REDOX



1. Which equation represents a redox reaction?

   **A.** $C+O_2 \rightarrow CO_2$

   **B.** $NH_3+HCl \rightarrow NH_4Cl$

   **C.** $2CrO_4^{2-}+2H^+ \rightarrow Cr_2O_7^{2-}+H_2O$

   **D.** $CaCO_3+2HCl \rightarrow CaCl_2+CO_2+H_2O$

   *Source: August 2003*

2. What is a typical characteristic of a strong oxidizing agent?

   **A.** It is readily oxidized.

   **B.** It easily loses electrons.

   **C.** It has a negative oxidation number.

   **D.** It has a positive reduction potential.

   *Source: August 2003*

3. Identify the oxidizing agent in the following equation:

   $Pb+PbO_2+4H^++2SO_4^{2-} \rightarrow 2PbSO_4+2H_2O$

   **A.** $H^+$ **B.** Pb

   **C.** $PbO_2$ **D.** $SO_4^{2-}$

   *Source: January 2004*

4. Which of the following is a redox equation?

   **A.** $2H_2+O_2 \rightarrow 2H_2O$

   **B.** $Ag_2CrO_4 \rightarrow 2Ag^++CrO_4^{2-}$

   **C.** $Ag(NH_3)_2^++2H^++Cl^- \rightarrow AgCl+2NH_4^+$

   **D.** $Mn(OH)_2+2HC_2H_3O_2 \rightarrow Mn^{2+}+2H_2O+2C_2H_3O_2^-$

   *Source: January 2004*

5. Which of the following is not a redox equation?

   **A.** $Zn+S \rightarrow ZnS$

   **B.** $Cl_2+2NaI \rightarrow 2NaCl+I_2$

   **C.** $Sn^{4+}+2Fe^{2+} \rightarrow 2Fe^{3+}+Sn^{2+}$

   **D.** $K_3PO_4+3AgNO_3 \rightarrow Ag_3PO_4+3KNO_3$

   *Source: April 2004*

6. Consider the following equation:

   $8H^++MnO_4^-+5Fe^{2+} \rightarrow 5Fe^{3+}+Mn^{2+}+4H_2O$

   Which ion is the reducing agent?

   **A.** $H^+$ **B.** $Fe^{2+}$

   **C.** $Fe^{3+}$ **D.** $MnO_4^-$

   *Source: April 2004*

7. When $U_3O_8$ (pitchblende) is dissolved in nitric acid, it changes into $UO_2(NO_3)_2$ (uranyl nitrate). What is the change in oxidation number for uranium?

   **A.** $+2\frac{2}{3}$ **B.** $+\frac{2}{3}$

   **C.** $-3\frac{1}{3}$ **D.** $-10$

   *Source: August 2003*

8. Which of the following contains molybdenum with its highest oxidation number?

   **A.** $MoCl_5$ **B.** $Mo_2S_3$

   **C.** $MoO_4^{2-}$ **D.** $Mo_6Cl_{12}$

   *Source: January 2004*

**9.** Which of the following skeletal half-reactions are not oxidations?

| | |
|---|---|
| **I.** | $ClO^- \rightarrow ClO_3^-$ |
| **II.** | $C_2H_5OH \rightarrow C_2H_4O_2$ |
| **III.** | $NO_2 \rightarrow N_2O_4$ |

**A.** I
**B.** II
**C.** III
**D.** I and II

*Source: January 2004*

**10.** What is the oxidation number of chromium in the complex ion $Cr(H_2O)_4Br_2^+$?

**A.** −1
**B.** +1
**C.** +2
**D.** +3

*Source: April 2004*

**11.** The metals Hg, Cd, Ga and Pd react as follows:

$3Pd^{2+} + 2Ga \rightarrow 2Ga^{3+} + 3Pd$

$Cd + Ga^{3+} \rightarrow$ no reaction

$Hg^{2+} + Pd \rightarrow Pd^{2+} + Hg$

Which of the following metals is the strongest reducing agent?

**A.** Pd
**B.** Ga
**C.** Cd
**D.** Hg

*Source: August 2003*

**12.** Consider the following half-reactions under standard conditions:

| | |
|---|---|
| **I.** | $ClO_2 + e^- \rightarrow ClO_2^-$ |
| **II.** | $PbSO_4 + 2e^- \rightarrow Pb + SO_4^{2-}$ |
| **III.** | $Fe^{3+} + 3e^- \rightarrow Fe$ |

In an experiment when $ClO_2$ and Fe were combined, they reacted. In a second experiment when $PbSO_4$ and Fe were combined, there was no observable change. Which of the following shows the reduction half-reactions I, II and III in order of decreasing E°?

**A.** I, II, III
**B.** I, III, II
**C.** II, III, I
**D.** III, II, I

*Source: January 2004*

**13.** Three elements (X, Y, and Z) and solutions of their anions were combined. The results are shown in the following data table:

| | $X^-_{(aq)}$ | $Y^-_{(aq)}$ | $Z^-_{(aq)}$ |
|---|---|---|---|
| **X** | — | No reaction | No reaction |
| **Y** | Reaction | — | Reaction |
| **Z** | Reaction | No reaction | — |

Which of the following best describes the strength of the ions acting as reducing agents?

**A.** $X^- > Y^- > Z^-$
**B.** $X^- > Z^- > Y^-$
**C.** $Y^- > Z^- > X^-$
**D.** $Z^- > Y^- > X^-$

*Source: April 2004*

**14.** Which of the following metals can be oxidized by 1.0 M $Fe^{2+}$?

**A.** Sn
**B.** Co
**C.** Cr
**D.** Ag

*Source: August 2003*

**15.** Which of the following combinations will react spontaneously?

**A.** $I_2 + Cu^{2+}$

**B.** $Pb^{2+} + Ag$

**C.** $Zn^{2+} + Mg$

**D.** $Sn^{2+} + Ni^{2+}$

*Source: January 2004*

**16.** When the skeletal equation $Cr_2O_7^{2-} \rightarrow Cr^{3+}$ is balanced in acidic solution $H^+$ and $e^-$ will appear. Which of the following best describes the $H^+$ and $e^-$ for the balanced half-reaction?

**A.** $7H^+$ on the left and $1e^-$ on the right

**B.** $14H^+$ on the left and $9e^-$ on the left

**C.** $14H^+$ on the left and $6e^-$ on the left

**D.** $14H^+$ on the left and $6e^-$ on the right

*Source: April 2004*

**17.** Consider the following skeletal redox equation for a reaction in basic solution:

$Zn + NO_3^- \rightarrow ZnO_2^{2-} + NH_3$ (basic)

Which of the following best represents the reduction half-reaction occurring in this solution?

**A.** $9H^+ + NO_3^- + 8e^- \rightarrow NH_3 + 3H_2O$

**B.** $3H_2O + NO_3^- + 5e^- \rightarrow NH_3 + 6OH^-$

**C.** $6H_2O + NO_3^- + 8e^- \rightarrow NH_3 + 9OH^-$

**D.** $4OH^- + Zn + 2e^- \rightarrow ZnO_2^{2-} + 2H_2O$

*Source: January 2004*

**18.** The titration of a 25.0 mL $SnCl_2$ sample, in acidic solution, requires 14.4 mL of 0.030 M $K_2Cr_2O_7$. The balanced equation for the reaction is shown below:

$Cr_2O_7^{2-} + 14H^+ + 3Sn^{2+} \rightarrow 3Sn^{4+} + 2Cr^{3+} + 7H_2O$

What is the number of moles of $SnCl_2$ in the original sample?

**A.** $1.4 \times 10^{-4}$ mol

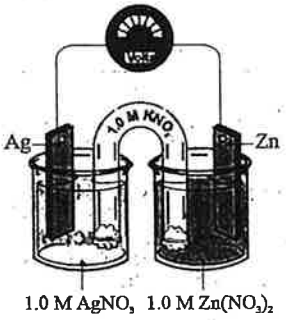
**B.** $4.3 \times 10^{-4}$ mol

**C.** $1.3 \times 10^{-3}$ mol

**D.** $5.2 \times 10^{-2}$ mol

*Source: April 2004*

*Use the following diagram to answer the next three questions.*




**19.** What is the equation for the half-reaction that occurs at the cathode?

**A.** $Ag \rightarrow Ag^+ + e^-$

**B.** $Ag^+ + e^- \rightarrow Ag$

**C.** $Zn \rightarrow Zn^{2+} + 2e^-$

**D.** $Zn^{2+} + 2e^- \rightarrow Zn$

*Source: August 2003*

**20.** What happens to the mass of each electrode as the cell operates?

**A.** $Ag_{(s)}$ increases, $Zn_{(s)}$ increases

**B.** $Ag_{(s)}$ decreases, $Zn_{(s)}$ decreases

**C.** $Ag_{(s)}$ decreases, $Zn_{(s)}$ increases

**D.** $Ag_{(s)}$ increases, $Zn_{(s)}$ decreases

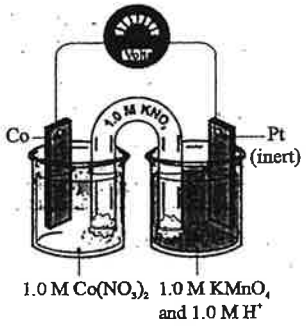
*Source: August 2003*

**21.** Which of the following is correct?

| | Electrons Flow Towards | Anions Move Towards |
|---|---|---|
| **A.** | Zn | Zn |
| **B.** | Zn | Ag |
| **C.** | Ag | Zn |
| **D.** | Ag | Ag |

*Source: August 2003*

**22.** Consider the following diagram:




Identify the anode reaction for the cell shown in the diagram.

