Past-paper walk-through

Standard electrode potentials E , standard cell potentials E cell and the ▪ 24.2

eXercise

Questions in this set

- 9701/41 N25 Q4 ▪ 11 marks
- 9701/41 N25 Q2 ▪ 8 marks
- 9701/42 N25 Q5 ▪ 15 marks

Reading the mark scheme

- **B1** a mark that stands on its own – the point is either made or not.
- **M1** a *method* mark: the right approach, even if the number is wrong.
- **A1** an *answer* mark, and it usually needs its M mark first.
- **C1** a working mark on the way to the answer.
- // separates answers the examiner accepts equally.
- **ecf** = error carried forward: a later part is marked on your own earlier answer.
- **owtte** = "or words to that effect" – the wording need not match.

Question 4

9701/41 N25 ▪ Q4 ▪ 11 marks

4 **(a)** Define standard cell potential, $E^{\ominus}_{\text{cell}}$. Include a description of standard conditions.

Question 4

(b) The Daniell cell is an electrochemical cell consisting of a $Cu^{2+}(aq)/Cu(s)$ electrode and a $Zn^{2+}(aq)/Zn(s)$ electrode.

(i) Draw a labelled diagram of this electrochemical cell.

Include all necessary substances and relevant pieces of apparatus needed to measure the $E^{\ominus}_{cell}$.

It is **not** necessary to state the conditions used.

Question 4

[3]

(ii) State the charge carriers that transfer current through the solutions and through the wire.

Question 4

(iii) The standard electrode potential, $E^{\ominus}$, for the $Zn^{2+}(aq)/Zn(s)$ electrode is –0.76 V.

Water is added to a standard $Zn^{2+}(aq)/Zn(s)$ electrode.

The new concentration of $Zn^{2+}(aq)$ is $0.25\,mol\,dm^{-3}$.

Use the Nernst equation to calculate the electrode potential, E, for this new $Zn^{2+}(aq)/Zn(s)$ electrode.

Question 4

(c) An electrochemical cell consists of a ZnO / Zn electrode and a MnO_2 / Mn_2O_3 electrode in an alkaline electrolyte.

The standard cell potential, $E^{\ominus}_{cell}$, for this cell is +1.47 V.

The half-equation at each electrode when this cell is discharging is shown.

$$Zn + 2OH^- \rightarrow ZnO + H_2O + 2e^-$$

$$2MnO_2 + H_2O + 2e^- \rightarrow Mn_2O_3 + 2OH^-$$

(i) Use this information to determine the change in oxidation state of manganese when this cell is discharging.

Question 4

(ii) Write the equation for the overall reaction that occurs when this cell is discharging.

Question 4

(iii) The $E^{\ominus}$ for the ZnO / Zn electrode is –1.28 V.

Calculate the standard electrode potential, $E^{\ominus}$, for the MnO_2 / Mn_2O_3 electrode.

Question 4

[Total: 11]

Solution 4

(a)

Mark scheme

- M1: potential difference / voltage / EMF AND between two half-cells / two electrodes (in a cell)
- M2: (at concentration of) 1 mol dm–3 AND (pressure of) 1 atm / 101 kPa AND (temperature of) 298 K / 25 °C
- 2
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Solution 4

(b)(i)

Mark scheme

- M1: salt bridge labelled AND voltmeter / V AND wire to electrode and liquid level shown [complete circuit]
- M2: Cu(s) AND Cu2+(aq)
- M3: Zn(s) AND Zn2+(aq)
- 3

Solution 4

(b)(ii)

Mark scheme

- ions AND electrons
- 1

Solution 4

(b)(iii)

Mark scheme

- M1: Nernst equation $E = Eo + (0.059/z) \log(Zn2+/Zn)$
- M2: $E = -0.76 + (0.059 / 2) \log(0.25) = -0.778$ V
- OR $E = -0.76 + (8.31 \times 298) / (96\,500 \times 2) \ln(0.25) = -0.778$ V min 2sf
- 2

Solution 4

(c)(i)

Mark scheme

- (from) +4 (to) +3
- 1

(c)(ii)

Mark scheme

- $Zn + 2MnO2 \rightarrow ZnO + Mn2O3$
- 1

Solution 4

(c)(iii)

Mark scheme

- Eo (MnO2 / Mn2O3) = 1.47 – 1.28 = (+)0.19 V
- 1
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Question 2

9701/41 N25 ▪ Q2 ▪ 8 marks

2 **(a)** Iron can form stable ions in the +2 and +3 oxidation states.

