ARANMORE CATHOLIC COLLEGE

YEAR 12 CHEMISTRY 3A3B - 2010

TOPIC TEST: ELECTROCHEMISTRY

NAME: SOLUTIONS

MARK: **/50**

INSTRUCTIONS:

- Answer ALL questions in the spaces provided.
- **Show ALL working** clearly to obtain maximum marks, as shown in brackets.
- Use of approved calculators is permitted.

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1. Write a balanced ionic equation for the reaction between the $Cr_2O_7^{2-}$ ion and SO_2 to produce SO_4^{2-} and Cr^{3+}.

[2 marks]

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

$$SO_2 + 2H_2O \rightarrow SO_4^{2-} + 4H^+ + 2e^-$$

$$Cr_2O_7^{2-} + 3SO_2 + 2H^+ \rightarrow 2Cr^{3+} + 3SO_4^{2-} + H_2O$$

2. Determine the oxidation number (or state) of the bracketed element in each of the following:

a)	$HS_2O_3^-$	**(S)**	+2	d)	$Mg_3(PO_4)_2$	**(P)**	+5
b)	Mn_2O_3	**(Mn)**	+3	e)	NaOCl	**(Cl)**	+1
c)	IO_3^-	**(I)**	+5	f)	NH_4Cl	**(N)**	−3

[3 marks]

3. Write balanced equations for the reactions which occur in the following experiments. Use ionic equations where appropriate. Describe observations such as colours, precipitates, evolution of gases, etc.

[7 marks]

a) Chlorine gas is bubbled through a potassium bromide solution.

(1) $Cl_{2(g)} + 2Br^-_{(aq)} \rightarrow Br_{2(aq)} + 2Cl^-_{(aq)}$

(1) C'LESS SOLN TURNS RED.

b) A strip of lead metal is added to a zinc nitrate solution.

(1) NO RXN.

c) A strip of copper metal is added to a **concentrated** nitric acid solution.

(1) $Cu_{(s)} + 2NO^-_{3(aq)} + 4H^+_{(aq)} \rightarrow Cu^{2+}_{(aq)} + 2NO_{2(g)} + 2H_2O_{(l)}$

(1) METAL DISSOLVES, C'LESS SOLN TURNS BLUE AND A BROWN GAS EVOLVES.

d) A 1.0 mol L^{-1} iron (II) sulfate solution is added to 1.0 mol L^{-1} hydrogen peroxide solution.

(1) $2Fe^{2+}_{(aq)} + H_2O_{2(aq)} + 2H^+_{(aq)} \rightarrow 2Fe^{3+}_{(aq)} + 2H_2O_{(l)}$.

(1) PALE GREEN SOLN TURNS ORANGE/BROWN.

4. When potassium iodate (KIO_3) reacts with oxalic acid ($H_2C_2O_4$), a colourless gas is observed to form and the solution changes from colourless to brown, indicating the presence of iodine. Write down the overall balanced redox equation for this reaction and identify the oxidising and reducing agents.

[3 marks]

$$H_2C_2O_4 \rightarrow 2CO_2 + 2H^+ + 2e^-$$

(1)

$$2IO_3^- + 12H^+ + 10e^- \rightarrow I_2 + 6H_2O$$

(1) $$\underset{\substack{\uparrow \\ \text{OXIDISING} \\ \text{AGENT}}}{2IO_{3\,(aq)}^-} + \underset{\substack{\uparrow \\ \text{REDUCING} \\ \text{AGENT}}}{5H_2C_2O_{4\,(aq)}} + 2H^+_{(aq)} \rightarrow I_{2\,(aq)} + 10CO_{2\,(g)} + 6H_2O_{(l)}$$

(1)

5. a) Write the overall reaction for the disproportionation of copper (I) ions.

[2 marks]

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

b) Determine the standard potential for the above reaction.

[1 mark]

$$E^\circ_{RXN} = E^\circ_c - E^\circ_a$$
$$= +0.52 - 0.16$$
$$= +0.36\text{ V.}$$

c) Is this reaction spontaneous under standard conditions (why or why not)? Comment on this value with respect to the stability of the copper (I) ion.