**A.** $H_2 \rightarrow 2H^+ + 2e^-$

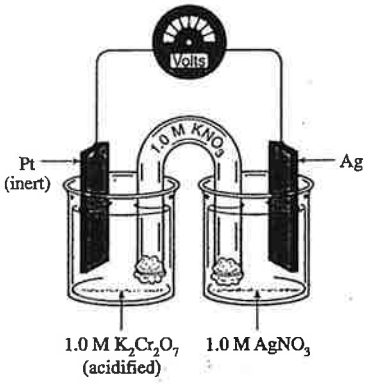
**B.** $Co \rightarrow Co^{2+} + 2e^-$

**C.** $Co^{2+} + 2e^- \rightarrow Co$

**D.** $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$

*Source: August 2003*

*Use the following diagram to answer the next three questions.*




**23.** Which of the following represents the overall cell reaction?

**A.** $Cr_2O_7^{2-} + H^+ + Ag \rightarrow Ag^+ + Cr^{3+} + H_2O$

**B.** $Cr_2O_7^{2-} + 14H^+ + 9Ag \rightarrow 9Ag^+ + Cr^{3+} + 7H_2O$

**C.** $Cr_2O_7^{2-} + 14H^+ + 6Ag \rightarrow 6Ag^+ + 2Cr^{3+} + 7H_2O$

**D.** $Cr_2O_7^{2-} + 14H^+ + 6Ag^+ \rightarrow 6Ag + 2Cr^{3+} + 7H_2O$

*Source: January 2004*

**24.** What happens to the pH at each electrode?

| | pH at Anode | pH at Cathode |
|---|---|---|
| **A.** | increases | decreases |
| **B.** | increases | increases |
| **C.** | stays the same | decreases |
| **D.** | stays the same | increases |

*Source: January 2004*

**25.** What is the cell voltage at equilibrium?

**A.** –0.43 V

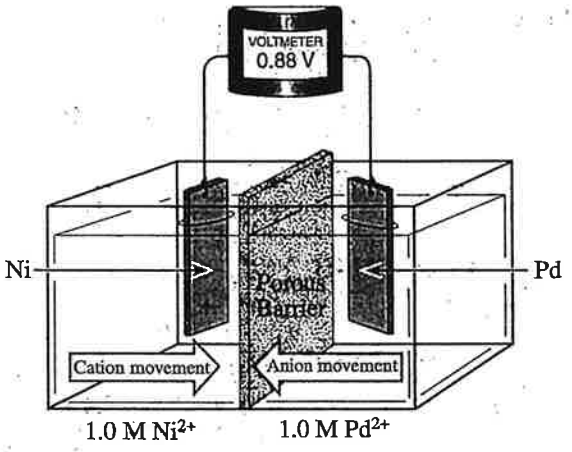
**B.** 0.00 V

**C.** +0.43 V

**D.** +2.03 V

*Source: January 2004*

*Use the following diagram to answer the next three questions.*




**26.** What is the half-cell reaction at the anode?

**A.** $Ni \rightarrow Ni^{2+} + 2e^-$

**B.** $Pd \rightarrow Pd^{2+} + 2e^-$

**C.** $Ni^{2+} + 2e^- \rightarrow Ni$

**D.** $Pd^{2+} + 2e^- \rightarrow Pd$

*Source: April 2004*

**27.** What best describes the flow of electrons?

**A.** from Ni to Pd

**B.** from Pd to Ni

**C.** from cathode to anode

**D.** into the solution around the Ni electrode

*Source: April 2004*

**28.** What is the voltage for the oxidation half-reaction of Pd?

**A.** –0.62 V

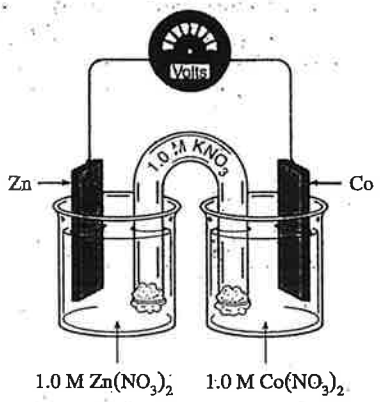
**B.** +0.62 V

**C.** +0.88 V

**D.** +1.14 V

*Source: April 2004*

**29.** Consider the following diagram:

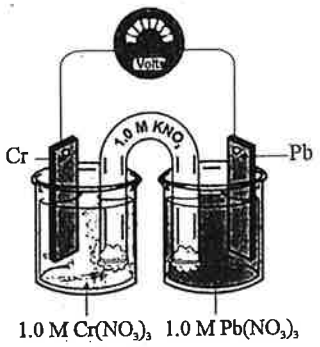



Which of the following best describes the $Co^{2+}$ ion movement and the mass of the zinc electrode as the cell operates?

| | $Co^{2+}$ movement | Mass of zinc |
|---|---|---|
| **A.** | toward the Co electrode | increases |
| **B.** | toward the Co electrode | decreases |
| **C.** | toward the Zn electrode | increases |
| **D.** | toward the Zn electrode | decreases |

*Source: January 2004*

**30.** Consider the following cell:




What is the initial cell voltage?

**A.** +0.87 V

**B.** +0.61 V

**C.** +0.54 V

**D.** +0.28 V

*Source: August 2003*

**31.** Which of the following would protect an iron pipeline from rusting?

**A.** connecting it to a solution of silver nitrate

**B.** connecting it to the positive terminal of a direct current power supply

**C.** connecting it to the negative terminal of a direct current power supply

**D.** connecting it to electrodes made of copper which are buried beside the pipeline

*Source: January 2004*

**32.** Which of the following best describes a car battery as it is being recharged?

**A.** It is an electrolytic cell.

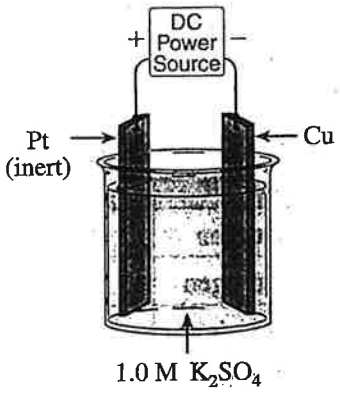
**B.** It is an electrochemical cell.

**C.** It is an example of a short circuit.

**D.** It is a system moving to a state of lower potential energy.

*Source: January 2004*

*Use the following diagram to answer the next two questions..*




**33.** What is the equation for the anode reaction?

**A.** $K \rightarrow K^+ + e^-$

**B.** $Cu \rightarrow Cu^{2+} + 2e^-$

**C.** $2SO_4^{2-} \rightarrow S_2O_8^{2-} + 2e^-$

**D.** $H_2O \rightarrow \frac{1}{2}O_2 + 2H^+\left(10^{-7}\,M\right) + 2e^-$

*Source: April 2004*

**34.** Which of the following best describes the mass of the copper electrode and the direction of cation movement as the cell operates?

| | Mass of the copper electrode | Cation movement |
|---|---|---|
| **A.** | increases | to the left |
| **B.** | stays the same | to the left |
| **C.** | stays the same | to the right |
| **D.** | decreases | to the right |

*Source: April 2004*

**35.** Which of the following are produced at the anode and the cathode in the electrolysis of molten lithium chloride using platinum inert electrodes?

| | Anode | Cathode |
|---|---|---|
| **A.** | oxygen | hydrogen |
| **B.** | hydrogen | oxygen |
| **C.** | chlorine | lithium |
| **D.** | lithium | chlorine |

*Source: August 2003*

**36.** Consider the following reaction:

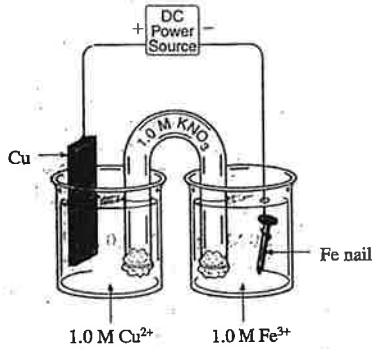
$Ni^{2+} + 2Ag \rightarrow 2Ag^{+} + Ni$

Which of the following is true?

| | $E^{\circ}$ | Reaction |
|---|---|---|
| **A.** | –1.06 V | non-spontaneous |
| **B.** | –0.54 V | non-spontaneous |
| **C.** | + 0.54 V | spontaneous |
| **D.** | +1.06 V | spontaneous |

*Source: April 2004*

**37.** Consider the following diagram:




Which of the following statements describes why the cell would not be used to electroplate an Fe nail with copper?

**A.** $Fe^{3+}$ will react with $Cu^{2+}$.

**B.** The cell reaction is non-spontaneous.

**C.** $Cu^{2+}$ ions do not collide with the nail.

**D.** The power supply is connected backwards.

*Source: April 2004*

**38.** In the electro-refining of an ore, the ore is dissolved and the solution placed in an electrolytic cell. A specific ore produces a solution which contains $Na^{+}$, $Cu^{2+}$, $Pb^{2+}$ and $Ag^{+}$. Which of these metals can be plated out from the solution by electrolysis?

