Week 34

A-Level Chemistry

Homework bundle for this week

本周作业合集

exercises.lol

Contents 目录

A-Level Chemistry Name: _____ Score: _____

1. At the cathode (negative electrode) in electrolysis:
   - ☐ positive ions gain electrons (reduction)
   - ☐ negative ions lose electrons
   - ☐ nothing happens
   - ☐ water is made
2. The standard hydrogen electrode is defined as having a potential of:
   - ☐ 0.00 V
   - ☐ 1.00 V
   - ☐ −1.00 V
   - ☐ 0.059 V
3. A redox reaction is feasible when the cell potential is:
   - ☐ positive
   - ☐ negative
   - ☐ zero only
   - ☐ equal to 0.059
4. Electrolysing an aqueous solution of a reactive metal salt gives, at the cathode:
   - ☐ hydrogen (rather than the reactive metal)
   - ☐ the reactive metal
   - ☐ oxygen
   - ☐ the halogen
5. The standard cell potential is:
   - ☐ E°(more positive) − E°(less positive)
   - ☐ E°(less positive) − E°(more positive)
   - ☐ the sum of the two E° values
   - ☐ always zero
6. During electrolysis, positive ions move to the negative electrode, the _____.

_____

7. A more negative standard electrode potential means a stronger reducing agent.

   ☐ True ☐ False

8. A reaction is feasible if the cell potential (Ecell) is positive.

   ☐ True ☐ False

9. The link between cell potential and free energy is:

   ☐ $\Delta G^\circ = -n\ E^\circ cell\ F$

   ☐ $\Delta G^\circ = n\ E^\circ cell\ F$

   ☐ $\Delta G^\circ = E^\circ cell\ /\ F$

   ☐ $\Delta G^\circ = F\ /\ E^\circ cell$

10. The charge passed during electrolysis is given by:

    ☐ Q = It (current × time)

    ☐ Q = I / t

    ☐ Q = F / e

    ☐ Q = m / M

24. Electrochemistry

Answers · 答案

A-Level Chemistry

1. **positive ions gain electrons (reduction)**

   Positive ions are attracted to the cathode and gain electrons (reduction); the anode does oxidation.

2. **0.00 V**

   It is the reference half-cell, set at exactly 0.00 V; all other E° values are measured against it.

3. **positive**

   A positive E°cell means the reaction can happen (analogous to a negative ΔG).

4. **hydrogen (rather than the reactive metal)**

   For a reactive metal, water is reduced to hydrogen at the cathode instead of the metal.

5. **E°(more positive) − E°(less positive)**

   Subtract the less positive electrode potential from the more positive one.

6. **cathode**

   Reduction occurs at the cathode.

7. **True**

   It loses electrons more readily.

8. **True**

   Ecell = E(reduced) − E(oxidised).

9. **$\Delta G^\circ = -n\ E^\circ cell\ F$**

   $\Delta G^\circ = -nE^\circ cellF$; a positive E°cell gives a negative ΔG° (feasible).

10. **Q = It (current × time)**

    Q = It; then moles of electrons = Q/F, and the half-equation gives the moles of product.

A-LEVEL CHEMISTRY • EXERCISE SHEET

24.1 Electrolysis

Name: _____ Class: _____ Date: _____ **Total: 11 marks**

KEY WORDS electrolysis 电解 • avogadro constant 阿伏伽德罗常量 • faraday constant 法拉第常量 • electrolyte 电解质 • cathode 阴极

Objective

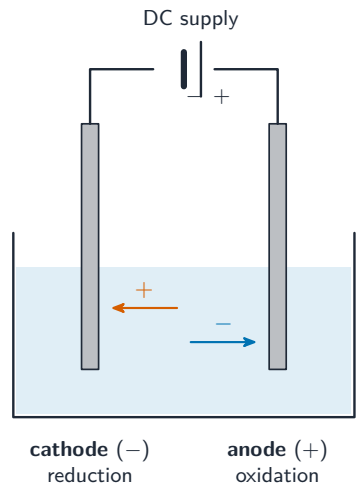
Build the skills to answer exam questions on **electrolysis** 电解 — predicting the products, the **Faraday constant** 法拉第常量, and electrolysis calculations.

