REDOX

PAST EXAM QUESTIONS

Section 1: Principles of oxidation and reduction

WACE 2AB 2013 Q19:

If a substance is oxidised, it means that

(a) electrons have been transferred to it
(b) it has lost electrons
(c) it has produced oxygen in a reaction
(d) it can now donate oxygen to another substance

WACE 2AB 2012 Q9:

The oxidation number of nitrogen in the series, N_2, N_2O, NO_3^-, and NH_3 respectively are:

(a)	0,	0,	+5,	–3
(b)	0,	–1,	–1,	+3
(c)	**0,**	**+1,**	**+5,**	**–3**
(d)	+2,	+2,	+1,	+1

WACE 2AB 2011 Q10:

Copper(II) oxide undergoes a redox reaction with carbon to produce copper metal and carbon monoxide gas, as shown in the equation below.

$$CuO(s) + C(s) \rightarrow Cu(s) + CO(g)$$

Which one of the following is the reductant in this reaction?

(a) copper(II) oxide
(b) carbon
(c) copper
(d) carbon monoxide

WACE 2AB 2011 Q11:

Consider the following reaction between manganese and lead ions.

$$Mn(s) + Pb^{2+}(aq) \rightarrow Mn^{2+}(aq) + Pb(s)$$

Which one of the following statements is correct?

(a) The manganese metal has removed electrons from the lead ions
(b) The lead ions have reduced the manganese metal
(c) Electrons have been transferred from the manganese metal to the lead ions
(d) The oxidation number of the lead ions has increased

Section 1: Principles of oxidation and reduction

WACE 2AB 2012 Q12:

Which one of the following reactions is **not** a redox reaction?

(a) $4\ K(s) + O_2(g) \rightarrow 2\ K_2O(s)$
(b) $2\ Na(s) + 2\ H_2O(\ell) \rightarrow 2\ NaOH(aq) + H_2(g)$
(c) $Li_2O(s) + H_2O(\ell) \rightarrow 2\ LiOH(aq)$
(d) $2\ Na(s) + H_2(\ell) \rightarrow 2\ NaH(s)$

WACE 2AB 2011 Q12:

Which one of the four elements (S, C, Cℓ, Cr) shown in **bold** below has the **lowest** oxidation number?

(a) $KH\mathbf{S}O_4$
(b) $Na_2C_2O_4$
(c) $HO\mathbf{C\ell}$
(d) $K_2\mathbf{Cr}_2O_7$

WACE 2AB 2010 Q20:

Zn, CO and H_2 are all reducing agents. Which one of the following statements **best** explains why this is so?

(a) They can all readily donate electrons
(b) They are all difficult to oxidise
(c) None of the them can be reduced
(d) They all react with oxygen

WACE 2AB 2010 Q21:

Which one of the following represents oxidation?

(a) $Cr_2O_7^{2-} \rightarrow CrO_4^{2-}$
(b) $Mn^{2+} \rightarrow MnO_4^-$
(c) $SO_4^{2-} \rightarrow SO_2$
(d) $NO_2 \rightarrow NO_2^-$

WACE 2AB 2010 Q22:

In which one of the following compounds is nitrogen in the highest oxidation state?

(a) N_2
(b) N_2O_5
(c) NO
(d) N_2O_4

WACE 2016 Sample Exam Q12: ***Originally from WACE 3AB 2011 Q21***

How many moles of electrons must be exchanged to oxidise 1 mole of hypophosphorous acid, H_3PO_2, to phosphoric acid, H_3PO_4?

(a) 2
(b) 3
(c) 4
(d) 5

WACE 2016 Sample Exam Q13: ***Originally from WACE 3AB 2011 Q22***

In which one of the following compounds is rhenium (Re) in the highest oxidation state?

(a) $NaReO_4$
(b) $ReC\ell O$
(c) Re_2O_3
(d) $ReC\ell_5$

WACE 2016 Sample Exam Q14: ***Originally from WACE 3AB 2010 Q12***

In which one of the following will a metal displacement reaction occur?

(a) A zinc rod is dipped in a 1.0 mol L^{-1} solution of sodium sulfate
(b) A copper rod is dipped in a 1.0 mol L^{-1} solution of cobalt(II) nitrate
(c) A silver rod is dipped in a 1.0 mol L^{-1} solution of gold(III) nitrate
(d) A tin rod is dipped in a 1.0 mol L^{-1} solution of manganese(II) sulfate

WACE 3AB 2014 Q17:

The equation below shows the reaction when sulfur dioxide gas (SO_2) is bubbled through a solution containing hypochlorite ions ($C\ell O^-$).

$$C\ell O^-(aq) + SO_2(g) + H_2O(\ell) \rightarrow C\ell^-(aq) + SO_4^{2-}(aq) + 2\,H^+(aq)$$

Which of the following statements about this reaction is correct?

i. The sulfur in sulfur dioxide is oxidised
ii. The hydrogen in water is reduced
iii. The hypochlorite ion is the reducing agent
iv. The water is the reducing agent

(a) i only
(b) ii and iii only
(c) iii and iv only
(d) i and iv only

Section 1: Principles of oxidation and reduction

WACE 2AB 2010 Q23:

Methanol, which can be used in fuel cells to provide power for laptops, can be produced by the following redox reaction:

$$CO(g) \; + \; 2\,H_2(g) \; \rightarrow \; CH_3OH(\ell)$$

Which one of the following statements is true?

(a) In this reaction hydrogen is the oxidising agent and carbon monoxide is the reducing agent
(b) The oxidation number of carbon changes from +2 to -2
(c) The oxygen has been reduced
(d) The hydrogen has been reduced

WACE 3AB 2015 Q1:

In which of the following compounds is the oxidation number of manganese the **lowest**?

(a) Mn_2O_3
(b) K_2MnO_4
(c) $NaMnO_4$
(d) MnO_2

WACE 3AB 2015 Q5:

Which one of the following reactions can occur spontaneously at 25.0 °C? (Assume the solutions have a concentration of 1.00 mol L^{-1}.)

(a) $C(s)$	+	$O_2(g)$	→	$CO_2(g)$		
(b) $2\,Fe^{2+}(aq)$	**+**	**$C\ell_2(g)$**	**→**	**$2\,Fe^{3+}(aq)$**	**+**	**$2\,C\ell^-(aq)$**
(c) $Zn^{2+}(aq)$	+	$Cu(s)$	→	$Zn(s)$	+	$Cu^{2+}(aq)$
(d) $Cu(s)$	+	$2\,H^+(aq)$	→	$Cu^{2+}(aq)$	+	$H_2(g)$

WACE 3AB 2013 Q9:

Consider the following reaction.

$$2\,H^+(aq) \; + \; 2\,NO_3^-(aq) \; + \; H_2S(g) \; \rightarrow \; 2\,NO_2(aq) \; + \; S(s) \; + \; 2H_2O(\ell)$$

Which one of the following statements is true for this reaction?

(a) The H^+ is oxidised and the H_2O is reduced
(b) The oxidation number of the nitrogen changes from +5 to +4 during the reaction
(c) The sulfur is reduced in the reaction
(d) This is an acid-base reaction, not a redox reaction

Section 1: Principles of oxidation and reduction

WACE 3AB 2014 Q18:

Consider the following mixtures:

i. Solid I_2 added to a solution of H_2S
ii. Liquid Br_2 is added to a solution of Fe^{2+}
iii. Freshly exposed Aℓ metal is added to a solution of HCℓ
iv. A piece of cobalt metal is placed in an aqueous solution of Cr^{3+}

Based on E° values, in which of the above mixtures will a chemical reaction occur?

(a) ii only
(b) i and iii only
(c) i, ii and iii only
(d) iii and iv only

WACE 3AB 2012 Q9:

Identify the oxidant in the following reaction:

$$2\ A\ell(s)\ +\ 2\ Cr_2O_3(s)\ \rightarrow\ A\ell_2O_3(s)\ +\ 2\ Cr(s)$$

(a) Aℓ
(b) Cr_2O_3
(c) O
(d) Cr

WACE 3AB 2012 Q11:

Which one of the reactions below is **most** likely to occur spontaneously?

(a) $H_2(g)$	+	$PbSO_4(s)$	→	$2\ H^+(aq)$	+	$Pb(s)$	+	$SO_4^{2-}(aq)$
(b) $Zn^{2+}(aq)$	+	$Fe(aq)$	→	$Zn(s)$	+	$Fe^{2+}(aq)$		
(c) $Cu^{2+}(aq)$	**+**	**$Ni(s)$**	**→**	**$Cu(s)$**	**+**	**$Ni^{2+}(aq)$**		
(d) $2\ Fe^{3+}(aq)$	+	$H_2O_2(aq)$	→	$O_2(g)$	+	$2\ Fe^{2+}(aq)$	+	$2\ H^+(aq)$

WACE 3AB 2011 Q23:

Corrosion is a redox process. Which one of the following explains why coating iron with nickel protects the iron from corrosion?

(a) Nickel accepts electrons from iron.
(b) Iron and nickel form an alloy that is particularly resistant to redox processes.
(c) Nickel is a stronger oxidising agent than iron.
(d) The thin coating of nickel prevents iron from reacting.

Use the table of standard reduction potentials in the Chemistry Data Sheet to answer WACE 2013 Questions 10 and 11.

WACE 3AB 2013 Q10:

Predict in which of the following a reaction would occur. Assume all solutions are 1.0 mol L^{-1}.

i. Acidified potassium permanganate is mixed with potassium iodide solution
ii. Chlorine gas is bubbled through hydrogen sulfide solution
iii. Acidified potassium dichromate is mixed with potassium fluoride
iv. An iron(II) sulfate solution is placed in a nickel container
v. A piece of copper metal is placed in a hydrochloric acid solution

(a) i and ii only
(b) i, ii and iii only
(c) ii and v only
(d) ii, iii and v only

Use the following additional table of standard reduction potentials to answer Question 11.

Half-reaction	**E° (V)**
$2\ HC\ell O_2 + 6\ H^+ + 6\ e^- \rightarrow C\ell_2(g) + 4\ H_2O$	1.64
$2\ HOC\ell + 2\ H^+ + 2e^- \rightarrow C\ell_2(g) + 2\ H_2O$	1.63
$2\ C\ell O_3^- + 12\ H^+ + 10\ e^- \rightarrow C\ell_2(g) + 6\ H_2O$	1.47
$2\ C\ell O_4^- + 16\ H^+ + 14\ e^- \rightarrow C\ell_2(g) + 8\ H_2O$	1.42
$C\ell O^- + H_2O + 2\ e^- \rightarrow C\ell^- + 2\ OH^-$	0.89
$C\ell O_2^- + 2\ H_2O + 4\ e^- \rightarrow C\ell^- + 4\ OH^-$	0.78

Which of the following <u>cannot</u> react with hydrogen peroxide to produce oxygen gas?

(a) $C\ell O^-$, $C\ell O_2^-$, $C\ell O_3^-$ and $C\ell O_4^-$
(b) $HC\ell O_2$ and $HOC\ell$
(c) $HC\ell O_2$, $HOC\ell$, $C\ell O_3^-$ and $C\ell O_4^-$
(d) All can react with hydrogen peroxide to produce oxygen gas

WACE 3AB 2010 Q10:

Which one of the four elements (Cℓ, Cr, P, Mn) underlined below has the **lowest** oxidation state?

