Name ______________________

Class ____________ Date ____________

CHAPTER 21 WORKSHEET

Part I

If the following electrolytes were electrolyzed, predict what half reactions would occur at the anode and cathode. (Remember, 1.0*M* NaCl implies an aqueous solution of NaCl.)

1. Molten $ZnCl_2$; inert electrodes.

cathode $Zn^{2+} + 2e^- \longrightarrow Zn$

anode $2Cl^- \longrightarrow Cl_2 + 2e^-$

2. 1.0*M* Na_2SO_4; inert electrodes.

anode $H_2O \longrightarrow \frac{1}{2}O_2 + 2H^+ + 2e^-$

cathode $2H_2O + 2e^- \longrightarrow H_2 + 2OH^-$

3. Molten CuI_2; inert electrodes.

cathode $Cu^{2+} + 2e^- \longrightarrow Cu$

anode $2I^- \longrightarrow I_2 + 2e^-$

4. 1.0*M* KI; inert electrodes.

anode $2I^- \longrightarrow I_2 + 2e^-$

cathode $2H_2O + 2e^- \longrightarrow H_2 + 2OH^-$

5. 1.0*M* $NiSO_4$; nickel electrodes.

Anode $Ni \longrightarrow Ni^{2+} + 2e^-$

cathode $Ni^{2+} + 2e^- \longrightarrow Ni$

6. 1.0*M* HI; lead electrodes.

anode $Pb_{(s)} \longrightarrow Pb^{2+} + 2e^-$

cathode $2H_2^+ + 2e^- \longrightarrow H_{2(g)}$

7. 1.0*M* $ZnSO_4$; inert electrodes.

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

cathode $Zn^{2+} + 2e^- \rightarrow Zn_{(s)}$ overpotential effect

8. 1.0*M* NaCl; inert electrodes.

anode $2Cl^- \rightarrow Cl_{2(g)} + 2e^-$ overpotential effect

cathode $2H_2O + 2e^- \rightarrow H_{2(g)} + 2OH^-$

Part II

9. When a 1.0*M* KCl solution is electrolyzed using silver electrodes, a precipitate forms at the anode. Explain this result.

anode: $Ag_{(s)} \rightarrow Ag^+ + e^-$

The Ag^+ ions react with the Cl^- ions to produce $AgCl_{(s)}$.

10. If a 1.0*M* HCl solution is electrolyzed using platinum electrodes (inert), what minimum voltage must be applied? Predict the anode and cathode half-reactions, as well as the overall cell reaction.

anode $2Cl^- \rightarrow Cl_{2(g)} + 2e^-$ $E° = -1.36\ v$

cathode $2H^+ + 2e^- \rightarrow H_{2(g)}$ $E° = 0.00\ v$

overall $2Cl^- + 2H^+ \rightarrow Cl_{2(g)} + H_{2(g)}$ $E°cell = -1.36\ v$

Therefore, 1.36 v must be applied.

11. a. In the electrorefining of lead, lead bullion is used as the anode and pure lead is used as the cathode in an electrolytic solution containing Pb^{2+} ions. Lead bullion is primarily lead, but it does contain impurities such as silver and gold. What happens to these three metals at the anode during electrolysis?

The Pb is oxidized to Pb^{2+}. The Ag + Au are "left behind". These metals can be recovered in subsequent electrolytic processes

11. b. **Lead bullion may also contain trace amounts of impurities such as zinc metal. Describe what happens to this zinc during electrolysis, and explain why the pure lead cathode does not become contaminated with zinc.**

Although both the $Zn_{(s)}$ and the $Pb_{(s)}$ are oxidized at the anode, the Zn^{2+} ions are "trapped" in solution. Since, at the cathode the Pb^{2+} ions have a higher reduction potential than the Zn^{2+}, only the Pb^{2+} ions are reduced to $Pb_{(s)}$

12. **Why can aluminum metal not be produced by electrowinning Al(s) from an aqueous solution containing Al^{3+} ions? Write the cathode half-reaction that would occur.**

