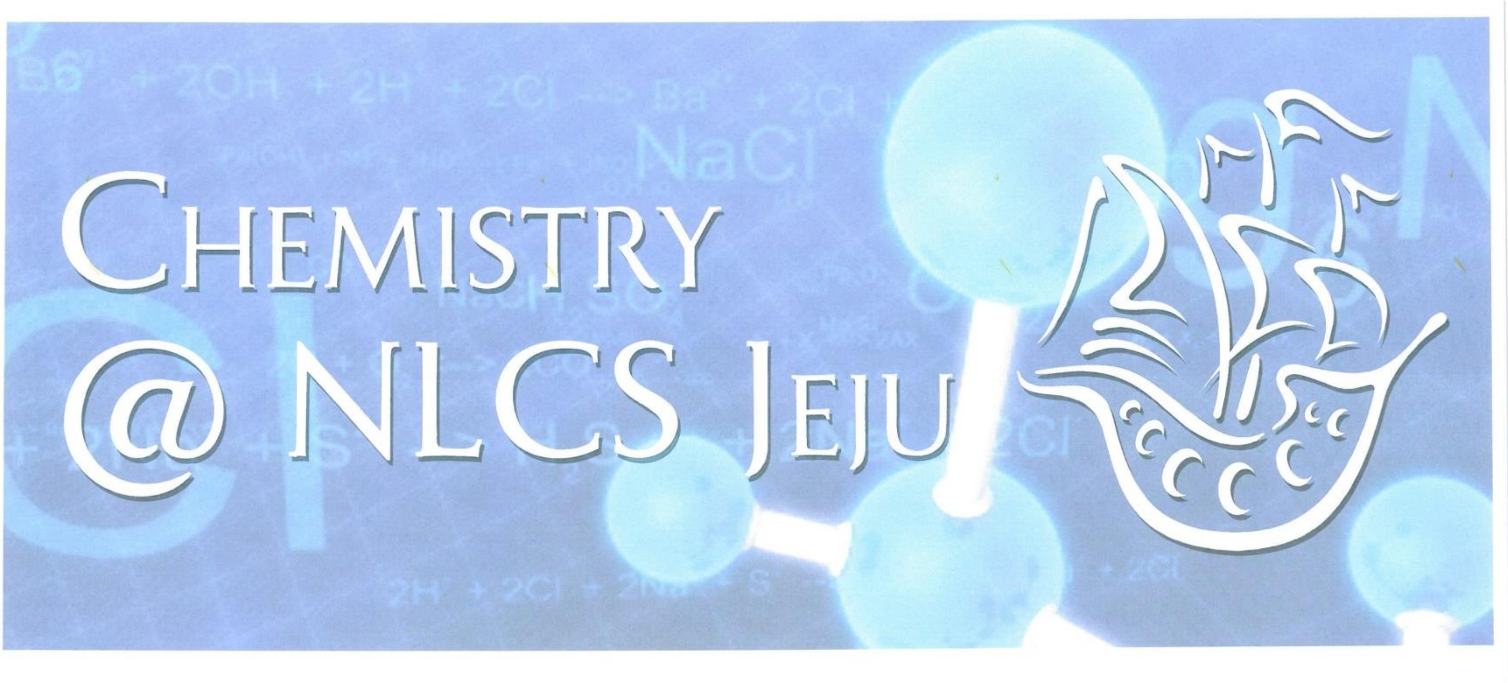
Topic 7

HIGHER Level



Summer & Winter Papers

Summer 1999 to Summer 2013

Name:________________________________

Topic Exam Statistics (Paper 2):

Section	Marks	% of All Marks	Last four exams marks	Last four exams %
A	66/1080	**6%**	/160	%
B	247/2700	**9%**	/400	%
TOTAL	313/3780	**8%**	/560	%

Total number of papers represented here is 27, each with 40 marks of Section A and 100 marks of section B (4 questions from which you chose to answer only 2)

4. Methanol may be produced by the exothermic reaction of carbon monoxide gas and hydrogen gas.

$$CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g) \qquad \Delta H^{\ominus} = -103 \text{ kJ}$$

(a) State the equilibrium constant expression, K_c, for the production of methanol. *[1]*

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(b) State and explain the effect of changing the following conditions on the amount of methanol present at equilibrium:

(i) increasing the temperature of the reaction at constant pressure. *[2]*

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(ii) increasing the pressure of the reaction at constant temperature. *[2]*

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(c) The conditions used in industry during the production of methanol are a temperature of 450 °C and pressure of up to 220 atm. Explain why these conditions are used rather than those that could give an even greater amount of methanol. *[2]*

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(d) A catalyst of copper mixed with zinc oxide and alumina is used in industry for this production of methanol. Explain the function of the catalyst. *[1]*

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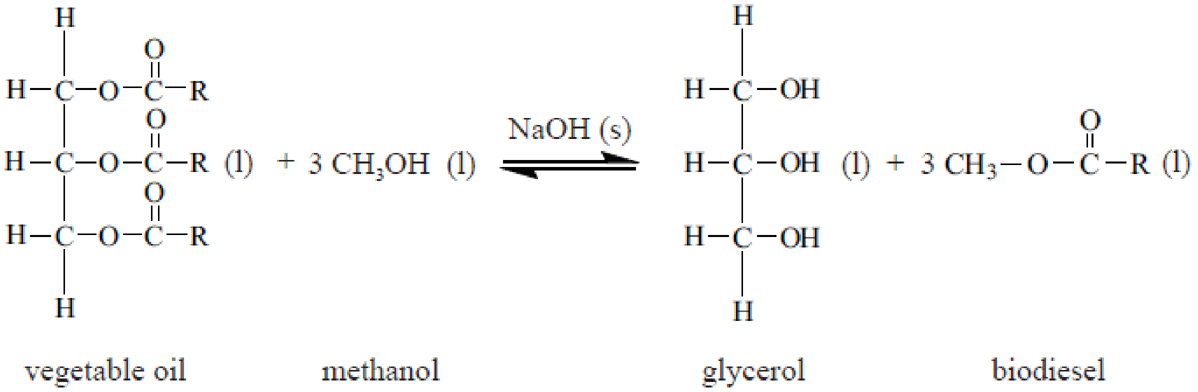
HL SECTION A 09s

1. Biodiesel makes use of plants' ability to fix atmospheric carbon by photosynthesis. Many companies and individuals are now using biodiesel as a fuel in order to reduce their carbon footprint. Biodiesel can be synthesized from vegetable oil according to the following reaction.



(c) The reversible arrows in the equation indicate that the production of biodiesel is an equilibrium process.

(i) State what is meant by the term *dynamic equilibrium*. *[1]*

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(ii) Using the abbreviations [vegetable oil], [methanol], [glycerol] and [biodiesel] deduce the equilibrium constant expression (K_c) for this reaction. *[1]*

(iii) Suggest a reason why excess methanol is used in this process. *[1]*

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(iv) State and explain the effect that the addition of the sodium hydroxide catalyst will have on the position of equilibrium. *[2]*

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HL SECTION A 09s

5. Sodium oxide, Na_2O, is a white solid with a high melting point.

(b) Molten sodium oxide is a good conductor of electricity. State the half-equation for the reaction occurring at the positive electrode during the electrolysis of molten sodium oxide. *[1]*

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HL SECTION A 08w

4. Much of the hydrogen used to manufacture ammonia is made using the following reaction between methane and steam.

$$CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3H_2(g) \qquad \Delta H^\ominus = +210\ \text{kJ}$$

(a) In an experiment using a nickel catalyst, 1.0 mol of methane and 2.0 mol of steam were added to a container of volume 20 dm^3 and heated to a constant temperature. When equilibrium was reached the mixture contained 0.50 mol of carbon monoxide.

(i) Calculate the amount, in moles, of each of the other substances present. *[2]*

Methane . . . .

Steam . . . .

Hydrogen . . . .

(ii) Deduce the equilibrium constant expression, K_c, for this reaction. *[1]*

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(iii) Calculate a value for K_c for this reaction and deduce its units. *[2]*

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HL SECTION A 04s

7. The equilibrium between nitrogen dioxide (dark brown) and dinitrogen tetroxide (colourless) is represented by the following equation.

$$2NO_2(g) \rightleftharpoons N_2O_4(g) \quad \Delta H = \text{negative} \quad K_c = 1 \text{ at } 328\text{ K}$$

(a) Write the equilibrium constant expression, K_c. *[1]*

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(b) State and explain the effect of an increase in temperature on the value of K_c. *[2]*

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(c) State and explain the visible change that takes place as a result of a decrease in pressure, after equilibrium is re-established. *[2]*

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(d) Two moles of $NO_2(g)$ and two moles of $N_2O_4(g)$ were placed in an empty 1 dm^3 container and allowed to come to equilibrium at 328 K. Predict, with reference to the value of K_c, whether the equilibrium mixture would contain more or less than two moles of $NO_2(g)$. *[2]*

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HL SECTION A 03s

3. (a) An industrial gas mixture is produced by the catalytic reforming of methane using steam.

$$CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3H_2(g) \qquad \Delta H = +206\text{ kJ}$$

By circling the appropriate letter(s) below, identify the change(s) that would shift the position of equilibrium to the right.

