2.5 Redox and Electrochemistry: 30-minute Pack

AQA 7405 · Calculator allowed · about 30 minutes

Total Marks

Name: _____ Date: ____ / ____ / _____

Answer ALL questions. Show all your working.

1 This question uses practical techniques from the required practical on redox titrations. A student wants to determine the percentage by mass of iron in an iron supplement tablet. One tablet, of total mass 2.50 g, is dissolved in dilute sulfuric acid and the resulting solution made up to exactly 250 cm^3 in a volumetric flask.
25.0 cm^3 portions of this solution are titrated against 0.0150 mol dm-3 potassium manganate(VII) solution. The mean titre required is 16.00 cm^3.
Equation: $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$

(a) State the colour change observed at the end point of this titration, and explain why no external indicator is required. **(2)**

(b) Calculate the amount, in mol, of MnO_4^- used in the titration of one 25.0 cm^3 sample. **(1)**

(c) Calculate the amount, in mol, of Fe^{2+} present in the 25.0 cm^3 sample. **(1)**

(d) Calculate the total amount, in mol, of Fe^{2+} in the whole 250 cm^3 solution, and hence the mass, in g, of iron in the tablet. **(3)**

(e) Calculate the percentage by mass of iron in the tablet. **(2)**

(Total for Question 1 is 9 marks)

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2 Explain, using ideas about standard electrode potentials and reaction kinetics, why a reaction between two half-cells that is predicted to be thermodynamically feasible from a positive E cell value may not actually be observed to occur, or may occur only extremely slowly, when the two solutions are mixed.

(Total for Question 2 is 6 marks)

3 This question is about how ion concentration affects electrode potential, applying Le Chatelier's principle to two half-equilibria.

(a) Using the equilibrium $Ag^+(aq) + e^- \rightleftharpoons Ag(s)$ and Le Chatelier's principle, predict and explain the effect on the electrode potential of the Ag^+/Ag half-cell of decreasing the concentration of $Ag^+(aq)$ (for example, by adding chloride ions to precipitate insoluble AgCl). **(3)**

(b) For the half-equation $MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$, predict and explain the effect on the electrode potential of increasing the concentration of $H^+(aq)$ (for example, by adding more acid). **(3)**

(Total for Question 3 is 6 marks)

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Mark scheme · 2.5 Redox and Electrochemistry: 30-minute Pack

Question 1

(a) colour changes from colourless (or very pale green) to a persistent pale pink/purple oe **B1**

(a) MnO_4^- is intensely coloured (purple) while Mn^{2+} is almost colourless, so $KMnO_4$ acts as its own indicator oe **B1**

(a) Answer: Colourless to persistent pale pink/purple; MnO_4^- (purple) is reduced to almost colourless Mn^{2+}, so the manganate(VII) is self-indicating

(b) 2.40 x 10^{-4} mol cao (awrt 2.40e-4) **B1**

(b) Answer: 2.40 x 10^{-4} mol

(c) 1.20 x 10^{-3} mol cao ft from (b) **B1**

(c) Answer: 1.20 x 10^{-3} mol

(d) total n(Fe^{2+}) = 1.20e-3 x (250/25.0) = 1.20 x 10^{-2} mol ft **M1**

(d) mass = 1.20e-2 x 55.8 **M1**

(d) 0.670 g cao (awrt 0.669-0.670 g) ft **A1**

(d) Answer: 0.670 g

(e) % = (0.6696/2.50) x 100 ft **M1**

(e) 26.8% cao (awrt 26.7-26.8%) ft **A1**

(e) Answer: 26.8%

Question 2

Level 3 A detailed and coherent explanation that links both the thermodynamic meaning of E cell (feasibility under standard conditions) and the kinetic reason (activation energy/rate) why a feasible reaction may not be observed, using correct chemical terminology throughout. **5-6**

Level 2 A explanation that covers at least one of the two key ideas (thermodynamic feasibility versus kinetics) in reasonable detail, or touches on both but with limited development or some inaccuracy. **3-4**

Level 1 Basic or isolated relevant statement(s) with little or no linkage between ideas; chemistry is underdeveloped. **1-2**

Level 0 No relevant chemistry given, or answer entirely incorrect/irrelevant. **0**

Indicative content:

- E cell values are calculated from standard electrode potentials measured under standard conditions (298 K, 100 kPa, 1.00 mol dm-3 solutions); if the actual conditions differ from standard, the real value (and occasionally even the sign) of E cell can differ from the predicted value.
- A positive E cell value shows only that a reaction is thermodynamically feasible (energetically favourable); it gives no information about how fast the reaction will occur.
- Many thermodynamically feasible reactions have a high activation energy, so the rate of reaction may be extremely slow (or immeasurably slow at room temperature) even though the reaction is energetically favourable.
- As a result the reaction may not be observed to occur within a reasonable timescale, even though it is predicted to be feasible.
- Some reactions that are feasible from E cell values only proceed at an observable rate in the presence of a suitable catalyst.
- In practice, concentrations and temperatures often differ from standard conditions, which changes the effective (non-standard) cell potential and can also affect whether the reaction is observed to proceed.

Question 3

(a) the electrode potential becomes less positive (decreases), oe **B1**

(a) decreasing [Ag^+(aq)] shifts the equilibrium $Ag^+ + e^- \rightleftharpoons Ag$ to the left, by Le Chatelier's principle, oe **B1**

(a) this reduces the tendency of the half-cell to be reduced (to gain electrons), which is measured as a less positive electrode potential, oe **B1**

(a) Answer: The electrode potential becomes less positive.

(b) the electrode potential becomes more positive (increases), oe **B1**

(b) increasing [H^+(aq)] shifts the equilibrium to the right, by Le Chatelier's principle, oe **B1**

(b) this increases the tendency of the half-cell to be reduced (to gain electrons), which is measured as a more positive electrode potential, oe **B1**

(b) Answer: The electrode potential becomes more positive.

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