You must be able to:

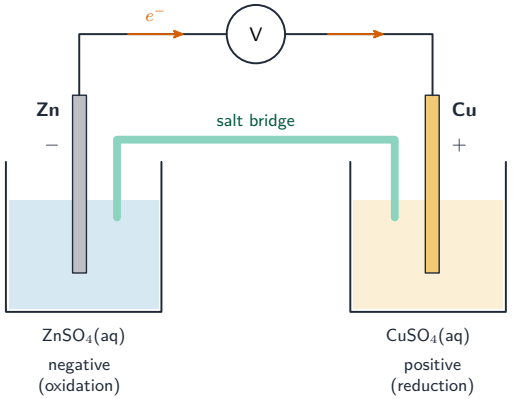
- predict the products of electrolysis (molten or aqueous, from the redox series)
- use $F = Le$ and $Q = It$
- calculate the mass or volume of substance liberated

1 Worked examples

■ **Electrolysis and its products**



Cathode (−): positive ions gain electrons (reduction). Anode (+): negative ions lose electrons (oxidation)



Electrolysis is the reverse of an electrochemical cell

A **molten** electrolyte gives the metal (cathode) and non-metal (anode). An **aqueous** one: a reactive metal gives H_2 instead at the cathode; the anode gives O_2 unless a concentrated halide gives the halogen.

■ Calculations

$F = Le$ (Faraday = Avogadro × electron charge). $Q = It$, then: moles of e^- = Q/F → moles of product (from the half-equation) → mass or volume.

Example. 2.0 A for 30 min through $CuSO_4$ ($Cu^{2+} + 2e^- \rightarrow Cu$): $Q = 2.0 \times 1800 = 3600$ C; mol $e^- = 3600/96500 = 0.0373$; mol Cu = 0.0187; $m = 0.0187 \times 64 = 1.2$ g.

2 Practice

2.1 State what is formed at the cathode and anode when **molten** sodium chloride is electrolysed. [2]

2.2 Write the relationship between the Faraday constant, the Avogadro constant and the electron charge. [1]

3 Exam-style questions

3.1 At the cathode during electrolysis, positive ions are: [1]

- **A** oxidised (lose electrons)
- **B** reduced (gain electrons)
- **C** neutralised
- **D** unchanged

3.2 What charge passes when a current of 1.5 A flows for 200 s? [1]

- **A** 130 C
- **B** 300 C
- **C** 0.0075 C
- **D** 3.0 C

3.3 A current of 0.50 A flows for 1 hour through silver nitrate solution: $Ag^+ + e^- \rightarrow Ag$. ($F = 96\,500$ C mol^{-1}, A_r Ag = 108.)

(a) Calculate the charge passed. [1]

(b) Calculate the mass of silver deposited at the cathode. [3]

3.4 Explain why electrolysing **aqueous** sodium chloride gives hydrogen at the cathode, not sodium. [2]

Solutions

2.1 cathode: sodium (metal); anode: chlorine.

2.2 $F = Le$.

3.1 B. Cations are reduced at the cathode.

3.2 B. $Q = It = 1.5 \times 200 = 300$ C.

3.3 (a) $Q = It = 0.50 \times 3600 = 1800$ C.
(b) mol $e^- = 1800/96500 = 0.01865$; $Ag^+ + e^- \rightarrow Ag$ so mol Ag $= 0.01865$; $m = 0.01865 \times 108 = 2.0$ g.

3.4 sodium is too reactive (very negative electrode potential), so the water/H^+ is reduced instead, releasing hydrogen.

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.

Co^{2+} = [Ar] .....

Co^{3+} = [Ar] .....

[1]

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

.....

..... [1]

(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.

The colour changes from ..... to ..... .

The state of the cobalt-containing product is ..... .

[2]

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

..... [1]

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

..... [1]

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

..... [1]

(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.

[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}.