A increasing the temperature
B decreasing the temperature
C increasing the pressure
D adding a catalyst
E decreasing the pressure
F increasing the concentration of H_2 *[2]*

(ii) After 19 s the concentrations of the reactant and product do not change. State what this indicates about the reaction. *[1]*

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HL SECTION A 01s

4. When 1.0 mole of ethanoic acid is mixed with 1.0 mole ethanol, and the mixture allowed to reach equilibrium, the following reaction occurs:

$$CH_3COOH(l) + C_2H_5OH(l) \rightleftharpoons CH_3COOC_2H_5(l) + H_2O(l)$$

The amounts of ethyl ethanoate and water at equilibrium are both 0.67 moles.

(a) (i) What is meant by the term *equilibrium*? *[2]*

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(ii) Write an expression for K_c for this reaction. *[1]*

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(iii) Calculate the value of K_c for this reaction. *[2]*

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(b) For the dissociation $H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$, the ionic product is given by $K_w = [H^+(aq)][OH^-(aq)]$. The value of K_w is 1.0×10^{-14} $mol^2\,dm^{-6}$ at 298 K and 2.4×10^{-14} $mol^2\,dm^{-6}$ at 310 K. Using Le Chatelier's principle, deduce whether the dissociation of water is exothermic or endothermic. *[3]*

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(c) For the system $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ state and explain the effect on the position of equilibrium of

(i) adding a catalyst. *[2]*

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(ii) adding some helium gas but keeping the total gas volume constant. *[2]*

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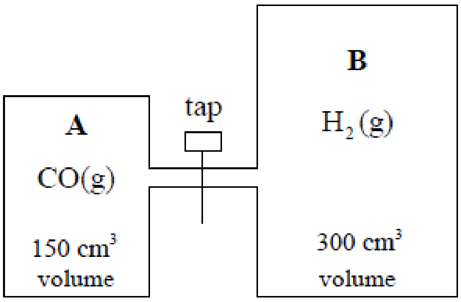
HL SECTION A 00w

4b

4. Methanol is an important industrial solvent and fuel. It can be produced from carbon monoxide and hydrogen according to the following equation:

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

The effect of different catalysts on this reaction is investigated using the following apparatus:



A contains 1 mole of carbon monoxide and **B** contains 2 moles of hydrogen. The gases in both containers are at the same temperature and pressure. The tap is closed at the start of the experiment.

(ii) What will happen to the concentration of methanol if the system is allowed to reach equilibrium at a lower temperature? *[1]*

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(c) (i) Write the equilibrium expression for the above reaction, and give the units for K_c. *[1]*

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(ii) Calculate a value for K_c if the maximum yield of methanol is 85 %. *[3]*

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(iii) When this reaction is carried out on an industrial scale, the yield is about 60 %. Suggest a reason for this. *[1]*

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(iv) Copper is a good catalyst for this reaction. What effect, if any, will the addition of copper have on the value of K_c? *[1]*

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HL SECTION A 00s

1. (a) Automobile engines produce a variety of air pollutants at high temperatures. One of these pollutants is nitrogen(II) oxide, NO, formed in the following reaction:

$$N_2(g) + O_2(g) \rightleftharpoons 2NO(g) \qquad \Delta H^{\ominus} \text{ is positive}$$

(i) What is the significance of the positive value for $\Delta H^{\ominus}$? *[1]*

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(ii) State and explain the effect of a decrease in temperature on the value of K_c for this reaction. *[2]*

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(iii) In an experiment, the initial concentrations are $N_2(g) = 1.6 \text{ mol dm}^{-3}$ and $O_2(g) = 1.6 \text{ mol dm}^{-3}$. Calculate the concentration of the NO(g) after equilibrium is established. ($K_c = 1.7 \times 10^{-3}$) *[3]*

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HL SECTION A .99s

3. Dinitrogen tetroxide, N_2O_4, decomposes endothermically according to the following equation:

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

(a) Give the equilibrium constant expression, K_c, for the above reaction. *[1]*

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(b) 1 mol of N_2O_4 was placed in a 1.00 dm^3 evacuated flask at 25° C. The flask was stoppered and equilibrium was allowed to be established at this temperature. State whether the yield of NO_2 will increase, decrease, or remain the same when equilibrium is re-established after each of the following independent changes. Give a brief explanation in each case.

(i) More N_2O_4 is added to the equilibrium mixture in the flask. *[2]*

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(ii) The pressure on the system is increased. *[2]*

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(iii) A catalyst is added to the system. *[2]*

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(c) The 1 mol of N_2O_4 in the flask at 25° C became 0.8 mol of N_2O_4 when equilibrium had been established.
Calculate;

(i) the mols of NO_2 present at equilibrium. *[1]*

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(ii) the numerical value of K_c with its units. *[2]*

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(d) State and explain the effect on the value of K_c if the temperature of the reaction were raised to 100° C. *[2]*

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HL B 13s

8

(c) The manufacture of gaseous methanol from CO and H_2 involves an equilibrium reaction.

$$\mathrm{CO(g)} + 2\mathrm{H_2(g)} \rightleftharpoons \mathrm{CH_3OH(g)} \qquad \Delta H^{\ominus} < 0$$

(i) Outline the characteristics of a chemical equilibrium. *[2]*

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(ii) Deduce the equilibrium constant expression, K_c, for this reaction. *[1]*

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(iii) 1.00 mol of CH_3OH is placed in a closed container of volume 1.00 dm^3 until equilibrium is reached with CO and H_2. At equilibrium 0.492 mol of CH_3OH are present. Calculate K_c. *[3]*

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(d) State and explain the effect of the following changes on the equilibrium position of the reaction in part (c).

(i) Increase in temperature. *[2]*

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(ii) Increase in pressure. *[2]*

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HL B 12w

5

NOT with 5c(iii)

Methanol reacts with carbon monoxide to form ethanoic acid, $CH_3COOH\,(l)$.

$$CH_3OH\,(l) + CO\,(g) \rightarrow CH_3COOH\,(l)$$

$$\Delta H^{\ominus} = (-485 + 111 + 239 =) -135\ (\text{kJ mol}^{-1})$$

(vi) In industry, this reaction is carried out at a temperature greater than 298 K. State and explain the effect of increasing the temperature on the value of the equilibrium constant, K_c. *[2]*

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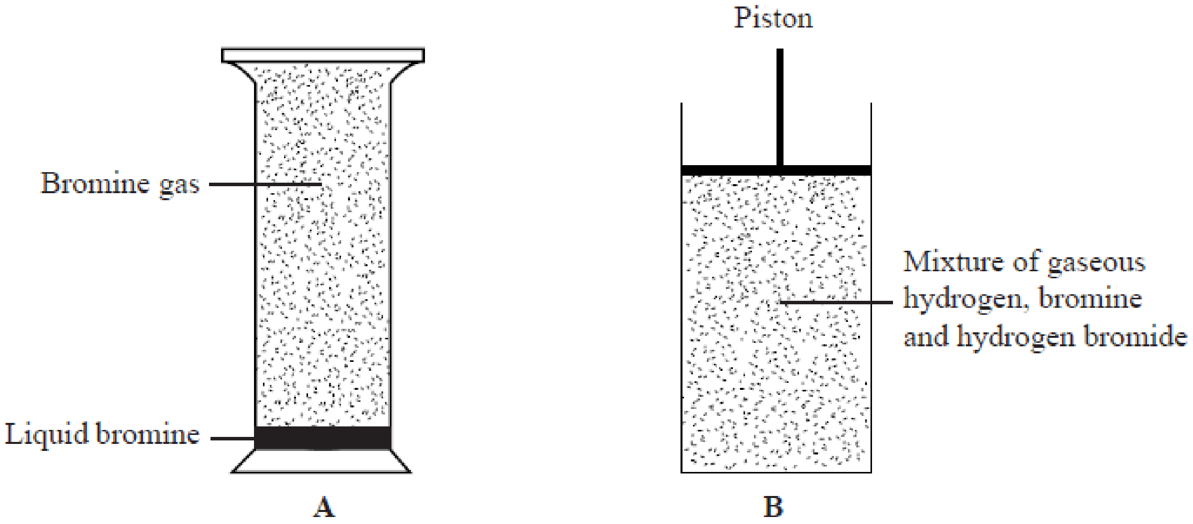
HL B 12s

6. Consider the two equilibrium systems involving bromine gas illustrated below.



(a) State equations to represent the equilibria in **A** and **B** with $Br_2(g)$ on the left-hand side in both equilibria. *[2]*

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(b) (i) Describe what you would observe if a small amount of liquid bromine is introduced into **A**. *[1]*