(a) $H\underline{C\ell}O_2$
(b) $K_2\underline{\mathbf{Cr}}O_4$
(c) $Na_3\underline{\mathbf{P}}O_4$
(d) $K\underline{\mathbf{Mn}}O_4$

Section 1: Principles of oxidation and reduction

WACE 3AB 2010 Q11:

Which of the following are redox reactions?

i. $N_2 + 3\,H_2 \rightarrow 2\,NH_3$
ii. $CaCO_3 + 2\,HC\ell \rightarrow CaC\ell_2 + H_2O + CO_2$
iii. $Zn + 2\,HC\ell \rightarrow ZnC\ell_2 + H_2$
iv. $S + O_2 \rightarrow SO_2$

(a) Equations i, ii, iii and iv
(b) Equations i, iii and iv only
(c) Equations ii, iii and iv only
(d) Equations iv only

WACE 3AB 2010 Sample Q14:

Which of the four elements (Fe, S, O or Cr) shown in bold below has the highest oxidation state?

(a) $\mathbf{Fe}C\ell_3$
(b) Na_2SO_3
(c) $K_2Cr_2\mathbf{O}_7$
(d) $\mathbf{Cr}_2O_3$

WACE 3AB 2010 Sample Q16:

Chlorine gas is bubbled through a solution of a salt, and the solution turns brown. A separate solution of the same salt is added to a solution of lead(II) nitrate, and a bright yellow precipitate is formed. Which of the following is the most likely identity of the ion causing the colour changes?

(a) Iron(III)
(b) Bromide
(c) Iodide
(d) Chromate

TEE 2009 Q20:

In which one of the following reactions is the manganese-containing species acting as a reducing agent?

(a) $MnO + Mg \rightarrow Mn + MgO$
(b) $MnC\ell_2 + 2\,H_2O + C\ell_2 \rightarrow MnO_2 + 4\,C\ell^- + 4H^+$
(c) $MnO_2 + 2\,Ag + 4\,H^+ \rightarrow Mn^{2+} + 2\,Ag^+ + 2\,H_2O$
(d) $MnO_4^- + 5\,Fe^{2+} + 8\,H^+ \rightarrow Mn^{2+} + 5\,Fe^{3+} + 4\,H_2O$

Section 1: Principles of oxidation and reduction

TEE 2007 Q17:

Consider the following equation:

$$2\ BrO_3^-(aq)\ +\ 10\ I^-(aq)\ +\ 12\ H^+(aq)\ \rightarrow\ 5\ I_2(aq)\ +\ Br_2(\ell)\ +\ 6\ H_2O(\ell)$$

For this reaction, which one of the following is true?

(a) BrO_3^- is the reducing agent
(b) H^+ is reduced
(c) I^- is the oxidising agent
(d) BrO_3^- is reduced

TEE 2007 Q18:

Consider the statements about the following reaction:

$$2\ H_2O_2(\ell)\ \rightarrow\ 2\ H_2O(\ell)\ +\ O_2(g)$$

i. H_2O_2 is reduced
ii. H_2O_2 is oxidised
iii. H_2O_2 acts as a reducing agent
iv. This is not a redox reaction

Which of the above statements are true?

(a) iv only
(b) ii and iii only
(c) i only
(d) i, ii and iii only

TEE 2007 Q19:

Consider the following reaction:

$$2\ VO_2^+\ +\ H_2O_2\ +\ 2\ H^+\ \rightarrow\ 2\ VO^{2+}\ +\ O_2\ +\ 2\ H_2O$$

Which one of the following statements is true for this reaction?

(a) The VO_2^+ is reduced and the H^+ is oxidised
(b) The H^+ is reduced and the H_2O_2 is oxidised
(c) The VO_2^+ is the oxidising agent and the H^+ is the reducing agent
(d) The VO_2^+ is reduced and the H_2O_2 is oxidised

Section 1: Principles of oxidation and reduction

TEE 2007 Q30:

Unlike structure made of iron, structures made of aluminium often do not need to be protected from corrosion. Why is aluminium more resistant to corrosion than iron?

(a) Aluminium is unreactive
(b) Aluminium has a high reduction potential
(c) Aluminium forms a protective oxide layer
(d) Iron can be alloyed with other elements

TEE 2005 Q16:

What is the change in the oxidation number of bromine in the following reaction?

$$2\,BrO_3^-(aq) \;+\; 10\,I^-(aq) \;+\; 12\,H^+(aq) \;\rightarrow\; 5\,I_2(aq) \;+\; Br_2(aq) \;+\; 6\,H_2O(\ell)$$

(a) +7 to 0
(b) +7 to -1
(c) +5 to 0
(d) +5 to -1

TEE 2005 Q17:

In which one of the following pairs do the underlined elements have the same oxidation number?

(a) $\underline{Cr}_2O_7^{2-}$ and $\underline{Cr}O_4^{2-}$

(b) $\underline{H}F$ and $Mg\underline{H}_2$

(c) $H_2\underline{O}$ and $H_2\underline{O}_2$

(d) $\underline{N}O_2$ and $H\underline{N}O_3$

TEE 2004 Q8:

For the reaction:

$$3\,Au^+(aq) \;\rightarrow\; Au^{3+}(aq) \;+\; 2\,Au(s)$$

which one of the following statements is **true**?

(a) $Au^+(aq)$ disproportionates
(b) $Au^{3+}(aq)$ is a reducing agent
(c) Au(s) is oxidised
(d) None of (a), (b) or (c) is true

Section 1: Principles of oxidation and reduction

TEE 2004 Q11:

Which one of the following pairs of substances will react together when mixed?

(a) $C\ell_2$ + Cu^{2+}
(b) Ni^{2+} + Cu
(c) Ni^{2+} + Zn
(d) Zn + $C\ell^-$

TEE 2003 Q15:

Which one of the following arranges the substances in order of strongest to weakest reducing agent?

(a) $C\ell_2$ > $A\ell^{3+}$ > Na^+
(b) $C\ell_2$ > Si > Mg
(c) Mg > $A\ell$ > $C\ell^-$
(d) Na^+ > $A\ell^{3+}$ > $C\ell^-$

TEE 2003 Q21:

Which of the following equations does **not** represent an oxidation-reduction reaction?

(a) $2\ MnO_4^- + 2\ H_2O + 3\ C_2O_4^{2-} \rightarrow 2\ MnO_2 + 6\ CO_3^{2-} + 4\ H^+$
(b) $Cr_2O_7^{2-} + H_2O \rightarrow 2\ CrO_4^{2-} + 2\ H^+$
(c) $2\ Br_2 + N_2H_5^+ \rightarrow N_2 + 5\ H^+ + 4\ Br^-$
(d) $6\ I^- + 14\ H^+ + Cr_2O_7^{2-} \rightarrow 3\ I_2 + 7\ H_2O + 2\ Cr^{3+}$

TEE 2003 Q30:

For the following equation

$$H_2SO_3 + H_2O_2 \rightarrow H_2SO_4 + H_2O$$

which one of the following statements is true?

(a) Hydrogen peroxide is acting as an acid
(b) Hydrogen peroxide is acting as an acid and a base
(c) Hydrogen peroxide is acting as an oxidising agent only
(d) Hydrogen peroxide is acting as an oxidising and reducing agent

<u>Section 1: Principles of oxidation and reduction</u>

VCE 2015 Q6:

In which one of the following compounds is sulfur in its lowest oxidation state?

(a) SO_3
(b) HSO_4^-
(c) SO_2
(d) $A\ell_2S_3$

VCE 2015 Q24:

The reaction between hydrogen peroxide and ammonium ions is represented by the following equation.

$$3\ H_2O_2(aq) \ + \ 2\ NH_4^+(aq) \ \rightarrow \ N_2(g) \ + \ 2\ H^+(aq) \ + \ 6\ H_2O(\ell)$$

Which one of the following is the correct half-equation for the reduction reaction?

(a) $H_2O_2(aq) \ + \ 2\ H^+(aq) \ + \ 2\ e^- \ \rightarrow \ 2\ H_2O(\ell)$
(b) $2\ NH_4^+(aq) \ \rightarrow \ N_2(g) \ + \ 8\ H^+(aq) \ + \ 6\ e^-$
(c) $2\ NH_4^+(aq) \ + \ 2\ e^- \ \rightarrow \ N_2(g) \ + \ 4\ H_2(g)$
(d) $H_2O_2(aq) \ + \ 2\ H_2O(\ell) \ \rightarrow \ 2\ O_2(g) \ + \ 6\ H^+(aq) \ + \ 6\ e^-$

VCE 2014 Q10:

Which one of the reactions of hydrochloric acid below is a redox reaction?

(a) $2\ HC\ell(aq) \ + \ Fe(s) \ \rightarrow \ H_2(g) \ + \ FeC\ell_2(aq)$
(b) $2\ HC\ell(aq) \ + \ Na_2S(s) \ \rightarrow \ H_2S(g) \ + \ 2\ NaC\ell(aq)$
(c) $2\ HC\ell(aq) \ + \ MgO(s) \ \rightarrow \ MgC\ell_2(aq) \ + \ H_2O(\ell)$
(d) $2\ HC\ell(aq) \ + \ K_2CO_3(s) \ \rightarrow \ CO_2(g) \ + \ 2\ KC\ell(aq) \ + \ H_2O(\ell)$

VCE 2014 Q26:

Consider the following experiments that are carried out under standard conditions.

Beaker I	A strip of nickel metal is placed into a 1.0 mol L^{-1} silver nitrate solution
Beaker II	A 1.0 mol L^{-1} copper(II) sulfate solution is added to a 1.0 mol L^{-1} sodium iodide solution
Beaker III	Chlorine gas is bubbled through a 1.0 mol L^{-1} potassium iodide solution

It would be predicted that a reaction will occur in:

(a) Beaker I only
(b) Beaker II only
(c) Beakers I and III only
(d) Beakers II and III only

VCE 2014 Q25:

Consider the following information about the reaction of Ru^{2+} with various reagents.

$Ru^{2+}(aq)$	+	$Fe^{2+}(aq)$	→	no observed reaction		
$Ru^{2+}(aq)$	+	$Ni(s)$	→	$Ru(s)$	+	$Ni^{2+}(aq)$
$Ru^{2+}(aq)$	+	$Ag(s)$	→	no observed reaction		
$Ru^{2+}(aq)$	+	$Cu(s)$	→	$Ru(s)$	+	$Cu^{2+}(aq)$

Where would the following reaction be placed in the electrochemical series if the above tests were carried out under standard conditions?