**A.** all four metals

**B.** Cu and Pb only

**C.** Na, Cu and Pb only

**D.** Cu, Pb and Ag only

*Source: April 2004*

## Written Response

**1.** A reaction occurs when copper metal is dropped into a solution of silver nitrate. Write the balanced formula equation and the balanced net ionic equation for this reaction. **(3 marks)**

Formula equation: _____

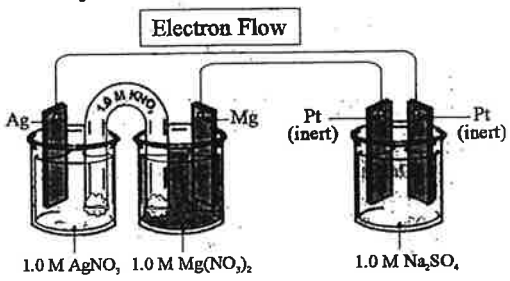
Net ionic equation: _____

*Source: August 2003*

**2.** When setting up the apparatus to electroplate a zinc object with copper, the object is suspended in a $Cu^{2+}$ solution. Explain why it is a good idea to turn on the power supply before immersing the electrodes in the solution. **(1 mark)**

*Source: August 2003*

**3.** Consider the following apparatus consisting of an electrochemical cell joined to an electrolytic cell:




**a)** On the diagram above, indicate the direction of electron flow in the top wire. **(1 mark)**

**b)** Which metal in cell A is the cathode? **(1 mark)**

Cathode: _____

**c)** Write the anode and cathode half-reactions for cell B. **(3 marks)**

Anode: _____

Cathode: _____

*Source: August 2003*

**4.** Balance the following skeletal redox equation in acidic solution:

$MnO_4^- + As_2O_3 \rightarrow Mn^{2+} + AsO_4^{3-}$ (acidic)

**(4 marks)**

*Source: January 2004*

**5.** Draw an electrolytic cell that could be used to plate an iron ring with gold. Be sure to include all of the necessary parts. In addition, label the anode, solution used and composition of the electrodes. **(3 marks)**

*Source: January 2004*

**6.** Balance the following redox reaction in basic solution:

$ClO_3^- + N_2H_4 \rightarrow NO_3^- + Cl^-$ (basic)

**(4 marks)**

*Source: April 2004*

**7.** A 1.0 M solution of $CoSO_4$ is electrolyzed using inert electrodes.

**a)** Write the anode and cathode half-reactions that would occur. **(2 marks)**

Anode: _____

Cathode: _____

**b)** What is observed when bromthymol blue is added to the solution around the anode? **(1 mark)**

*Source: April 2004*

# UNIT TEST 5 – OXIDATION–REDUCTION

**1.** Which of the following represents an oxidation?

**A.** $2H^+ + S \rightarrow H_2S$

**B.** $2SO_4^{2-} \rightarrow S_2O_8^{2-}$

**C.** $Na^+ + Cl^- \rightarrow NaCl$

**D.** $SO_2 + H_2O \rightarrow 2H^+ + SO_3^{2-}$

*Source: June 2003*

**2.** Identify the oxidation number for manganese in $MnO_4^-$.

**A.** –7 **B.** +7

**C.** +8 **D.** +9

*Source: June 2003*

**3.** Which of the following is more difficult to reduce than the $H^+_{(aq)}$ ion?

**A.** $I_2$

**B.** $Ag^+$

**C.** $Zn^{2+}$

**D.** $Cu^{2+}$

*Source: June 2003*

**4.** Nitric oxide (NO) can be prepared by the oxidation of Cu with $NO_3^-$ in acidic solution. Copper is oxidized to $Cu^{2+}$ and $NO_3^-$ is reduced to NO. Which of the following equations correctly describes this process?

**A.** $Cu + NO_3^- + 4H^+ \rightarrow Cu^{2+} + NO + 2H_2O$

**B.** $Cu + NO_3^- + 2H^+ \rightarrow Cu^{2+} + NO_2 + H_2O$

**C.** $Cu + 4NO_3^- + 4H^+ \rightarrow Cu^{2+} + 2NO_2 + 2H_2O$

**D.** $3Cu + 2NO_3^- + 8H^+ \rightarrow 3Cu^{2+} + 2NO + 4H_2O$

*Source: June 2003*

**5.** Acidified potassium permanganate $(KMnO_4)$ solution is often used in redox titrations. Permanganate reacts with $Sn^{2+}$ as follows:

$$2MnO_4^- + 5Sn^{2+} + 16H^+ \rightarrow 2Mn^{2+} + 5Sn^{4+} + 8H_2O$$

A 10.0 mL solution containing $Sn^{2+}$ is titrated with 19.3 mL of 0.10 M $KMnO_4$. What is the $\left[Sn^{2+}\right]$ in the solution?

**A.** 0.077 M

**B.** 0.19 M

**C.** 0.25 M

**D.** 0.48 M

*Source: June 2003*

**6.** Given the following half-cell reactions:

$$PbO_2 + 4H^+ + SO_4^{2-} + 2e^- \rightleftarrows PbSO_4 + 2H_2O \quad E° = +1.69\,V$$
$$PbSO_4 + 2e^- \rightleftarrows Pb + SO_4^{2-} \quad E° = -0.36\,V$$

Which of the following best describes the overall reaction and the standard cell voltage in a lead acid storage battery?

**A.** $Pb + 2H_2O \rightarrow PbO_2 + 4H^+ + 4e^-$
$E°_{cell} = +1.33$

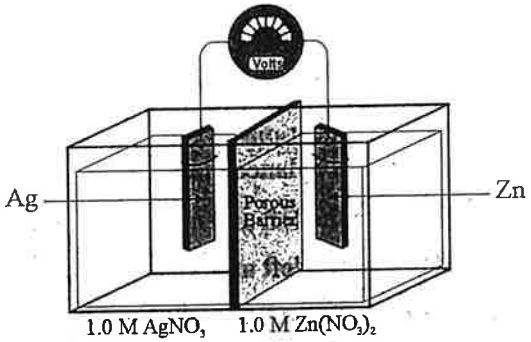
**B.** $PbO_2 + 4H^+ + 4e^- \rightarrow Pb + 2H_2O$
$E°_{cell} = +1.33$

**C.** $Pb + PbO_2 + 2SO_4^{2-} + 4H^+ \rightarrow 2PbSO_4 + 2H_2O$
$E°_{cell} = +2.05$

**D.** $2PbSO_4 + 2H_2O \rightarrow Pb + PbO_2 + 2SO_4^{2-} + 4H^+$
$E°_{cell} = +2.05$

*Source: June 2003*

7. Consider the following electrochemical cell:




What is the anode half-reaction?

**A.** $Ag \rightarrow Ag^+ + e^-$

**B.** $Ag^+ + e^- \rightarrow Ag$

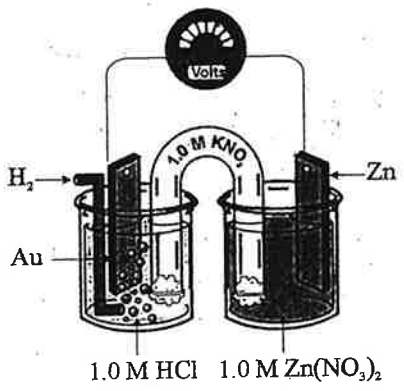
**C.** $Zn \rightarrow Zn^{2+} + 2e^-$

**D.** $Zn^{2+} + 2e^- \rightarrow Zn$

*Source: June 2003*

**CHALLENGE QUESTION** 48.9

8. Consider the following cell:




What is the value of the standard cell potential?

**A.** –0.76 V

**B.** +0.76 V

**C.** +2.12 V

**D.** +2.26 V

*Source: June 2003*

9. Two separate reactions involved in the refining of copper ore are:

| | |
|---|---|
| Reaction I | $2Cu_2S + 3O_2 \rightarrow 2Cu_2O + 2SO_2$ |
| Reaction II | $Cu_2S + 2Cu_2O \rightarrow 6Cu + SO_2$ |

What happens to the copper ions during this process?

**A.** They are reduced in Reaction I.

**B.** They are reduced in Reaction II.

**C.** They are oxidized in Reaction I.

**D.** They are oxidized in Reaction II.

*Source: June 2003*

10. Why is aluminum a good choice for the manufacture of outdoor structures?

**A.** Pure aluminum is easily reduced.

**B.** Pure aluminum is not easily oxidized.

**C.** Pure aluminum is easily reduced, but forms a protective coating.

**D.** Pure aluminum is easily oxidized, but forms a protective coating.

*Source: June 2003*

**CHALLENGE QUESTION** 52.1

11. Which of the following are produced at the anode and the cathode during the electrolysis of aqueous calcium iodide using carbon electrodes?

| | Anode | Cathode |
|---|---|---|
| **A.** | Iodine | Calcium |
| **B.** | Hydrogen | Oxygen |
| **C.** | Oxygen | Hydrogen |
| **D.** | Iodine | Hydrogen |

*Source: June 2003*

**12.** What is the oxidation number of carbon in the ethanoate ion $C_2H_3O_2^-$?

**A.** $-1$

**B.** $0$

**C.** $+\frac{1}{2}$

**D.** $+1$

*Source: June 2004*

**13.** What reaction will occur when a solution containing 1.0 M $MgSO_4$ and 1.0 M $CoCl_2$ is stored in a galvanized (Zn coated) bucket?