Even with hydrogen's high overpotential, H_2O is reduced before Al^{3+}.
Cathode: $2H_2O + 2e^- \longrightarrow H_{2(g)} + 2OH^-$

Answer Key

<u>Practice Exercise: Electrochemical Cells</u> TO HAND IN

1. Sketch and label an electrochemical cell that makes use of the following spontaneous redox reaction:

$Mg(s) + FeSO_4(aq) \longrightarrow MgSO_4(aq) + Fe(s)$

a. Label the anode.
Write the half-reaction that occurs at the anode. $Mg \rightarrow Mg^{2+} + 2e^-$

b. Label the cathode.
Write the half-reaction that occurs at the cathode. $Fe^{2+} + 2e^- \rightarrow Fe$

c. This electrochemical cell is joined using a salt bridge containing KNO_3.

What is the purpose of a salt bridge? To connect the two halves of the cell, allowing ions to migrate through it.

d. There are spectator ions involved in this process.

They are SO_4^{2-}, K^+, NO_3^-.

On your diagram, show the direction in which these 3 ions travel. $\left.\begin{matrix} SO_4^{2-} \\ NO_3^- \end{matrix}\right\}$ move toward anode K^+ moves toward cathode

e. In which direction do Mg^{2+} ions travel?
In which direction do Fe^{2+} ions travel?
} towards the cathode.

2. Aluminum will displace tin from solution according to the following equation:

$2Al(s) + 3Sn^{2+}(aq) \longrightarrow 2Al^{3+}(aq) + 3Sn(s)$

What would be the individual half-cell reactions if this were the cell reaction in an electrochemical cell? $Sn^{2+} + 2e^- \rightarrow Sn$ $Al \rightarrow Al^{3+} + 3e^-$

Which metal would be the anode and which the cathode?

Al would be anode ; Sn would be cathode.

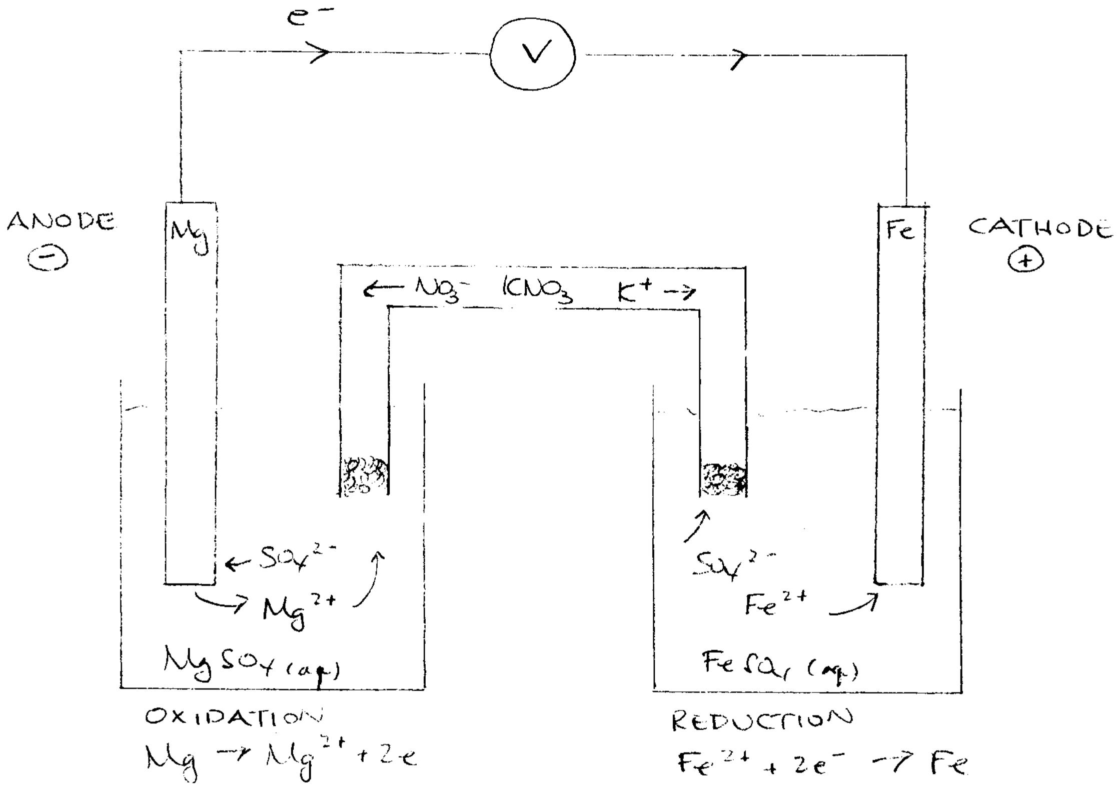
3. Design an electrochemical cell using Ag as the cathode.