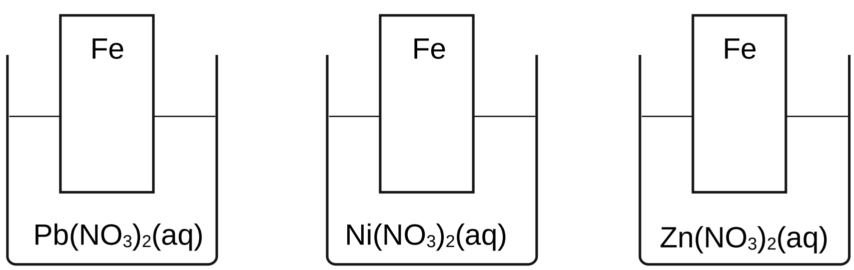
$$Ru^{2+}(aq) + 2e^- \rightarrow Ru(s)$$

(a) Below –0.23 V
(b) Between –0.44 V and –0.23 V
(c) Between 0.77 V and 0.34 V
(d) Above 0.77 V

$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$ $E° = 0.80$ V
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$ $E° = 0.77$ V
$\mathbf{Ru^{2+}(aq) + 2e^- \rightarrow Ru(s)}$ $\mathbf{E° = ?}$
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$ $E° = 0.34$ V
$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$ $E° = –0.23$ V

VCE 2013 Q24:

Three beakers, each containing an iron strip and a 1.0 M solution of a metal salt, were set up as following:



A reaction will occur in beaker(s):

(a) I and II only
(b) I and III only
(c) II and III only
(d) III only

WACE 2AB 2014 Q37:

A student placed an iron (Fe) nail into an aqueous solution of copper(II) sulfate ($CuSO_4$) in a beaker. Her observations are recorded below.

	Original appearance	**Appearance after 1 minute**	**Appearance after 10 minutes**
Iron nail	shiny grey solid metal	the part of the nail in the solution turns a darker colour	the part of the nail in the solution has a salmon pink coloured coating
Copper(II) sulfate solution	bright blue solution	no change	there are some salmon pink to brown coloured solid lumps at the bottom of the beaker. The blue colour of the solution has faded slightly.

(a) Name the substance that forms on the surface of the iron nail. (1 mark)

Copper metal

(b) Write a half-equation showing the reaction of Cu^{2+} ions that occurs at the surface of the nail. Include state symbols. (2 marks)

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

(c) Use this half-equation to explain why the blue colour of the solution fades. (2 marks)

The blue colour is caused by Cu^{2+} ions (1 mark)
The concentration of Cu^{2+} ions decreases as they are consumed in the reaction (1 mark)

(d) The student conducting the investigation had a second beaker of copper(II) sulfate that she used as a control so she did not add a nail to this beaker. Explain why a control was used. (2 marks)

So a comparison can be made with the only difference being the presence of the nail (1 mark)

It allows for the identification of the colour change of the solution. Any change in appearance can then be conclusively attributed to the presence of the iron nail (1 mark)

WACE 3AB 2015 Q31:

(b) Describe a chemical test that can be used to distinguish between magnesium solid and cobalt solid. State the observations expected for each of the solids when test. (3 mark)

Example Answer #1:

- **Add 1.0 mol/L HCℓ to each metal**
- **Observation with Mg: bubbles of colourless gas are formed**
- **Observation with Co: bubbles of colourless gas are formed, solution turns pink**

Example Answer #1:

- **Add a 1.0 mol/L solution of $Cr(NO_3)_3(aq)$**
- **Observation with Mg: green colour of solution fades**
- **Observation with Co: no visible reaction**

Notes:

- **Accept any chemical valid tests. e.g. metal displacement**
- **The test must be described. Just stating the name of a chemical (e.g. hydrochloric acid) does not justify the description of a test and so no mark should be awarded**

Examiner's comments: Nearly 5% of candidates did not attempt to answer part (b). Many candidates were aware that the magnesium and cobalt ions produce different coloured solutions but chose to apply an unsuitable chemical test and so could not be awarded marks. In describing the observations candidates tended to focus on the distinguishing observation rather than provide the full observation as required by the question. There is a large range of responses possible for this type of question. Some candidates used acronyms (COG, NVR, etc.); unless they are stated in the syllabus, acronyms are not accepted.

WACE 3AB 2015 Q37:

Sulfur compounds in sewerage and industrial processes can cause problems due to their odours, often because they eventually form dihydrogen sulfide gas, also known as rotten egg gas.

One class of sulfur compounds that need to be removed from sewerage is the thiosulfates. One step in their removal is the reaction of tetrathionate ions, $S_4O_6^{2-}$, with hydrogen peroxide, H_2O_2. The tetrathionate produces trithionate ions, $S_3O_6^{2-}$, and sulfate ions.

(a)

i. Complete the table below by writing balanced half-equations and the final redox equation for the reaction of tetrathionate and hydrogen peroxide. (6 marks)

Half-equation one $4\ H_2O + S_4O_6^{2-} \rightarrow S_3O_6^{2-} + SO_4^{2-} + 8\ H^+ + 6\ e^-$
Half-equation two $2e^- + 2\ H^+ + H_2O_2 \rightarrow 2\ H_2O$
Redox $S_4O_6^{2-} + 3\ H_2O_2 \rightarrow S_3O_6^{2-} + SO_4^{2-} + 2\ H_2O + 2\ H^+$

Examiner's comments: Part (a)(i) was done poorly by many candidates. Candidates should be encouraged to practice the skills required to constant and combine half-equations. The half-equation for the reduction of hydrogen peroxide could be either determined from first principles or copied from the data sheet.

ii. Which substance is being oxidised? (1 mark)

Tetrathionate or $S_4O_6^{2-}$

Examiner's comments: Part (a)(ii) was done well generally but some candidates provided the atom found within the reacting species as their answer rather than the reacting substance as asked.

WACE 3AB 2014 Q27:

Write balanced ionic equations to represent the reactions described below.

(a) Chlorine gas is bubbled through an aqueous solution of sodium bromide. (2 marks)

$C\ell_2(g) + 2\ Br^-(aq) \rightarrow 2\ C\ell^-(aq) + Br_2(aq)$

Examiner's comments: For part (a) some candidates are still writing molecular equations rather than ionic equations.

17

WACE 3AB 2014 Q32:

Nitrogen gas from the atmosphere undergoes a series of redox reactions to transform it into nitrate ions that are absorbed by plants. The process can be simplified into the following three steps.

Step 1 – Nitrogen-fixing soil bacteria reduce nitrogen gas to ammonium ions
Step 2 – Nitrifying bacteria then oxidise ammonium ions to nitrite ions
Step 3 – Nitrifying bacteria then oxidise nitrite ions to nitrate ions

Write the half-equations for each of these steps. Assume acidic conditions.

Step 1	$N_2 + 8H^+ + 6e^- \rightarrow 2NH_4^+$
Step 2	$NH_4^+ + 2H_2O \rightarrow NO_2^- + 8H^+ + 6e^-$
Step 3	$NO_2^- + H_2O \rightarrow NO_3^- + 2H^+ + 2e^-$

2 marks per correct half-equation. For an incorrect half-equation, award 1 mark if reactants and products are correct but electrons or balancing incorrect. If atoms are consistently balanced for all half-equations but electrons incorrect award 4 marks.

Examiner's comments: Some candidates could not balance equation even if the formulae for the reactants and products were correct. Some equations contained no electrons at all.

WACE 3AB 2013 Q38:

Ilmenite, a titanium-iron oxide mineral, is used to produce titanium dioxide and metallic titanium. The ilmenite is first reacted with sulfuric acid to convert iron(II) oxide to iron(II) sulfate according to the equation below.

Equation 1: $FeO(s) + H_2SO_4(aq) \rightarrow FeSO_4(aq) + H_2O(\ell)$

The iron(II) sulfate is crystallised and filtered off to give synthetic rutile which still contains some iron combined with the titanium. This synthetic rutile is treated with carbon and chlorine gas to give titanium tetrachloride, as shown in the following equation.

Equation 2: $2FeTiO_3(s) + 7C\ell_2(g) + 6Cs) \rightarrow 2TiC\ell_4(\ell) + 2FeC\ell_3(s) + 6CO(g)$

(a) Given that the titanium remains in the +4 oxidation state throughout the reaction in Equation 2, identify the following. (2 marks)

Substance/s oxidised in Equation 2:

- **Fe^{2+} or $FeTiO_3$ and C**

Substance/s reduced in Equation 2:

- **$C\ell_2$**

Examiner's comments: Although candidates were able to identify iron as being oxidised in (a), they did not specifically state Fe^{2+} or $FeTiO_3$. Candidates must be encouraged to identify the chemical being oxidised or reduced, e.g. $C\ell_2$ not $C\ell$.

WACE 3AB 2012 Q33:

(a) Write the oxidation and reduction half-equations, and the overall redox equation, for this reaction. (6 marks)

Oxidation half-equation	$Mn^{2+}(aq) + 4\,H_2O(\ell) \rightarrow MnO_4^-(aq) + 8\,H^+(aq) + 5\,e^-$
Reduction half-equation	$BiO_3^-(aq) + 6\,H^+(aq) + 2\,e^- \rightarrow Bi^{3+}(aq) + 3\,H_2O(\ell)$
Overall redox equation	$2\,Mn^{2+}(aq) + 5\,BiO_3^-(aq) + 14\,H^+(aq) \rightarrow 2\,MnO_4^-(aq) + 5\,Bi^{3+}(aq) + 7\,H_2O(\ell)$

Examiner's comments: The majority of candidates were able to obtain the oxidation half-equation form the data sheet. The reduction half-equation, which required an interpretation of the written description, was more difficult. Most candidates were able to obtain a full equation form their half-equations. Candidates should be encouraged to check the balance of charge in their final equation.

WACE 3AB 2010 Q35:

Concentrated sulfuric acid can behave as an oxidising agent. Depending upon conditions, it can react in one of three ways to form […]

(a) Write half-equations showing each of these possible three reactions. (3 marks)

(i) $H_2SO_4 + 2\,H^+ + 2e^- \rightarrow SO_2 + 2\,H_2O$
(ii) $H_2SO_4 + 8\,H^+ + 8\,e^- \rightarrow H_2S + 4\,H_2O$
(iii) $H_2SO_4 + 6\,H^+ + 6\,e^- \rightarrow S + 4\,H_2O$

Answers could should reactions of SO_4^{2-} or HSO_4^- instead of H_2SO_4 so long as they are correctly balanced

(b) Write half-equations and an overall redox equation for the reaction between concentrated sulfuric acid and hydrogen iodide to form hydrogen sulfide, iodine and water. (2 marks)

Reduction half-equation: $H_2SO_4 + 8\,H^+ + 8\,e^- \rightarrow H_2S + 4\,H_2O$

Oxidation half-equation: $2\,HI \rightarrow I_2 + 2\,H^+ + 2\,e^-$ OR $2\,I^- \rightarrow I_2 + 2\,e^-$

Overall redox equation: $H_2SO_4 + 8\,HI \rightarrow H_2S + 4\,I_2 + 4\,H_2O$

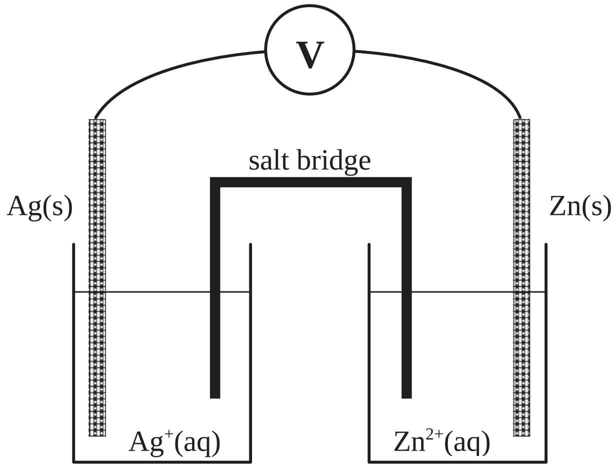
Full marks if incorrect equation form previous question used, but overall equation balanced correctly for that equation

Examiner's comments: This question was very poorly done. Although the reactants and products of the reaction were given in both parts of the question, a very large number of candidates were unable to convert these words to formulae and equations. Candidates also, apparently, experienced difficulty balancing half equations.