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(ii) Predict what happens to the position of equilibrium if a small amount of hydrogen is introduced into **B**. *[1]*

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(iii) State and explain the effect of increasing the pressure in **B** on the position of equilibrium. *[2]*

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(c) (i) Deduce the equilibrium constant expression, K_c, for the equilibrium in **B**. *[1]*

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(ii) State the effect of increasing $[H_2]$ in **B** on the value of K_c. *[1]*

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(e) When bromine dissolves in water, 1 % of the original bromine molecules react according to the following equation.

$$Br_2\,(aq) + H_2O\,(l) \rightleftharpoons HOBr\,(aq) + HBr\,(aq)$$

(iii) Estimate the magnitude of K_c for this reaction. Choose your value from the following options: [1]

$K_c = 0$ $K_c < 1$ $K_c = 1$ $K_c > 1$

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HL B 11w

8

(d) Nitrogen reacts with hydrogen to form ammonia in the Haber process, according to the following equilibrium.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \qquad \Delta H^{\ominus} = -92.6\text{ kJ}$$

(i) Define the term *rate of reaction*. [1]

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(ii) A high pressure such as 1000 atm and a low temperature such as 300 K can produce a high yield of ammonia. Discuss how these conditions compare with the actual conditions of pressure and temperature used in the Haber process. [4]

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HL B 11s

7

(c) In an experiment conducted at 25.0°C, the initial concentration of propanoic acid and methanol were 1.6 mol dm^{-3} and 2.0 mol dm^{-3} respectively. Once equilibrium was established, a sample of the mixture was removed and analysed. It was found to contain 0.80 mol dm^{-3} of compound **X**.

(i) Calculate the concentrations of the other three species present at equilibrium. *[3]*

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(ii) State the equilibrium constant expression, K_c, and calculate the equilibrium constant for this reaction at 25.0°C. *[2]*

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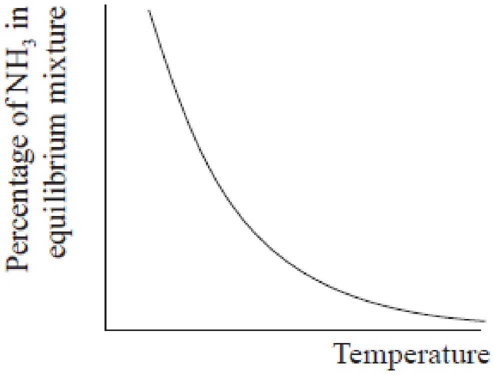
HL B 10w

6. The Haber process enables the large-scale production of ammonia needed to make fertilizers.

(a) The equation for the Haber process is given below.

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

The percentage of ammonia in the equilibrium mixture varies with temperature.



(i) Use the graph to deduce whether the forward reaction is exothermic or endothermic and explain your choice. *[2]*

(ii) State and explain the effect of increasing the pressure on the yield of ammonia. *[2]*

(iii) Deduce the equilibrium constant expression, K_c, for the reaction. *[1]*

(iv) A mixture of 1.00 mol N_2 and 3.00 mol H_2 was placed in a 1.0 dm^3 flask at 400 °C. When the system was allowed to reach equilibrium, the concentration of NH_3 was found to be 0.062 mol dm^{-3}. Determine the equilibrium constant, K_c, of the reaction at this temperature. *[3]*

(v) Iron is used as a catalyst in the Haber process. State the effect of a catalyst on the value of K_c. *[1]*

HL B 10s

5. (a) The production of ammonia is an important industrial process.

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

(b) The reaction used in the production of ammonia is an equilibrium reaction. Outline the characteristics of a system at equilibrium. *[2]*

(c) Deduce the equilibrium constant expression, K_c, for the production of ammonia. *[1]*

(d) (i) 0.20 mol of $N_2(g)$ and 0.20 mol of $H_2(g)$ were allowed to reach equilibrium in a 1 dm^3 closed container. At equilibrium the concentration of $NH_3(g)$ was 0.060 mol dm^{-3}. Determine the equilibrium concentrations of $N_2(g)$ and $H_2(g)$ and calculate the value of K_c. *[3]*

(ii) Predict and explain how increasing the temperature will affect the value of K_c. *[2]*

(e) Describe how increasing the pressure affects the yield of ammonia. *[2]*

(f) In practice, typical conditions used in the Haber process are a temperature of 500 °C and a pressure of 200 atmospheres. Outline why these conditions are used rather than those that give the highest yield. *[2]*

(g) A catalyst of iron is used in the Haber process. State and explain how the catalyst affects K_c and the position of equilibrium. *[3]*

HL B 09w

6

(e) Consider the following equilibrium reaction.

$$Cl_2(g) + SO_2(g) \rightleftharpoons SO_2Cl_2(g) \qquad \Delta H^{\ominus} = -84.5 \text{ kJ}$$

In a 1.00 dm^3 closed container, at 375 °C, 8.60×10^{-3} mol of SO_2 and 8.60×10^{-3} mol of Cl_2 were introduced. At equilibrium, 7.65×10^{-4} mol of SO_2Cl_2 was formed.

(i) Deduce the equilibrium constant expression, K_c, for the reaction. *[1]*

(ii) Determine the value of the equilibrium constant, K_c. *[3]*

(iii) If the temperature of the reaction is changed to 300 °C, predict, stating a reason in each case, whether the equilibrium concentration of SO_2Cl_2 and the value of K_c will increase or decrease. *[3]*

(iv) If the volume of the container is changed to 1.50 dm^3, predict, stating a reason in each case, how this will affect the equilibrium concentration of SO_2Cl_2 and the value of K_c. *[3]*

(v) Suggest, stating a reason, how the addition of a catalyst at constant pressure and temperature will affect the equilibrium concentration of SO_2Cl_2. *[2]*

HL B 08w

6. (a) The reaction between chlorine and bromide ions is a redox reaction.

$$Cl_2(g) + 2Br^-(aq) \rightarrow Br_2(aq) + 2Cl^-(aq)$$

Define the term *oxidation* in terms of electron transfer and identify the species that is oxidized in this reaction. *[2]*

(b) The oxidation number of oxygen is −2 in most compounds containing oxygen. Identify the oxidation numbers of all the other elements in both reactants and products in the following equation.

$$TiO_2(s) + 2Cl_2(g) + C(s) \rightarrow TiCl_4(l) + CO_2(g)$$ *[3]*

(c) By referring to oxidation numbers, deduce what happens, if anything, in terms of oxidation and reduction, to the named element in each of these reactions.

(i) Chromium in
$2K_2CrO_4(aq) + 2HCl(aq) \rightarrow K_2Cr_2O_7(aq) + 2KCl(aq) + H_2O(l)$ *[2]*

(ii) Chlorine in
$Cl_2(g) + H_2O(l) \rightarrow HCl(aq) + HClO(aq)$ *[2]*

(d) The table shows some reactions involving the metals **W**, **X**, **Y** and **Z**.

Reaction	Reactants	Products
1	$W + Z(NO_3)_2$	$Z + W(NO_3)_2$
2	$X + YCl_2$	no reaction
3	$Y + ZSO_4$	no reaction
4	$Z + XO$	$X + ZO$

(i) Use the information to arrange the four metals in a reactivity series, starting with the most reactive. Explain with reference to each of the metals how you decided which metal was the least reactive. *[4]*

(ii) Metal **V** forms compounds in which it has an oxidation number of +3. It is more reactive than any of the metals in the table. Predict the equation for the reaction between metal **V** and the oxide of metal **X**. *[1]*

HL B 08s

7

(c) The following equilibrium is involved in the industrial production of nitric acid from ammonia.

$$4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(g) \qquad \Delta H^{\ominus} = -910 \text{ kJ mol}^{-1}$$

Describe the effect, if any, of each of the following changes on the equilibrium concentration of nitrogen monoxide in a particular equilibrium mixture, give a reason in each case.

(i) Increasing the pressure, at constant temperature *[2]*

(ii) Increasing the temperature, at constant pressure *[2]*

(iii) Addition of a heterogeneous catalyst, at constant pressure and temperature *[2]*

(d) Deduce the equilibrium constant expression, K_c, including units for the forward reaction in part (c). *[2]*

(e) Identify which of the changes in part (c) will affect the value of K_c and predict whether the value of K_c will increase or decrease. *[2]*

HL B 07w

8

(b) In the gaseous state, methane and steam react to form hydrogen and carbon dioxide.