[2 marks]

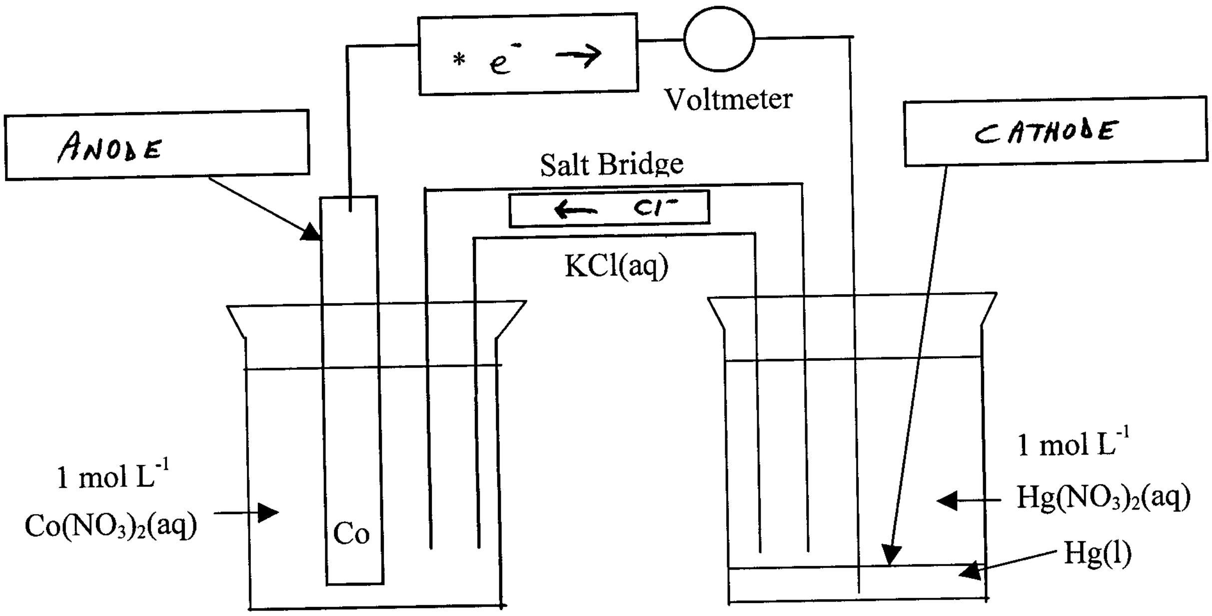
(1) YES, DUE TO +ve E°.

THIS SUGGESTS (UNDER STD CONDITIONS AT LEAST) THAT THE

(1) Cu(I) ION IS NOT STABLE AND IS LIKELY TO SPONTANEOUSLY DISPROPORTIONATE INTO $Cu_{(s)}$ AND Cu(II) IONS.

6. A demonstration electrochemical cell is made by assembling two half cells consisting of beakers with cobalt metal in contact with cobalt(II) nitrate solution and a layer of pure liquid mercury in contact with mercury(II) nitrate. The half cells are joined as shown in the diagram.

[8 marks]



(a) **In the boxes** provided on the above diagram:

1 MARK EACH.

(i) Label the ANODE.

(ii) Label the CATHODE.

(iii) Draw an arrow to indicate the direction of electron flow in the external circuit. (*)

(iv) Draw an arrow to indicate the direction of flow of the chloride ions through the salt bridge.

(b) Write balanced equations for the:

(1) (i) ANODE REACTION $Co \rightarrow Co^{2+} + 2e^-$ $-0.28V$

(1) (ii) CATHODE REACTION $Hg^{2+} + 2e^- \rightarrow Hg$ $+0.85V$

(1) (iii) OVERALL REACTION $Co_{(s)} + Hg^{2+}_{(aq)} \rightarrow Co^{2+}_{(aq)} + Hg_{(l)}$

(1) (c) The expected reading on the voltmeter is: ......+1.13...... volts.