Section 2: Electrochemical cells

Use the following information to answer VCE 2012 (Part 2) Questions 16-18

A galvanic cell set up under standard conditions is shown below.



VCE 2012 Part 2 Q16:

Which one of the following is correct?

As the cell discharges:

	electrons would flow from the	**in the salt bridge**
(a)	zinc electrode to the silver electrode	anions migrate to the Ag^+/Ag half-cell
(b)	silver electrode to the zinc electrode	cations migrate to the Zn^{2+}/Zn half-cell
(c)	silver electrode to the zinc electrode	cations migrate to the Ag^+/Ag half-cell
(d)	**zinc electrode to the silver electrode**	**anions migrate to the Zn^{2+}/Zn half-cell**

VCE 2012 Part 2 Q17:

In this cell:

(a) Ag^+(aq) is reduced and the Zn(s) is oxidised
(b) Ag(s) is oxidised and the Zn^{2+}(aq) is reduced
(c) Ag(s) is reduced and the Zn^{2+}(aq) is oxidised
(d) Ag^+(aq) is oxidised and the Zn(s) is reduced

VCE 2012 Part 2 Q18:

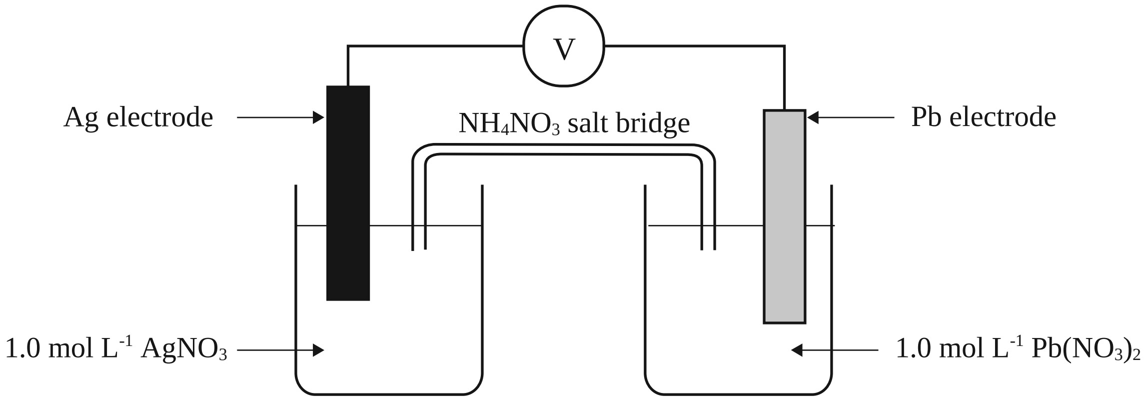
The cathode in this cell and the maximum voltage produced by this cell, under standard conditions, are:

(a) Ag and 0.16 V
(b) Ag and 1.56 V
(c) Zn and 0.16 V
(d) Zn and 1.56 V

<u>Section 2: Electrochemical cells</u>

WACE 3AB 2010 Sample Q15:

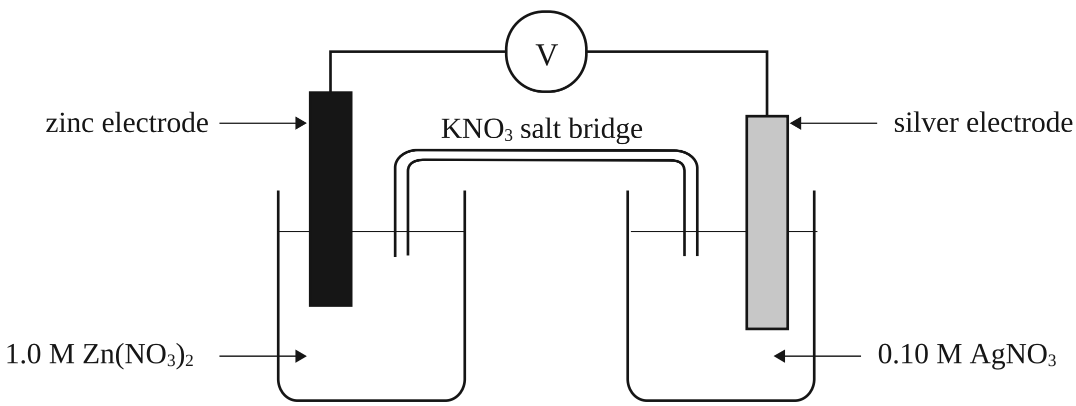
A short rod of silver metal dips into a 1.0 mol L^{-1} solution of silver nitrate to create an Ag/Ag^+ half-cell. Similarly, a rod of lead metal dips into a 1.0 mol L^{-1} solution of lead(II) nitrate. The two rods are joined by a piece of copper wire. A salt bridge of ammonium nitrate, as shown in the diagram below, joins the two solutions.



Which one of the following will occur?

(a) Lead deposits from solution onto the lead rod
(b) Electrons flow through the wire from the lead rod to the silver rod
(c) Nitrate ions migrate through the salt bridge form the Pb/Pb^{2+} half-cell to the Ag/Ag^+ half-cell
(d) The silver rod starts to dissolve

Use the following information to answer VCE 2014 Questions 27 and 28



Section 2: Electrochemical cells

VCE 2014 Q27:

Which one of the following statements about the cell above is true as the cell discharges?

(a) The silver electrode is the anode
(b) The concentration of Zn^{2+} ions will increase
(c) The maximum voltage delivered by this cell will be 1.56 V
(d) Electrons in the external circuit will flow from the silver electrode to the zinc electrode

VCE 2014 Q28:

What should be observed at the zinc electrode as the cell discharges?

(a) No change will be observed at this electrode
(b) The electrode will become thinner and pitted
(c) Crystals will form over the surface of the electrode
(d) Bubbles of gas will form over the surface of the electrode

Use the following information to answer VCE 2013 Questions 26 and 27.

Four standard galvanic cells are set up as indicated below.

Cell I	a Br_2/Br^- standard half-cell connected to a Cu^{2+}/Cu standard half-cell
Cell II	an Sn^{2+}/Sn standard half-cell connected to a Zn^{2+}/Zn standard half-cell
Cell III	a Br_2/Br^- standard half-cell connected to an I_2/I^- standard half-cell
Cell IV	a Co^{2+}/Co standard half-cell connected to an Fe^{3+}/Fe^{2+} standard half-cell

VCE 2013 Q26:

Which cell would be expected to develop the largest potential difference?

(a) I
(b) II
(c) III
(d) IV

VCE 2013 Q27:

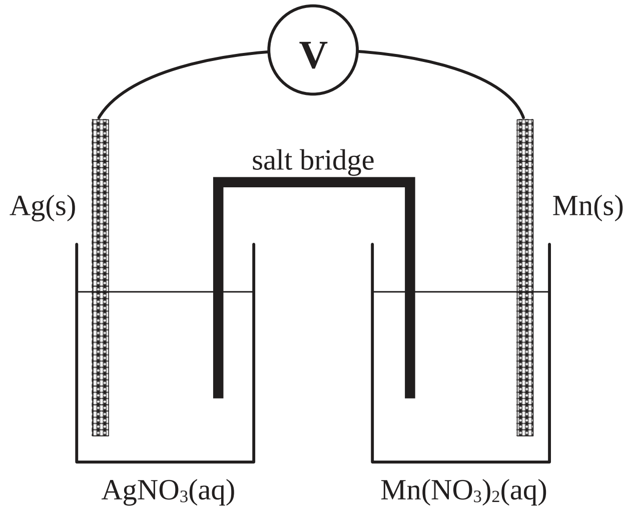
The reaction occurring at the cathode as cell IV is discharged is:

(a) $Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$
(b) $Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$
(c) $Co(s) \rightarrow Co^{2+}(aq) + 2e^-$

(d) $Co^{2+}(aq) + 2\,e^{-} \rightarrow Co(s)$

WACE 2016 Sample Exam Q15:

An electrochemical cell consisting of $Ag//Ag^+$ and $Mn//Mn^{2+}$ couples is constructed as represented by the diagram below. All solutions are 1.0 mol L^{-1} and the temperature is 25 °C.



Which one of the following gives the predicted emf for this cell in volts?

(a) 0.42
(b) 0.80
(c) 1.98
(d) 2.78

TEE 2003 Q23:

An electrochemical cell based on the following reaction has an E° = 1.03 V.

$$C\ell_2 + 2\,V^{3+} + 2\,H_2O \rightarrow 2\,VO^{2+} + 4\,H^+ + 2\,C\ell^-$$

What is the standard reduction potential for the reduction of VO^{2+} to V^{3+}?

(a) -3.05 V
(b) -0.33 V
(c) +0.33 V
(d) +3.05 V

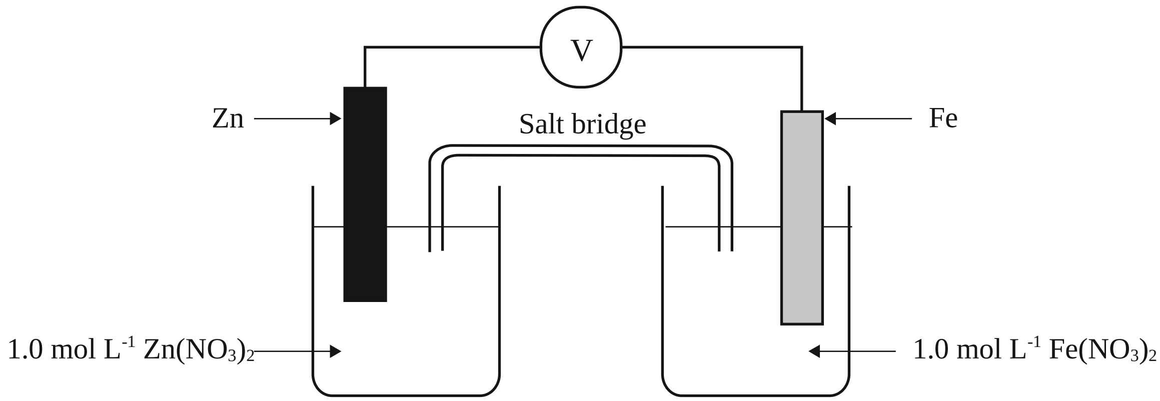
VCE 2013 Q28:

The main reason an aqueous solution of potassium nitrate, KNO_3, is used in salt bridges is:

(a)	K^+(aq) is a strong oxidant	NO_3^-(aq) is a weak reductant
(b)	K^+(aq) is a weak reductant	NO_3^-(aq) is a strong oxidant
(c)	**K^+(aq) salts are solution in water**	**NO_3^-(aq) salts are solution in water**
(d)	K^+(aq) ions will migrate to the anode half-cell	NO_3^-(aq) ions will migrate to the cathode half-cell

Section 2: Electrochemical cells

Use the following information to answer TEE 2009 Questions 22 and 23.