Ag as cathode $Ag^+ + e^- \rightarrow Ag$ with $AgNO_3$ (aq)

Choose any metal reducing agent stronger than Ag (i.e. lower on right side of table)

eg. Cu as anode with $Cu(NO_3)_2$ (ag)

Ques. 1 a) to d) $Mg(s) + FeSO_4(aq) \rightarrow MgSO_4(aq) + Fe(s)$



Key.

Chemistry 12 Electrochemistry Worksheet No. 1

1. Balance the following half reactions and state whether oxidation or reduction is taking place.

a. BrO_3^- ----> Br^-	(basic)	$BrO_3^- + 3H_2O + 6e^- \rightarrow Br^- + 6OH^-$	reduction
b. $HClO_2$ ----> ClO_3^-	(acidic)	$HClO_2 + H_2O \rightarrow ClO_3^- + 3H^+ + 2e^-$	oxidation
c. Ga -------> $H_2GaO_3^-$	(basic)	$Ga + 4OH^- \rightarrow H_2GaO_3^- + H_2O + 3e^-$	oxidation
d. ReO_4^- ----> ReO_2	(acidic)	$ReO_4^- + 4H^+ + 3e^- \rightarrow ReO_2 + 2H_2O$	reduction
e. MnO_2 -----> MnO_4^-	(acidic)	$MnO_2 + 2H_2O \rightarrow MnO_4^- + 4H^+ + 3e^-$	oxidation
f. N_2O_4 -----> NO	(acidic)	$N_2O_4 + 4H^+ + 4e^- \rightarrow 2NO + 2H_2O$	reduction
g. CrO_4^{2-} ---> $Cr(OH)_3$	(basic)	$CrO_4^{2-} + 4H_2O + 3e^- \rightarrow Cr(OH)_3 + 5OH^-$	reduction
h. $Sn(OH)_6^{2-}$ -> $HSnO_2^-$	(basic)	$Sn(OH)_6^{2-} + 2e^- \rightarrow HSnO_2^- + H_2O + 3OH^-$	reduction

2. Use half reactions to balance the following redox reactions and underline the oxidizing agent.

a. $Cl_2 + ClO_3^-$ -----> ClO^- (acidic) $2Cl_2 + \underline{ClO_3^-} + 2H_2O \rightarrow 5ClO^- + 4H^+$

b. $MnO_4^- + C_2O_4^{2-}$ -----> $Mn(OH)_3 + CO_3^{2-}$ (basic) $\underline{MnO_4^-} + 2C_2O_4^{2-} + 3OH^- \rightarrow Mn(OH)_3 + 4CO_3^{2-}$

c. $IO_3^- + I^-$ -----> I_3^- (acidic) $\underline{IO_3^-} + 8I^- + 6H^+ \rightarrow 3I_3^- + 3H_2O$

d. $ClO_4^- + Mn^{2+}$ ------> $Cl^- + MnO_2$ $\underline{ClO_4^-} + 4Mn^{2+} + 4H_2O \rightarrow Cl^- + 4MnO_2 + 8H^+$