(i) Write an equation for the endothermic equilibrium reaction. Deduce the equilibrium expression for the reaction and state its units. *[4]*

(ii) Deduce and explain the conditions of temperature and pressure under which the forward reaction is favoured. *[4]*

(iii) Explain, at the molecular level, why the reaction is carried out at high pressure in industry. *[2]*

(iv) 1.0 mol methane and 3.0 mol steam are placed in a 1.0 dm^3 container and after equilibrium is reached, 2.0 mol hydrogen gas are present. Calculate the amount of each reactant and product at equilibrium, and thus determine the value of K_c for the reaction. *[4]*

HL B 07s

6. Consider the following reaction where colourless bromide ions react with colourless hydrogen peroxide to form a red-brown bromine solution.

$$2\text{Br}^-(\text{aq}) + \text{H}_2\text{O}_2(\text{aq}) + 2\text{H}^+(\text{aq}) \rightleftharpoons \text{Br}_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l}) \qquad \Delta H = \text{negative}$$

(a) Predict and explain the effect on the **position of equilibrium** when

(i) a small amount of sodium bromide solution is added. *[2]*

(ii) a small amount of sodium hydroxide solution is added. *[2]*

(iii) a catalyst is added. *[2]*

(b) State and explain the effect on the value of the **equilibrium constant** when the temperature of the reaction is increased. *[2]*

(c) State and explain the colour change when hydrochloric acid is added to the reaction solution at equilibrium. *[3]*

HL B 06w

8. (a) The equation for a reaction used in industry is

$$CH_4(g) + H_2O(g) \rightleftharpoons 3H_2(g) + CO(g) \qquad \Delta H^{\ominus} = +210\ \text{kJ}$$

Deduce the equilibrium constant expression, K_c, for this reaction. *[1]*

(b) Use Le Chatelier's principle to predict the effect of each of the following changes on the position of equilibrium and the value of K_c for the reaction in (a).

(i) Increasing the temperature at constant pressure. *[2]*

(ii) Increasing the pressure at constant temperature. *[2]*

(c) The equation for another reaction used in industry is

$$CO(g) + H_2O(g) \rightleftharpoons H_2(g) + CO_2(g) \qquad \Delta H^{\ominus} = -42\ \text{kJ}$$

(i) Under certain conditions of temperature and pressure, 2.0 mol of carbon monoxide and 3.2 mol of steam were left to reach equilibrium. At equilibrium, 1.6 mol of both hydrogen and carbon dioxide were present.
Calculate the amounts of carbon monoxide and steam at equilibrium and the value of K_c. *[3]*

(ii) Under the same conditions of temperature and pressure, 2.0 mol of carbon monoxide and 2.0 mol of steam were left to reach equilibrium.
Calculate the amounts of each reactant and product at equilibrium.
(If you were unable to calculate a value for K_c in (c)(i) use the value 9.0, although this is not the correct value.) *[2]*

(d) Many reversible reactions in industry use a catalyst. State and explain the effect of a catalyst on the position of equilibrium and the value of K_c. *[4]*

HL B 06s

6. Consider the following reaction in the Contact process for the production of sulfuric acid for parts (a) to (f) in this question.

$$2SO_2 + O_2 \rightleftharpoons 2SO_3$$

(a) Write the equilibrium constant expression for the reaction. *[1]*

(b) (i) State the catalyst used in this reaction of the Contact process. *[1]*

(ii) State and explain the effect of the catalyst on the value of the equilibrium constant and on the rate of the reaction. *[4]*

(d) Using Le Chatelier's principle explain the effect on the position of equilibrium of

(i) increasing the pressure at constant temperature. *[2]*

(ii) removing sulfur trioxide. *[2]*

(e) Using the following data, explain whether the above reaction is exothermic or endothermic. *[2]*

Temperature / K	**Equilibrium constant K_c / $dm^3\ mol^{-1}$**
298	9.77×10^{25}
500	8.61×10^{11}
700	1.75×10^{6}

(f) 1.50 mol of SO_2 and 2.00 mol of O_2 are placed in a 1.50 dm^3 flask and heated at 400 °C until equilibrium is reached. The final equilibrium mixture contained 0.500 mol of SO_3. Calculate the equilibrium constant, K_c of the reaction at this temperature, include units in your answer. *[5]*

(g) (i) Using the data below, state and explain the relationship between enthalpy of vaporization and intermolecular forces. *[3]*

Substance	**Pentane**	**Propanoic acid**
Standard enthalpy of vaporization $\Delta H_v^{\ominus}$ / kJ mol^{-1}	27	57

(ii) Draw a graph to show how the vapour pressure of pentane changes as it is heated from its melting point to its boiling point. Explain this change in terms of the kinetic theory. *[3]*

HL B 05w

6. (a) The equation for the decomposition of hydrogen iodide is

$$2HI(g) \rightleftharpoons H_2(g) + I_2(g) \qquad \Delta H^{\ominus} = +52\text{ kJ}$$

Predict and explain the effect on the position of equilibrium of

(i) increasing the pressure, at constant temperature. *[2]*

(ii) increasing the temperature, at constant pressure. *[2]*

(iii) adding a catalyst, at constant temperature and pressure. *[2]*

(b) Deduce the expression for K_c for the forward reaction. *[1]*

(c) The equilibrium formed during this reaction was investigated in two experiments carried out at different temperatures. The results are shown in the table below.

Experiment number	Initial concentration / mol dm^{-3}			Equilibrium concentration / mol dm^{-3}		
	[HI]	[H_2]	[I_2]	[HI]	[H_2]	[I_2]
1	0.06	0.00	0.00		0.01	
2	0.00	0.04	0.04	0.04		

(i) For each experiment, deduce the concentrations of the other species present at equilibrium. Calculate the values of K_c for the forward reaction for each experiment. *[6]*

(ii) Use the two calculated values of K_c to deduce which of the two experiments was carried out at the higher temperature, and explain your choice. (If you were not able to calculate the values of K_c in (c)(i), assume that the values are 0.1 for experiment 1 and 0.2 for experiment 2, although these are not the correct values.) *[2]*

HL B 03w

8

(b) The following equilibrium is established at 1700 °C.

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

If only carbon dioxide gas and hydrogen gas are present initially, sketch on a graph a line representing rate against time for

- the forward reaction
- the reverse reaction until shortly after equilibrium is established.

Explain the shape of each line. *[7]*

(c) K_c for the equilibrium reaction is determined at two different temperatures. At 850 °C, $K_c = 1.1$ whereas at 1700 °C, $K_c = 4.9$.

On the basis of these K_c values explain whether the reaction is exothermic or endothermic. *[3]*

HL B 02w

7. When 1 mole of hydrogen iodide is allowed to dissociate into its elements at 440 °C in a 1.0 dm^3 vessel, only 0.78 moles of hydrogen iodide are present at equilibrium.

(a) Write an equation for the equilibrium reaction, including state symbols. *[2]*

(b) (i) Write the expression for the equilibrium constant K_c for the above reaction and state its units. *[2]*

(ii) Calculate the equilibrium constant for the reaction at 440 °C using the information given above. *[4]*

(c) If $K_c = 0.04$ in the gas phase at 600 °C, deduce whether the dissociation of hydrogen iodide is exothermic or endothermic and explain your answer. *[3]*

(d) Predict and explain the effect on the position of equilibrium of

(i) increasing the total pressure
(ii) adding more hydrogen iodide at constant pressure
(iii) adding a catalyst. *[6]*

(e) The equilibrium constants for the dissociation of hydrogen chloride and hydrogen bromide are given below.

$$K_c(\text{HCl}) = 1.0 \times 10^{-17}$$
$$K_c(\text{HBr}) = 1.0 \times 10^{-9}$$

Explain what information these values give about the extent of each dissociation compared to that of hydrogen iodide. *[1]*

(f) Ammonia is manufactured by the Haber Process, for which the reaction is shown below.

$$\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g}) \quad \Delta H = -92 \text{ kJ}$$

A *heterogeneous* catalyst of iron is used.

(i) Sketch an enthalpy level diagram showing the enthalpy change and the activation energy of the reaction. *[3]*

(ii) State and explain the effect of the catalyst on the activation energy of the reaction. *[2]*

(iii) Explain the term *heterogeneous* and outline how such catalysts work. *[2]*

HL B 01w

8. When steam is passed over heated carbon, the following equilibrium is established:

$$C(s) + H_2O(g) \rightleftharpoons H_2(g) + CO(g)$$

The value of the equilibrium constant (K_c) at various temperatures is given in the table below:

Temperature / K	700	1000	1300
Equilibrium constant (K_c)	4.88×10^{-5}	4.51×10^{-2}	1.88

(a) Write the expression for the equilibrium constant and give its units. Deduce whether the forward reaction is endothermic or exothermic and state how you reached this conclusion. *[4]*

(b) Predict and explain the effect of an increase in temperature, total pressure and surface area of carbon on the:

(i) rate of the forward reaction; *[6]*

(ii) $[H_2O]:[H_2]$ ratio; *[6]*

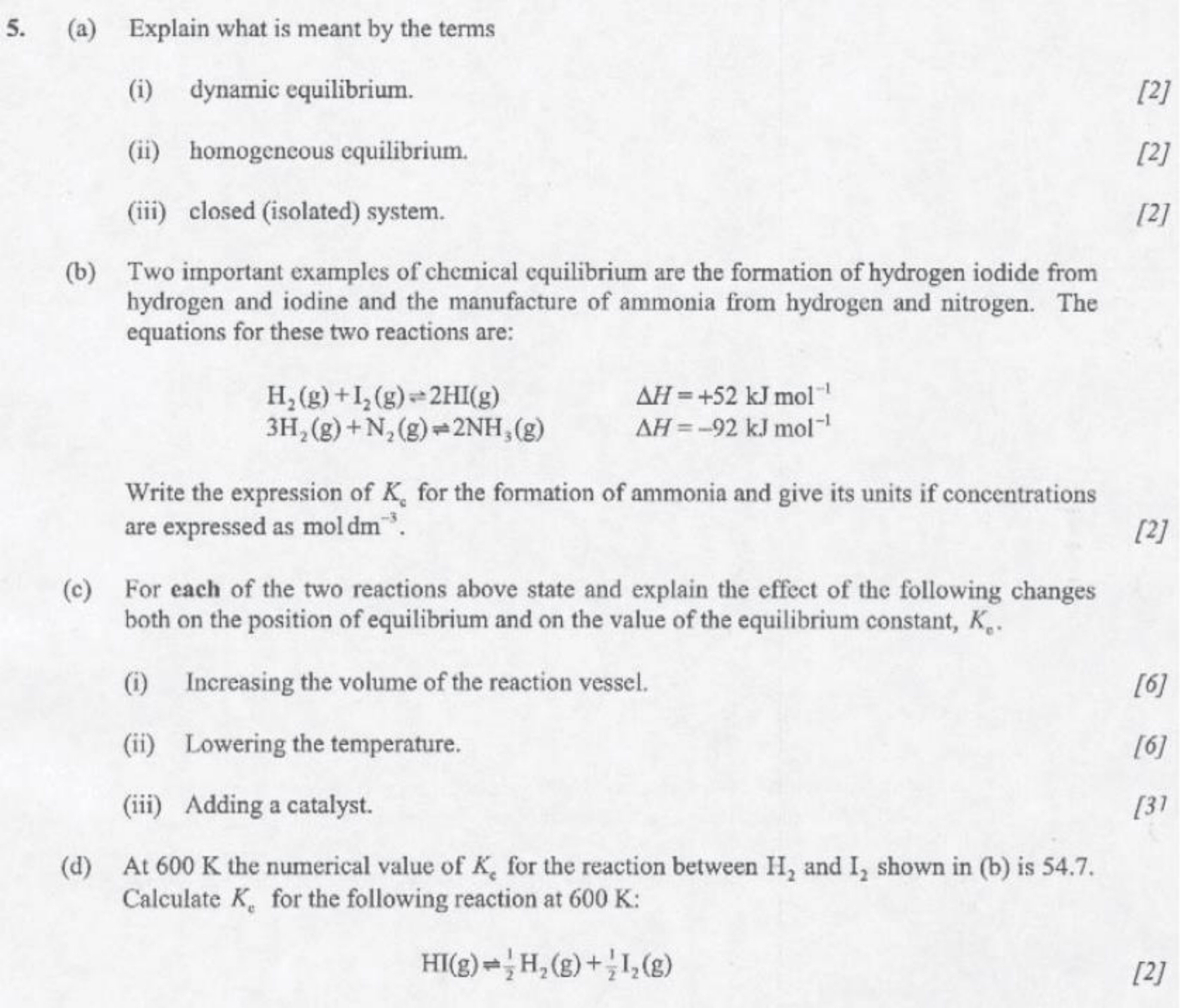
(iii) value of the equilibrium constant. *[3]*

HL B .99w

$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ $\Delta H = +52$ kJ mol^{-1}

$3H_2(g) + N_2(g) \rightleftharpoons 2NH_3(g)$ $\Delta H = -92$ kJ mol^{-1}

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



5. (a) Explain what is meant by the terms

(i) dynamic equilibrium. *[2]*

(ii) homogeneous equilibrium. *[2]*

(iii) closed (isolated) system. *[2]*

(b) Two important examples of chemical equilibrium are the formation of hydrogen iodide from hydrogen and iodine and the manufacture of ammonia from hydrogen and nitrogen. The equations for these two reactions are:

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g) \qquad \Delta H = +52\ \text{kJ mol}^{-1}$$
$$3H_2(g) + N_2(g) \rightleftharpoons 2NH_3(g) \qquad \Delta H = -92\ \text{kJ mol}^{-1}$$

Write the expression of K_c for the formation of ammonia and give its units if concentrations are expressed as mol dm^{-3}. *[2]*

(c) For **each** of the two reactions above state and explain the effect of the following changes both on the position of equilibrium and on the value of the equilibrium constant, K_c.

(i) Increasing the volume of the reaction vessel. *[6]*

(ii) Lowering the temperature. *[6]*

(iii) Adding a catalyst. *[3]*

(d) At 600 K the numerical value of K_c for the reaction between H_2 and I_2 shown in (b) is 54.7. Calculate K_c for the following reaction at 600 K:

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

IB HL 7 EQ Paper 2 s99 to s13 incl W Mark Scheme

HL SECTION A 11s

4. (a) $(K_c =)\frac{[CH_3OH]}{[CO][H_2]^2}$; *[1]*

Do not award mark if incorrect brackets are used or brackets are missing.

(b) (i) amount (of methanol)/product decreases / less methanol;
(forward reaction) exothermic / reverse reaction endothermic / *OWTTE*; *[2]*

(ii) amount (of methanol)/product increases / more methanol;
3 gas molecules/mol → 1 / decrease in volume / fewer gas molecules on right hand side/products / more gas molecules on left hand side/reactants; *[2]*

(c) high pressure expensive / greater cost of operating at high pressure;
lower temperature – lower (reaction) rate; *[2]*

(d) increases rate of forward and reverse reactions (equally) / lowers activation energy/E_a (of both forward and reverse reaction equally) / provides alternative path with lower activation energy/E_a; *[1]*

Accept reactants adsorb onto the catalyst surface and bonds weaken resulting in a decrease in activation energy.

HL SECTION A 09s

1

(c) (i) rate of the forward reaction is equal to the rate of the reverse reaction / forward and reverse reactions occur **and** the concentrations of the reactants and products do not change / *OWTTE*; *[1]*

(ii) $K_c = \frac{[\text{glycerol}]\times[\text{biodiesel}]^3}{[\text{vegetable oil}]\times[\text{methanol}]^3}$; *[1]*

(iii) to move the position of equilibrium to the right/product side / increase the yield of biodiesel; *[1]*

(iv) no effect (on position of equilibrium);
increases the rate of the forward and the reverse reactions equally (so equilibrium reached quicker) / it lowers E_a for both the forward and reverse reactions by the same amount / *OWTTE*; *[2]*

No ECF for explanation.

HL SECTION A 09s

5. (a) in the solid state ions are in fixed positions / there are no moveable ions / *OWTTE*; *[1]*

Do not accept answer that refers to atoms or molecules.

(b) $2O^{2-} \rightarrow O_2 + 4e^-$ / $O^{2-} \rightarrow \frac{1}{2}O_2 + 2e^-$; *[1]*

Accept e instead of e^-.

(c) (i) basic; *[1]*

Allow alkaline

(ii) $Na_2O + H_2O \rightarrow 2NaOH$ / $Na_2O + H_2O \rightarrow 2Na^+ + 2OH^-$; *[1]*

Do not accept $\rightleftharpoons$

HL SECTION A 08w

4. (a) (i) (methane) 0.50 (mol)
(steam) 1.5 (mol)
(hydrogen) 1.5 (mol) ***[2]***
*Three correct award **[2]**, two correct award **[1]**, one or zero correct award **[0]**.*

(ii) $K_c = \frac{[CO][H_2]^3}{[CH_4][H_2O]}$; ***[1]***

(iii) $K_c = \frac{(0.50 \div 20) \times (1.5 \div 20)^3}{(0.50 \div 20) \times (1.5 \div 20)} = 5.6 \times 10^{-3}$;
$mol^2\ dm^{-6}$ / equilibrium constants have no units; ***[2]***
Apply ECF from (a)(i) and (ii).

