2.5 Redox and Electrochemistry: Test Pack

AQA 7405 · Calculator allowed · about 60 minutes

Total Marks

Name: _____ Date: ____ / ____ / _____

Answer ALL questions. Show all your working.

1 Potassium manganate(VII), $KMnO_4$, is a strong oxidising agent that is widely used in redox titrations. In acidified solution it reacts with iron(II) ions according to the equation:
$MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$

(a) State the oxidation state of manganese in the MnO_4^- ion. **(1)**

(b) State the oxidation state of chromium in the dichromate(VI) ion, $Cr_2O_7^{2-}$. **(1)**

(c) The half-equation for the reduction of MnO_4^- is: **(3)**
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$
Use this half-equation and the half-equation for the oxidation of Fe^{2+} to Fe^{3+} to show that the overall balanced ionic equation is:
$MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$

(d) Identify the oxidising agent in this reaction and explain your answer in terms of electron transfer. **(2)**

(Total for Question 1 is 7 marks)

2 This question tests key definitions used throughout the redox and electrochemistry topic.

(a) Define oxidation in terms of electron transfer. **(1)**

(b) Define an oxidising agent. **(1)**

(c) State the conditions that define a standard electrode potential. **(2)**

(Total for Question 2 is 4 marks)

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3 Copper(I) ions, $Cu^+(aq)$, are unstable in aqueous solution and undergo disproportionation:
$2Cu^+(aq) \rightarrow Cu^{2+}(aq) + Cu(s)$
Standard electrode potentials: $E(Cu^{2+}/Cu^+) = +0.15$ V; $E(Cu^+/Cu) = +0.52$ V

(a) Write the two half-equations that combine to give this disproportionation reaction, and confirm that they combine to give the overall equation shown above. **(2)**

(b) Calculate the E cell value for this disproportionation reaction. **(2)**

(c) State, with a reason, whether this disproportionation reaction is thermodynamically feasible under standard conditions. **(2)**

(Total for Question 3 is 6 marks)

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4 This question is about the redox reaction between acidified potassium dichromate(VI) and iron(II) ions.

(a) Write the balanced half-equation for the reduction of dichromate(VI) ions, $Cr_2O_7^{2-}$, to Cr^{3+} ions in acidic solution. **(2)**

(b) Write the half-equation for the oxidation of Fe^{2+} ions to Fe^{3+} ions. **(1)**

(c) Combine the two half-equations to give the overall ionic equation for the reaction between acidified dichromate(VI) ions and Fe^{2+} ions, ensuring the electrons cancel. **(2)**

(d) State the colour change observed during this reaction, and identify the oxidising agent and the reducing agent. **(3)**

(Total for Question 4 is 8 marks)

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5 Standard electrode potentials: $E(Cl_2/Cl^-) = +1.36$ V; $E(Br_2/Br^-) = +1.09$ V; $E(I_2/I^-) = +0.54$ V; $E(Fe^{3+}/Fe^{2+}) = +0.77$ V

(a) Use the E values to deduce, with a calculation, whether chlorine gas will oxidise bromide ions to bromine under standard conditions. **(2)**

(b) Use the E values to deduce, with a calculation, whether iodide ions will reduce Fe^{3+} ions to Fe^{2+} under standard conditions. **(3)**
(Overall equation: $2Fe^{3+}(aq) + 2I^-(aq) \rightarrow 2Fe^{2+}(aq) + I_2(aq)$)

(c) A student mixes potassium iodide solution with iron(III) chloride solution and observes a colour change. Describe the colour change and explain how it supports your answer to part (b). **(2)**

(Total for Question 5 is 7 marks)

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6 A hydrogen-oxygen fuel cell operating under alkaline conditions is used to power a small electric vehicle.
$E(O_2/OH^-, \text{alkaline}) = +0.40$ V; $E(H_2O/H_2, \text{alkaline}) = -0.83$ V

(a) Write the half-equation at the negative electrode of this fuel cell. **(2)**

(b) Write the half-equation at the positive electrode of this fuel cell. **(2)**

(c) Combine your two half-equations to give the overall equation for the reaction in this fuel cell. **(1)**

(d) Calculate the standard emf of this fuel cell. **(2)**

(e) State two advantages and one disadvantage of using a hydrogen-oxygen fuel cell, rather than a rechargeable lithium-ion cell, to power a vehicle. **(3)**

(Total for Question 6 is 10 marks)

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Mark scheme · 2.5 Redox and Electrochemistry: Test Pack

Question 1

(a) +7 cao **B1**

(a) Answer: +7

(b) +6 cao **B1**

(b) Answer: +6

(c) Fe^{2+} -> Fe^{3+} + e- written as the oxidation half-equation **M1**

(c) oxidation half-equation multiplied by 5 so that electrons balance with the reduction half-equation (5e-) **M1**

(c) half-equations added and electrons cancelled to give the printed overall equation cso **A1**

(c) Answer: MnO_4^- + $8H^+$ + $5Fe^{2+}$ -> Mn^{2+} + $5Fe^{3+}$ + $4H_2O$

(d) MnO_4^- (or potassium manganate(VII)) identified as the oxidising agent **B1**

(d) correct explanation: MnO_4^- gains electrons (is reduced, Mn goes from +7 to +2) and in doing so causes Fe^{2+} to lose electrons (be oxidised to Fe^{3+}) oe **B1**

(d) Answer: MnO_4^- (manganate(VII) ion); it is reduced (gains 5 electrons per Mn) and so oxidises Fe^{2+} to Fe^{3+}

Question 2

(a) loss of electrons, oe **B1**

(a) Answer: Loss of electrons.