e. $IO_3^- + H_2O_2$ -------> $I_2 + O_2$ $\underline{2IO_3^-} + 5H_2O_2 + 2H^+ \rightarrow I_2 + 5O_2 + 6H_2O$

f. $MnO_2 + Sn^{4+}$ -------> $MnO_4^- + Sn^{2+}$ $2MnO_2 + \underline{3Sn^{4+}} + 4H_2O \rightarrow 2MnO_4^- + 3Sn^{2+} + 8H^+$

g. $NO_3^- + Br_2$ -------> $NO + BrO_3^-$ $\underline{10NO_3^-} + 3Br_2 + 4H^+ \rightarrow 10NO + 6BrO_3^- + 2H_2O$

h. $H_2O_2 + I^-$ -------> $H_2O + I_2$ (acidic) $\underline{H_2O_2} + 2I^- + 2H^+ \rightarrow 2H_2O + I_2$

i. $Cr_2O_7^{2-} + Fe^{2+}$ ---> $Cr^{3+} + Fe^{3+}$ (acidic) $\underline{Cr_2O_7^{2-}} + 6Fe^{2+} + 14H^+ \rightarrow 2Cr^{3+} + 6Fe^{3+} + 7H_2O$

j. $Cu + NO_3^-$ -------> $Cu^{2+} + NO$ (acidic) $3Cu + \underline{2NO_3^-} + 8H^+ \rightarrow 3Cu^{2+} + 2NO + 4H_2O$

k. $HBr + H_2SO_4$ -----> $SO_2 + Br_2 + H_2O$ $2HBr + \underline{H_2SO_4} \rightarrow Br_2 + SO_2 + 2H_2O$

l. $HNO_3 + HCl$ ------> $NO + Cl_2 + H_2O$ $\underline{2HNO_3} + 6HCl \rightarrow 2NO + 3Cl_2 + 4H_2O$

m. $Zn + HNO_3$ -------> $Zn(NO_3)_2 + NH_4NO_3 + H_2O$ $4Zn + 10\underline{HNO_3} \rightarrow 4Zn(NO_3)_2 + NH_4NO_3 + 3H_2O$

n. HBrO -----> HBr + $HBrO_3$ $\underline{3HBrO} \rightarrow 2HBr + HBrO_3$

o. $ClO_2^- + MnO_2$ -----> $ClO^- + MnO_4^-$ (basic) $\underline{3ClO_2^-} + 2MnO_2 + 2OH^- \rightarrow 3ClO^- + 2MnO_4^- + H_2O$

p. $V^{2+} + H_2SO_3$ ------> $V^{3+} + S_2O_3^{2-}$ $4V^{2+} + \underline{2H_2SO_3} + 2H^+ \rightarrow 4V^{3+} + S_2O_3^{2-} + 3H_2O$

q. $Mn + IrCl_6^{3-}$ -----> $Mn^{2+} + Ir + Cl^-$ $3Mn + \underline{2IrCl_6^{3-}} \rightarrow 3Mn^{2+} + 2Ir + 12Cl^-$

3. Iodine is recovered from iodates in Chilean saltpeter ($NaIO_3$) by the reaction described in this unbalanced equation:

$$HSO_3^- + IO_3^- \text{ -----> } I_2 + SO_4^{2-} \quad \text{(acidic solution)}$$

What mass of iodine is produced when a 20 kg sample of $NaIO_3$ reacts with excess HSO_3^- ?

$$5HSO_3^- + 2IO_3^- \rightarrow I_2 + 5SO_4^{2-} + H_2O + 3H^+$$

$$20\,000\text{ g}\left(\frac{1\text{ mole}}{198\text{ g}}\right) = 101 \text{ moles}$$

$$101 \text{ moles } IO_3^- \left(\frac{1 \text{ mole } I_2}{2 \text{ mole } IO_3^-}\right) = 50.5 \text{ moles } I_2 \left(\frac{254\text{ g}}{1\text{ mole}}\right)$$

$$= 13.0 \text{ kg.}$$

Answer key.