7. Two methods of preventing the rusting of iron are galvanizing and tin plating.

a) What do these two methods have in common with respect to the way in which they protect the iron from corroding?

[1 mark]

– BOTH EXCLUDE O_2 AND MOISTURE FROM CONTACTING THE IRON SURFACE.

b) Describe the main advantage of galvanizing over tin plating with a clear explanation of what happens in each case.

[2 marks]

(1) – CONTINUES TO PROTECT IRON AFTER SURFACE SCRATCHED

(1) – ZINC ACTS AS SACRIFICIAL ANODE DUE TO LOWER E°_r.

(TIN WOULD ACCELERATE RUSTING IF SCRATCHED DUE TO HIGHER E°_r.)

c) Describe one other method of corrosion protection which is quite different to the above.

[2 marks]

(1) – e.g. APPLIED POT'L

(1) – DESCRIPTION / DIAGRAM

8. Write down the following chemical equations involved in the eventual rusting of iron:

a) The initial oxidation and reduction half-equations and associated overall equation.

[2 marks]

$Fe \rightarrow Fe^{2+} + 2e^-$ (ox)

(1) $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$ (RED)

(1) $2Fe_{(s)} + O_{2(g)} + 2H_2O_{(l)} \rightarrow 2Fe^{2+}_{(aq)} + 4OH^-_{(aq)}$

b) The formation of iron (II) hydroxide.

[1 mark]

$Fe^{2+}_{(aq)} + 2OH^-_{(aq)} \rightarrow Fe(OH)_{2(s)}$

c) The subsequent oxidation to iron (III) hydroxide.

[1 mark]

$4Fe(OH)_{2(s)} + O_{2(g)} + 2H_2O_{(l)} \rightarrow 4Fe(OH)_{3(s)}$

d) The formation of rust from the iron (III) hydroxide.

[1 mark]

$2Fe(OH)_{3(s)} \rightarrow Fe_2O_3 \cdot H_2O_{(s)} + 2H_2O_{(l)}$

9. Fuel cells are starting to be used in vehicles in order to save fossil fuels and to reduce greenhouse gas emissions. In one fuel cell, hydrogen gas and oxygen gas are bubbled across a platinum catalyst in acidic solution to produce the only product, water:

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$$

a) Write the anode and cathode half reactions by referring to the E° table of values.

[2 marks]

$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$ CATHODE

$H_2 \rightarrow 2H^+ + 2e^-$ ANODE

b) What would the output voltage of the cell be? (UNDER STD CONDITIONS)

[1 mark]

$$E^\circ_{CELL} = E^\circ_C - E^\circ_A$$
$$= 1.23 - 0$$
$$= 1.23\ V.$$

10. A 1.00 g sample of impure iron wire was dissolved in sulfuric acid to produce an acidified solution of iron (II) sulfate and was made up to 250.0 mL with distilled water. 20.00 mL samples of this solution were then titrated with a 0.0194 mol L^{-1} solution of $KMnO_4$ with the following results:

	Rough trial (mL)	Trial 1 (mL)	Trial 2 (mL)	Trial 3 (mL)
Final	13.96	26.84	39.66	14.02
Initial	0.60	13.96	26.84	1.16
Titre	13.36	12.88	12.82	12.86

(2)

Calculate the (percentage) purity of the iron wire.

[9 marks]

(1) AVERAGE TITRE = 12.85 mL (LAST 3 ONLY)

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

(1) $n(MnO_4^-) = c \times V = 0.0194 \times 0.01285$
$= 2.494 \times 10^{-4}$ mol.

(1) $n(Fe^{2+}) = 5\,n(MnO_4^-)$
$= 1.247 \times 10^{-3}$ mol

(1) IN 20.00 mL: $m(Fe) = n \times M$
$= 0.0696$ g

IN TOTAL (250.0 mL)

(1) $m(Fe) = \frac{250}{20} \times 0.0696$
$= 0.870$ g

(1) % PURITY $= \frac{0.870}{1.00} \times 100\%$
$= 87.0\%$.

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END OF TEST