TEE 2009 Q22:

Which of the following will be true for this cell?

(a) The mass of the zinc electrode will increase
(b) Electrons will flow form the zinc electrode to the iron electrode
(c) The concentration of Fe^{2+} ions in the electrolyte will increase
(d) Iron will be deposited on the zinc electrode

TEE 2009 Q23:

Which one of the following will be the closest to the cell EMF, at 25 °C?

(a) 0.32 V
(b) 1.20 V
(c) 1.53 V
(d) 1.59 V

WACE 2AB 2013 Q20:

In an electrolytic cell, the electrons flow

(a) clockwise from the positive electrode to the negative electrode
(b) from the anode to the cathode
(c) through the solution to balance the charge build-up
(d) from the oxidizing agent to the reducing agent

WACE 2AB 2013 Q25:

A difference between electro-winning and electro-refining is that one of the processes

(a) uses electrolysis while the other uses a redox reaction
(b) involves electroplating the desired metal onto the cathode while the other does not
(c) uses an inert anode while the other uses an anode made of the unrefined impure metal
(d) has the metal forming on the anode while the other has the metal forming on the cathode

WACE 2AB 2012 Q14:

Which combination of anode, cathode and electrolyte could be used to silver-plate a nickel knife?

	Anode	**Cathode**	**Electrolyte**
(a)	Knife	Ag(s)	$AgNO_3$(aq)
(b)	Knife	Ag(s)	$Ni(NO_3)_2$(aq)
(c)	**Ag(s)**	**Knife**	**$AgNO_3$(aq)**
(d)	Ag(s)	Knife	$Ni(NO_3)_2$(aq)

TEE 2006 Q16:

An electrochemical cell has a positive value of E°. Which one of the following statements about the two half cells forming the cell is true?

(a) Both half cells must have positive standard electrode potentials
(b) The cathode half cell must have a positive standard reduction potential, while the anode half cell must have a negative standard reduction potential
(c) At least one of the half cells must have a positive standard reduction potential
(d) Both half cells may have a negative standard reduction potential

WACE 2016 Sample Exam Q16:

Consider the following statements about fuel cells.

i. A fuel cell is a device that converts chemical energy to electrical energy via a redox reaction
ii. Fuel cell technology involves the continuous supply of reactants to the cells and the continuous removal of products
iii. A fuel cell can be recharged by reversing the direction of current flow through the cell
iv. Fuel cells are considered a low-emission technology

Which of the above statements about fuel cells are true?

(a) i only
(b) i and ii only
(c) i, iii and iv only
(d) i, ii and iv only

Section 2: Electrochemical cells

VCE 2013 Q29:

The lead-acid battery used in cars consists of secondary galvanic cells.

The following equations relate to the lead acid battery.

$PbSO_4(s) + 2e^- \rightleftharpoons Pb(s) + SO_4^{2-}(aq)$ $\quad E° = –0.36$ V

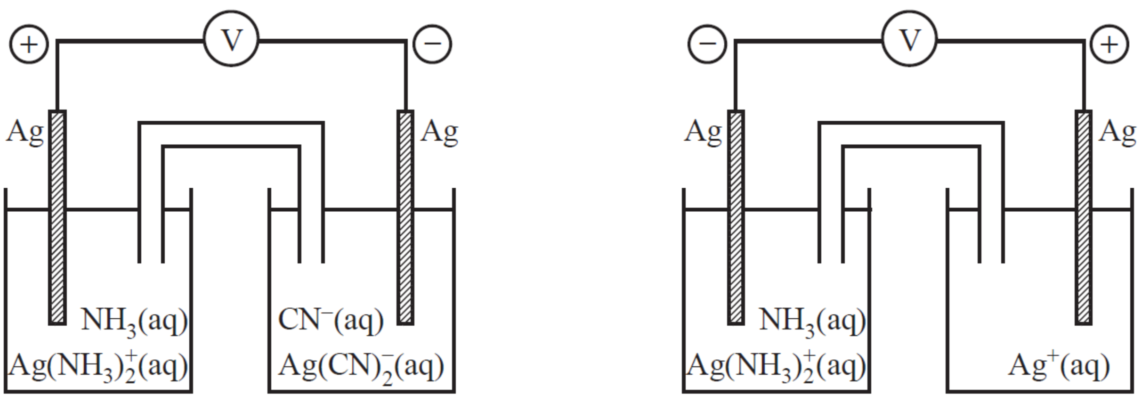
$PbO_2(s) + SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightleftharpoons PbSO_4(s) + 2H_2O(\ell)$ $\quad E° = +1.69$ V

When an external power source is used to recharge a flat lead acid battery:

(a) the concentration of sulfuric acid decreases
(b) $PbSO_4$ is both oxidised and reduced
(c) The mass of metallic lead decreases
(d) PbO_2 is oxidised to Pb

VCE 2011 Q10:

Two galvanic cells were constructed under standard conditions in an experiment to determine the relative positions in the electrochemical series of the standard electrode potential, E°, for the following reactions. Both cells generate a potential difference.



$Ag(NH_3)_2^+(aq) + e^- \rightleftharpoons Ag(s) + 2NH_3(aq)$ $\quad E°_1$

$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$ $\quad E°_2$

$Ag(CN)_2^- + e^- \rightleftharpoons Ag(s) + 2CN^-(aq)$ $\quad E°_3$

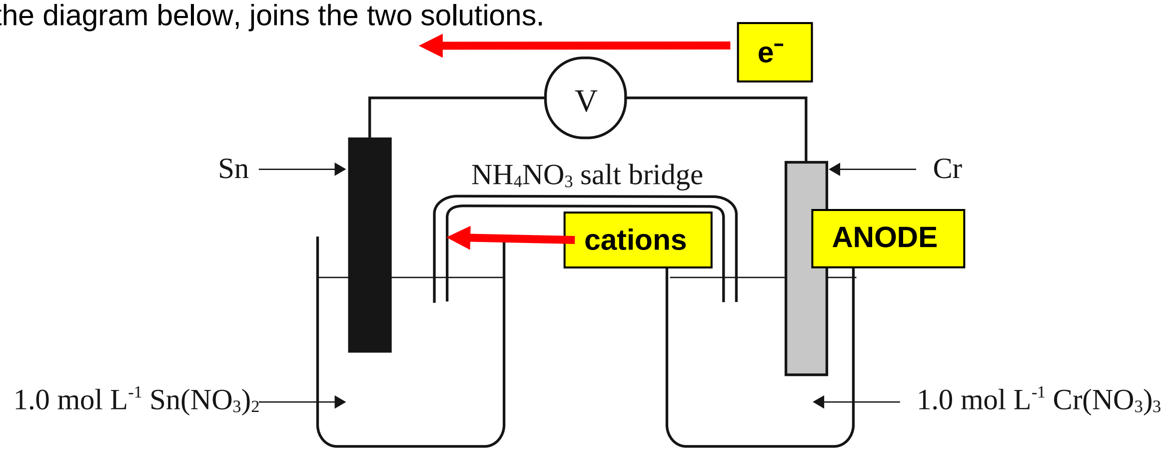
The values of the electrode potentials in order from highest to lowest would be:

(a) $E°_1$, $E°_2$, $E°_3$
(b) $E°_1$, $E°_3$, $E°_2$
(c) $E°_2$, $E°_1$, $E°_3$
(d) $E°_3$, $E°_2$, $E°_1$

$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$ $\quad E°_2$
$Ag(NH_3)_2^+(aq) + e^- \rightleftharpoons Ag(s) + 2NH_3(aq)$ $\quad E°_1$
$Ag(CN)_2^-(aq) + e^- \rightleftharpoons Ag(s) + 2CN^-(aq)$ $\quad E°_3$

WACE 3AB 2010 Q36:

An electrochemical cell consists of a tin electrode in a solution of 1.0 mol L^{-1} tin(II) nitrate, to create a Sn/Sn^{2+} half-cell, and a similarly constructed half-cell composed of a chromium electrode in a solution of 1.0 mol L^{-1} chromium(III) nitrate. The two electrodes are joined by a piece of copper wire. A salt bridge, as shown in the diagram below, joins the two solutions.



(a) On the diagram, label: (3 marks)

i. the anode

ii. the direction of electron flow

iii. the direction of cation flow in the salt bridge

(b) Write balanced anode and cathode reactions. (2 marks)

Anode reaction	**$Cr(s) \rightarrow Cr^{3+} + 3\ e^-$**
Cathode reaction	**$Sn^{2+} + 2\ e^- \rightarrow Sn(s)$**

(c) Would sodium carbonate be suitable as a salt for the salt bridge? Explain. (2 marks)

- **No (1 mark)**
- **A precipitate would form with Cr^{3+} and Sn^{2+}, preventing ion flow (1 mark)**

(d) Why does the rate of production of electrical current from an electrochemical cell decrease as it operates? (1 mark)

The rate of reaction decreases as reactants are consumed. Therefore, the rate of production of electrical current decreases

(e) During the operation of an electrochemical cell, why is it important that the anode and cathode do not come into contact with each other? (1 mark)

If the anode and cathode come into contact with each other, the system will be short circuited and electron current (and ions) will not flow. i.e. current will not be forced to flow through an external circuit. Instead, electron exchange will only occur at the metal surface.

WACE 3AB 2014 Q31:

(a) State the role of the standard hydrogen half-cell in determining the table of Standard Reduction Potentials. (2 marks)

Hydrogen half-cell is the reference half-cell (1 mark) against which the reduction potential of all other half-cells are measured (1 mark)

(b) State **three** limitations of Standard Reduction Potential tables. (3 marks)

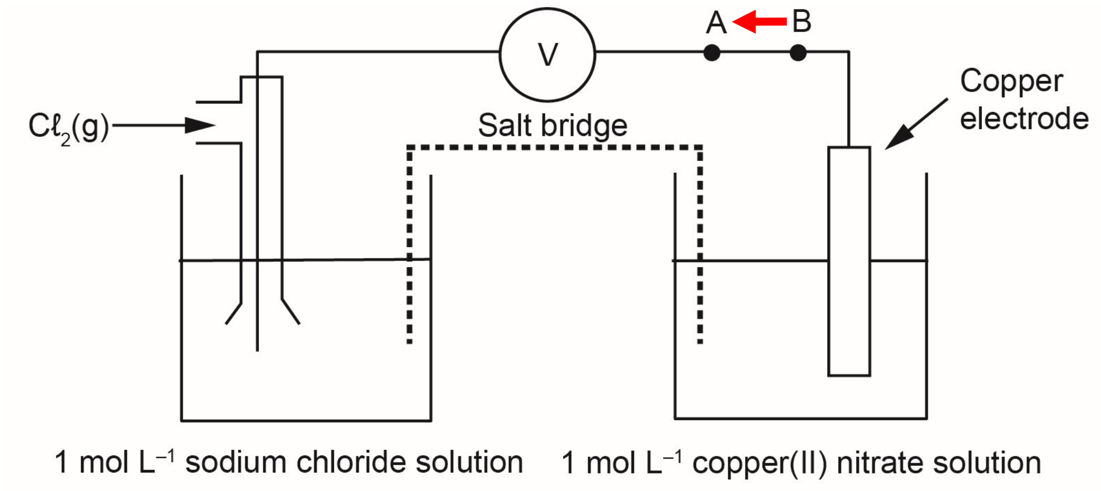
Three of following:

- **E° values depend on concentration (1 mol L^{-1})**
- **Applies only to aqueous solutions**
- **Depends on temperature**
- **Values of E° give no indication of reaction rate / activation energy**
- **Predictive tool – reaction may not occur**
- **All gases at 100 kPa (standard pressure)**

Examiner's comments: For part (a) candidates could not express themselves clearly enough although they had the idea of the role of the hydrogen half-cell. Part (b) was done quite well but many mentioned that the values were obtained at 100 kPa rather than the gases being at 100 kPa.