E_{cell} = ... [1]

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

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

..... [1]

(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)**.

The ..... electrode is the negative electrode.

Electrons flow from the ..... electrode to the ..... electrode. [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.

time = ... min [3]

[Total: 15]

A-LEVEL CHEMISTRY • EXERCISE SHEET

24.2 Standard electrode potentials and the Nernst equation

Name: _____ Class: _____ Date: _____ **Total: 34 marks**

KEY WORDS electrode potential 电极电势 • standard electrode potential 标准电极电势 • nernst equation 能斯特方程 • standard cell potential 标准电池电势 • standard hydrogen electrode 标准氢电极

Objective

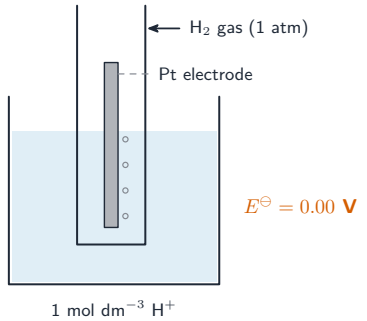
Build the skills to answer exam questions on **standard electrode potentials** 标准电极电势 — the **standard hydrogen electrode** 标准氢电极, **cell potentials** 电池电势, **feasibility** 可行性, and the **Nernst equation** 能斯特方程.

You must be able to:

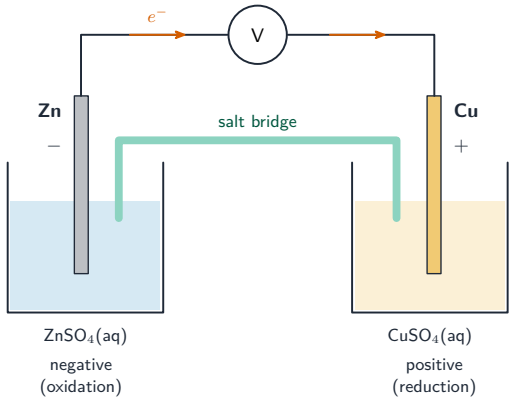
- describe the standard hydrogen electrode and how $E^\ominus$ is measured
- calculate a cell potential and deduce polarity, electron flow and feasibility
- use $E^\ominus$ values to compare oxidising/reducing agents and construct redox equations

1 Worked examples

■ Electrode and cell potentials



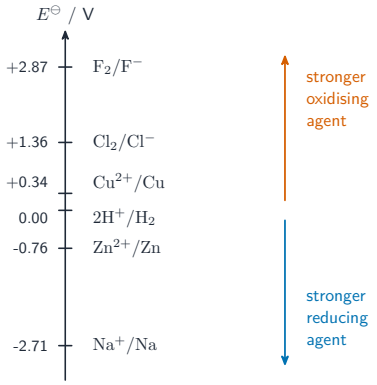
$E^\ominus$ is measured against the standard hydrogen electrode (defined as 0.00 V), under standard conditions, as a reduction



$E^\ominus_{cell} = E^\ominus(\textit{more positive}) - E^\ominus(\textit{less positive})$. *Electrons flow from the more negative electrode*

Example. Zn^{2+}/Zn (−0.76 V) + Cu^{2+}/Cu (+0.34 V): $E^\ominus_{cell} = (+0.34) - (-0.76) = +1.10$ V.

■ Feasibility and the series



A more positive $E^\ominus$ = better oxidising agent; more negative = better reducing agent

A reaction is feasible when $E^\ominus_{cell}$ is **positive**. ($\Delta G^\ominus = -nE^\ominus_{cell}F$.) The **Nernst equation** corrects for non-standard concentrations: $E = E^\ominus + \dfrac{0.059}{z}\log\dfrac{[ox]}{[red]}$.

■ Reading a table of $E^\ominus$ values

A **standard electrode potential** $E^\ominus$ is measured against the **standard hydrogen electrode** ($E^\ominus$ = 0.00 V) under standard conditions: **298 K, 1 bar, and all solutions at 1 mol dm^{-3}**. Quoting those three conditions is routinely a mark.