Chemistry 12 Electrochemistry Worksheet No. 2

1. Determine the oxidation number of phosphorus in each of the following:
   1 a. phosphorus pentoxide, P_2O_5 +5
   2 b. phosphorus trioxide, P_2O_3 +3
   3 c. hypophosphoric acid, $H_4P_2O_6$ +4
   4 d. hydrogen diphosphide, P_2H_4 -2
   5 e. hypophosphorus acid, H_3PO_2 1
   6 f. phosphine, PH_3 -3
   7 g. phosphite, PO_3^{3-} +3
   8 h. phosphorus acid, H_3PO_3 +3
   9 i. metaphosphoric acid, HPO_3 +5
   10 j. white phosphorus, P_4 0

2. Indicate the change in oxidation number for each of the following conversions:
   a. gallium III, Ga^{3+} is converted to $H_2GaO_3^-$ $+3 \rightarrow +3$ no change
   b. americium III, Am^{3+} is converted to AmO_2^{2+} $+3 \rightarrow +6$ increases by 3
   c. selenate, SeO_4^{2-} is converted to selenous acid, H_2SeO_3 $+6 \rightarrow +4$ decreases by 2
   d. thiosulphate, $S_2O_3^{2-}$ is converted to tetrathionate, $S_4O_6^{2-}$ $+2 \rightarrow +5/2$ increases by 1/2
   e. magnetite, Fe_3O_4 is converted to iron, Fe $+8/3 \rightarrow 0$ decreases by 8/3

3. For each of the following compounds find the oxidation number of the indicated atom:

a. potassium	K	K	0
b. potassium oxide	K_2O	K	+1
		O	-2
c. chlorine	Cl_2	Cl	0
d. magnesium chloride	$MgCl_2$	Mg	+2
		Cl	-1
e. hydrogen peroxide	H_2O_2	H	+1
		O	-1
f. sodium sulphate	Na_2SO_4	Na	+1
		S	+6
		O	-2
g. ammonia	NH_3	N	-3
		H	+1
h. potassium permanganate	$KMnO_4$	K	+1
		Mn	+7
		O	-2

4. Balance the following reactions using either half reactions or the oxidation number method:

error → a. $Cr_2O_7^{2-} + HNO_2 \text{ -----> } Cr^{3+} + NO_3^-$ (acidic) $Cr_2O_7^{2-} + 3HNO_2 + 5H^+ \rightarrow 2Cr^{3+} + 3NO_3^- + 4H_2O$
b. $IO_3^- + N_2O \text{ ------> } I_2 + NO$ (acidic) $2IO_3^- + 5N_2O + 2H^+ \rightarrow I_2 + 10NO + H_2O$
c. $MnO_4^- + Te \text{ ------> } MnO_2 + TeO_3^{2-}$ (basic) $4MnO_4^- + 3Te + 2OH^- \rightarrow 4MnO_2 + 3TeO_3^{2-} + H_2O$
d. $P_4 + NO_2^- \text{ ------> } H_2PO_2^- + N_2O$ (basic) $P_4 + 2NO_2^- + 3H_2O + 2OH^- \rightarrow 4H_2PO_2^- + N_2O$
e. $HPO_2^- \text{ -----> } PO_4^{3-} + P_4$ (basic) $20HPO_3^- + 4OH^- \rightarrow 8PO_4^{3-} + 3P_4 + 12H_2O$
f. $N_2O \text{ -----> } N_2H_4 + NO_3^-$ (basic) $7N_2O + 5H_2O + 6OH^- \rightarrow 4N_2H_4 + 6NO_3^-$

Worksheet No. 2B

Name Key

Class ____________ Date ____________

1. Balance the following equations by adding half-equations provided in Appendix D of the *Heath Chemistry* text. Then decide on the basis of the positions of the half-equations in the table whether or not the reaction may be expected to proceed.

