C6bD Reversible Reactions and Equilibrium: Exam Drill

AQA 8464 · Calculator allowed · about 90 minutes

Total Marks

Name: _____ Date: ____ / ____ / _____

Answer ALL questions. Show all your working.

1 Define a reversible reaction and give one example using chemical formulae.

(a) Give a concise definition of a reversible reaction. **(1)**

(Total for Question 1 is 1 mark)

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2 Consider the reversible reaction: N2(g) + 3H2(g) <-> 2NH3(g). State what is meant by dynamic equilibrium in this closed system.

(a) State one feature of a dynamic equilibrium. **(1)**

(b) Explain what happens to the concentrations of reactants and products at dynamic equilibrium in this closed system. **(1)**

(Total for Question 2 is 2 marks)

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3 Le Chatelier's principle predicts how equilibrium shifts when conditions change. For the equilibrium: CO(g) + 2H2(g) <-> CH3OH(g), predict and explain the effect of each change.

(a) Predict what happens to the position of equilibrium if the pressure is increased, and explain your answer. **(1)**

(b) Predict the effect of adding more CO gas and give the reason. **(1)**

(c) Predict the effect of increasing temperature and give the reason, given that the formation of CH3OH is exothermic. **(1)**

(Total for Question 3 is 3 marks)

4 A student investigates the effect of a catalyst on an equilibrium. The reaction is A(g) <-> B(g). The student measures the time to reach equilibrium without and with a catalyst. Explain what changes the catalyst causes to the equilibrium position, the equilibrium constant, and the rates of the forward and backward reactions.

(a) State the effect of adding a catalyst on the position of equilibrium and the equilibrium constant Kc. **(1)**

(b) Explain how a catalyst affects the rates of the forward and backward reactions and why equilibrium is reached faster. **(2)**

(Total for Question 4 is 3 marks)

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5 Required practical style: a student investigates the equilibrium between iron(III) thiocyanate and its ions: Fe3+ + SCN- <-> FeSCN2+. The student mixes 10.0 cm3 of 0.100 mol dm^{-3} Fe3+ with 10.0 cm3 of 0.0100 mol dm^{-3} SCN- and allows equilibrium to form. The equilibrium concentration of FeSCN2+ is determined spectrophotometrically as 8.00 x 10^{-4} mol dm^{-3} in the final mixture.

(a) Calculate the initial concentrations of Fe3+ and SCN- in the mixture immediately after mixing but before reaction proceeds. Give your working and units. **(2)**

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(b) The student then adds a few more drops of concentrated Fe3+ solution to the equilibrium mixture and observes the red colour become darker. Explain, using Le Chatelier's principle, why the colour becomes darker and what this shows about the position of equilibrium. **(2)**

(Total for Question 5 is 4 marks)

6 The gases nitrogen dioxide (NO_2, brown) and dinitrogen tetroxide (N_2O_4, colourless) exist in equilibrium: $2NO_2(g) \rightleftharpoons N_2O_4(g)$. The forward reaction is exothermic. Predict and explain what would be observed if the pressure on the equilibrium mixture is increased at constant temperature.

(a) Explain, using Le Chatelier's principle and the number of gas moles on each side, what happens to the position of equilibrium and the colour of the mixture. **(3)**

(Total for Question 6 is 3 marks)

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7 Error analysis: a student uses the equilibrium reaction between hydrated cobalt(II) ions (pink) and chloride ions to form a blue complex ion, where the forward reaction (forming the blue complex) is endothermic: $[Co(H_2O)_6]^{2+}$ (pink) + $4Cl^-(aq) \rightleftharpoons [CoCl_4]^{2-}$ (blue) + $6H_2O(l)$. The student adds concentrated hydrochloric acid to a pink solution and heats it gently, expecting the colour change to show the effect of temperature alone. However, the flask is left open to the air, so some water also evaporates from the mixture as it is heated. Explain how (a) the increase in temperature and (b) the loss of water by evaporation each affect the position of this equilibrium, and state whether the two effects act in the same direction or opposite directions.