What the sign tells you. A **more positive** $E^\ominus$ means the species on the LEFT of the half-equation is a **better oxidising agent** (more readily reduced). A **more negative** $E^\ominus$ means the species on the RIGHT is a **better reducing agent**.

Will a reaction happen? Combine the two half-cells:

$$E^\ominus_{cell} = E^\ominus(\text{reduction, the more positive}) - E^\ominus(\text{oxidation, the more negative})$$

A **positive** $E^\ominus_{cell}$ means the reaction is **feasible** — thermodynamically. △ Feasible is not the same as fast: a positive value says nothing about rate, and a reaction with a large activation energy may not be observed at all. That caveat is a mark in almost every "suggest why no reaction was seen" question.

Non-standard conditions, by Le Chatelier. Write the half-equation and treat it as an equilibrium. Increasing the concentration of the **oxidised** species (the left-hand side) shifts it right, so the electrode potential becomes **more positive**; increasing the **reduced** species makes it more negative.

Worked example. Given $E^\ominus(Zn^{2+}/Zn) = -0.76$ V and $E^\ominus(Cu^{2+}/Cu) = +0.34$ V, deduce the cell reaction and $E^\ominus_{cell}$.

1. The more positive electrode is reduced: $Cu^{2+} + 2e^- \rightarrow Cu$.
2. The more negative is oxidised, so its equation is reversed: $Zn \rightarrow Zn^{2+} + 2e^-$.
3. Overall: $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$.
4. $E^\ominus_{cell} = +0.34 - (-0.76) = +1.10$ V, positive, so the reaction is feasible.

2 Practice

2.1 Describe the standard hydrogen electrode. [2]

2.2 State the condition on $E^\ominus_{cell}$ for a reaction to be feasible. [1]

2.3 State the three standard conditions for measuring $E^{\ominus}$. [3]

2.4 Explain why a positive $E^{\ominus}_{\text{cell}}$ does not guarantee that a reaction is observed. [2]

2.5 State the effect on $E^{\ominus}(\text{Fe}^{3+}/\text{Fe}^{2+})$ of increasing $[\text{Fe}^{3+}]$, and give your reasoning. [2]

3 Exam-style questions

3.1 A more **positive** standard electrode potential indicates a: [1]

- **A** stronger reducing agent
- **B** stronger oxidising agent
- **C** more reactive metal
- **D** higher concentration

3.2 Increasing the concentration of the oxidised species in a half-cell makes its electrode potential: [1]

- **A** more positive
- **B** more negative
- **C** unchanged
- **D** zero

3.3 Use these $E^{\ominus}$ values: $\text{Fe}^{2+}/\text{Fe} = -0.44$ V; $\text{Ag}^{+}/\text{Ag} = +0.80$ V.

(a) Calculate the standard cell potential of a cell made from these half-cells. [2]

(b) Deduce which electrode is the negative terminal and the direction of electron flow. [2]

3.4 Using the values above, state whether silver can oxidise iron metal, and explain using $E^{\ominus}_{\text{cell}}$. [2]

3.5 Use the standard electrode potentials below.

half-equation	$E^{\ominus}$ / **V**
$\text{MnO}_4^- + 8\text{H}^+ + 5e^- \rightleftharpoons \text{Mn}^{2+} + 4\text{H}_2\text{O}$	+1.52
$\text{Fe}^{3+} + e^- \rightleftharpoons \text{Fe}^{2+}$	+0.77
$\text{Sn}^{4+} + 2e^- \rightleftharpoons \text{Sn}^{2+}$	+0.15
$\text{V}^{3+} + e^- \rightleftharpoons \text{V}^{2+}$	−0.26

(a) Identify the strongest oxidising agent and the strongest reducing agent in the table, giving your reasoning. [3]

(b) Deduce whether Fe^{3+} will oxidise Sn^{2+}. Write the overall equation and calculate $E^{\ominus}_{cell}$. [4]

(c) Write the overall ionic equation for the reaction of acidified MnO_4^- with Fe^{2+}, and calculate $E^{\ominus}_{cell}$. [4]

