ALL SAINTS'
COLLEGE

Name: **ANSWERS**   Mark = _____ / 48

Part One: Multiple Choice Section   **8 marks**

*Answer by placing a cross through, or a circle around, the letter of the most correct answer.*

1.   In which of the following species does platinum have the lowest oxidation number?

   A. $H_2PtC\ell_6$ **+4**

   B. $NaPtC\ell_4$ **+3**

   C. PtO_2 **+4**

   D. $PtCr_2O_7$ **+2**

2.   In which of the following does the oxidation number of an element decrease by 3?

   A. NO $\rightarrow$ NO_3^-

   B. MnO_2 $\rightarrow$ MnO_4^-

   C. PH_4^+ $\rightarrow$ P

   D. CrO_4^{2-} $\rightarrow$ Cr^{3+}
           (+6)   **(+3)**

3.   Consider the following unbalanced equation.

$$I^-(aq) \ + \ IO_3^-(aq) \ + \ H^+(aq) \ \rightarrow \ I_2(s) \ + \ H_2O(\ell)$$
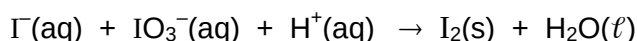
   Which one of the following statements is true?

   A. H^+ is reduced.

   B. IO_3^- is not the oxidising agent.

   C.   The oxidising agent is I_2.

   D.   **I^- is the reducing agent.** $I^- \rightarrow I_2$
                                                      (-1)   (0)
   **I^- is oxidised, $\therefore$ it is the reductant**
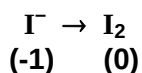
4.   A 20 mL sample a of 0.1 mol L−1 solution of metal ions completely reacted with 20 mL of a 0.1 mol L−1 solution of iodine. The iodine reacted according to the equation.

$$I_2(aq) + 2\,H_2O(l) \rightarrow 2\,HOI(aq) + H^+(aq) + 2\,e^-$$

If the original oxidation number of the metal ions was +3, then their oxidation number after the reaction would be

A.   +1 $n(M^{3+}) = n(I_2)$, ∴ $M^{3+} + 2\,e^- \rightarrow M^+$

B.   +4

C.   +5

D.   +2

5.   The diagram below represents a Daniell Cell, a battery from the mid 1800's.
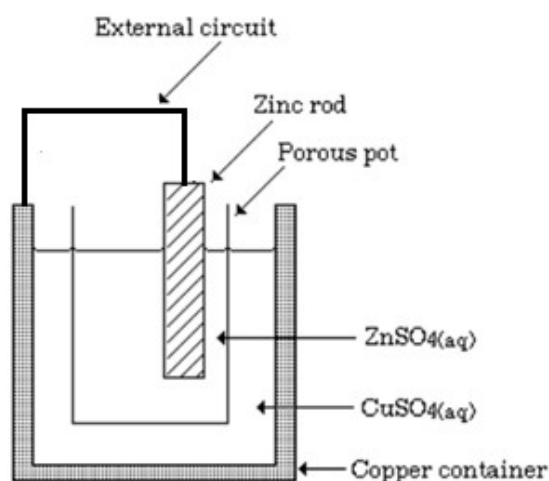
When the zinc rod and the copper container are connected as part of a completed electrical circuit, a current flows in this circuit. When the cell is operating, which one of the following statements is true?

A.   The copper container gradually dissolves.

B.   Electrons flow from the copper to the zinc through the external circuit.

C.   Zinc is deposited around the zinc rod.

D.   **Sulfate ions migrate through the porous pot from the copper compartment to the zinc compartment.**

Zn(s) is anode: **$Zn(s) \rightarrow Zn^{2+} + 2\,e^-$**
Cu(s) is the cathode: **$Cu^{2+}(aq) + 2\,e^- \rightarrow Cu(s)$**

6.   The EMF of a cell composed of a Ga^+/Ga^{3+} half-cell and a $C\ell_2/C\ell^-$ half-cell is 2.06 V under standard conditions. A cell composed of an I_2/I^- half-cell and a $C\ell_2/C\ell^-$ half-cell has an EMF of 0.82 V under standard conditions. In each cell the $C\ell_2$ acts as the oxidising agent.

If a cell were formed from a I_2/I^- half-cell and a Ga^+/Ga^{3+} half-cell under standard conditions, then its EMF would be:

A.   0.16 V

B.   0.70 V

C.   **1.24 V**   **Difference between Ga^+/Ga^{3+} and I_2/I^- in comparison to $C\ell_2/C\ell^-$**

D.   2.88 V   **is 2.06 – 0.82 = 1.24 V**

Questions 7 and 8 refer to the following standard reduction potentials.