   a. $ClO_4^- + Mn^{2+} \rightarrow Cl^- + MnO_2$

   b. $IO_3^- + H_2O_2 \rightarrow I_2 + O_2$

   c. $MnO_2 + Sn^{4+} \rightarrow MnO_4^- + Sn^{2+}$

   d. $NO_3^- + Br_2 \rightarrow NO + BrO_3^-$

a) $ClO_4^- + 4Mn^{2+} + 4H_2O \rightarrow Cl^- + 4MnO_2 + 8H^+$

yes, proceeds

b) $2IO_3^- + 2H^+ + 5H_2O_2 \rightarrow I_2 + 6H_2O + 5O_2$

yes. proceeds

c) $2MnO_2 + 8OH^- + 3Sn^{4+} \rightarrow 2MnO_4^- + 4H_2O + 3Sn^{2+}$

no. doesn't proceed

d) $5NO_3^- + 2H^+ + 1\frac{1}{2}Br_2 \rightarrow 5NO + 3BrO_3^- + H_2O$

no. doesn't proceed.

2. Balance each of the following half-reactions, in either acidic or basic solution, as requested. Then state whether the process is oxidation or reduction.

   a. $BrO_3^- \rightarrow Br^-$ (basic)

   b. $HClO_2 \rightarrow ClO_3^-$ (acidic)

   c. $Ga \rightarrow H_2GaO_3^-$ (basic)

   d. $ReO_4^- \rightarrow ReO_2$ (acidic)

$6e^- + 3H_2O + BrO_3^- \rightarrow Br^- + 6OH^-$ reduction

$H_2O + HClO_2 \rightarrow ClO_3^- + 3H^+ + 2e^-$ oxidation

$4OH^- + Ga \rightarrow H_2GaO_3^- + H_2O + 3e^-$ oxidation

$3e^- + 4H^+ + ReO_4^- \rightarrow ReO_2 + 2H_2O$ reduction

3. Balance each of the following redox equations. Use either the half-reaction equation method or the oxidation number method.

a. $Cr_2O_7^{2-} + HNO_2 \rightarrow Cr^{3+} + NO_3^-$ (acidic)
b. $IO_3^- + N_2O \rightarrow I_2 + NO$ (acidic)
c. $MnO_4^- + Te \rightarrow MnO_2 + TeO_3^{2-}$ * (basic)
d. $P_4 + NO_2^- \rightarrow H_2PO_2^- + N_2O$ (basic)
e. $ClO_4^- + I^- \rightarrow Cl^- + IO_3^-$ (acidic)
f. $IO_4^- + PH_3 \rightarrow I^- + P_4$ (basic)
g. $C_2H_5OH + Cr_2O_7^{2-} \rightarrow CH_3COOH + Cr^{3+}$ (acidic)
h. $MnO_4^- + CH_3CH(OH)CH_3 \rightarrow Mn^{2+} + CH_3COCH_3$ (acidic)
i. $HPO_2^- \rightarrow PO_4^{3-} + P_4$ (basic)
j. $N_2O \rightarrow N_2H_4 + NO_3^-$ (basic)

a. $Cr_2O_7^{2-} + 5H^+ + 3HNO_2 \rightarrow 2Cr^{3+} + 4H_2O + 3NO_3^-$

b. $2IO_3^- + 2H^+ + 5N_2O \rightarrow I_2 + H_2O + 10NO$

c. $4MnO_4^- + 3Te + 2OH^- \rightarrow 4MnO_2 + H_2O + 3TeO_3^{2-}$

d. $2NO_2^- + P_4 + 3H_2O + 2OH^- \rightarrow N_2O + 4H_2PO_2^-$

e. $3ClO_4^- + 4I^- \rightarrow 3Cl^- + 4IO_3^-$

f. $3IO_4^- + 8PH_3 \rightarrow 3I^- + 2P_4 + 12H_2O$

g. $3C_2H_5OH + 2Cr_2O_7^{2-} + 16H^+ \rightarrow 3CH_3COOH + 4Cr^{3+} + 11H_2O$

h. $2MnO_4^- + 5CH_3CH(OH)CH_3 + 6H^+ \rightarrow 2Mn^{2+} + 5CH_3COCH_3 + 8H_2O$