HL SECTION A 04s

7. *No ECF throughout this question.*

(a) $K_c = \frac{[N_2O_4]}{[NO_2]^2}$; ***[1]***

(b) K_c decreases;
forward reaction is exothermic / ΔH is negative / equilibrium moves to left / *OWTTE*; ***[2]***

(c) (mixture will get) darker / darker than expected;

equilibrium position moves to the left / towards reactants as there is an increase in the number of moles of gas from right to left; ***[2]***

(d) (equilibrium mixture contains) less (than 2 moles NO_2);
given values make $\frac{[N_2O_4]}{[NO_2]^2} = \frac{1}{2}$ *i.e.* too much NO_2 / *OWTTE*; ***[2]***

HL SECTION A 03s

3. (a) A;
E; ***[2]***
*If 3 choices shown **[1 max]**, if 4 choices shown **[0]**.*

(b) (i) after 15s (product) $= 0.37\ (mol\ dm^{-3})$;
rate $= \frac{0.37}{15} = 0.025$;
2 sig figs – if penalized here do not penalize in Q6 (c) (iv) or Q8 (d) (ii)
$mol\ dm^{-3}\ s^{-1}$ / $M\ s^{-1}$ / $\frac{M}{sec}$; ***[3]***

(ii) at equilibrium / rates of forward and reverse reactions are equal / $\Delta G = 0$; ***[1]***

HL SECTION A 01s

4. (a) (i) *Any two of the following, [1] each*
Some mention of a **reversible** reaction / $\rightleftharpoons$ *etc.*
Indication that rate of left to right = rate right to left
At equilibrium no net change in concentration of reactants and products
Closed system *[2 max]*

(ii) $K_c = \dfrac{[CH_3CO_2C_2H_5][H_2O]}{[CH_3CO_2H][C_2H_5OH]}$ *[1]* *[1]*

(iii) $[C_2H_5OH] = [CH_3CO_2H] = 0.33$ *[1]*

$K_c = 4.1$ (accept 4.0) *[1]* *[2 max]*

(Allow for ECF from (ii))

(b) K_w increases with temperature rise / $[H^+]$ or $[OH^-]$ rises as temperature rises *[1]*
Equilibrium shifts to right (as temperature increases) *[1]* *[3 max]*
Endothermic reaction *[1]*

(c) (i) No effect *[1]*
Some description of role of catalyst *e.g.* increases rate of forward **and** reverse reaction (equally) / increases rate of attainment of equilibrium *[1]* *[2 max]*

(ii) No effect *[1]*
Concentration of reactants the same *[1]* *[2 max]*

HL SECTION A 00w

4b

(ii) The methanol concentration will increase. *[1]*

(c) (i) $K_c = \dfrac{[CH_3OH]}{[CO][H_2]^2}\ \text{mol}^{-2}\ \text{dm}^6$ *[1]*

(Need units to score mark.)

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

n:	1.00	2.00	–	
neq:	$1.00 - 0.85$	$2.00 - 1.70$	0.85	*[1]*
[]eq:	$\dfrac{0.15}{0.45}$	$\dfrac{0.30}{0.45}$	$\dfrac{0.85}{0.45}$	*[1]*
	$= 0.333$	$= 0.667$	$= 1.889$	

$K_c = \dfrac{1.889}{0.333 \times 0.667^2} = 12.8\ (\text{mol}^{-2}\ \text{dm}^6)$ *[1]*

(iii) Side reactions **OR** leaks in the system **OR** not operating under equilibrium conditions **OR** operating at a **higher** temperature **OR** the product might be collected before equilibrium is reached. *[1]*

(iv) No effect on K_c (it just speeds up the reaction). *[1]*

HL SECTION A 00s

1. (a) (i) Endothermic/heat absorbed/energy absorbed / increase in enthalpy / needs (a lot of) energy *[1]*

(ii) K_c is decreased / OWTTE. *[1]*
Since heat is removed / since equilibrium moves to left / reverse reaction favoured. *[1]*

(iii)
$$\begin{array}{ccccc} N_2 & + & O_2 & \rightleftharpoons & 2NO \\ (1.6-x) & & (1.6-x) & & 2x \end{array}$$ *[1]*

$K_c = \dfrac{[NO]^2}{[N_2][O_2]}$ *[1]*

$[NO] = 0.065\ \text{mol dm}^{-3}$ *[1]*
(If candidate uses 1.6 *instead of* $(1.6 - x)$*, which gives an answer of* 0.066*, award* ***[2]*** *unless it is specifically stated that an approximation has been made or* $1.6 \gg x$*.)*

HL SECTION A .99s

3. (a) $K_c = \frac{[NO_2]^2}{[N_2O_4]}$ insist on [ ] ***[1 mark]***

(b) (i) $[NO_2]$ increases or yield increases ***[1 mark]***
since $[N_2O_4]$ increases and K_c is constant ***[1 mark]***
OR equilibrium moves to the right

(ii) $[NO_2]$ decreases or yield decreases ***[1 mark]***
since increased pressure pushes equilibrium to the left ***[1 mark]***
OR by Le Chatelier's Principle the smaller volume is favoured

(iii) $[NO_2]$ unchanged or yield unchanged ***[1 mark]***
catalyst does not affect position of equilibrium ***[1 mark]***
OR both forward and backward reaction rates affected equally

(c) (i) $N_2O_4 \rightleftharpoons 2NO_2$ ***[1 mark]***
0.8 0.4

(ii) $\frac{(0.4)^2}{0.8} = \underset{[1]}{0.2}\ \underset{[1]}{\text{mol dm}^{-3}}$ ***[2 marks]***

Note: Carry forward error from (a) in both numerical answer and units
Error carried forward in (c) (ii), *e.g.* if $[NO_2]$ is given as 0.2 then $K_c = 0.05$

(d) K_c increased ***[1 mark]***
since forward reaction is endothermic ***[1 mark]***
OR equilibrium moves to right as temperature increases

Total [14 marks]

HL B 13s

8

(c) (i) rate of forward reaction equals rate of backward reaction;
concentrations of reactants and products do not change / constant macroscopic properties; ***[2]***

(ii) $K_c = \frac{[CH_3OH]}{[CO][H_2]^2}$; ***[1]***

Do not award mark if incorrect brackets are used or brackets omitted.

(iii) $n(CO) = 0.508\,(mol)$;
$n(H_2) = 2 \times 0.508\,(mol)$;
$K_c \left(= \frac{0.492}{0.508 \times (2 \times 0.508)^2} \right) = 0.938$; ***[3]***

Accept answer in range between 0.930 and 0.940.
Award [3] for correct final answer.
Award [2] for $K_c = 1.066$ if (c)(ii) is correct.

(d) (i) shifts to left/reactants;
to endothermic side / (forward) reaction is exothermic; ***[2]***

(ii) shifts to the right/products;
to the side with fewer gas molecules/moles of gas; ***[2]***

HL B 12w

5

(vi) (K_c) decreases;
exothermic / $\Delta H < 0$; ***[2]***
M1 and M2 depend on sign of ΔH in (iii).

HL B 12s

6. (a) $Br_2(g) \rightleftharpoons Br_2(l)$;
$H_2(g) + Br_2(g) \rightleftharpoons 2HBr(g)$; ***[2]***

(b) (i) increase volume of liquid / no change of colour of vapour; ***[1]***

(ii) shifts to right/toward products/forward reaction favoured; ***[1]***
Accept reverse statement if process written the other way around.
Answer must match stated equation.

(iii) no effect;
same amounts/number of (gaseous) moles/molecules on both sides; ***[2]***

(c) (i) $(K_c =) \frac{[HBr]^2}{[H_2][Br_2]}$; ***[1]***

(ii) no effect (only depends on the temperature); ***[1]***

6e

(iii) $K_c < 1$; ***[1]***

HL B 11w

8

(d) (i) change in concentration of reactant/product with time / rate of change of concentration; *[1]*

Increase can be used instead of change for product or decrease can be used instead of change for reactant.
Allow mass/amount/volume instead of concentration.
Do not accept substance.

(ii) pressure is lower/moderate **and** temperature is higher in Haber process / ~ 200 atm (pressure) **and** ~ 700 K (temperature) used in Haber process;

Pressure:
high pressure shifts equilibrium to right;
high pressure (faster rate but) expensive/dangerous / greater capital and running costs;

Temperature:
low temperature shifts equilibrium (even further) to right;
low temperature gives slower rate (but high yield);

high pressure increases yield **and** lower temperature decreases rate;
Accept converse argument.

(not possible to have high yield and fast rate simultaneously therefore) compromise needed / *OWTTE*; ***[4 max]***

HL B 11s

7

(c) (i) *[CH_3CH_2COOH]:*
(1.6 – 0.80 =) 0.8 (mol dm^{-3});
[CH_3OH]:
(2.0 – 0.80 =) 1.2 (mol dm^{-3});
[H_2O]:
0.80 (mol dm^{-3}); *[3]*

(ii) $(K_c =)\dfrac{[CH_3CH_2COOCH_3]\ [H_2O]}{[CH_3CH_2COOH]\ [CH_3OH]}$;

$(K_c = \dfrac{[(0.80)^2]}{[(1.2 \times 0.8)]} =)0.7$; *[2]*

Allow 0.67.
*Award **[1 max]** for 0.83.*

HL B 10w

6. (a) (i) exothermic;

Accept either of the following for the second mark.
increasing temperature favours endothermic/reverse reaction;
as yield decreases with increasing temperature; *[2 max]*

(ii) yield increases / equilibrium moves to the right / more ammonia;
increase in pressure favours the reaction which has fewer moles of gaseous products; *[2]*

(iii) $(K_c =)\dfrac{[NH_3]^2}{[N_2][H_2]^3}$; *[1]*

(iv) $[N_2]$: (at equilibrium $= 1.00 - 0.031 =$) 0.969 (mol dm^{-3});
$[H_2]$: (at equilibrium $= 3.00 - 3(0.031) =$) 2.91 (mol dm^{-3});
$\left(K_c = \dfrac{(0.062)^2}{(0.969)(2.91)^3}\right) = 1.6(1)\times10^{-4}$; *[3]*
Ignore units.
Award [1] for $K_c = 1.4\times10^{-4}$

(v) no effect; *[1]*

HL B 10s

5

(b) macroscopic properties remain constant / concentrations remain constant / no change to copper solution seen;
rate of reverse/backwards reaction = rate of forward reaction; *[2]*

(c) $(K_c=)\ \dfrac{[NH_3]^2}{[N_2][H_2]^3}$; *[1]*
Do not award mark if [ ] missing or round brackets used.