Half-reaction	E^o
$Co^{3+}(aq) + e^- ⇌ Co^{2+}(aq)$	+ 1.30 V
$Sn^{4+}(aq) + 2 e^- ⇌ Sn^{2+}(aq)$	+ 0.15 V
$Sn^{2+}(aq) + 2 e^- ⇌ Sn(s)$	– 0.14 V
$Cr^{3+}(aq) + 3 e^- ⇌ Cr(s)$	– 0.74 V
$Be^{2+}(aq) + 2 e^- ⇌ Be(s)$	– 1.85 V

7.   Of the species listed below, the strongest reducing agent is:

A. Sn^{4+}

B. Be^{2+}

C.   Sn

D.   **Be**   **Be has the highest 'standard oxidation potential' (+1.85 V)**

8.   Which of the following reactions would occur under standard conditions?

A. $2 Co^{2+}(aq) + Sn^{2+}(aq) ⇌ 2 Co^{3+}(aq) + Sn(s)$ **– 1.44 V**

B. $Sn^{2+}(aq) + Be^{2+}(aq) ⇌ Sn^{4+}(aq) + Be(s)$ **– 2.00 V**

C.   **$Be(s) + Sn^{2+}(aq) ⇌ Be^{2+}(aq) + Sn(s)$**   **+ 1.71 V**

D. $2 Cr^{3+}(aq) + 3 Sn^{2+}(aq) ⇌ 2 Cr(s) + 3 Sn^{4+}(aq)$ **– 0.89 V**


End of Part One

Part Two: Short Answer Section   **40 marks**
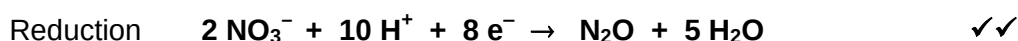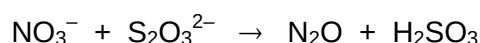
Write all answers in the spaces provided. When calculating numerical answers, show your working or reasoning clearly unless instructed otherwise. Final answers to calculations should be expressed to three significant figures and include appropriate units where applicable.

Question 9   **(3 marks)**

Assign oxidation numbers to the element in bold type in each of following:

(a) $Na_2\mathbf{S}_2O_3$ **+2** ✓   (b) $H_2\mathbf{C}_2O_4$ **+3** ✓   (c) $\underline{\mathbf{Cr}}(OH)_4^{-}$ **+3** ✓
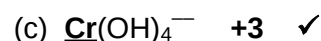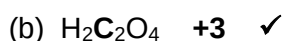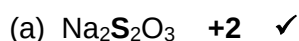
Question 10   **(5 marks)**

Construct oxidation and reduction half-equations and then the write the full redox equation for the following:

$$NO_3^- + S_2O_3^{2-} \rightarrow N_2O + H_2SO_3$$

Oxidation $\mathbf{S_2O_3^{2-} + 3\,H_2O \rightarrow 2\,H_2SO_3 + 2\,H^+ + 4\,e^-}$ ✓✓

Reduction $\mathbf{2\,NO_3^- + 10\,H^+ + 8\,e^- \rightarrow N_2O + 5\,H_2O}$ ✓✓

Overall $\mathbf{2\,NO_3^- + 2\,S_2O_3^{2-} + H_2O + 6\,H^+ \rightarrow N_2O + 4\,H_2SO_3}$ ✓

Question 11   **(2 marks)**

(a)   Which one of the following is unlikely to be produced by the reduction of nitrous acid, HNO_2? Circle your answer.

                 NH_4^+ NO   **NO_2** N_2

                                               ✓   (1 mark)

(b)   Explain your answer to (a).

         The oxidation state of nitrogen in HNO_2 is +3.

         The oxidation state of nitrogen in NO_2 is +4, since this is higher it could not be formed by the reduction of HNO_2.   ✓
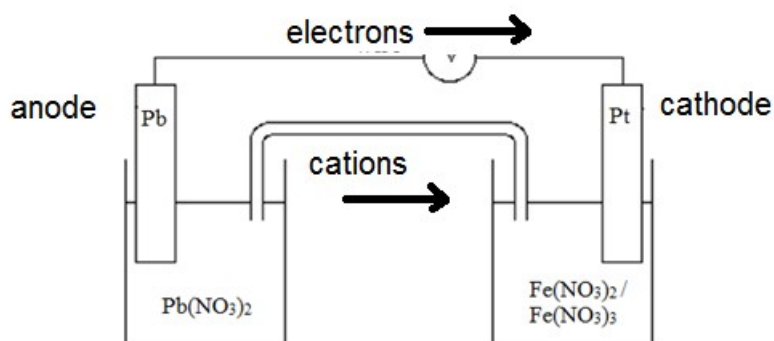
                                               (1 mark)

Question 12 (9 marks)