**A.** $Mg_{(s)} + Cl_{2(g)} \rightarrow MgCl_2$

**B.** $Co^{2+} + SO_4^{2-} \rightarrow CoSO_{4(s)}$

**C.** $Co^{2+} + Zn_{(s)} \rightarrow Zn^{2+} + Co_{(s)}$

**D.** $Mg^{2+} + Zn_{(s)} \rightarrow Zn^{2+} + Mg_{(s)}$

*Source: June 2004*

**14.** When the skeletal equation $Br_2 \rightarrow BrO_3^-$ is balanced in acidic solution, $H_2O$, $H^+$ and $e^-$ will appear. Which of the following are the correct balancing coefficients?

| | $H_2O$ | $H^+$ | $e^-$ |
|---|---|---|---|
| **A.** | 3 | 3 | 2 |
| **B.** | 6 | 6 | 4 |
| **C.** | 6 | 6 | 5 |
| **D.** | 6 | 12 | 10 |

*Source: June 2004*

**15.** Consider the following balanced redox equation in acidic solution:

$$5H_2O + 2CoCl_2 + OCl^- \rightarrow 2Co(OH)_3 + 5Cl^- + 4H^+$$

Which of the following describes the amounts and locations of $OH^-$ and $H_2O$ if the equation is balanced in basic solution?

**A.** $1H_2O$ on the left and no $OH^-$

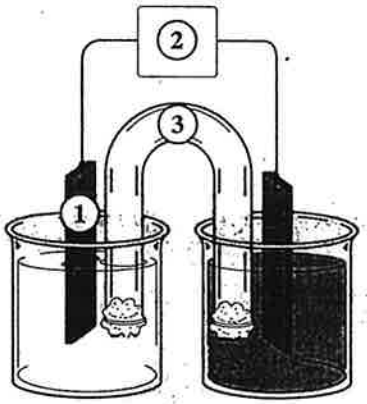
**B.** $1H_2O$ on the left and $4OH^-$ on the left

**C.** $5H_2O$ on the left and $4OH^-$ on the left

**D.** $1H_2O$ on the left and $4OH^-$ on the right

*Source: June 2004*

**16.** Consider the numbered components in the following diagram:

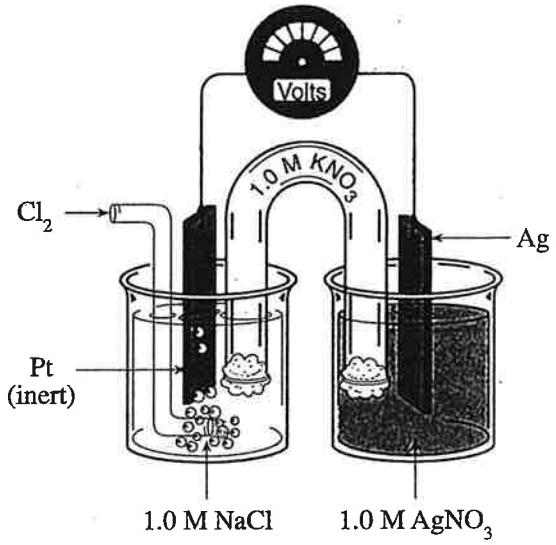



Which of the following would best describe the components of this electrochemical cell?

| | Component 1 | Component 2 | Component 3 (Content) |
|---|---|---|---|
| **A.** | non-metal | power supply | $NaNO_{3(aq)}$ |
| **B.** | metal | light bulb | |
| **C.** | metal | voltmeter | $CH_3OH_{(aq)}$ |
| **D.** | metal | power supply | $CH_3OH_{(aq)}$ |

*Source: June 2004*

*Use the following diagram to answer the next three questions.*




**17.** Which of the following represents the anode half-cell reaction?

**A.** $Ag \rightarrow Ag^+ + e^-$

**B.** $Ag^+ + e^- \rightarrow Ag$

**C.** $Cl_2 + 2e^- \rightarrow 2Cl^-$

**D.** $2Cl^- \rightarrow Cl_2 + 2e^-$

*Source: June 2004*

**18.** What changes in mass occur to the anode and cathode?

| | Anode Mass | Cathode Mass |
|---|---|---|
| **A.** | decreases | increases |
| **B.** | decreases | no change |
| **C.** | increases | decreases |
| **D.** | increases | no change |

*Source: June 2004*

**19.** What is the voltage for this cell under standard conditions?

**A.** –0.24 V

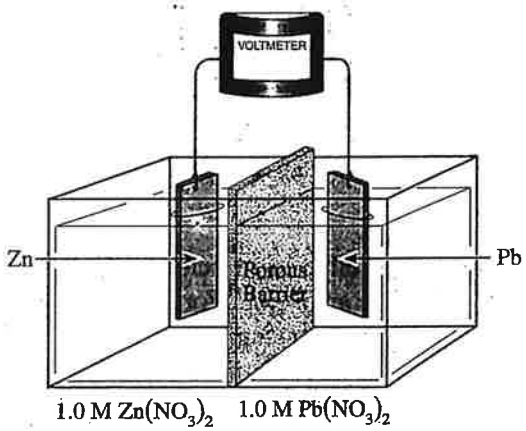
**B.** –0.56 V

**C.** +0.56 V

**D.** +2.16 V

*Source: June 2004*

**20.** Consider the following diagram:




As the cell operates, the voltage gradually changes. Which of the following is responsible for this change?

**A.** The $[Pb^{2+}]$ is increasing.

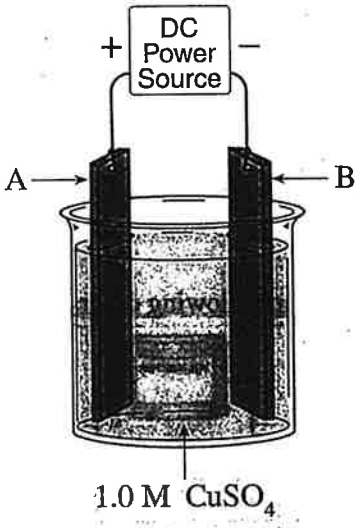
**B.** The $[Pb^{2+}]$ is decreasing.

**C.** The $[Zn^{2+}]$ is decreasing.

**D.** The mass of the $Pb_{(s)}$ electrode is decreasing.

*Source: June 2004*

*Use the following diagram to answer the next two questions.*




**21.** The above cell is constructed in order to copper plate an object. For best results, which of the following should be used for electrodes A and B?

| | Electrode A | Electrode B |
|---|---|---|
| **A.** | object | pure copper |
| **B.** | pure copper | object |
| **C.** | object | any conductor |
| **D.** | any conductor | object |

*Source: June 2004*

**22.** A student tries to use the above apparatus to copper plate a zinc object. What will happen if the student places the zinc object at A and the copper electrode at B?

| | Electrode A | Electrode B |
|---|---|---|
| **A.** | $Cu_{(s)}$ forms | Cu dissolves |
| **B.** | Zn dissolves | $Zn_{(s)}$ forms |
| **C.** | Zn dissolves | $Cu_{(s)}$ forms |
| **D.** | Bubbles form | Bubbles form |

*Source: June 2004*

**23.** Which of the following occurs during the electrolysis of molten KCl?

**A.** Oxygen forms at the anode.

**B.** Potassium forms at the anode.

**C.** Chlorine forms at the cathode.

**D.** Potassium forms at the cathode.

*Source: June 2004*

## Written Response

**1.** Consider the following experimental results:

$Ce^{4+} + Pd \rightarrow Pd^{2+} + Ce^{3+}$

$In^{3+} + Cd \rightarrow$ no reaction

$Pd^{2+} + In^{2+} \rightarrow In^{3+} + Pd$

$Cd^{2+} + Pd \rightarrow$ no reaction

Use these results to complete the table of reduction half-reactions below. **(3 marks)**

| Oxidizing Agents | | Reducing Agents |
|---|---|---|
| $Ce^{4+} + e^-$ | $\rightleftarrows$ | $Ce^{3+}$ |
| $Pd^{2+} + 2e^-$ | $\rightleftarrows$ | $Pd$ |
| $Cd^{2+} + 2e^-$ | $\rightleftarrows$ | $Cd$ |
| $In^{3+} + e^-$ | $\rightleftarrows$ | $In^{2+}$ |

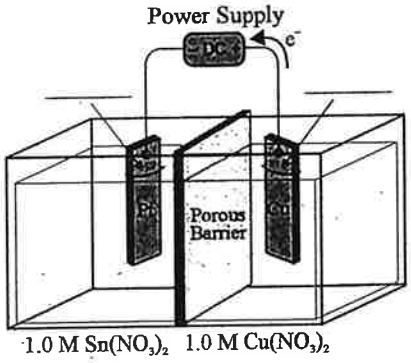
*Source: April 2003*

**2.** Balance the following equation. **(3 marks)**

$Cr_2O_7^{2-} + C_2O_4^{2-} \rightarrow Cr^{3+} + CO_2$ (acidic)

*Source: April 2003*

**3.** Consider the following electrolytic cell which contains a porous barrier to prevent general mixing of solutions.




**a)** Label the anode and cathode in the space provided on the diagram above. **(1 mark)**

**b)** Write an equation for the overall cell reaction. **(2 marks)**

**c)** Calculate the minimum theoretical voltage required for this reaction under standard conditions. **(1 mark)**

*Source: April 2003*

**4.** Balance the following redox equation: **(4 marks)**

$H_2S + CrO_4^{2-} \rightarrow S_8 + Cr^{3+}$ (acidic)

*Source: June 2003*

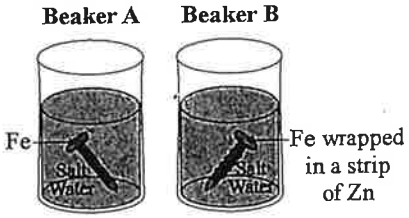
**5.** An excess of copper solid is dropped into a solution which contains

$AgNO_3$, $Fe(NO_3)_3$ and $Zn(NO_3)_2$.

