InkChemistry

ANSWER BOOK

The electrochemical series and predicting feasibility

Electrochemistry · Unit C10

AQA 3.1.11.1 · OCR A 5.2.3

18 questions · 58 marks

NAVY EDITION

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SECTION B · Standard, 3 to 5 marks

B1 ESP-Q5

M1 dichromate(VI) is reduced, so its half-cell is on the right: E°cell = 1.33 - 0.77. **B1** the dichromate half-equation takes six electrons, so six Fe^{2+} ions are oxidised for each $Cr_2O_7^{2-}$. **B1** $Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6Fe^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + 7H_2O(l) + 6Fe^{3+}(aq)$; ignore state symbols. **A1** E°cell = +0.56 V. [4]

B2 ESP-Q7

B1 $E^\circ(Al^{3+}/Al)$ is more negative than $E^\circ(H^+/H_2) = 0.00$ V, so the reduction of H^+ by aluminium is feasible, E°cell = +1.66 V. **B1** the aluminium is covered by a thin, unreactive layer of aluminium oxide that keeps the acid off the metal. **B1** so the rate is negligible; E° predicts whether a reaction can happen, not how fast it goes. **B1** the conditions may not be standard, so the actual electrode potentials differ from the E° values used; allow a high activation energy for a different reaction. [4]

B3 ESP-Q9

B1 (a) chloride precipitates silver chloride, so $[Ag^+]$ falls and $Ag^+(aq) + e^- \rightleftharpoons Ag(s)$ shifts to the left. **B1** (a) the silver electrode, on the right, becomes less positive, so the EMF falls. **B1** (b) $[Cu^{2+}]$ falls, so $Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$ shifts to the left and the copper electrode becomes less positive. **B1** (b) copper is the left-hand electrode, so E right - E left increases and the EMF rises. [4]

B4 ESP-Q17

B1 E°cell = 1.33 - 1.36 = -0.03 V, so the reaction is not feasible under standard conditions and the calculation is correct. **B1** the value is very close to zero, so a small change in conditions can reverse the prediction. **B1** with concentrated hydrochloric acid and heat, higher $[H^+]$ and $[Cl^-]$ make the dichromate electrode more positive and the chlorine electrode less positive. **B1** chlorine is then released, so 'never' is not justified; the prediction applies only under standard conditions. [4]

B5 ESP-Q3

M1 iron(III) is reduced, so the iron half-cell is on the right: E°cell = 0.77 - 0.15. **A1** E°cell = +0.62 V. **B1** E°cell is positive, so the reaction is feasible under standard conditions. [3]

B6 ESP-Q4

M1 silver ions are reduced, so the silver half-cell is on the right: E°cell = 0.80 - 0.34. **M1** the E° of silver is not doubled when its half-equation is doubled, so the student's value is wrong. **A1** E°cell = +0.46 V. [3]

B7 ESP-Q6

B1 the manganate(VII) E° is more positive, so MnO_4^- oxidises Cl^- to Cl_2, with E°cell = +0.15 V. **B1** some manganate(VII) would then react with chloride ions as well as with the substance being analysed. **B1** so the titre would be too large; sulfate ions are not oxidised by manganate(VII). [3]

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B8 ESP-Q8 **B1** the half-cell equilibrium is $Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$. **B1** increasing $[Cu^{2+}]$ shifts the position of equilibrium to the right, so the electrode takes electrons more readily. **B1** the electrode potential becomes more positive than +0.34 V. [3]

B9 ESP-Q10 **B1** the very negative E° of sodium means sodium is a strong reducing agent, readily oxidised to Na^+ by oxygen and water. **B1** the very positive E° of gold means Au^{3+} is readily reduced and gold atoms are very hard to oxidise. **B1** oxygen's E° of +1.23 V is less positive than gold's +1.50 V, so oxidising gold with oxygen is not feasible and gold stays as the metal. [3]

B10 ESP-Q11 **B1** Cu^+ is reduced in the first half-equation and oxidised in the reverse of the second, E°cell = 0.52 - 0.15 = +0.37 V, so the reaction is feasible. **B1** $2Cu^+(aq) \rightarrow Cu(s) + Cu^{2+}(aq)$, a disproportionation. **B1** a brown or pink solid of copper forms in a blue solution of copper(II) ions. [3]

B11 ESP-Q12 **B1** the oxygen half-equation is more positive, so dissolved oxygen oxidises Fe^{2+} to Fe^{3+}, E°cell = +0.46 V. **B1** $4Fe^{2+}(aq) + O_2(g) + 4H^+(aq) \rightarrow 4Fe^{3+}(aq) + 2H_2O(l)$. **B1** Fe^{3+} ions are yellow-brown, and the colour builds up slowly because the reaction is slow at room temperature. [3]

B12 ESP-Q13 **B1** $[Cu^{2+}]$ falls, so $Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$ shifts left and the copper electrode becomes less positive. **B1** $[Zn^{2+}]$ rises, so $Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$ shifts right and the zinc electrode becomes less negative. **B1** the difference between the two potentials falls, and is zero when the cell reaction reaches equilibrium. [3]

B13 ESP-Q14 **B1** the half-equations are written as reductions with the more negative, Zn^{2+}/Zn, above the more positive, Cu^{2+}/Cu. **B1** an anticlockwise loop runs forwards along the lower equation, Cu^{2+} to Cu, and backwards along the upper, Zn to Zn^{2+}. **B1** so Cu^{2+} oxidises Zn: $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$. [3]

B14 ESP-Q16 **B1** the more negative the E°, the stronger the reducing agent, so zinc is strongest, then iron, then copper. **B1** zinc and iron have E° values below 0.00 V, so each reduces H^+ to hydrogen, E°cell +0.76 V and +0.44 V. **B1** copper's E° is above 0.00 V, so its reaction with H^+ would have E°cell = -0.34 V and is not feasible. [3]

B15 ESP-Q18 **M1** z = 1, so $E = 0.80 + 0.059 \times \log_{10}(0.010) = 0.80 + 0.059 \times (-2)$. **A1** $E(Ag^+/Ag)$ = +0.68 V. **A1** the silver electrode is still the more positive, so EMF = 0.68 - 0.34 = +0.34 V; allow error carried forward. [3]

B16 ESP-Q19 **M1** two electrons are transferred in the equation as written, so n = 2. **M1** $\Delta G° = -nFE°cell = -2 \times 96\,500 \times 0.27 = -52\,110$ J mol^{-1}. **A1** $\Delta G°$ = -52 kJ mol^{-1}, to the two significant figures of E°cell; allow -52.1; a positive value is rejected. [3]

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B17 ESP-Q20

B1 $\Delta G^\circ = -nFE^\circ_{cell}$, and ΔG° refers to the amounts in the equation as written. **B1** doubling the equation doubles n, the number of moles of electrons transferred, so ΔG° doubles. **B1** E°_{cell} is energy per unit of charge, a property of the two couples that does not depend on how many electrons are written, so it is unchanged. [3]

B18 ESP-Q21

B1 ln K, and the total entropy change, are proportional to E°_{cell}. **B1** a large positive E°_{cell} gives a very large K, so at equilibrium almost no reactant remains. **B1** an E°_{cell} close to zero gives K close to 1, so reactants and products are both present in significant amounts at equilibrium. [3]

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