(a) Explain qualitatively how each effect (temperature increase and water loss) shifts the position of equilibrium, and whether they act in the same or opposite directions. **(4)**

(Total for Question 7 is 4 marks)

8 Hydrated copper sulfate is involved in a reversible reaction: CuSO4.5H2O(s) [blue] <-> CuSO4(s) [white] + 5H2O(l), where the forward reaction (dehydration) is endothermic. A student heats blue hydrated copper sulfate crystals in an open test tube until they turn white, then adds a few drops of water to the white solid. Explain, in terms of the reversible reaction and Le Chatelier's principle, why (a) heating causes the crystals to turn white and (b) adding water afterwards turns the solid blue again. State any assumption about whether this counts as a true equilibrium in the open test tube.

(a) Explain the effect of heating and of adding water on the position of the equilibrium, and state the assumption you make about the open test tube. **(4)**

(Total for Question 8 is 4 marks)

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9 Hydrogen and chlorine react reversibly to form hydrogen chloride gas: H2(g) + Cl2(g) <-> 2HCl(g). A student says that increasing the pressure on this equilibrium mixture will increase the yield of HCl. Explain why the student is incorrect.

(a) Compare the number of gas moles on each side of the equation and explain the effect of pressure on this equilibrium. **(3)**

(Total for Question 9 is 3 marks)

10 A reversible reaction A(g) + B(g) <-> C(g) is set up in a sealed container at a fixed temperature. A chemist starts with 0.400 mol of A and 0.600 mol of B, and no C. At equilibrium, 0.100 mol of C has been formed. Calculate the number of moles of A and B remaining at equilibrium, then calculate the percentage of the original A that has reacted. Finally, state and explain whether increasing the pressure on this system would shift the position of equilibrium.

(a) Show each step: use the 1:1:1 reaction stoichiometry to find the moles of A and B remaining, calculate the percentage of A reacted, then explain the effect of increasing pressure. **(5)**

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(Total for Question 10 is 5 marks)

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11 Extended response (6 marks): Ethanoic acid and ethanol react to form ethyl ethanoate and water in a reversible esterification: CH3COOH(aq) + C2H5OH(aq) <-> CH3COOC2H5(aq) + H2O(l). A student investigates the effect of removing water on the yield of ester at constant temperature. Describe and explain what happens to the equilibrium position and to the value of Kc when water is removed, and suggest one practical way the student could remove water during the experiment. Include consideration of whether the change is temporary or permanent and any experimental limitations.

(Total for Question 11 is 6 marks)

12 Stretch question (10 marks): Consider the industrial Haber process for ammonia synthesis: $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$. An industrial chemist must choose operating conditions that balance yield, rate and cost. Discuss how each of the following factors affects the yield and rate of the reaction and the practical considerations in an industrial setting: temperature, pressure, and the use of a catalyst. Conclude with a justified choice of operating temperature and pressure for an industrial plant, recognising trade-offs. You should include references to Le Chatelier's principle, collision theory, and economic or safety constraints.

(Total for Question 12 is 10 marks)

Mark scheme · C6bD Reversible Reactions and Equilibrium: Exam Drill

Question 1

- (a) **B1** a reaction that can proceed in both the forward and backward directions so products can change back into reactants, oe
- **(a) Answer: A reaction that can proceed in both directions so products can change back into reactants.**

Question 2

- (a) **B1** the rates of the forward and backward reactions are equal, oe
- **(a) Answer: The rates of the forward and backward reactions are equal.**
- (b) **B1** their concentrations remain constant over time even though both forward and backward reactions continue to occur, oe
- **(b) Answer: Concentrations of reactants and products remain constant over time while both reactions continue.**

Question 3

- (a) **B1** shifts to the right (towards CH3OH) because there are fewer gas moles on the product side (1) than on the reactant side (3), oe
- **(a) Answer: Shifts to the right towards CH3OH because there are fewer gas moles on the product side.**
- (b) **B1** shifts to the right (towards products) because adding a reactant is opposed by the system producing more product, oe
- **(b) Answer: Shifts to the right to make more CH3OH because the system counteracts the added CO.**
- (c) **B1** shifts to the left (towards reactants) because increasing temperature favours the endothermic direction, oe
- **(c) Answer: Shifts to the left towards CO and H2 because the forward reaction is exothermic so higher temperature favours the endothermic reverse.**