WACE 3AB 2015 Q35:

The following electrochemical cell was set up under standard conditions.



(a) Draw an **arrow** between **A** and **B** on the diagram to indicate the direction of electron flow. (1 mark)

(b) Write a balanced equation to represent the overall reaction occurring in this cell. (2 marks)

$Cℓ_2(g) + Cu(s) \rightarrow Cu^{2+}(aq) + 2\ Cℓ^-(aq)$

(c) State the reason for the reactants being kept in separate half-cells. (1 mark)

The electrons which are transferred must pass through an external circuit rather than being transferred through direction contact.

Also accept:

- **prevent direct exchange of electrons**
- **to create a current**
- **to create a potential different across the half-cells**

(d) State the observation predicted to occur in the $Cℓ_2$/NaCℓ half-cell. (1 mark)

Greenish-yellow gas dissolves

(e) Predict a metal/metal ion cell that could be used in place of the Cu/Cu^{2+} cell to give a higher emf (volts). (1 mark)

Any metal/metal ion below 0.34 V excluding Na, Ca, Sr, Ba and K.
Example: Zn/Zn^{2+}

Examiner's comments: Most candidates answered part (a) correctly. Most of the more capable candidates answered part (b) correctly. In part (c) many candidates discussed a short circuit occurring and missed that the purpose of a galvanic cell is to produce a useable electric current. In part (d) candidates often stated that the solution would go yellow-green rather than a yellow-green gas dissolves. Little more than half of the candidates who attempted part (e) were able to effectively use the *Standard Reduction Potentials* table in the provided data sheet.

WACE 3AB 2014 Q42:

Using the diagram below, explain the role of the following in the operation of a galvanic cell. (8 marks)

- Anode process
- Cathode process
- Lead(II) nitrate electrolyte
- Salt bridge
- Electron flow in external circuit

Anode:

- **Oxidation occurs at the anode (1 mark)**
- **Mg loses electrons to form Mg ions.**
  $Mg \rightarrow Mg^{2+} + 2e^-$ (1 mark)

Cathode:

- **Reduction occurs at the cathode (1 mark)**
- **Pb^{2+} gains electrons to form Pb(s)**
  $Pb^{2+} + 2e^- \rightarrow Pb$ (1 mark)

Electrolyte:

- **Lead(II) nitrate is the source of Pb^{2+} ions for reaction (1 mark)**

Salt bridge:

- **Allows for movement of ions between two half-cells (1 mark)**
- **Allows electrical neutrality to be maintained in each half-cell (1 mark)**

Electron flow

- **Provides energy to do work (e.g. cause reading on voltmeter) (1 mark)**

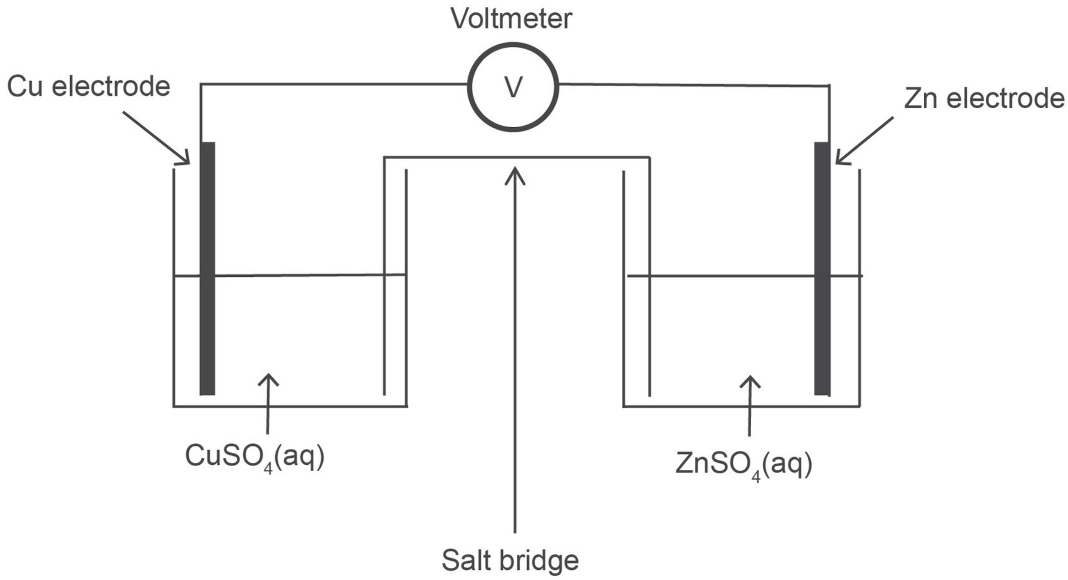
Examiner's comments: Many candidates think that the salt bridge involves a transfer of electrons and few could relate the role of the electron flow in the external circuit to providing energy to do work. Candidates were able to discuss the role of the anode and cathode and provide the appropriate half-equations. A large number of candidates analysed the cell in full which indicates a failure to read and comprehend what the question was asking. This question relates directly to the syllabus dot points.

WACE 3AB 2014 Q36:

A student investigated the effect of concentration of the electrolyte on the electrical potential of electrochemical cells. The student measured the electrical potential for the cell shown in the diagram below. In all trials the volumes of solutions and their temperatures were the same and the surface areas of the copper and zinc electrodes were the same.

The overall cell reaction is as follow:

$$Zn(s) + Cu^{2+}(aq) \rightleftharpoons Zn^{2+}(aq) + Cu(s)$$



(a) What is the independent variable in this investigation? (1 mark)

Concentrations of solution (accept concentration of Cu^{2+} or concentration of Zn^{2+})

(b) What is the dependent variable in this investigation? (1 mark)

Electric potential / voltage / volts

(c) Why did the volumes and temperatures of solutions and surface areas of the electrodes need to be the same in each trial? (1 mark)

To be confident that any changes in electrical potential are due to concentration changes only

Examiner's comments: For part (c) candidates should be discouraged from simply stating 'to ensure a fair test' or 'as control variables'. If they make these statements, they should expand on it and explain what they mean.

Section 2: Electrochemical cells

The student found that as the concentration of Cu^{2+} ions increased, electrical potential increased, but as concentration of Zn^{2+} ions increased, electrical potential decreased, as shown in the table below.

Trial	Cu^{2+} concentration (mol L^{-1})	Zn^{2+} concentration (mol L^{-1})	Electrical potential, E (V)
1	0.00001	1.00	1.01
2	0.010	1.00	1.04
3	1.00	1.00	1.09
4	1.00	0.010	1.16
5	1.00	0.00001	1.26

(d) Explain the increase in electrical potential as the concentration of Cu^{2+} ions increased and the decrease in electrical potential as the concentration of Zn^{2+} ions increased. (2 marks)

- **The rate of forward reaction increases as concentration of Cu^{2+} ions increases *(forward reaction is favoured)* (1 mark)**

- **The rate of reverse reaction increases as concentration of Zn^{2+} ions increases *(reverse reaction is favoured)* (1 mark)**

Examiner's comments: In part (d) many candidates simply rephrased the question and hence were awarded zero marks. Candidates were unable to apply the concept of reaction rates to electrochemical cells to account for the change in electrical potential.

(e) The student also observed that as the cells were allowed to run for a while their electrical potential slowly decreased from its maximum value. Why did this happen? (2 marks)

- **As the cell operates the concentration of reactants decreases (1 mark)**

- **…meaning over time the forward reaction rate decreases**
  OR
  …meaning over time the system approaches equilibrium (1 mark)

Examiner's comments: In part (e) candidates had some idea that reactant concentration was decreasing but were unable to express themselves clearly enough (many candidates referred to the concentration of electrolytes decreasing but which one?). Few followed through to relate the decrease in concentration to a decrease in reaction rate.

The student concluded:

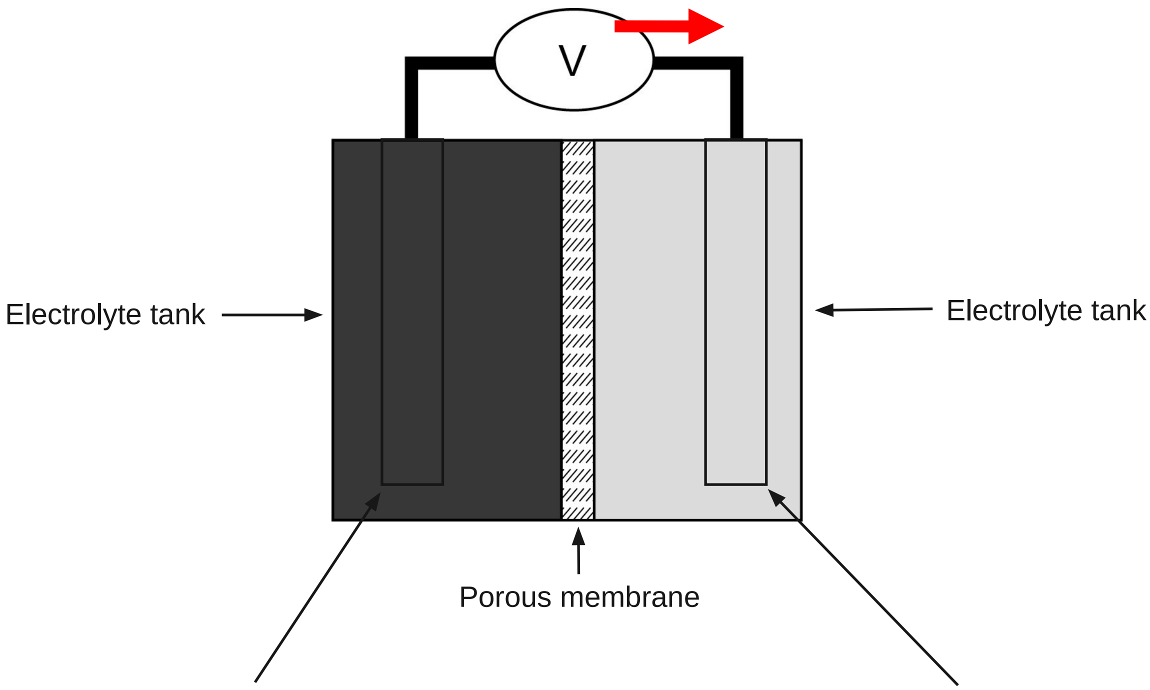
'As the concentration of the oxidant increases, so does the cell voltage (electrical potential).'