Explain why transition elements have variable oxidation states.

Question 2

(b) Aqueous solutions of iron(II) salts contain the complex ion $[Fe(H_2O)_6]^{2+}$.

Define complex ion.

Question 2

(c) $[Fe(H_2O)_6]^{2+}$ can be converted into $[Fe(H_2O)_4(OH)_2]$.

(i) Suggest a suitable reagent for this conversion. State the type of reaction.

Question 2

Question 2

[1]

(ii) $[Fe(H_2O)_4(OH)_2]$ is a green precipitate that turns brown on standing in air.

Table 2.1 shows electrode potentials for some electrode reactions.

Table 2.1

electrode reaction	$E^{\ominus}$ / V
$Fe(H_2O)_3(OH)_3 + H_2O + e^- \rightleftharpoons Fe(H_2O)_4(OH)_2 + OH^-$	–0.56
$O_2 + 2H_2O + 4e^- \rightleftharpoons 4OH^-$	+0.40

Use the information in Table 2.1 to explain why $[Fe(H_2O)_4(OH)_2]$ turns brown on standing in air.
Include an equation for this reaction.

Question 2

(d) The complex $[Co(NH_3)_6]^{2+}$ reacts with hydrogen peroxide as shown.

reaction 1 $2[Co(NH_3)_6]^{2+} + H_2O_2 \rightarrow 2[Co(NH_3)_6]^{3+} + 2OH^-$ $E^{\ominus}_{cell} = +1.67\,V$

Calculate $\Delta G^{\ominus}$, in $kJ\,mol^{-1}$, for reaction 1.

Question 2

[Total: 8]

Solution 2

(a)

Mark scheme

- the (3)d and (4)s sub-shells / orbitals / electrons are close / similar in energy
- 1

Solution 2

(b)

Mark scheme

- species or ion formed by a central metal atom/ion AND surrounded by/bonded to (one or more) ligands
- 1

Solution 2

(c)(i)

Mark scheme

- NaOH / OH–(aq) AND precipitation / ligand exchange / deprotonation / acid–base
- 1

Solution 2

(c)(ii)

Mark scheme

- M1: $Fe(H_2O)_4(OH)_2$ is oxidised (from + 2 to + 3 on standing)
- M2: Eo of $Fe(H_2O)_3(OH)_3$ is more negative than Eo of O_2
- OR Eocell = 0.40 – (–0.56) = (+)0.96 V
- M3: $4Fe(H_2O)_4(OH)_2 + O_2 \rightarrow 4Fe(H_2O)_3(OH)_3 + 2H_2O$
- 3

Solution 2

(d) Mark scheme

- $G^o = -2 \times 1.67 \times 96\,500 = -322\,310$ J mol–1
- M1: $G^o = -nE^o_{cell}$ IF AND $n = 2$
- OR $-2 \times 1.67 \times 96\,500$
- M2: $G^o = -322.3$ kJ mol–1
- 2
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Question 5

9701/42 N25 ▪ Q5 ▪ 15 marks

5 Cobalt is a transition element which forms compounds containing Co^{2+} and Co^{3+} ions. Cobalt(II) sulfate dissolves in water to form a solution containing the $[Co(H_2O)_6]^{2+}$ complex ion.

(a) (i) Complete the electronic configurations of a Co^{2+} ion and a Co^{3+} ion.

Question 5

Question 5

[1]

(ii) Explain why transition elements can form complex ions.

Question 5

(iii) An excess of concentrated HC*l* is added to a solution containing $[Co(H_2O)_6]^{2+}$.

Describe the colour change observed and the state of the cobalt-containing product.