Question 4

- (a) **B1** no change to the position of equilibrium or to the equilibrium constant Kc, oe
- **(a) Answer: No change to the equilibrium position and no change to Kc.**
- (b) **M1** states catalyst provides an alternative lower-energy pathway so increases the rate(s)
- (b) **A1** both forward and backward reaction rates increase so equilibrium is reached faster, oe
- **(b) Answer: A catalyst provides an alternative lower-energy pathway, increasing both forward and backward rates so equilibrium is reached faster.**

Question 5

- (a) **M1** calculates dilution: initial moles or concentration after mixing for each species, e.g. Cfinal = Cinitial x (Vinitial / Vtotal) or Cfinal = (C x 10.0) / 20.0, oe
- (a) **A1** Fe3+: 0.0500 mol dm^{-3} and SCN-: 0.00500 mol dm^{-3} cao
- **(a) Answer: Fe3+: 0.0500 mol dm^{-3}; SCN-: 0.00500 mol dm^{-3}**
- (b) **M1** identifies that adding more Fe3+ (a reactant) disturbs the equilibrium, oe
- (b) **A1** the equilibrium shifts to the right (towards FeSCN2+) to oppose the increase and use up some of the added Fe3+, producing more of the red FeSCN2+ complex, which darkens the colour, oe
- **(b) Answer: Adding more Fe3+ disturbs the equilibrium; by Le Chatelier's principle, the equilibrium shifts to the right to oppose the change, producing more red FeSCN2+ and darkening the colour.**

Question 6

- (a) **B1** there are 2 moles of gas on the reactant side (NO2) but only 1 mole of gas on the product side (N2O4), oe
- (a) **B1** by Le Chatelier's principle, increasing pressure shifts the equilibrium towards the side with fewer gas moles, i.e. towards N2O4, oe
- (a) **B1** so more N2O4 forms and the mixture becomes paler (less brown), oe
- **(a) Answer: Increasing pressure shifts equilibrium towards N2O4 (the side with fewer gas moles), so more colourless N2O4 forms and the brown colour fades.**

Question 7

- (a) **M1** increasing temperature favours the endothermic direction (forward, formation of the blue complex) so equilibrium shifts right, towards blue, oe
- (a) **M1** loss of water reduces the concentration of water (a product), so by Le Chatelier's principle the equilibrium shifts to the right, towards blue, to replace the water removed, oe
- (a) **A1** both effects act in the same direction (both shift equilibrium towards the blue complex), so the solution turns a stronger blue than temperature alone would cause, oe
- (a) **A1** concludes that the evaporation of water is an uncontrolled variable that adds to the intended effect of temperature, making it hard to know how much of the colour change is due to heating alone, oe
- **(a) Answer: Increasing temperature shifts equilibrium right (towards the blue complex, the endothermic direction). Loss of water (a product) also shifts equilibrium right, replacing the water lost. Both effects act in the same direction, so the uncontrolled evaporation adds to the intended temperature effect, making the colour change larger than temperature alone would explain.**

Question 8

- (a) **M1** heating is an endothermic input of energy, and the forward (dehydration) reaction is endothermic, so heating favours the forward reaction, oe
- (a) **M1** water vapour escapes from the open tube, removing a product, which by Le Chatelier's principle also shifts the reaction further to the right, towards white anhydrous CuSO4, oe
- (a) **M1** adding water increases the concentration of a product (water), so by Le Chatelier's principle the equilibrium shifts back to the left, towards blue hydrated CuSO4, oe
- (a) **A1** notes that because the tube is open, water vapour can escape rather than being retained, so this is not a true closed-system equilibrium; in a sealed container both reactions would occur simultaneously to reach a dynamic equilibrium, oe
- **(a) Answer: Heating supplies energy to the endothermic forward (dehydration) reaction, and water vapour escaping the open tube removes a product, both driving the reaction to the white anhydrous form. Adding water increases the concentration of a product, shifting the equilibrium back to the blue hydrated form. Because the tube is open, this is not a true closed-system equilibrium.**