Consider the following electrochemical cell:

(a)   On the diagram above

    (i)   identify the anode and cathode.   ✓ **(both needed)**

    (ii)   indicate the direction of flow of electrons in the wire   ✓

    (iii)   indicate the direction of flow of cations within the salt-bridge   ✓

                                                                         (3 marks)

(b)   Write equations for the reactions occurring at the anode and cathode.

    anode   **$Pb(s) \rightarrow Pb^{2+}(aq) + 2\ e^{-}$**   ✓

    cathode   **$Fe^{3+}(aq) + e^{-} \rightarrow Fe^{2+}(aq)$**   ✓

                                                                         (2 marks)

(c)   Assuming standard conditions, what will be the reading on the voltmeter? **0.90 V** ✓

                                                                         (1 mark)

(d)   Describe an observation that would be made in each half-cell.

    $Pb \mid Pb^{2+}$ **decrease in size of the anode (electrode)** ✓

    $Pt \mid Fe^{2+}, Fe^{3+}$ **solution becomes less brown / more green**   ✓

                                                                         (2 marks)

(e)   State the function of the salt bridge.

   **To maintain electrical neutrality in each half-cell**   ✓
    (also accept 'allows for the flow of ions')

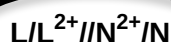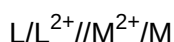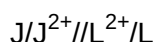   (1 mark)

Question 13 **(4 marks)**

Pieces of four different metals labelled J, L, M and N were separately placed into solutions containing J^{2+}, L^{2+}, M^{2+} and N^{2+} ions respectively and the observations listed below:

I   Metal L remained unchanged in all four solutions.

II   Displacement reactions were observed when metal M was placed in solutions of L^{2+} and J^{2+}, but not in N^{2+}.

(a)   Which of the following combination of half-cells would produce an electrochemical cell with the greatest cell potential? Circle your answer.

$J/J^{2+}//L^{2+}/L$ $L/L^{2+}//M^{2+}/M$ **$L/L^{2+}//N^{2+}/N$** $M/M^{2+}//J^{2+}/J$
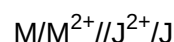
                                                                     ✓   (1 mark)

(b)   Explain your answer to (a).

The greatest cell potential is produced by combining the strongest oxidant (i.e. the greatest reduction potential) and the strongest reductant (i.e. greatest oxidation potential) ✓

Metal L could not displace any other metal ions from solution, suggesting that $L^{2+}(aq) + 2\,e^- \rightarrow L(s)$ has the highest reduction potential and that $L^{2+}(aq)$ is the strongest oxidant. ✓

Metal M displaces all other metal ions from solution apart from N^{2+}, suggesting that $N(s) \rightarrow N^{2+}(aq) + 2\,e^-$ has the highest oxidation potential and that $N(s)$ is the strongest reductant. ✓

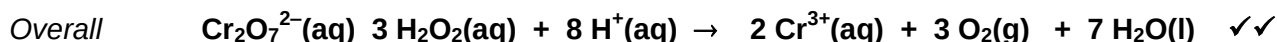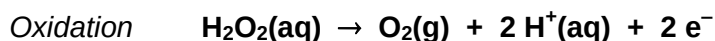   (3 marks)

Question 14 **(9 marks)**

A solution of acidified potassium dichromate may be used in a titration to analyse a hydrogen peroxide solution.

(a)   Use the table of Standard Reduction Potentials to obtain the oxidation and reduction half equations and then write an overall equation for the reaction that occurs when potassium dichromate solution is added to a solution containing hydrogen peroxide and sulfuric acid.

Oxidation $H_2O_2(aq) \rightarrow O_2(g) + 2 H^+(aq) + 2 e^-$

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

Overall $Cr_2O_7^{2-}(aq)\ 3 H_2O_2(aq) + 8 H^+(aq) \rightarrow 2 Cr^{3+}(aq) + 3 O_2(g) + 7 H_2O(l)$ ✓✓

(2 marks)

(b)   A solution of hydrogen peroxide can be used to bleach hair. A sample of a commercial product containing hydrogen peroxide, which is used by hairdressers, was analysed in the following way.
- a pipette was used to transfer 20.00 mL of commercial hairdressers' hydrogen peroxide into a 250.0 mL volumetric flask and made up to the graduation mark with distilled water.
- another pipette was used to place 20.00 mL portions of the diluted solution in a titration vessel along with 5.00 mL of 2.00 mol L^{-1} sulfuric acid.
- the solution was titrated against with 0.02056 mol L^{-1} potassium dichromate, with the following results:

Final reading (mL)	29.50	29.53	29.35	28.74	28.62
Initial reading (mL)	0.11	1.55	0.41	0.81	0.66
Titre volume (mL)	~~29.39~~	27.98	~~28.94~~	27.93	27.96

Calculate the concentration, in mol L^{-1}, of hydrogen peroxide in the commercial hairdressers' product. Express your answer to three significant figures.