Write the equations for any reduction half-reactions that occur over time under standard conditions. **(2 marks)**

*Source: June 2003*

**6.** Consider the following diagrams:




**a)** Predict what should happen to the Fe in Beaker **A.** **(1 mark)**

**b)** Predict what should happen to the Fe in Beaker **B.** Explain. **(2 marks)**

*Source: June 2003*

**7.** An impure sample of $CaC_2O_4$ weighing 0.803 g is titrated with 15.70 mL of 0.101 M $KMnO_4$. The net reaction is:

$2MnO_4^- + 5C_2O_4^{2-} + 16H^+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$

What is the percent by mass of the $CaC_2O_4$ in the original sample? **(4 marks)**

*Source: June 2004*

**8.** A sample of Zn corrodes in moist air.

**a)** Write the reduction half-reaction. **(1 mark)**

**b)** What metal could be attached to the sample to prevent the corrosion of the zinc? Explain. **(2 marks)**

*Source: June 2004*

# ANSWERS AND SOLUTIONS
# UNIT REVIEW – OXIDATION–REDUCTION

| | | | | | |
|---|---|---|---|---|---|
| 1. A | 8. C | 15. C | 22. B | 29. B | 36. A |
| 2. D | 9. C | 16. C | 23. C | 30. B | 37. C |
| 3. C | 10. D | 17. C | 24. D | 31. C | 38. D |
| 4. A | 11. B | 18. C | 25. B | 32. A | WR1–7. See Solution |
| 5. D | 12. B | 19. B | 26. A | 33. D | |
| 6. B | 13. B | 20. D | 27. A | 34. C | |
| 7. B | 14. C | 21. C | 28. A | 35. C | |

**1. A**

In a redox reaction oxidation states must change. At least one element experiences an increase in oxidation state and least one experiences a decrease in oxidation state. In reaction A, carbon goes from 0 to +4 and oxygen goes from 0 to –2.

**2. D**

Strong oxidizing agents are up high on the left side of the Standard Reduction Potentials of Half-Cells table in the Data Booklet (page A8). Half-reactions in this area have large positive reduction potentials.

**3. C**

The oxidizing agent can be identified by finding the species containing the element which is reduced (undergoes a decrease in oxidation state). Oxidation states are shown above the given equation:
oxidation

**states:** 0 +4 –2 +1 +6 –2 +2 +6 –2 +1 –2
**equation:** $Pb + PbO_2 + 4\,H^+ + 2\,So_4^{2-} \rightarrow 2\,PbSO_4 + 2\,H_2O$

The Pb in $PbO_2$ goes down from +4 to +2 in $PbSO_4$. Therefore $PbO_2$ is the oxidizing agent.

**4. A**

Redox equations will be characterized by reactions in which at least one element's oxidation state increases and at least one element's oxidation state decreases. If no oxidation states change the reaction is not redox. This is true only for reaction A.

oxidation
**states:** 0 0 + 1 –2
**equation:** $2H_2 + O_2 \rightarrow 2H_2O$

A useful generalization which may save time is that dissociation equations and double replacement reactions are never redox.

**5. D**

In equation D, the oxidation number of each atom in each compound remains the same. Therefore the reaction is not redox.

**6. B**

$Fe^{2+}$ is oxidized to $Fe^{3+}$. The oxidation number of iron goes from +2 to +3. An increase in oxidation number is oxidation. The reducing agent is oxidized as it causes something else to be reduced.

**7. B**

In $U_3O_8$, uranium has an oxidation number of $^{+}\frac{16}{3}$ or $^{+}5\frac{1}{3}$. In $UO_2(NO_3)_2$ its oxidation number is +6. The difference is $^{+}6 - {}^{+}\frac{16}{3} = +\frac{2}{3}$.

**8. C**

Using the oxidation state rules given in your text book, the oxidation state of molybdenum is +5 in $MoCl_5$, +3 in $Mo_2S_3$, + 6 in $MoO_4^{2-}$, and +2 in $Mo_6Cl_{12}$.

**9. C**

In oxidations the oxidation state of at least one element in the compound must increase. Consider each of the equations with their oxidation state changes.

| | Reaction | Oxidation state change |
|---|---|---|
| I | $ClO^- \rightarrow ClO_3^-$ | Cl from +1 to +5 |
| II | $C_2H_5OH \rightarrow C_2H_4O_2$ | C from -2 to 0 |
| III | $NO_2 \rightarrow N_2O_4$ | No change |

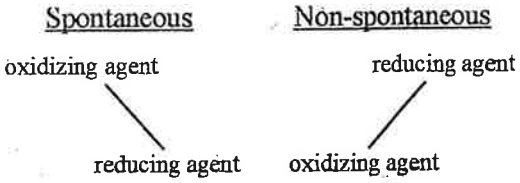
III is the only one which is not redox.

**10. D**

Using the standard rules for determining oxidation numbers, the following assignments are made and the value for chromium is determined by subtraction from the ionic charge of +1.
$H = {}^{+}1 \times 8 = +8$, $O = -2 \times 4 = -8$, $Br = -1 \times 2 = -2$ $\quad {}^{+}1(\text{total}) - {}^{+}8 - {}^{-}8 - {}^{-}2 = {}^{+}3$

**11. B**

The reducing agents are the metal elements, not the metal ions. Following the format of the Standard Reduction Potentials of Half-Cells table in the Data Booklet (page A8):

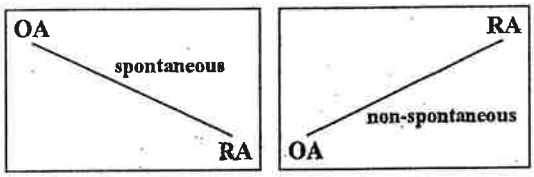



A workable table would be as follows:

| Oxidizing agents | Reducing agents |
|---|---|
| $Hg^{2+}$ | Hg |
| $Pd^{2+}$ or $Cd^{2+}$ | Pd or Cd |
| $Ga^{3+}$ | Ga |

Regardless of whether Pd is above or below Cd, Ga is the strongest reducing agent.

**12. B**




Using the spontaneity rule, a table of redox half-reactions in correct order can be constructed based on the spontaneity of the described reactions:

I $ClO_2 + e^- \rightarrow ClO_2^-$

III $Fe_3^+ + 3\ e^- \rightarrow Fe$

II $PbSO_4 + 2\ e^- \rightarrow Pb + SO_4^{2-}$

The higher the half-reaction on the table, the greater the $E°$ value.

**13 B**

Non-metal anions are reducing agents, while non-metal elements are oxidizing agents. The strongest reducing agent, $X^-_{(aq)}$, reacts with each of the other 2 elements. The next strongest, $Z^-_{(aq)}$, reacts with 1 of the other 2 elements. The weakest, $Y^-_{(aq)}$, does not react with either of the other elements.

**14. C**

Using the same spontaneity rule as in question 11, and the Standard Reduction Potentials of Half-Cells table in the Data Booklet (page A8), Cr, a reducing agent, is the only reducing agent below $Fe^{2+}$, the oxidizing agent, on the chart.

**15. C**

Using the Standard Reduction Potentials of Half-Cells chart on page A8 of your Data Booklet, and the spontaneity rule from the last question, only the combination of $Zn^{2+}$ and Mg will react spontaneously.

**16. C**

$Cr_2O_7^{2-} \rightarrow Cr^{3+}$ To balance, proceed as shown in this order:

Balance Cr's:
$Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$

Balance O's:
$Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$

Balance H's:
$Cr_2O_7^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O$

Balance total charge:
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$

**17. C**

Constructing the $NO_3^-$ to $NH_3$ half-reaction:
$NO_3^- \rightarrow NH_3$

Balance O by adding $3H_2O$ to the right side of the equation: $NO_3^- \rightarrow NH_3 + 3H_2O$

Balance H by adding 9 $H^+$ to the left side of the equation: $NO_3^- + 9H^+ \rightarrow NH_3 + 3H_2O$

Balance charge by adding 8 electrons to the left side of the equation:
$NO_3^- + 9H^+ + 8e^- \rightarrow NH_3 + 3H_2O$

Since the reaction occurs in a basic environment, add 9 $OH^-$ to both sides of the equation to convert $H^+$ into $H_2O$:
$NO_3^- + 9H_2O + 8e^- \rightarrow NH_3 + 3H_2O + 9OH^-$

Cancel the $H_2O$ from the left with the $H_2O$ from the right to yield the final equation:
$NO_3^- + 6H_2O + 8e^- \rightarrow NH_3 + 9OH^-$

**18. C**

$Cr_2O_7^{2-} + 14H^+ + 3Sn^{2+} \rightarrow 3Sn^{4+} + 2Cr_3^+ + 7H_2O$

| $n_1$ | $n_2$ |
|---|---|
| 0.030 M | $c = ?$ |
| 14.4 mL | 25.0 mL |

$n_1 = 0.030\ M \times 0.014\ 4\ L = 4.3 \times 10^{-4}\ mol$

$$n_2 = 4.3 \times 10^{-4}\ mol \times \frac{3}{1} = 1.3 \times 10^{-3}\ mol$$

moles of $Sn^{2+}$ = moles of $SnCl_2$
since $SnCl_{2(s)} \rightarrow Sn^{2+}_{(aq)} + 2\ Cl^-_{(aq)}$

**19. B**

Ignoring the $KNO_3$ in the salt bridge, which is not involved in the cell reaction, the chemical species present are: Ag, $Ag^+$, $NO_3^-$, $Zn^{2+}$, Zn, and $H_2O$. Of these $Ag^+$ is the strongest oxidizing agent and Zn is the strongest reducing agent. By definition, the cathode is the electrode at which reduction occurs, where the strongest oxidizing agent ($Ag^+$ in this case) reacts.