(d) (i) $[H_2] = 0.11 / 0.11$ (mol dm^{-3});
$[N_2] = 0.17 / 0.17$ (mol dm^{-3});
$K_c = 16$; *[3]*
Ignore units.
Allow ECF from incorrect equilibrium expression and incorrect concentrations for third mark.

(ii) decrease;
heat is a product/reaction is exothermic so equilibrium moves to left / *OWTTE*; *[2]*

(e) yield increases / equilibrium moves to the right / more ammonia;
4 gas molecules → 2 / decrease in volume / fewer gas molecules on right hand side; *[2]*

(f) high pressure expensive / greater cost of operating at high pressure / reinforced pipes *etc*. needed;
Do not accept "high pressure is dangerous" without further explanation.

lower temperature – greater yield, but lowers rate; *[2]*
Do not award a mark just for the word "compromise".

(g) K_c unaffected;
position of equilibrium unaffected;
rate of forward and reverse reactions are increased (equally); *[3]*

HL B 09w

6

(e) (i) $(K_c) = \frac{[SO_2Cl_2]}{[Cl_2][SO_2]}$; *[1]*
Ignore state symbols.
Square brackets [ ] required for the equilibrium expression.

(ii) 7.84×10^{-3} mol of SO_2 and 7.84×10^{-3} mol of Cl_2;
7.84×10^{-3} mol dm^{-3} of SO_2, 7.84×10^{-3} mol dm^{-3} of Cl_2 and
7.65×10^{-4} mol dm^{-3} of SO_2Cl_2;
12.5; *[3]*
Award [1] for 10.34
Award [3] for the correct final answer

(iii) value of K_c increases;
$[SO_2Cl_2]$ increases;
decrease in temperature favours (forward) reaction which is exothermic; *[3]*
Do not allow ECF.

(iv) no effect on the value of K_c / depends only on temperature;
$[SO_2Cl_2]$ decreases;
increase in volume favours the reverse reaction which has more gaseous moles; *[3]*
Do not allow ECF.

(v) no effect;
catalyst increases the rate of forward and reverse reactions (equally) / catalyst decreases activation energies (equally); *[2]*

HL B 08w

6. (a) (oxidation is) loss of electron(s);
Br$^-$ / bromide (ion); *[2]*

(b) Ti +4 and +4;
Cl 0 and −1;
C 0 and +4; *[3]*
Penalize missing +, or answers written as 4+ once only.
If no marks scored allow [1] if all oxidation numbers for reactants or for products are correct.

(c) (i) Cr oxidation number +6/same on both sides/does not change;
neither oxidation nor reduction occur; *[2]*
No ECF

(ii) Cl oxidation number 0 on left and −1 and +1 on right;
both oxidation and reduction occur / disproportionation; *[2]*
No ECF

(d) (i) $W > Z > Y > X$;
Award [1] mark for correct order.

X below Y because of reaction 2/because X will not displace Y;
X below Z because of reaction 4/because X displaced by Z;
X below W because of reaction 1/because Z displaced by W **and** Z displaces X;
Y below Z because of reaction 3/because Y will not displace Z; *[4 max]*
Give credit for OWTTE in this part.
Any three of last four score [1] each.

(ii) $2V + 3XO \rightarrow V_2O_3 + 3X$; *[1]*

HL B 08s

7

(c) (i) (equilibrium shifted to the left) equilibrium concentration of NO/[NO] is reduced;
more gas molecules on the right hand side than on the left; *[2]*

(ii) (equilibrium shifted to the left) equilibrium concentration of NO/[NO] is reduced;
forward reaction is exothermic; *[2]*

(iii) no effect on the equilibrium concentration of NO/[NO];
catalyst increases the rate of forward and reverse reaction equally; *[2]*

(d) $K_c = \dfrac{[NO]^4[H_2O]^6}{[NH_3]^4[O_2]^5}$;
mol dm^{-3}; *[2]*

(e) increase in temperature;
K_c decreases; *[2]*

HL B 07w

8

(b) (i) $CH_4(g) + 2H_2O(g) \rightleftharpoons 4H_2(g) + CO_2(g)$;

States not required. Award [1] for balanced equation and [1] for equilibrium sign.

$$K_c = \frac{[H_2]^4[CO_2]}{[CH_4][H_2O]^2};$$

ECF

units: $mol^2\ dm^{-6}$ / $mol^2\ L^{-2}$ / $mol^2\ l^{-2}$; *do not accept: M^2* *[4]*

ECF

(ii) (endothermic reaction) increase in temperature (favours the forward reaction);
absorbs (some of) the heat supplied / *OWTTE*;
Award no marks for saying: "because of Le Chatelier's principle".
low pressure (will allow system to occupy more volume);
$V_{product} > V_{reactant}$ / reaction proceeds to greater number of gaseous moles / molecules / more moles of gases on right / *OWTTE*;
ECF from (b) (i) *[4]*

(iii) at high pressure concentration increases / reaction rate faster;
more frequent collisions; *[2]*

(iv) $(CH_4(g) + 2H_2O(g) \rightleftharpoons 4H_2(g) + CO_2(g))$

(2.0 mol H_2 produced at equilibrium) 0.50 mol CO_2 formed;

(0.50 mol CH_4 reacted) $[CH_4]_{eq} = 1.00 - 0.50 = 0.50$ mol;

ECF

(1.0 mol H_2O reacted) $[H_2O]eq = 3.0 - 1.0 = 2.0$ mol;

ECF

$$K_c = \frac{[H_2]^4[CO_2]}{[CH_4][H_2O]^2} = \frac{[2.0]^4[0.50]}{[0.50][2.0]^2} = 4.0;$$ *[4 max]*

ECF

HL B 07s

6. (a) (i) shifts to the right/toward products / forward reaction favoured;
to consume excess Br^- added; *[2]*
Do not accept "due to Le Chatelier's principle".

(ii) shifts to the left/toward reactants / reverse reaction favoured;
NaOH reacts to consume H^+ / an increase in the amount of H_2O resulting from neutralization; *[2]*
Do not accept "due to Le Chatelier's principle".

(iii) no effect;
catalyst increases the rate of the forward and backward reactions equally / lowers the activation energy of both forward and backward reaction equally / lowers E_A so rate of forward and backward reactions increase equally; *[2]*

(b) equilibrium constant decreases;
forward reaction is exothermic/produces heat / reverse reaction is endothermic/absorbs heat; *[2]*

(c) colour change from red-brown to darker red-brown of Br_2 / red-brown colour intensifies / OWTTE;
equilibrium position shifts to the right/products;
to consume H^+; *[3]*

HL B 06w

8. (a) $K_c = \dfrac{[H_2]^3[CO]}{[CH_4][H_2O]}$; *[1]*

(b) (i) shifts position of equilibrium to the right;
increases value of K_c; *[2]*

(ii) shifts position of equilibrium to the left;
does not change value of K_c; *[2]*

(c) (i) $CO = 0.4$ (mol);
$H_2O = 1.6$ (mol);
$K_c\ (= 1.6^2 \div 0.4 \times 1.6) = 4.0 / 4$; *[3]*
Apply ECF from K_c expression.
Ignore units.

(ii) H_2 and CO_2 / products $= 1.33 / 1.3$ (mol);
CO and H_2O / reactants $= 0.67 / 0.7$ (mol); *[2]*
Using $K_c = 9.0$, values for H_2 and CO_2 are 1.5 and values for CO and H_2O are 0.5.

(d) no effect on position of equilibrium;
forward and reverse reactions speeded up equally / affects the rate of reaction but not the extent of the reaction;
no effect on value of K_c;
no change in concentrations of reactants or products / K_c only changes if temperature alters; *[4]*

HL B 06s

6. (a) $K / K_c = [SO_3]^2 \div [SO_2]^2[O_2]$; *[1]*

Exactly as written.

Accept correct K_p expression.

(b) (i) vanadium(V) oxide / (di)vanadium pentaoxide / V_2O_5 /Pt; *[1]*

Allow just vanadium oxide but not incorrect formula.