$V(Cr_2O_7^{2-}) = (27.98 + 27.93 + 27.96)/3 = 27.96$ **mL**   ✓

$n(Cr_2O_7^{2-}) = c.V = 0.02056 \times 0.02796 = 0.0005748$ **mol**   ✓

$n(H_2O_2)_{20mL\ dilute} = 3.n(Cr_2O_7^{2-}) = 0.001724$ **mol**   ✓

$n(H_2O_2)_{250mL\ dilute} = n(H_2O_2)_{20mL\ dilute} \times 250/20 = 0.02155$ **mol**   ✓

$n(H_2O_2)_{20\ mL\ conc} = n(H_2O_2)_{250mL\ dilute} = 0.02155$ **mol**   ✓

$[H_2O_2]_{conc} = n/V = 0.02155/0.0200 = 1.08$ **mol L^{-1}**   ✓

3 sig figs   ✓

(7 marks)

Question 15 **(8 marks)**

In the late 1980's a new type of electrochemical cell was developed, with some of the research conducted at Murdoch University. The cell is filled with zinc bromide solution and divided in half by a plastic membrane. One half of the cell also contains liquid bromine. When the cell is operating, metallic zinc is converted to zinc ions in one half, while bromine is converted to bromide ions in the other half.

(a) Write the oxidation and reduction half equations for the reactions occurring in the "zinc bromide battery".

Oxidation $Zn(s) \rightarrow Zn^{2+}(aq) + 2\,e^{-}$ ✓

Reduction $Br_2(l) + 2\,e^{-} \rightarrow 2\,Br^{-}(aq)$ ✓
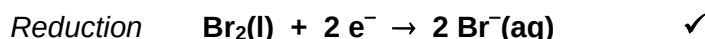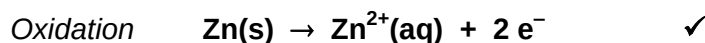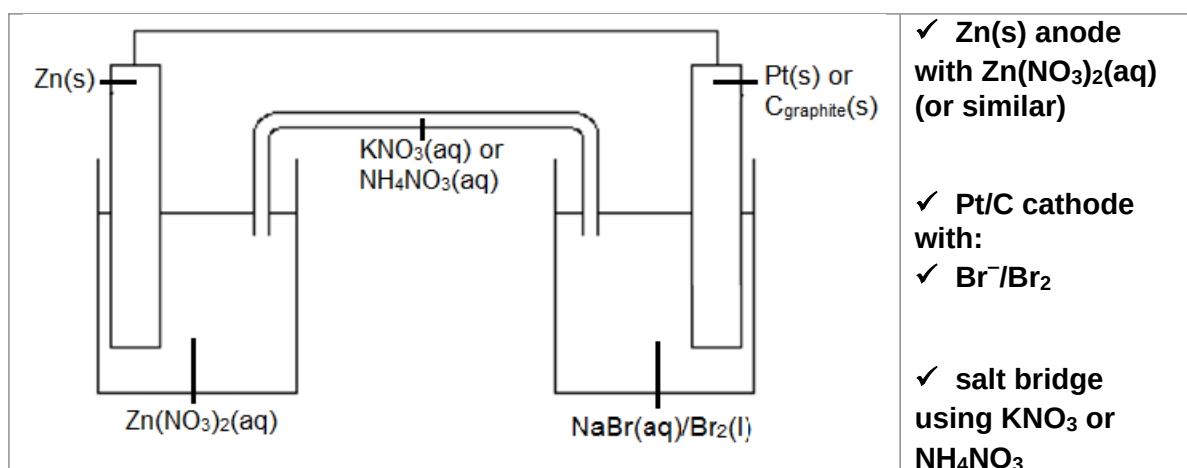
(2 marks)

(b) Suggest a non-metallic material that would be suitable to use as the cathode.

graphite ✓

(1 mark)

(c) In the space below, draw a diagram to show a version of the "zinc bromide battery" that you could construct using equipment typically found in a school laboratory. To complicate matters, the school does not have any zinc bromide.

On your diagram, clearly label, using either name or formula, all materials used in the construction of the cell.

✓ **Zn(s) anode with $Zn(NO_3)_2(aq)$ (or similar)**

✓ **Pt/C cathode with:**
✓ **Br^{-}/Br_2**

✓ **salt bridge using KNO_3 or NH_4NO_3**

(4 marks)

(d) If the above cell was operated under standard conditions,
then what voltage could it be expected to produce? **1.84 V** ✓

(1 mark)

End of Test

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