**20. D**

Referring to question 19, the cathode reaction is B – the mass of the cathode (Ag) increases. The anode reaction is C – the mass of the anode (Zn) decreases.

**21. C**

Two useful generalizations:

- electrons flow in the wire from anode to cathode (Ag in this case)
- anions in the solution migrate towards the anode (Zn in this case)

**22. B**

Ignoring the $KNO_3$ in the salt bridge and the inert Pt electrode, chemical species present are: Co, $Co^{2+}$, $NO_3^-$, $K^+$, $MnO_4^-$, $H^+$, and $H_2O$. Consulting the Standard Reduction Potentials of Half-Cells table in the Data Booklet (page A8), the strongest reducing agent is Co. It undergoes oxidation at the anode.

**23. C**

List of all entities present in both half-cells (the salt bridge and the inert platinum electrode can be ignored): $K^+$, $Cr_2O_7^{2-}$, $H^+$, $Ag^+$, $NO_3^-$, $Ag_{(s)}$, $H_2O$

Consulting Standard Reduction Potentials of Half-Cells chart on page A8 of your Data Booklet, shows $Cr_2O_7^{2-}$ with $H^+$ to be the strongest oxidizing agent and $Ag_{(s)}$ to be the strongest reducing agent. Half-reactions can be added to produce the overall cell reaction:

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$$
$$6(Ag \rightarrow Ag^+ + e^-)$$

$$Cr_2O_7^{2-} + 14H^+ + 6Ag \rightarrow 6Ag^+ + 2Cr^{3+} + 7H_2O$$

**24. D**

Oxidation occurs at the anode (silver half-reaction); reduction occurs at the cathode (dichromate half-reaction). At the anode neither $H^+_{(aq)}$ or $OH^-_{(aq)}$ is produced or consumed so pH remains constant. At the cathode $H^+_{(aq)}$ is consumed meaning that pH will increase.

**25. B**

At equilibrium the cell is "dead"; cell voltage is 0.00 V.

**26. A**

In the diagram anions are moving towards the Ni electrode meaning it is the anode, since in all cells, anions migrate towards the anode. The anode is the electrode where oxidation (loss of electrons) occurs.

**27. A**

Electrons move through the wire from anode to cathode in all cells. In the previous question Ni was identified as the anode.

**28. A**

$$E^\circ_{net} = E^\circ_{r\ cathode} - E^\circ_{r\ anode}$$
$$+0.88\,V = E^\circ_{r\ cathode} - {}^-0.26\,V$$
$$E^\circ_{r\ cathode(Pd)} = {}^+0.62\,V$$
$$E^\circ_{ox\,Pd} = {}^-0.62\,V$$

**29. B**

List of potentially reacting entities: $Zn_{(s)}$, $Zn^{2+}$, $NO_3^-$, $Co_{(s)}$, $Co^{2+}$, $H_2O$. Consulting the Standard Reduction Potentials of Half-Cells chart on page A8 of your Data Booklet shows $Co^{2+}$ to be the strongest oxidizing agent and $Zn_{(s)}$ to be the strongest reducing agent. The strongest reducing agent, $Zn_{(s)}$, is oxidized at the anode and loses mass. The $Co^{2+}$, a cation, moves towards the Co electrode which is the cathode since cations move towards the cathode.

**30. B**

Ignoring the $KNO_3$ in the salt bridge, the chemical species present are: Cr, $Cr^{3+}$, $Pb^{2+}$, Pb, and $H_2O$. The strongest oxidizing agent is $Pb^{2+}$; the strongest reducing agent is Cr, leading to the following half-reactions with their reduction potentials taken from the Standard Reduction Potentials of Half-Cells table in the Data Booklet (page A8):

$Pb^{2+} + 2e^- \rightarrow Pb_{(s)}$ $\quad E^\circ_r = {}^-0.13\,V$

$Cr_{(s)} \rightarrow Cr^{3+} + 3e^-$ $\quad E^\circ_r = {}^-0.74\,V$

$$E^\circ_{net} = {}^-0.13\,V - {}^-0.74\,V = +0.61\,V$$

**31. C**

The iron, in an iron pipeline, will normally be oxidized by $O_2$, dissolved in water in the soil. By attaching the pipeline to the negative terminal of a direct current power supply, the power supply's negative terminal becomes the anode and the iron becomes the inert cathode on which the $O_2$ is reduced. Because of this iron will not be oxidized (allowed to rust). This is called impressed current cathodic protection.

**32. A**

An electrolytic cell has a non-spontaneous reaction driven by an applied current. This is what happens when you recharge a car's battery.

**33. D**

The anode is the electrode where oxidation occurs (where the strongest reducing agent reacts). The anode is $Pt_{(s)}$, since it is connected to the positive terminal of the power supply. Of the contents of the cell electrolyte, $H_2O$, $K^+$, and $SO_4^{2-}$, water is the only reducing agent present.

**34. C**

$H_2O$ is also the strongest oxidizing agent present, and will react at the copper cathode by the half-reaction: $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ ($10^{-7}$ M)

The copper electrode will not change in mass, and cation movement will be to the right because cations migrate to the cathode in all cells.

**35. C**

Chemical species present: $Li^+_{(l)}$, $Cl^-_{(l)}$, $Pt_{(s)}$.
Strongest oxidizing agent is $Li^+_{(l)}$. Strongest reducing agent is $Cl^-_{(l)}$.
Half-reactions:

| | |
|---|---|
| Cathode (reduction) | $Li^+_{(l)} + e^- \rightarrow Li_{(l)}$ |
| Anode (oxidation) | $2Cl^-_{(l)} \rightarrow Cl_{2(g)} + 2e^-$ |

**36. A**

To determine net voltage for the cell, the equation is

$$\Delta E^\circ_{net} = E_r^\circ(\text{cathode}) - E_r^\circ \text{ (anode)}$$
$$= -0.26\text{ V} - (+0.80\text{ V}) = -1.06\text{ V}.$$

If the net voltage for the cell is negative the cell reaction is non-spontaneous.

If you look at the relationship between the oxidizing agent $Ni^{2+}$ and the reducing agent $Ag_{(s)}$ on the Data Book Table of Selected Standard Electrode Potentials, you have to go "uphill" to go from oxidizing agent to reducing agent. That is also an indicator of a non-spontaneous reaction.

**37. C**

The Fe nail is the cathode since it is connected to the negative terminal of the power supply. Reduction will occur there, but it will be the reduction of $Fe^{3+}$ to $Fe^{2+}$, not the reduction of $Cu^{2+}$ to $Cu_{(s)}$, since there are no $Cu^{2+}$ ions in that half-cell.

**38. D**

$Na_{(s)}$ cannot be plated out from the solution since $H_2O$ is a stronger oxidizing agent than $Na^+$, and will therefore react in its place. $Na_{(s)}$ could never exist in contact with water since it is a very reactive metal.

## Written Response

**1.** *A reaction occurs when copper metal is dropped into a solution of silver nitrate. Write the balanced formula equation and the balanced net ionic equation for this reaction.* **(3 marks)**

Formula equation:

$$Cu_{(s)} + 2AgNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + 2Ag_{(s)}$$

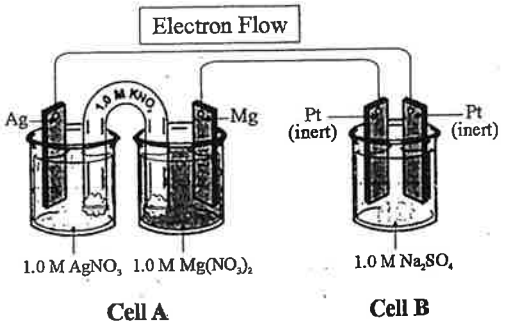
Net ionic equation:

$$Cu_{(s)} + 2Ag^{+}_{(aq)} \rightarrow Cu^{2+}_{(aq)} + 2Ag_{(s)}$$

**2.** *When setting up the apparatus to electroplate a zinc object with copper, the object is suspended in a $Cu^{2+}$ solution. Explain why it is a good idea to turn on the power supply before immersing the electrodes in the solution.* **(1 mark)**

If you did not turn on the power supply before immersing the electrodes in the solution, the $Cu^{2+}$ would react spontaneously with the zinc to be plated, oxidizing the zinc.

