	−460

(2)

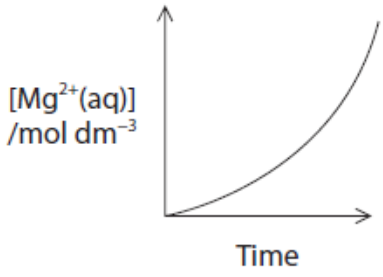
(iv) Which graph shows the change in the concentration of the $Mg^{2+}(aq)$ ions when some solid magnesium hydroxide is shaken with water and left to reach equilibrium?

(1)

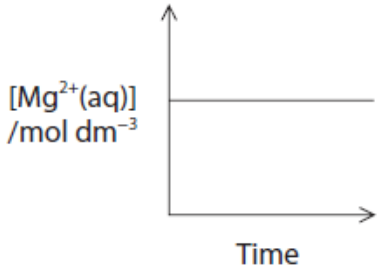
☒ **A**



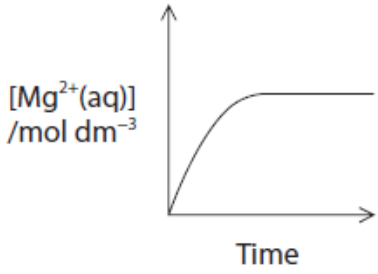
☒ **B**



☒ **C**



☒ **D**



(v) Predict the effect, if any, of adding each of the following to a saturated solution of magnesium hydroxide in contact with solid magnesium hydroxide. Justify your answers in terms of the effect on the equilibrium.

$$Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2OH^-(aq)$$

(4)

Magnesium sulfate solution

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Dilute hydrochloric acid

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(Total for question = 9 marks)

Q22.

This question is about the reversible reaction below.

$$2NO_2(g) \rightleftharpoons N_2O_4(g)$$

(a) A chemist investigating this reaction started with 10 moles of NO_2 and allowed the system to reach equilibrium. If 3 moles of N_2O_4 are formed, the number of moles of NO_2 at equilibrium is

(1)

☒ **A** 8.5

☒ **B** 7

☒ **C** 6

☒ **D** 4

(b) Under different conditions, 40% of the moles present at equilibrium is N_2O_4. If the total pressure of the system is 2.0 atm, the numerical value of the equilibrium constant, K_p is

(1)

☒ **A** 0.56

☒ **B** 0.67

☒ **C** 1.5

☒ **D** 1.8

(Total for question = 2 marks)

Q23.

The equation for the synthesis of methanol is

$$CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$$

At equilibrium, when the temperature is 340 K, the total pressure is 20 atm. The moles of each component present at equilibrium are shown in the table below.

Formula	Equilibrium moles / mol	Mole fraction
CO	0.15	0.23
H_2	0.32	
CH_3OH	0.18	0.28

(a) The mole fraction of hydrogen in the equilibrium mixture is

(1)

☒ **A** 0.23

☒ **B** 0.46

☒ **C** 0.49

☒ **D** 0.92

(b) The numerical value for the equilibrium partial pressure of the carbon monoxide, in atmospheres, is

(1)

☒ **A** 3.0

☒ **B** 4.6

☒ **C** 5.0

☒ **D** 9.2

(c) Units for the equilibrium constant, K_p, for this reaction are

(1)

☒ **A** no units

☒ **B** atm

☒ **C** atm^{-1}

☒ **D** atm^{-2}

(Total for question = 3 marks)

Q24.

Hydrogen is produced on a large scale by several different processes.

One process for producing hydrogen involves reacting white-hot carbon with steam.

$$C(s) + H_2O(g) \rightleftharpoons H_2(g) + CO(g) \qquad \Delta H = +131 \text{ kJ mol}^{-1}$$

The expression for the equilibrium constant, K_p, is

$$K_p = \frac{p(H_2)\, p(CO)}{p(H_2O)}$$

(i) Give a reason why the partial pressure of carbon is not included in the expression.

(1)

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(ii) Explain the effect of an increase in pressure on the equilibrium position of this reaction.

(2)

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(iii) Explain, by reference to any change in the value of K_p , the effect of an increase in temperature on the equilibrium position of this reaction.

(2)

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(iv) At 1000 K and a total pressure of 2.0 atm, 1.00 mol of steam reacted with excess carbon.

At equilibrium, 0.81 mol of hydrogen was present.

Calculate the value of K_p at 1000 K, stating any units.

(4)

(Total for question = 9 marks)

Q25.

This is a question about water.

An equation for the ionisation of water is

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

The expression for the ionic product of water is

$$K_w = [H^+(aq)]\,[OH^-(aq)]$$

The value of K_w at 310 K is 2.40×10^{-14} mol^2 dm^{-6}

(i) Calculate the pH of water at 310 K.

Give your answer to **two** decimal places.

(2)

(ii) Predict, with a reason, whether water is acidic, alkaline or neutral at 310 K.

(2)

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(iii) Predict, with a reason, the sign of the enthalpy change for the ionisation of water.

(1)

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(Total for question = 5 marks)

Q26.

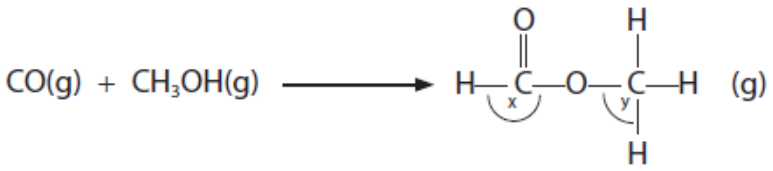
Answer the questions with a cross in the boxes you think are correct ☒. If you change your mind about an answer, put a line through the box ☒ and then mark your new answer with a cross ☒.

A mixture of ethanoic acid, ethanol and a catalyst was left for several days to reach equilibrium.

$$CH_3COOH(l) + CH_3CH_2OH(l) \rightleftharpoons CH_3COOCH_2CH_3(l) + H_2O(l)$$

The equilibrium constant, K_c, **under these conditions**, was 0.28.

Another ester, methyl methanoate, can be formed by the reaction between methanol and carbon monoxide in the gaseous phase.



(i) The two O−C−H bond angles, x and y, in the ester are approximately

(1)

☒ **A** 180° and 90°

☒ **B** 120° and 90°

☒ **C** 120° and 109.5°

☒ **D** 109.5° and 109.5°

(ii) The reaction often forms an equilibrium mixture.

Which could be the units for the equilibrium constant, K_p?

(1)

☒ **A** mol dm^{-3}

☒ **B** dm^3 mol^{-1}

☒ **C** atm

☒ **D** atm^{-1}

(iii) Describe what effect, if any, increasing the pressure would have on the equilibrium constant, K_p. Justify your answer.

(2)

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(Total for question = 4 marks)

Q27.

A mixture of ethanoic acid, ethanol and a catalyst was left for several days to reach equilibrium.

$$CH_3COOH(l) + CH_3CH_2OH(l) \rightleftharpoons CH_3COOCH_2CH_3(l) + H_2O(l)$$

The equilibrium constant, K_c, **under these conditions**, was 0.28.

(i) Write the expression for the equilibrium constant, K_c.

(1)

(ii) The initial amounts of ethanol and ethanoic acid used were 1.2 mol of each reactant.

Use this information, your expression for the equilibrium constant, K_c, and the value for K_c, to find the amounts of each product at equilibrium, in moles.

(3)