(b) a species that accepts electrons from another species (causing it to be oxidised) and is itself reduced, oe **B1**

(b) Answer: A species that accepts electrons from another species and is itself reduced.

(c) 298 K (25°C), 100 kPa pressure for any gases, and 1 mol/dm3 concentration for solutions, oe **B1**

(c) measured relative to a standard hydrogen electrode, oe **B1**

(c) Answer: 298 K, 100 kPa, 1 mol/dm3, measured against a standard hydrogen electrode.

Question 3

(a) oxidation: Cu^+ -> Cu^{2+} + e- and reduction: Cu^+ + e- -> Cu **B1**

(a) adding the two half-equations and cancelling one electron on each side gives $2Cu^+$ -> Cu^{2+} + Cu cao **B1**

(a) Answer: Cu^+ -> Cu^{2+} + e- ; Cu^+ + e- -> Cu ; sum = $2Cu^+$ -> Cu^{2+} + Cu

(b) E cell = E(Cu^+/Cu) - E(Cu^{2+}/Cu^+) = 0.52 - 0.15 **M1**

(b) +0.37 V cao **A1**

(b) Answer: +0.37 V

(c) yes, feasible **B1**

(c) because E cell is positive ft from (b), so the reaction is thermodynamically favourable/feasible oe **B1**

(c) Answer: Yes; E cell (+0.37 V) is positive, so the reaction is thermodynamically feasible

Question 4

(a) correct atom balance: $Cr_2O_7^{2-}$ + $14H^+$ -> $2Cr^{3+}$ + $7H_2O$ **B1**

(a) correct electron balance: + 6e- on the left, giving matching charge (+6) on both sides **B1**

(a) Answer: $Cr_2O_7^{2-}$ + $14H^+$ + 6e- -> $2Cr^{3+}$ + $7H_2O$

(b) Fe^{2+} -> Fe^{3+} + e- **B1**

(b) Answer: Fe^{2+} -> Fe^{3+} + e-

(c) multiply the Fe^{2+} half-equation by 6 so that 6 electrons are transferred in each half-equation **M1**

(c) $Cr_2O_7^{2-}$ + $14H^+$ + $6Fe^{2+}$ -> $2Cr^{3+}$ + $7H_2O$ + $6Fe^{3+}$ **A1**

(c) Answer: $Cr_2O_7^{2-}$ + $14H^+$ + $6Fe^{2+}$ -> $2Cr^{3+}$ + $7H_2O$ + $6Fe^{3+}$

(d) colour changes from orange (dichromate(VI)) to green (Cr^{3+}), oe **B1**

(d) oxidising agent = dichromate(VI) ions, $Cr_2O_7^{2-}$ **B1**

(d) reducing agent = Fe^{2+} ions **B1**

(d) Answer: Orange to green; oxidising agent $Cr_2O_7^{2-}$; reducing agent Fe^{2+}.

Question 5

(a) E cell = E(Cl_2/Cl^-) - E(Br_2/Br^-) = 1.36 - 1.09 **M1**

(a) +0.27 V, positive, so chlorine will oxidise bromide ions to bromine (feasible) cao **A1**

(a) Answer: +0.27 V; yes, chlorine oxidises bromide to bromine

(b) identifies Fe^{3+}/Fe^{2+} as the reduction half-reaction and I_2/I^- as the oxidation (reversed) half-reaction **M1**

(b) E cell = E(Fe^{3+}/Fe^{2+}) - E(I_2/I^-) = 0.77 - 0.54 **M1**

(b) +0.23 V, positive, so iodide ions will reduce Fe^{3+} to Fe^{2+} (feasible) cao **A1**

(b) Answer: +0.23 V; yes, iodide ions reduce Fe^{3+} to Fe^{2+}

(c) solution turns brown/dark orange-brown (accept blue-black if starch indicator is also mentioned) oe **B1**

(c) this colour is due to iodine, I_2, being formed, confirming the reaction predicted as feasible in (b) has occurred ft **B1**

(c) Answer: The solution turns brown/dark orange-brown due to iodine formation, confirming the feasible reaction predicted in (b) has taken place

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Question 6

(a) correct species: H_2, OH^-, H_2O, e- present **B1**

(a) $H_2 + 2OH^- \rightarrow 2H_2O + 2e^-$ fully balanced cao **B1**

(a) Answer: $H_2 + 2OH^- \rightarrow 2H_2O + 2e^-$

(b) correct species: O_2, H_2O, e-, OH^- present **B1**

(b) $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$ fully balanced cao **B1**

(b) Answer: $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$

(c) $2H_2 + O_2 \rightarrow 2H_2O$ cao ft from (a) and (b) **B1**

(c) Answer: $2H_2 + O_2 \rightarrow 2H_2O$

(d) E cell = $E(O_2/OH^-) - E(H_2O/H_2) = 0.40 - (-0.83)$ **M1**

(d) +1.23 V cao **A1**

(d) Answer: +1.23 V

(e) advantage: the vehicle can be refuelled quickly rather than waiting for a lithium-ion cell to recharge oe **B1**

(e) advantage: the fuel cell can continue operating for as long as fuel and oxygen are supplied, without needing to recharge the cell oe; credit a distinct point from refuelling speed **B1**

(e) disadvantage: hydrogen is difficult/expensive to store and transport safely (low energy density by volume; needs high pressure or cryogenic storage), and/or hydrogen is often manufactured from fossil fuels so is not always a 'clean' fuel overall oe **B1**

(e) Answer: Advantages: quick refuelling instead of waiting for recharging; continued operation while hydrogen and oxygen are supplied, without needing to recharge the cell.
Disadvantage: hydrogen storage and transport are difficult and expensive

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