**3.** *Consider the following apparatus consisting of an electrochemical cell joined to an electrolytic cell:*




**a)** *On the diagram above, indicate the direction of electron flow in the top wire.* **(1 mark)**

**b)** *Which metal in cell A is the cathode?* **(1 mark)**

**Cathode:** Ag

**c)** *Write the anode and cathode half-reactions for cell B.* **(3 marks)**

**Anode:** $H_2O \rightarrow 2H^+ + \frac{1}{2}O_2 + 2e^-$

**Cathode:** $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$

**4.** *Balance the following skeletal redox equation in acidic solution:* **(4 marks)**

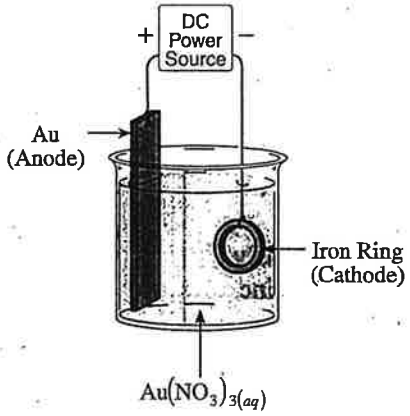
$$MnO_4^- + As_2O_3 \rightarrow Mn^{2+} + AsO_4^{3-} \quad \text{(acidic)}$$

$$\left(5e^- + 8H^+ + MnO_4^- \rightarrow Mn^{2+} + 4H_2O\right) \times 4$$
$$\left(5H_2O + As_2O_3 \rightarrow 2AsO_4^{3-} + 10H^+ + 4e^-\right) \times 5$$

$$\cancel{20e^-} + 32H^+ + 4MnO_4^- + 25H_2O + 5As_2O_3 \rightarrow 4Mn^{2+} + 16H_2O + 10AsO_4^{3-} + 50H^+ + \cancel{20e^-}$$

$$4MnO_4^- + 9H_2O + 5As_2O_3 \rightarrow 4Mn^{2+} + 10AsO_4^{3-} + 18H^+$$

**5.** *Draw an electrolytic cell that could be used to plate an iron ring with gold. Be sure to include all of the necessary parts. In addition, label the anode, solution used and composition of the electrodes.* **(3 marks)**




**6.** *Balance the following redox reaction in basic solution:* **(4 marks)**

$$ClO_3^- + N_2H_4 \rightarrow NO_3^- + Cl^- \quad \text{(basic)}$$

$$\left(6e^- + 6H^+ + ClO_3^- \rightarrow Cl^- + 3H_2O\right) \times 7$$
$$\left(6H_2O + N_2H_4 \rightarrow 2NO_3^- + 16H^+ + 14e^-\right) \times 3$$

$$\cancel{42e^-} + 42H^+ + 7ClO_3^- + 18H_2O + 3N_2H_4 \rightarrow 7Cl^- + 21H_2O + 6NO_3^- + 48H^+ + \cancel{42e^-}$$

$$7ClO_3^- + 3N_2H_4 \rightarrow 7Cl^- + 3H_2O + 6NO_3^- + 6H^+$$
$$6OH^- \qquad\qquad 6OH^-$$

$$7ClO_3^- + 6OH^- + 3N_2H_4 \rightarrow 7Cl^- + 9H_2O + 6NO_3^-$$

**7.** **a)** *Write the anode and cathode half-reactions that would occur.* **(2 marks)**

Anode: $H_2O \rightarrow 2H^+ + \frac{1}{2}O_{2(g)} + 2e^-$ **(1 mark)**

Cathode: $Co^{2+} + 2e^- \rightarrow Co_{(s)}$ **(1 mark)**

**b)** *What is observed when bromthymol blue is added to the solution around the anode?* **(1 mark)**

Bromthymol blue will turn yellow.

# ANSWERS AND SOLUTIONS
# UNIT TEST 5 – OXIDATION–REDUCTION

| | | | | | |
|---|---|---|---|---|---|
| 1. B | 5. D | 9. B | 13. C | 17. A | 21. B |
| 2. B | 6. C | 10. D | 14. D | 18. B | 22. C |
| 3. C | 7. C | 11. D | 15. B | 19. C | 23. D |
| 4. D | 8. B | 12. B | 16. B | 20. B | WR1–8. See Solution |

**1. B**

The oxidation number of S in $SO_4^{2-}$ is +6, its oxidation number in $S_2O_8^{2-}$ is +7. This is oxidation.

**2. B**

In $MnO_4^-$, O has an oxidation number of –2. $4 \times -2 = -8$. Since the total charge of the ion is –1, Mn is +7.

**3. C**

"More difficult to reduce" means *weaker oxidizing agent.* The only oxidizing agent from the list given which is weaker than $H^+$ is $Zn^{2+}$. (See *Data Booklet*, page A8).

**4. D**

The half-reactions are $Cu \rightarrow Cu^{2+} + 2e^-$ and

$NO_3^- + 4H^+ + 3e^- \rightarrow NO_{(g)} + 2H_2O$

Balancing electrons and adding leads to the correct answer.

**5. D**

$2MnO_4^- + 5Sn^{2+} + 16H^+ \rightarrow 2Mn^{2+} + 5Sn^{4+} + 8H_2O$

| $n1$ | $n2$ |
|---|---|
| 19.3 mL | 10.0 Ml |
| 0.10 M | $c = ?$ |

$n_1 = 0.10\ \text{M} \times 19.3\ \text{mL} = 1.9\ \text{mmol}$

$n_2 = 1.9\ \text{mmol} \times \frac{5}{2} = 4.8\ \text{mmol}$

$\left[Sn^{2+}\right] = \frac{4.8\ \text{mmol}}{10.0\ \text{mL}} = 0.48\ \text{M}$

**6. C**

A cell in a car battery **must** have a *positive* $E^\circ_{cell}$. Therefore $PbO_2$ must react at the cathode, and Pb at the anode as shown below:

cathode:
$PbO_2 + 4H^+ + SO_4^{2-} + 2e^- \rightarrow PbSO_4 + 2H_2O$ $E^\circ_{reduction} = +1.69$ V

anode:
$Pb + SO_4^{2-} \rightarrow PbSO_4 + 2e^-$ $E^\circ_{reduction} = -0.36$ V

cell:
$Pb + PbO_2 + 2SO_4^{2-} + 4H^+ \rightarrow 2PbSO_4 + 2H_2O$ $E^\circ_{cell} = (+1.69 - {}^-0.36)\text{V} = +2.05\text{V}$

**7. C**

The anode half-reaction is an oxidation half-reaction which will involve the strongest reducing agent. When all the cell reagents are considered: $Ag_{(s)}$, $Ag^{+}_{(aq)}$, $NO_3^{-}{}_{(aq)}$, $Zn_{(s)}$, $Zn^{2+}_{(aq)}$, and $H_2O_{(l)}$ are considered, the strongest reducing agent is $Zn_{(s)}$.

**8. B**

Considering the cell reagents: $Au_{(s)}$, $H_{2(g)}$, $H^{+}_{(aq)}$, $Cl^{-}_{(aq)}$, $Zn^{2+}_{(aq)}$, $Zn_{(s)}$, $NO_3^{-}{}_{(aq)}$, and $H_2O_{(l)}$, the strongest oxidizing agent is $H^{+}_{(aq)}$ while the strongest reducing agent is $Zn_{(s)}$. Using $E^{\circ}_{reduction}$ from the *Data Booklet* page 8 leads to:

$$E^{\circ}_{cell} = E^{\circ}_{reduction_{cathode}} - E^{\circ}_{reduction_{anode}}$$
$$= (0.00 - {}^{-}0.76)\,V = +0.76\,V$$

**9. B**

In Reaction I, copper has an oxidation number of +1 before and after the reaction. Neither oxidation nor reduction of copper has occurred. In Reaction II, copper has an oxidation number of +1 before the reaction and 0 after the reaction. Reduction has occurred.

**10. D**

"Pure aluminum" is a very strong reducing agent and thus is not usually pure aluminum, it is already in an oxidized state as evidenced by its greyish colour. This oxidized aluminum protects the *truly pure* aluminum underneath.

**11. D**

Considering the cell reagents: $Ca^{2+}_{(aq)}$, $I^{-}_{(aq)}$, $C_{(s)}$, and $H_2O_{(l)}$, the strongest oxidizing agent is $H_2O_{(l)}$ and the strongest reducing agent is $I^{-}_{(aq)}$. The following half-reactions occur:

cathode (reduction):
$2H_2O_{(l)} + 2e^{-} \rightarrow H_{2(g)} + 2OH^{-}_{(aq)}$

anode (oxidation): $2\,I^{-}_{(aq)} \rightarrow I_{2(s)} + 2\,e^{-}$

**12. B**

The oxidation number of H is +1 in most compounds. The oxidation number of O is –2 in most compounds. To determine the oxidation number of C it is a matter of adding oxidation numbers multiplied by their coefficients to add to 0 in compounds or to the ion charge in ions.

In $C_2H_3O_2^{-}$, $3\times1^{+}$ for H = $^{+}3$ and $2\times{}^{-}2$ for O is $4^{-}$.

$4^{-} + 3^{+} = 1^{-}$, which is the charge of the ion. Carbon's oxidation number is therefore 0.

**13. C**

$Zn_{(s)}$ is the strongest reducing agent, and $CO^{2+}_{(aq)}$ is the strongest oxidizing agent of all the chemical species present, using the *Standard Reduction* Potentials *of Half–Cells* table on page A8 of your *Data Booklet*. Half-reactions can be added to produce the net-ionic reaction equation.

**14. D**

Following the steps to produce a balanced half-reaction in acidic solution:

$Br_2 \rightarrow BrO_3^{-}$

$Br_2 + 6\,H_2O \rightarrow 2\,BrO_3^{-}$

$Br_2 + 6\,H_2O \rightarrow 2\,BrO_3^{-} + 12\,H^{+}$

$Br_2 + 6\,H_2O \rightarrow 2\,BrO_3^{-} + \underline{12}\,H^{+} + \underline{10}\,e^{-}$

**15. B**

Starting with the given redox equation, add $OH^{-}$ to both sides to change all of the $H^{+}$ to $H_2O$. Cancel the $H_2O$'s on the 2 sides to get the final equation.

$$\text{Given: } 5\,H_2O + 2\,CoCl_2 + OCl^{-} \longrightarrow 2\,Co(OH)_3 + 5\,Cl^{-} + 4\,H^{+}$$
$$+4\,OH^{-} \qquad\qquad +4\,OH^{-}$$
$$4\,OH^{-} + H_2O + 2\,CoCl_2 + OCl^{-} \longrightarrow 2\,Co(OH)_3 + 5\,Cl^{-}$$

**16. B**

The electrode, component 1, must be either carbon or a metal. The salt bridge solution, component 3 must be an ionic compound, and the cell output is running to component 2 which could be either a voltmeter or a light bulb.