(ii) catalyst does not affect the value of K_c;

forward and reverse rate increase equally/by the same factor;

catalyst increases the rate of the reaction;

(by providing an alternative path for the reaction with) lower activation energy; *[4]*

(c) more energetic collisions / more molecules have energy greater than activation energy;

more frequent collisions; *[2]*

Do not accept more collisions without reference to time.

(d) (i) shifts equilibrium position to the products/right;

to the side with least gas molecules or moles / lower volume of gas; *[2]*

(ii) shifts equilibrium position to the products/right;

to compensate for loss of SO_3 / produce more SO_3; *[2]*

(e) exothermic;

K_c decreases with increasing temperature / back reaction favoured / heat used up / *OWTTE*; *[2]*

(f) $n(SO_2)_{\text{at equilibrium}} = 1.50 - 0.50 = 1.00$ mol;

$n(O_2)_{\text{at equilibrium}} = 2.00 - 0.250 = 1.75$ mol;

$[SO_2] = 1.00 \div 1.50 = 0.667\ \text{mol dm}^{-3}$, $[O_2] = 1.75 \div 1.50 = 1.17\ \text{mol dm}^{-3}$

$[SO_3] = 0.500 \div 1.50 = 0.333\ \text{mol dm}^{-3}$;

$K_c = (0.333)^2 \div 1.17 \times (0.667)^2$;

$= 0.213\ \text{dm}^3\ \text{mol}^{-1} / 0.214\ \text{dm}^3\ \text{mol}^{-1}$; *[5]*

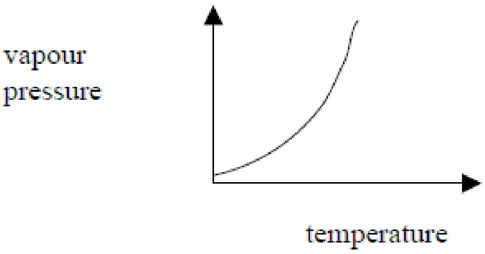
Allow ECF.

*If 0.202 dm^3 mol^{-1} is given award **[4]**, this is obtained by premature rounding.*

*Award **[5]** for correct answer with units.*

(g) (i) the greater the strength of the intermolecular forces the greater the enthalpy of vaporization/*OWTTE*;
pentane has only van der Waals' forces between molecules;
propanoic acid has H-bonding (as well as van der Waals' forces); *[3]*

(ii)



;

1st mark: graph goes upwards with T;
2nd mark: curve as shown;
as temperature increases (more) molecules have enough energy to overcome intermolecular / attractive forces; *[3]*

HL B 05w

6. (a) (i) no effect;
equal <u>gas</u> moles on each side; *[2]*

(ii) shift to right;
forward reaction absorbs heat/endothermic/*OWTTE*; *[2]*

(iii) no effect;
catalyst speeds up both forward and reverse reactions <u>equally</u>; *[2]*

(b) $K_c = \frac{[H_2][I_2]}{[HI]^2}$; *[1]*

Ignore state symbols.

(c) (i) *experiment 1* $[HI] = 0.04\ (\text{mol dm}^{-3})$;

$[I_2] = 0.01\ (\text{mol dm}^{-3})$;

$$K_c = \frac{(0.01)^2}{(0.04)^2} = 6.25 \times 10^{-2};$$

ECF from above values.

experiment 2 $[H_2] = 0.02\ (\text{mol dm}^{-3})$;

$[I_2] = 0.02\ (\text{mol dm}^{-3})$;

$$K_c = \frac{0.02^2}{0.04^2} = 0.25;$$ *[6]*

ECF from above values.

(ii) experiment 2 (at higher temperature);
higher K_c value/equilibrium shifted to right; *[2]*

HL B 03w

8

(b) Time

two curves - one labled "forward" starting up high up y-axis and one labeled "reverse" starting from zero;
curves merge and become horizontal;
No penalty for failing to label axes.

forward reaction:
highest concentration, thus rate high to begin with;
as reaction proceeds, concentrations decrease, so does rate;

reverse reaction:
zero rate initially / at t = 0 (since no products present);
rate increases as concentration of products increases;

equilibrium established when rate of forward reaction = rate of reverse reaction; *[7]*

(c) (reaction is) endothermic;
K_c increases with (increasing) temperature;
forward reaction favoured / heat used up / OWTTE; *[3]*

HL B 02w

7. (a) $2HI(g) \rightleftharpoons H_2(g) + I_2(g)$ ***[2]***

*Award **[1]** for correct equation with $\rightleftharpoons$ and **[1]** for correct state symbols. No mark for reversed equation, but can still achieve **[1]** for correct state symbols.*

(b) (i) $K_c = \frac{[H_2][I_2]}{[HI]^2}$;

no units;
ECF for reverse equation ***[2]***

(ii) number of moles HI reacted $= (1 - 0.78) = 0.22$;
therefore number of moles H_2/I_2 formed $= 0.11$;

$K_c = \frac{0.11 \times 0.11}{(0.78)^2}$;

$K_c = 0.02$; *(ECF)* ***[4]***

(c) endothermic; (*ECF - if* $K_c = 0.080$, *then exothermic*)
K_c increases with increasing temperature; *(ECF)*
favours products / equilibrium shifts to the right; *(ECF)* ***[3]***

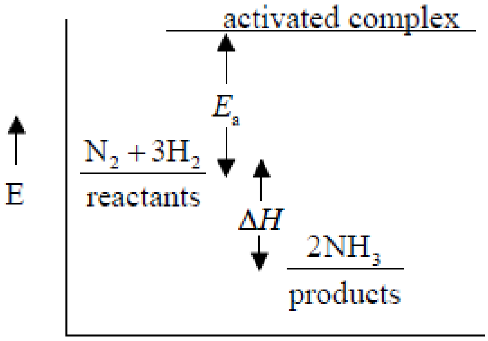
(d) (i) no change in equilibrium position;
same number of moles of **gas** on both sides of the equation / *OWTTE*; ***[2]***

(ii) equilibrium position moves from left to right;
system tries to remove added HI (according to LeChatelier) / *OWTTE*; ***[2]***

(iii) no change in position of equilibrium;
catalyst increases both forward and backward reaction rate(equally) / *OWTTE*; ***[2]***

(e) both HCl(g) and HBr(g) are more weakly dissociated; ***[1]***

(f) (i)



reactants / $N_2 + 3H_2$ higher than products / $2NH_3$;
ΔH / enthalpy change as the difference between reactants and products;
E_a / activation energy as the difference between reactants and intermediate / activated complex / transition state; *[3]*

(ii) lowers E_a;
provides (lower energy) alternative route / allows for more successful collisions; *[2]*

(iii) catalyst is in a different phase / state from the reactants;
reactants are held / adsorbed / held on surface / in sites / active sites (on catalyst); *[2]*

HL B 01w

8. (a) $\frac{[H_2(g)][CO(g)]}{[H_2O(g)]}$ *(state symbols not needed)* *[1]*

$mol\,dm^{-3}$ *(ECF if K_c expression incorrect)* *[1]*

Endothermic. *[1]*
The value of K_c increases with temperature (forward reaction is favoured). *[1]*

(b) (i) The rate of the forward reaction:

Increase in temperature:
increases rate of the forward reaction *[1]*
since it increases the number of collisions with $E \geq E_a$ *[1]*
(thus rate of forward and reverse reaction increases).

*(Award **[0]** for "more frequent collisions" as this is not the main reason.)*

Increase in pressure:
Increases the rate of the forward reaction. *[1]*
An increase in pressure increases the concentration / the same amount of molecules in a smaller volume (thus rate increases) / more frequent collisions. *[1]*

Increase in surface area of C(s):
Increases rate of forward reaction. *[1]*
The reactive surface area of C(s) increases / concentration of $H_2O(g)$ in contact with C(s) increases (thus rate increases) / more frequent collisions. *[1]*

(ii) $[H_2O]:[H_2]$ ratio:

Increase in T: decreases ratio / more $[H_2]$ compared to $[H_2O]$. *[1]*
Favours forward / endothermic reaction by using up (some of the) heat supplied *(no mark for saying "because of Le Chatelier's principle").* *[1]*

Increase in P:
Increases the ratio / more $[H_2O]$ compared to $[H_2]$. *[1]*
1 mol / volume of gas in the reactants, 2 mol / volume of gases in the products / increasing pressure will move position of equilibrium to the left to relieve pressure due to more volumes of gas (in the products). *[1]*

Increase in surface area of C(s):
No effect. *[1]*
Because [C(s)] is constant / position of equilibrium is unchanged. *[1]*

(iii) The value of the equilibrium constant:

Increase in T: increases value of K_c. *[1]*

Increase in P: no effect on K_c (because K_c is only temperature dependent). *[1]*

Increase in surface area: no effect (K_c is only temperature dependent). *[1]*

HL B .99w

Mark scheme unavailable