2.17 Reversible Reactions and Equilibrium: Test Pack

AQA 8464 · Calculator allowed · about 35 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)

2 Consider the reversible reaction: $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$. 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: $CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$, 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 CH_3OH is exothermic. **(1)**

(Total for Question 3 is 3 marks)

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4 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 4 is 3 marks)

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5 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) ⇌ $[CoCl_4]^{2-}$ (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 the increase in temperature and 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 and what this means for the student's attempt to show the effect of temperature alone.

(Total for Question 5 is 4 marks)

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6 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)**

(Total for Question 6 is 5 marks)

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7 Required practical style: a student investigates the equilibrium between iron(III) thiocyanate and its ions: $Fe^{3+} + SCN^- \rightleftharpoons FeSCN^{2+}$. The student mixes 10.0 cm^3 of 0.100 mol dm^{-3} Fe^{3+} with 10.0 cm^3 of 0.0100 mol dm^{-3} SCN^- and allows equilibrium to form. The equilibrium concentration of $FeSCN^{2+}$ is determined spectrophotometrically as 8.00×10^{-4} mol dm^{-3} in the final mixture.

(a) Calculate the initial concentrations of Fe^{3+} and SCN^- in the mixture immediately after mixing but before reaction proceeds. Give your working and units. **(2)**

(b) The student then adds a few more drops of concentrated Fe^{3+} 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 7 is 4 marks)

8 Hydrated copper sulfate is involved in a reversible reaction: $CuSO_4.5H_2O(s)$ [blue] $\rightleftharpoons$ $CuSO_4(s)$ [white] + $5H_2O(l)$, 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 heating causes the crystals to turn white and adding water afterwards turns the solid blue again. State any assumption about whether this counts as a true equilibrium in the open test tube.

(Total for Question 8 is 4 marks)

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Mark scheme · 2.17 Reversible Reactions and Equilibrium: Test Pack

Question 1

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

(a) Answer: A reaction that can proceed in both directions so products can change back into reactants.

Question 2

(a) the rates of the forward and backward reactions are equal, oe **B1**

(a) Answer: The rates of the forward and backward reactions are equal.

(b) their concentrations remain constant over time even though both forward and backward reactions continue to occur, oe **B1**

(b) Answer: Concentrations of reactants and products remain constant over time while both reactions continue.

Question 3

(a) shifts to the right (towards CH_3OH) because there are fewer gas moles on the product side (1) than on the reactant side (3), oe **B1**

(a) Answer: Shifts to the right towards CH_3OH because there are fewer gas moles on the product side.

(b) shifts to the right (towards products) because adding a reactant is opposed by the system producing more product, oe **B1**

(b) Answer: Shifts to the right to make more CH_3OH because the system counteracts the added CO.

(c) shifts to the left (towards reactants) because increasing temperature favours the endothermic direction, oe **B1**

(c) Answer: Shifts to the left towards CO and H_2 because the forward reaction is exothermic so higher temperature favours the endothermic reverse.

Question 4

(a) there are 2 moles of gas on the reactant side (NO_2) but only 1 mole of gas on the product side (N_2O_4), oe **B1**

(a) by Le Chatelier's principle, increasing pressure shifts the equilibrium towards the side with fewer gas moles, i.e. towards N_2O_4, oe **B1**

(a) so more N_2O_4 forms and the mixture becomes paler (less brown), oe **B1**

(a) Answer: Increasing pressure shifts equilibrium towards N_2O_4 (the side with fewer gas moles), so more colourless N_2O_4 forms and the brown colour fades.

Question 5

increasing temperature favours the endothermic direction (forward, formation of the blue complex) so equilibrium shifts right, towards blue, oe **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 **M1**

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 **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 **A1**

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 6

(a) 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 **M1**

(a) moles remaining at equilibrium: A = 0.400 - 0.100 = 0.300 mol; B = 0.600 - 0.100 = 0.500 mol, cao **A1**

(a) percentage of A reacted = (amount reacted / initial amount) x 100 = (0.100 / 0.400) x 100 **M1**

(a) = 25.0% cao **A1**

(a) 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 **B1**

(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 7

(a) 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 **M1**

(a) Fe^{3+}: 0.0500 mol dm^{-3} and SCN^-: 0.00500 mol dm^{-3} cao **A1**

(a) Answer: Fe^{3+}: 0.0500 mol dm^{-3}; SCN^-: 0.00500 mol dm^{-3}

(b) identifies that adding more Fe^{3+} (a reactant) disturbs the equilibrium, oe **M1**

(b) the equilibrium shifts to the right (towards $FeSCN^{2+}$) to oppose the increase and use up some of the added Fe^{3+}, producing more of the red $FeSCN^{2+}$ complex, which darkens the colour, oe **A1**

(b) Answer: Adding more Fe^{3+} disturbs the equilibrium; by Le Chatelier's principle, the equilibrium shifts to the right to oppose the change, producing more red $FeSCN^{2+}$ and darkening the colour.

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Question 8

heating is an endothermic input of energy, and the forward (dehydration) reaction is endothermic, so heating favours the forward reaction, oe **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 $CuSO_4$, oe **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 $CuSO_4$, oe **M1**

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 **A1**

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.

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