(f) List **two** ways to improve the investigation. (2 marks)
One: **Test other cells**
Two: **Repeat trials**

WACE 3AB 2012 Q32:

The vanadium redox battery is an electrochemical cell that is being developed to store electricity produced by solar or wind power on a large scale.

The general structure of the vanadium redox battery is shown below. In this battery, $VO^{2+}(aq)$ is converted to $VO_2^+(aq)$ at one electrode, while $V^{3+}(aq)$ is converted to $V^{2+}(aq)$ at the other.



$H_2O(\ell) + VO^{2+}(aq) \rightarrow VO_2^+(aq) + 2\,H^+(aq) + e^-$ | $V^{3+}(aq) + e^- \rightarrow V^{2+}(aq)$

Anode	**Cathode**

(a) In the boxes above, identify and label both the anode and cathode. (1 mark)
(b) Draw an arrow on the diagram to indicate the direction of electron flow. (1 mark)
(c) State briefly how the porous membrane functions to complete the circuit. (1 mark)
The porous membrane enables the flow of ions

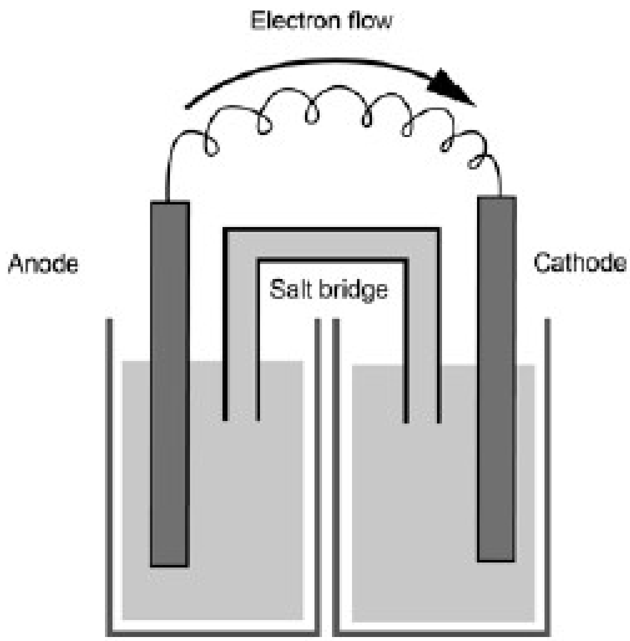
Examiner's comments: This question was well done generally. However, too many candidates indicated that the porous membrane was a source of ions (confusing it with a salt bridge) in part (c).

WACE 3AB 2013 Q41:

(a) The overall battery reaction during discharge is given below. Write and balance the anode and cathode reactions for the lead-acid storage battery. (2 marks)

Anode reaction	**$Pb + SO_4^{2-} \rightarrow PbSO_4 + 2e^-$**
Cathode reaction	**$PbO_2 + SO_4^{2-} + 4\,H^+ + 2\,e^- \rightarrow PbSO_4 + 2\,H_2O$**
Overall reaction	$Pb(s) + PbO_2(s) + 4\,H^+(aq) + 2\,SO_4^{2-}(aq) \rightarrow 2\,PbSO_4(s) + 2\,H_2O(\ell)$

(b) Draw a schematic diagram of the lead-acid battery showing the two half-cells. Label the anode, cathode and salt bridge, and indicate the direction of electron flow with an arrow. (4 marks)



Description	Marks
Anode	1
Cathode	1
Salt bridge	1
Direction of electron flow	1
Total	**4**

(c) i. With reference to the 'electrical potential' of a galvanic cell, describe how the lead-acid storage battery produces current. (2 marks)

In a galvanic cell there is a positive electrical potential difference which causes a flow of electrons from anode to cathode

ii. What determines the magnitude of the electrical potential of a cell? (1 mark)

The magnitude of difference between the E° values

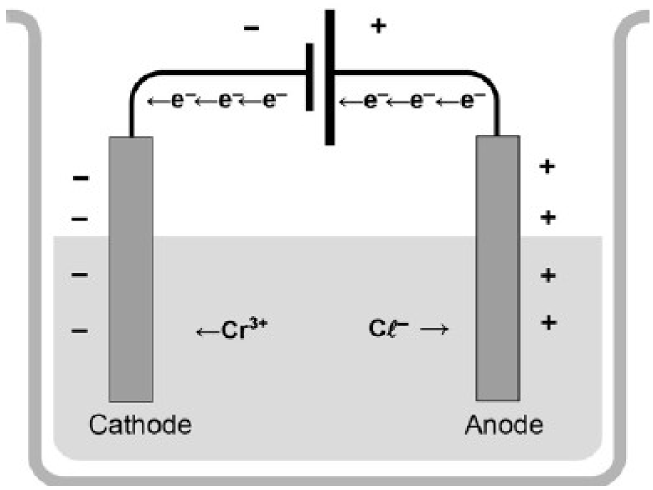
(also accept: concentration of reactants, the chemicals used, the half-equations, temperature)

Examiner's comments: In part (a) many candidates omitted the sulfate from the equations even though these reactions are on the Data Sheet. Part (b) was generally well done but (c)(i) was a difficult question for many candidates to answer and a significant number did not attempt it. Too many simply described what occurred in an electrochemical cell which does not answer the question. Use of the correct terminology and expressing the answer clearly was the main concern in (c)(ii).

WACE 2AB 2013 Q42:

(a) A quantity of pure chromium chloride ($CrC\ell_3$) is melted and placed in a heatproof vessel. Two inert electrodes are inserted as shown below and a current flow through the molten liquid. Complete and label the diagram below, showing the:

- Anode
- Cathode
- Direction of electron current
- Ions present and the direction in which they are flowing (5 marks)



(b) Write the oxidation half-equation. (2 marks)

$2\ C\ell^-(\ell) \rightarrow C\ell_2(g) + 2\ e^-$

(c) Write the reduction half-equation. (2 marks)

$Cr^{3+}(\ell) + 3\ e^- \rightarrow Cr(\ell)$

(d) Write the overall redox equation. (2 marks)

$2\ Cr^{3+}(\ell) + 6\ C\ell^-(\ell) \rightarrow 2\ Cr(\ell) + 3\ C\ell_2(g)$

Section 2: Electrochemical cells

(e) Indicate (by circling) which process occurs at the electrode that is connected to the negative terminal. (1 mark)

oxidation **reduction** redox

(f) What is acting as the oxidant in this reaction? (1 mark)

Cr^{3+} ion / chromium(III) ion

(g) What is produced at the anode? (1 mark)

$C\ell_2(g)$ / Chlorine gas

(h) State the oxidation number of: (4 marks)

i. Cr in solid chromium chloride **+3**

ii. Cℓ in molten chromium chloride **-1**

iii. Cr in chromium metal **0**

iv. Cℓ in chlorine gas **0**

WACE 2AB 2012 Q41:

A traditional technique for cleaning the tarnish, $Ag_2S(s)$, off silverware without using an abrasive cleaner (which wears away the precious metal) involves setting up a simple electrolytic cell.

The piece of silver to be cleaned can be placed on the bottom of a glass or enamel pan and covered with aluminium foil. A solution of baking soda, $NaHCO_3(aq)$, and table salt, $NaC\ell(aq)$, can be added a brought to boil in the pan. The piece of silver must be in contact with the aluminium foil. The process takes some time but at the end the silverware is sparkling clean.

This process works because tarnish on silver, $Ag_2S(s)$, is caused by sulfide ions. The salt and baking soda solution make an ion-carrying solution that transfer the sulfide ions from the silver to the aluminium foil.

(a) Write the half-equation where the silver sulfide, $Ag_2S(s)$, is converted to pure silver, Ag(s). (2 marks)

$\mathbf{Ag_2S \ + \ 2\,e^- \ \rightarrow \ 2\,Ag \ + \ S^{2-}}$

(b) Write the half-equation where aluminium forms aluminium ions. (2 marks)

$\mathbf{A\ell \ \rightarrow \ A\ell^{3+} \ + \ 3\,e^-}$

(c) Combine the two reaction half-equations to produce a balanced overall redox equation for the process. (2 marks)

$\mathbf{3\,Ag_2S \ + \ 2\,A\ell \ \rightarrow \ 6\,Ag \ + \ 2\,A\ell^{3+} \ + \ 3\,S^{2-}}$

Product can also be written as $A\ell_2S_3(s)$

(d) Explain the purpose of the salt and baking soda solution. (1 mark)

To allow the flow of ions

Examiner's comments: Most candidates were able to construct the half-equation for aluminium forming aluminium ions, but were unable to accurately construct half-equation involving Ag_2S. They often neglected to take into account the anion component of silver sulfide.

(e) Circle the substance that is acting as the cathode. (1 mark)

the tarnished silver or the aluminium foil

(f) Circle the substance at which reduction is occurring. (1 mark)

the tarnished silver or the aluminium foil

(g) Circle the arrow showing the correct direction in which electrons will flow during this process. (1 mark)

the tarnished silver ← the aluminium foil

(h) State the oxidation number of silver: (2 marks)

i. in the Ag_2S tarnish. **+1**

ii. when it has been converted to silver metal: **0**..

WACE 2AB 2011 Q33:

In industry, sodium is extracted from molten sodium chloride by electrolysis. Likewise, potassium can be extracted from its molten salts.