Question 5

Question 5

[2]

(iv) Write an equation for the reaction occurring in **(a)(iii)**.

Question 5

(v) Name the type of reaction occurring in **(a)(iii)**.

Question 5

(vi) Write an equation for the reaction that occurs when an excess of NaOH(aq) is added to a solution containing $[Co(H_2O)_6]^{2+}$.

Question 5

(b) Cobalt metal can be oxidised by acidified $K_2Cr_2O_7$. The relevant half-equations, and their $E^\ominus$ values, are shown.

$$Co^{2+} + 2e^- \rightleftharpoons Co \qquad E^\ominus = -0.28\,V$$
$$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightleftharpoons 2Cr^{3+} + 7H_2O \qquad E^\ominus = +1.33\,V$$

(i) A Co^{2+}/Co electrode is constructed in which $[Co^{2+}]$ is $0.020\,mol\,dm^{-3}$ at 298 K.

Use the Nernst equation to show that the E value for this Co^{2+}/Co electrode is –0.33 V.

Question 5

[2]

(ii) An electrochemical cell is constructed using the Co^{2+}/Co electrode described in **(b)(i)** and a $Cr_2O_7^{2-}/Cr^{3+}$ electrode in which all conditions are standard.

Calculate the value of E_{cell}.

Question 5

(iii) A current is drawn from the electrochemical cell described in **(b)(ii)**.

Write an equation for the reaction taking place in the cell.

Question 5

(iv) Complete the sentences to identify the negative electrode and the direction of electron flow when a current is drawn from the cell described in **(b)(ii)**.

Question 5

[1]

(c) A molten Co^{2+} salt is electrolysed using a current of 0.500 A.

0.547 g of cobalt metal forms at the cathode. Under the conditions used no other reduction reaction occurs at the cathode.

Calculate the time in minutes for which the current flows to produce this mass of cobalt.

Give your answer to **three** significant figures.

Question 5

[Total: 15]

Solution 5

(a)(i)

Mark scheme

- Co2+ = [Ar] 3d7 AND Co3+ = [Ar] 3d6
- 1

(a)(ii)

Mark scheme

- has vacant d orbitals that are energetically accessible
- 1

Solution 5

(a)(iii) Mark scheme

- ▪
- pink
- ▪
- blue
- ▪
- aqueous
- Any two [1], all three [2]
- 2

Solution 5

(a)(iv)

Mark scheme

Co(H2O)6 2+ + 4HCl → [CoCl4]2– + 6H2O + 4H+ OR [Co(H2O)6]2+ + 4Cl– → [CoCl4]2– + 6H2O

- 1
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Solution 5

(a)(v)

Mark scheme

- ligand exchange
- 1

Solution 5

(a)(vi)

Mark scheme

$Co(H_2O)_6^{2+} + 2OH^- \rightarrow Co(OH)_2(H_2O)_4 + 2H_2O$ OR

$Co(H_2O)_6^{2+} + 2OH^- \rightarrow Co(OH)_2 + 6H_2O$

- 1

Solution 5

(b)(i)

Mark scheme

- M1: Nernst equation $E = Eo + (0.059 / z) \times \log(\frac{[\text{oxidised}]}{[\text{reduced}]})$
- M2: $E = -0.28 + (0.059 / 2) \times \log(0.02)$
- 2

Solution 5

(b)(ii)

Mark scheme

- 1.66 V
- 1

(b)(iii)

Mark scheme

- $3Co + Cr2O72- + 14H+ \rightarrow 3Co2+ + 2Cr3+ + 7H2O$
- 1

Solution 5

(b)(iv)

Mark scheme

- Co^{2+} / Co, Co^{2+} / Co, $Cr_2O_7^{2-}$ / Cr^{3+}
- 1

Solution 5

(c)

Mark scheme

- M1: moles of cobalt $= 0.547 / 58.9 = 0.00929$
- M2: coulombs required $= 0.00929 \times 96500 \times 2 = 1792$
- M3: time required $= 1792 / 0.50 = 3584.8\,\text{s} = 59.7$ minutes 3sf
- 3