**17. A**

Listing all the components of the cell: (Ag, $Ag^+$, $NO_3^-$, $Cl_2$, $Cl^-$, $Na^+$, $H_2O$), Ag is found to be the strongest reducing agent. The reducing agent gets oxidized and oxidation occurs at the anode of the cell.

**18. B**

$Ag_{(s)} \rightarrow Ag^+_{(aq)} + e^-$ occurs at the anode, so its mass decreases. The strongest oxidizing agent is $Cl_{2(g)}$. It reacts at the cathode according to the following half-reaction: $Cl_{2(g)} \rightarrow 2\,e^- \rightarrow 2\,Cl^-_{(aq)}$. The $Cl^-_{(aq)}$ becomes part of the solution; there is no physical change to the cathode itself.

**19. C**

Referring to the Standard *Reduction Potentials of Half-Cells* table on page 8 of your *Data Booklet* for $E_r^\circ$ values,

$$E^\circ_{net} = E^\circ_{r\ cathode} - E^\circ_{r\ anode} = +1.36\,V - (+0.80\,V) = +0.56\,V$$

**20. B**

Start by listing all the chemical species present in the cell:

$Zn^{2+}_{(aq)}, Zn_{(s)}, NO^-_{3(aq)}, Pb^{2+}_{(aq)}, PB_{(s)}, H_2O_{(l)}$.

The strongest oxidizing agent is $Pb^{2+}_{(aq)}$. It undergoes reduction according to the following half-reaction:

$Pb^{2+}_{(aq)} + 2\,e^- \rightarrow Pb_{(s)}$.

Therefore $\left[Pb^{2+}_{(aq)}\right]$ is decreasing.

**21. B**

Plating will occur at the cathode: $Cu^{2+}_{(aq)} + 2\,e^- \rightarrow Cu^{2+}_{(aq)} + 2\,e^-$. The anode should be copper metal. It will undergo oxidation: $Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2\,e^-$, and will replenish the $Cu^{2+}_{(aq)}$ lost at the cathode as it is used in plating the object.

**22. C**

$Cu_{(s)}$ will still form at the cathode by the same half-reaction, but this time it will be $Cu_{(s)}$ plating on top of $Cu_{(s)}$. The half-reaction is unchanged: $Cu^{2+}_{(aq)} + 2\,e^- \rightarrow Cu_{(s)}$. At the anode it will now be zinc undergoing oxidation: $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2\,e^-$

**23. D**

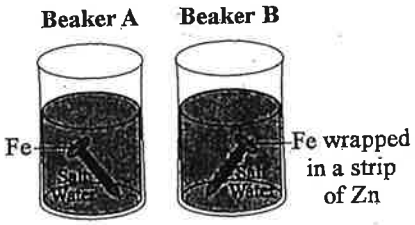
In molten KCl the only chemical species present besides the electrodes are $K^+_{(l)}$ and $Cl^-_{(l)}$. The half-reactions are:

cathode: $K^+_{(l)} + e^- \rightarrow K_{(l)}$

anode: $2Cl^-_{(l)} + e^- \rightarrow Cl_{2(g)} + 2\,e^-$

## Written Response

**1.** *Consider the following experimental results:*




$Ce^{4+} + Pd \rightarrow Pd^{2+} + Ce^{3+}$

$In^{3+} + Cd \rightarrow$ no reaction

$Pd^{2+} + In^{2+} \rightarrow In^{3+} + Pd$

$Cd^{2+} + Pd \rightarrow$ no reaction

*Use these results to complete the table of reduction half-reactions below.* **(3 marks)**

*Solution:*

| Oxidizing Agents | | Reducing Agents |
|---|---|---|
| $Ce^{4+} + e^-$ | $\rightleftarrows$ | $Ce^{3+}$ |
| $Pd^{2+} + 2e^-$ | $\rightleftarrows$ | $Pd$ |
| $Cd^{2+} + 2e^-$ | $\rightleftarrows$ | $Cd$ |
| $In^{3+} + e^-$ | $\rightleftarrows$ | $In^{2+}$ |

STRONGEST ↑ WEAKEST

WEAKEST ↓ STRONGEST

**2.** *Balance the following equation.* **(3 marks)**

$Cr_2O_7^{2-} + C_2O_4^{2-} \rightarrow Cr^{3+} + CO_2$ (acidic)

*Solution:*

$$\left(6e^- + 14H^+ + Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O\right) \times 1$$

$$\left(C_2O_4^{2-} \rightarrow 2CO_2 + 2e^-\right) \times 3$$

$$14H^+ + Cr_2O_7^{2-} + 3C_2O_4^{2-} \rightarrow 2Cr^{3+} + 7H_2O + 6CO_2$$

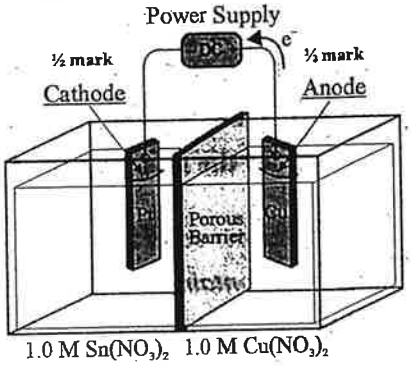
**(3 marks)**

**3.** *Consider the following electrolytic cell which contains a porous barrier to prevent general mixing of solutions.*

**a)** *Label the anode and cathode in the space provided on the diagram above.* **(1 mark)**

*Solution:*

See diagram below.




**b)** *Write an equation for the overall cell reaction.* **(2 marks)**

*Solution:*

$Sn^{2+} + Cu_{(s)} \rightarrow Cu^{2+} + Sn_{(s)}$

**c)** *Calculate the minimum theoretical voltage required for this reaction under standard conditions.* **(1 mark)**

*Solution:*

0.48 V

**4.** *Balance the following redox equation:*

$H_2S + CrO_4^{2-} \rightarrow S_8 + Cr^{3+}$ (acidic)

**(4 marks)**

*Solution:*

***Example:***

$$(8H_2S \rightarrow S_8 + 16H^+ + 16e^-) \times 3$$
$$(3e^- + 8H^+ + CrO_4^{2-} \rightarrow Cr^{3+} + 4H_2O) \times 16$$

$$24H_2S + \cancel{128H^+} + 16CrO_4^{2-} \rightarrow 3S_8 + \cancel{48H^+} + 16Cr^{3+} + 64H_2O$$

$$24H_2S + 80H^+ + 16CrO_4^{2-} \rightarrow 3S_8 + 16Cr^{3+} + 64H_2O$$

**2 marks (1 mark for each half-reaction)**

**1 mark for the correct electron ratio**

**1 mark for the final balanced equation**

**5.** *An excess of copper solid is dropped into a solution which contains* $AgNO_3$, $Fe(NO_3)_3$ *and* $Zn(NO_3)_2$. *Write the equations for any reduction half-reactions that occur over time under standard conditions.* **(2 marks)**

*Solution:*

***Example:***

$Ag^+ + e^- \rightarrow Ag$ **(1 mark)**

$Fe^{3+} + e^- \rightarrow Fe^{2+}$ **(1 mark)**

**6. a)** *Predict what should happen to the Fe in Beaker A.* **(1 mark)**

*Solution:*

***Example:***

***Prediction***: The iron is oxidized.

**(1 mark)**

**b)** *Predict what should happen to the Fe in Beaker B. Explain.* **(2 marks)**

*Solution:*

***Example:***

Prediction: Nothing happens to the Fe. ← 1 mark

Explanation: Zn is oxidized and protects the Fe.

**OR**

The Fe is cathodically protected by the Zn. } ← 1 mark

**7.** *What is the percent by mass of the* $CaC_2O_4$ *in the original sample?* **(4 marks)**

***For Example:***

$$\text{Moles of } MnO_4^- = 0.01570\,\text{L} \times 0.101\,\text{mol/L}$$
$$= 1.5857 \times 10^{-3}\,\text{mol}$$
$$\text{Moles of } C_2O_4^{2-} = 1.5857 \times 10^{-3}\,\text{mol}\,MnO_4^- \times \frac{5\,\text{mol}\,C_2O_4^{2-}}{2\,\text{mol}\,MnO_4^-}$$
$$= 3.9643 \times 10^{-3}\,\text{mol}$$
$$\text{Moles}\,CaC_2O_4 = \text{Moles}\,C_2O_4^{2-}$$
$$\text{Mass of } CaC_2O_4 = 3.9643 \times 10^{-3}\,\text{mol} \times \frac{128.1\,\text{g}}{1\,\text{mol}}$$
$$= 5.0782 \times 10^{-1}\,\text{g}\,CaC_2O_4$$
$$\%\,CaC_2O_4 = \frac{0.50782\,\text{g}}{0.803\,\text{g}} \times 100\%$$
$$= 63.2\%$$

(Deduct $\frac{1}{2}$ mark for incorrect significant figures.)

**8.** **a)** *Write the reduction half-reaction.* **(1 mark)**

***For Example:***

$$\frac{1}{2}O_{2(g)} + 2H^{+}(10^{-7}\ M) + 2e^{-} \rightarrow H_2O$$

**b)** *What metal could be attached to the sample to prevent the corrosion of the zinc? Explain.* **(2 marks)**

***For Example:***

Metal: Magnesium (Mg)

Explanation: The magnesium is more easily oxidized than the zinc.