(d) A student mixes V^{3+} and Sn^{2+} and observes no change. Use the data to explain whether a reaction is feasible, and give one reason a feasible reaction might still not be observed. [3]

(e) The Fe^{3+}/Fe^{2+} half-cell is made with $[Fe^{3+}] = 2.0\,\text{mol dm}^{-3}$ and $[Fe^{2+}] = 1.0\,\text{mol dm}^{-3}$. State and explain the effect on its electrode potential. [2]

Solutions

2.1 hydrogen gas at 1 atm bubbled over a platinum electrode in 1 mol dm^{-3} H^+, defined as 0.00 V.

2.2 it must be positive.

2.3 298 K; a pressure of 1 bar for any gas; all solutions at a concentration of 1 mol dm^{-3}.

2.4 $E^{\ominus}_{cell}$ describes only the thermodynamic feasibility, not the rate; a large activation energy can make the reaction immeasurably slow, so nothing is seen.

2.5 It becomes more positive; Fe^{3+} is on the left of the half-equation, so raising its concentration shifts the equilibrium to the right by Le Chatelier's principle.

3.1 B. A more positive $E^{\ominus}$ is a stronger oxidising agent.

3.2 A. More oxidised species makes the potential more positive.

3.3 (a) $E^{\ominus}_{cell} = (+0.80) - (-0.44) = +1.24$ V.
(b) the Fe^{2+}/Fe electrode (more negative) is the negative terminal; electrons flow from the iron to the silver electrode in the external circuit.

3.4 yes; $E^{\ominus}_{cell}$ for Ag^+ oxidising Fe is positive (+1.24 V), so the reaction is feasible.

3.5 (a) The strongest oxidising agent is MnO_4^-, because it has the most positive $E^{\ominus}$ and is most readily reduced; the strongest reducing agent is V^{2+}, because its half-cell has the most negative $E^{\ominus}$.
(b) $E^{\ominus}_{cell} = +0.77 - 0.15 = +0.62$ V, which is positive, so the reaction is feasible; $2Fe^{3+} + Sn^{2+} \rightarrow 2Fe^{2+} + Sn^{4+}$.
(c) $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$; $E^{\ominus}_{cell} = +1.52 - 0.77 = +0.75$ V.
(d) For Sn^{2+} to reduce V^{3+}: $E^{\ominus}_{cell} = -0.26 - 0.15 = -0.41$ V, which is negative, so that reaction is not feasible. Even where a value is positive, a high activation energy can make the reaction too slow to observe.
(e) It becomes more positive; the oxidised species is in excess, so by Le Chatelier's principle the equilibrium shifts to the right, favouring reduction.

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

Explain why transition elements have variable oxidation states.

.....

..... [1]

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

Define complex ion.

.....

..... [1]

(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.

reagent .....

type of reaction .....
[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.

.....

.....

.....

.....

.....

..... [3]

(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.

$\Delta G^{\ominus}$ = ... $kJ\,mol^{-1}$ [2]

[Total: 8]

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

.....

.....

..... [2]

(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.

[3]

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

the solutions ..... the wire ..... [1]

(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.

E $(Zn^{2+}(aq)/Zn(s))$ = ... V [2]

(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.

from ..... to ..... [1]

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

..... [1]

(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.

$E^{\ominus}$ (MnO_2/Mn_2O_3) = ... V [1]

[Total: 11]

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.

Co^{2+} = [Ar] .....

Co^{3+} = [Ar] .....

[1]

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

.....

..... [1]

(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.

The colour changes from ..... to ..... .

The state of the cobalt-containing product is ..... .

[2]

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

..... [1]

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

..... [1]

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

..... [1]

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

$Co^{2+} + 2e^- \rightleftharpoons Co$ $\quad E^\ominus = -0.28\,V$

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

(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.

[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}.

E_{cell} = ... [1]

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

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

..... [1]

(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)**.

The ..... electrode is the negative electrode.

Electrons flow from the ..... electrode to the ..... electrode.

[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.

time = ... min [3]

[Total: 15]