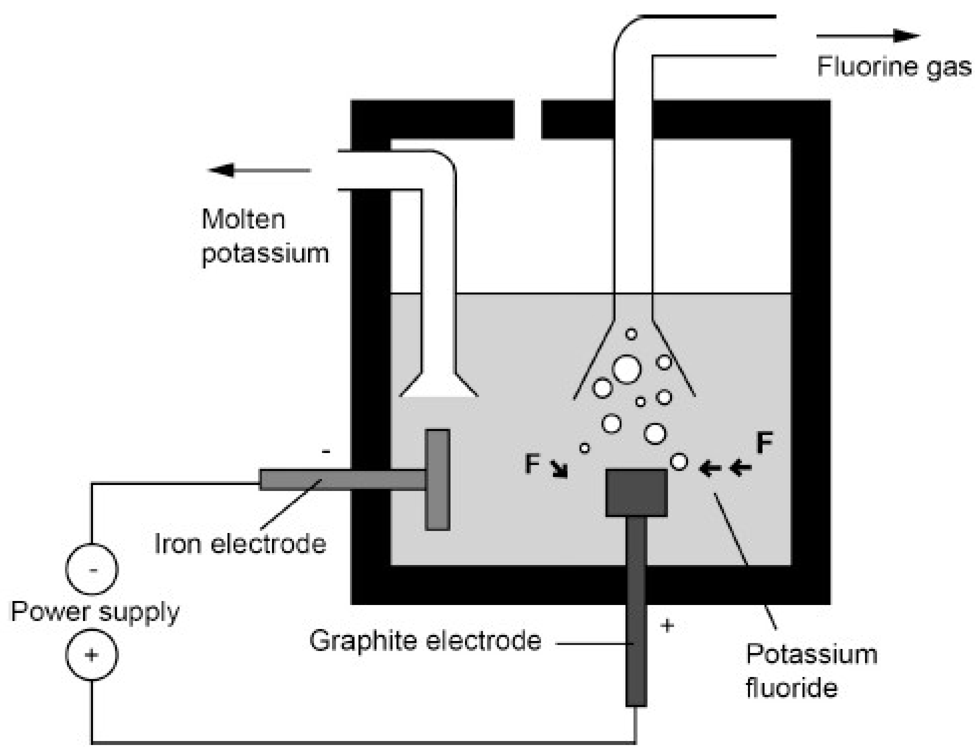
(a) When potassium fluoride is melted, it dissociates to release ions. Write the equation, including appropriate state symbols, to show this process. (2 marks)

$KF(s) \rightarrow K^+(\ell) + F^-(\ell)$

(b) Write the oxidation and reduction half equations for the electrolysis of molten potassium fluoride. (2 marks)

Oxidation	**$2 F^-(\ell) \rightarrow F_2(g) + 2 e^-$**
Reduction	**$K^+(\ell) + e^- \rightarrow K(\ell)$**

In the extraction of potassium from molten potassium fluoride, an apparatus similar to that shown in the diagram below is used. Refer to the diagram to answer the parts of the question that follow.



(c) Complete the table below by writing the electrode (anode or cathode) for which each of the substances shown in used. (2 marks)

Graphite	Iron
Anode	**Cathode**

(d) On the diagram, show clearly the direction of flow of fluoride ions. (1 mark)
Towards graphite electrode

(e) Name the substance that forms at the anode. (1 mark)

Fluorine gas

(f) Explain why solid potassium fluoride cannot undergo electrolysis. (2 marks)

Electrolysis requires ions taking part in reaction to be mobile. (1 mark)

In the solid state, K^+ and F^- ions are in fixed positions within the ionic lattice and are not free to move (1 mark)

(g)
i. List three potential hazards associated with the electrolysis of molten potassium fluoride. (3 marks)

Three of following:

- **High temperature**
- **Use of electricity**
- **Use of high current flows for 'good' reaction rate**
- **Toxic fluorine gas**
- **Molten metal**
- **Formation of very reactive potassium metal (potentially reacts with moisture in air)**

ii. State one safety precaution that workers using electrolysis to extract potassium from molten potassium fluoride should take. (1 mark)

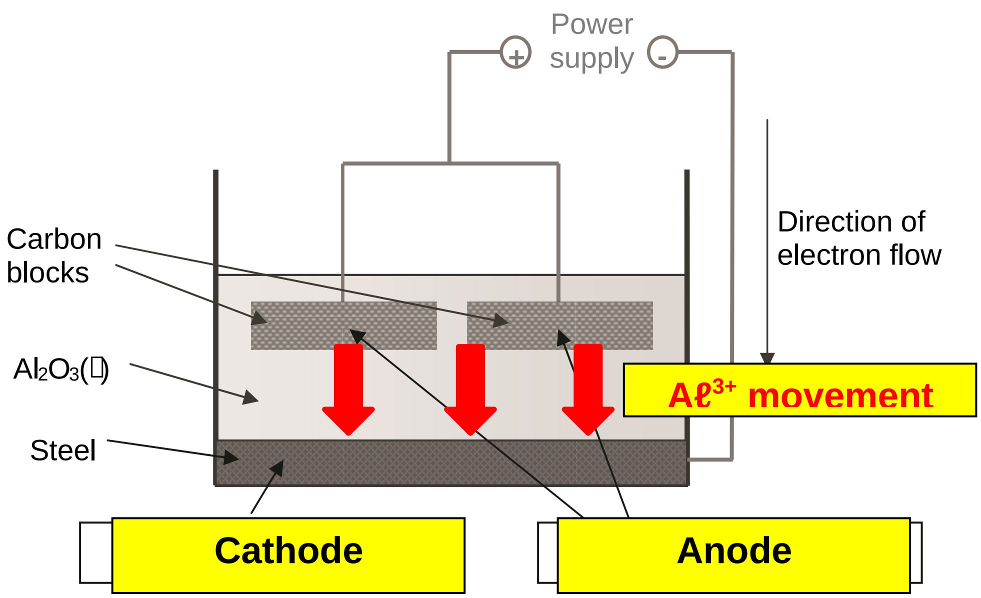
Answer may include

- **Breathing apparatus, gloves, protective clothing**
- **Ventilation**
- **Dry environment**
- **Extraction / safe storage of fluorine gas**
- **Removal / safe storage of potassium metal**

Examiner's comments: The first two parts of this question were poorly handled by most candidates, many of whom wrote equations in terms of aqueous solutions instead of molten liquids. Parts (c), (d) and (f) were generally well done. The safety question was well-answered by those candidates who realized that their answers had to be specifically about this particular chemical system rather than about general safety issues.

WACE 2AB 2010 Q29:

Below is a diagram of the electrolytic cell used to produce aluminium from molten aluminium oxide ($A\ell_2O_3$).



The reactions occurring at the electrodes are shown below.

$$A\ell^{3+}(\ell) \; + \; 3\,e^- \; \rightarrow \; A\ell(s)$$

$$2\,O^{2-}(\ell) \; + \; C(s) \; \rightarrow \; CO_2(g) \; + \; 4\,e^-$$

(a) Complete the diagram by:

i. Adding the words **anode** and **cathode** to the correct boxes above. (1 mark)

ii. Showing the direction of the movement of $A\ell^{3+}$ ions in the cell. (1 mark)

(b) Combine the two half-equations to give the overall redox reaction for the process occurring. (2 marks)

$4\,A\ell^{3+} \; + \; 6\,O^{2-} \; + \; 3\,C \; \rightarrow \; 4\,A\ell \; + \; 3\,CO_2$

(c) Explain why the $A\ell_2O_3$ needs to be in the liquid state. (2 marks)

Charged particles need to be mobile for conducts of electricity (1 mark)

Ions are free to move in liquid state but are in fixed positions in the solid state (1 mark)

Section 2: Electrochemical cells

(d) An early method of producing aluminium used the reaction of aluminium oxide with sodium metal to produce aluminium metal and sodium oxide.

i. Write a balanced equation for this reaction. (2 marks)

$Aℓ_2O_3 + 6\ Na \rightarrow 3\ Na_2O + 2\ Aℓ$

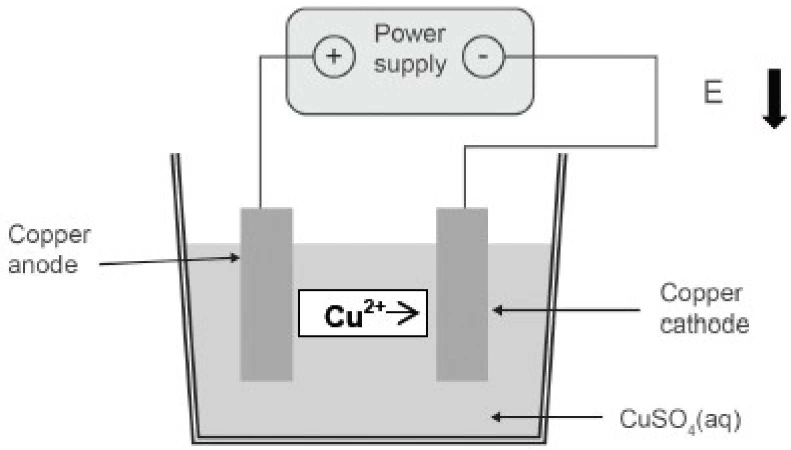
ii. Using the idea of oxidation numbers, explain how sodium is acting as a reducing agent in this process. (3 marks)

Na oxidation number changes from 0 to +1
Na donates an electron / is oxidised
Therefore, Na causes the reduction of the other species and is a 'reducing agent'

Section 2: Electrochemical cells

WACE 2AB 2014 Q40:

Look at the diagram below, which represents an electrolytic cell that can be used to produce pure copper from a solution of copper(II) sulfate ($CuSO_4$).



(c)

i. Draw an arrow on the diagram to show the direction of the movement of electrons in the external circuit at point E. (1 mark)

ii. Draw an arrow on the diagram to show the direction of the movement of copper ions (Cu^{2+}) in solution. (1 mark)

(d) Why must the power supply be direct current (DC) and not alternating current (AC)? (3 marks)

DC stands for direct current which means a current constantly moving in one direction around the circuit resulting in the same reaction occurring. (1 mark)

If AC (alternating current) is used, the electron flow is in a continual state of reversing direction (1 mark) and consequently the reaction would alternate between the forward and reverse reaction, resulting in no net reaction (1 mark)

(e) Explain why the copper(II) sulfate needs to be dissolved in water for the electrolysis process to work. (3 marks)

Any three of following:

- **Electrolysis requires the movement of charge around the complete circuit, including ions in the electrolyte**
- **In the solid state, the ions within copper(II) sulfate are not free to move so no current can flow**
- **When dissolved in water the copper(II) ions and sulfate ions are now free to move through the electrolyte and so complete the circuit**
- **Copper(II) sulfate is the source of the Cu^{2+} ions required for the reaction**

Section 2: Electrochemical cells

WACE 2016 Sample Exam Q41:

You are supplied with strips of three unknown metals, **A**, **B** and **C**, and are required to determine the order in which they are reduced, from most easily to least easily.

Using a voltmeter, electrical leads and clips, standard laboratory glassware and the typical range of chemical found in most laboratories, design an investigation and describe the procedures to be followed to determine the order of reduction for the metals. Use a labelled diagram to support your description. Ensure that you explain the purposes of substances or equipment (excluding beakers or other glassware) used.

Indicate the data you will collect and explain how this data gives the order of reduction. (10 marks)

Description	Marks
recognition that 3 Galvanic cells with A and B as electrodes, then A and C as electrodes and B and C as electrodes need to be constructed	1
recognition that the cells will need an electrolyte (eg. NaCℓ dissolved in water or 0.1 mol L^{-1} HCℓ)	1
recognition of role of ions as electrolyte to complete the circuit	1
recognition that voltmeter needs to be connected in the external circuit to measure potential difference between electrodes	1
recognition E(A/B), E(A/C) and E(B/C) is the data that needs to be collected	1
recognition that direction of current flow can be used to identify the anode and cathode in each of the cell couples and the magnitude of voltage can be used to order the 3 metals from most easily to least easily reduced	1–2
labelled diagram to show typical Galvanic cell (electrodes, voltmeter, salt bridge if 2 beakers used)	1–2
Recognition of variables to control. any 2 variables e.g. temp, conc. of electroyle soln, surface area of electrodes in contact with electrolyte	1
question incorrectly answered or not attempted	0
Total	**10**