Question 9

- (a) **B1** there are 2 moles of gas on the reactant side (H2 + Cl2) and 2 moles of gas on the product side (2HCl), oe
- (a) **B1** the number of gas moles is equal on both sides of the equation, oe
- (a) **B1** so changing the pressure does not favour either side and does not shift the position of equilibrium or change the yield of HCl, oe
- **(a) Answer: There are equal numbers of gas moles on both sides of the equation (2 on each side), so changing the pressure does not shift the equilibrium; the student is incorrect because the yield of HCl is unaffected by pressure.**

Question 10

- (a) **M1** uses 1:1:1 stoichiometry to deduce that 0.100 mol of A and 0.100 mol of B react to form 0.100 mol of C
- (a) **A1** moles remaining at equilibrium: A = 0.400 - 0.100 = 0.300 mol; B = 0.600 - 0.100 = 0.500 mol, cao
- (a) **M1** percentage of A reacted = (amount reacted / initial amount) x 100 = (0.100 / 0.400) x 100
- (a) **A1** = 25.0% cao
- (a) **B1** notes there are 2 mol of gas on the reactant side and 1 mol on the product side, so increasing pressure would shift the equilibrium further towards C, increasing the yield of C, oe
- **(a) Answer: At equilibrium: A = 0.300 mol, B = 0.500 mol. Percentage of A reacted = (0.100 / 0.400) x 100 = 25.0%. There are 2 mol of gas reactants but only 1 mol of gas product, so increasing pressure would shift the equilibrium further towards C, increasing the yield.**

Question 11

- **Level 3** (5-6): Detailed description and explanation: clearly explains that removing water shifts equilibrium to the right to form more ester, links to Le Chatelier and states Kc is unchanged, gives a suitable practical method for removing water and discusses temporary or permanent nature and limitations.
- **Level 2** (3-4): Partial explanation: states equilibrium shifts towards ester and that this increases yield, states Kc unchanged but with limited justification, and suggests a practical method without discussing limitations.
- **Level 1** (1-2): Simple statement: e.g. removing water increases ester yield, or suggests a method, with little or no explanation.
- **Indicative content:**
  - Removing water shifts equilibrium to the right (towards products) according to Le Chatelier, increasing the yield of ethyl ethanoate.
  - Kc is a constant at fixed temperature and does not change because only concentrations are changed; removing water changes the position of equilibrium but not Kc.
  - Practical methods: use a drying agent (e.g. anhydrous sodium sulfate) to remove water, use a Dean-Stark apparatus to continuously remove water under reflux, or distil off the water if boiling points allow.
  - Discuss temporary vs permanent: if water is removed continuously, the shift is maintained while removal continues; if drying agent becomes saturated the effect ends; removing water until completion may not drive reaction to 100% due to equilibrium constraints and side reactions.
  - Experimental limitations: drying agents may absorb some ester or cause side reactions; Dean-Stark requires suitable solvent and heating and may not remove all water; temperature must remain constant because Kc depends on temperature.

Question 12

- **M1** explains effect of increasing temperature: forward reaction is exothermic so increasing temperature shifts equilibrium left (reduces yield) but increases rate via collision theory, oe
- **M1** explains effect of increasing pressure: shifts equilibrium towards fewer gas moles (towards ammonia) increasing yield, and increases collision frequency so increases rate, oe
- **M1** states catalyst increases rate by providing an alternative pathway with lower activation energy and does not change equilibrium position or Kc, oe
- **M1** discusses economic/safety constraints: high pressure increases equipment cost and risk, high temperature reduces yield and may require compromise, oe
- **M1** links theory to industrial practice: justifies compromise of moderate-high pressure and moderate temperature with efficient catalyst to get acceptable rate and yield, oe
- **M1** gives a justified numerical or qualitative choice for operating conditions (e.g. around 150-250 bar and 400-500 degrees C in practice) and explains trade-offs between yield, rate and cost, oe
- **M1** evaluates benefit of recycling unreacted gases to improve overall yield and reduce required pressure or temperature, oe
- **M1** discusses catalyst lifetime and poisoning as practical constraints and mitigation (e.g. purification of feed gases), oe
- **M1** comments on energy requirements and environmental considerations (e.g. source of H2), oe
- **M1** provides a clear conclusion weighing the trade-offs and recommending operating range with justification, oe