i. $20HPO_2^- + 4OH^- \rightarrow 8PO_4^{3-} + 3P_4 + 12H_2O$

j. $7N_2O + 5H_2O + 6OH^- \rightarrow 4N_2H_4 + 6NO_3^-$

Electrochemistry #3

1. $I_2 + 2S_2O_3^{2-} \longrightarrow 2I^- + S_4O_6^{2-}$

mols $S_2O_3^{2-}$ = 20.0 mL × 0.20 M = 4.00 mmol

mols I_2 = 4.00 mmol × $\frac{1 \text{ mol } I_2}{2 \text{ mol } S_2O_3^{2-}}$ = 2.00 mmol

$[I_2] = \frac{2.00 \text{ mmol}}{35.0 \text{ mL}} = 0.057 \text{ M}$

2. $6H^+ + 5H_2O_2 + 2MnO_4^- \longrightarrow 5O_2 + 2Mn^{2+} + 8H_2O$

mols MnO_4^- = 50.0 mL × 0.25 M = 12.5 mmols

mols H_2O_2 = 12.5 mmols × $\frac{5 \text{ mol } H_2O_2}{2 \text{ mol } MnO_4^-}$ = 31.25 mmol (or 3.1×10^{-2} mol)

3. $2Cu^{2+} + 2I^- \longrightarrow 2Cu^+ + I_2$

$I_2 + 2S_2O_3^{2-} \longrightarrow 2I^- + S_4O_6^{2-}$

mols $S_2O_3^{2-}$ = 35.0 mL × 0.20 M = 7.00 mmols

mols I_2 = 7.00 mmols × $\frac{1 \text{ mol } I_2}{2 \text{ mol } S_2O_3^{2-}}$ = 3.50 mmols

mols Cu^{2+} = 3.50 mmols × $\frac{2 \text{ mol } Cu^{2+}}{1 \text{ mol } I_2}$ = 7.00 mmols

$[Cu(NO_3)_2] = \frac{7.00 \text{ mmols}}{25.0 \text{ mL}} = 0.28 \text{ M}$

4. $2MnO_4^- + SO_3^{2-} + H_2O \longrightarrow 2MnO_4^{2-} + SO_4^{2-} + 2H^+$

mols MnO_4^- = 35.0 mL × 0.020 M = 0.700 mmols

mols SO_3^{2-} = 0.700 mmols × $\frac{1 \text{ mol } SO_3^{2-}}{2 \text{ mol } MnO_4^-}$ = 0.350 mmols

$[H_2SO_3] = \frac{0.350 \text{ mmols}}{25.0 \text{ mL}} = 0.014 \text{ M}$

Key.

Electrochemistry #4

Use the following unbalanced equation to answer questions 1 - 9.

$$CrO_2^- + ClO^- \longrightarrow Cl^- + CrO_4^{2-}$$ (acid solution)

1. What is the highest oxidation number of chromium in the equation?

   +6

2. What is the highest oxidation number of chlorine in the equation?

   +1

3. Write the balanced reduction half reaction in an acid solution.

   $2e^- + 2H^+ + ClO^- \longrightarrow Cl^- + H_2O$

4. Write the balanced oxidation half reaction in an acid solution.

   $2H_2O + CrO_2^- \longrightarrow CrO_4^{2-} + 4H^+ + 3e^-$

5. Write the balanced full reaction in an acid solution.

   $H_2O + 2CrO_2^- + 3ClO^- \longrightarrow 3Cl^- + 2CrO_4^{2-} + 2H^+$

6. Write the formula for the oxidizing agent.

   ClO^-

7. Write the formula of the species that is getting reduced.

   ClO^-

8. What effect would the addition of $NaOH_{(SOLID)}$ have on the equilibrium?

   shift reaction to products $OH^- + H^+ \rightarrow H_2O$ $[H^+]\downarrow$ eq^m shifts $\longrightarrow$

9. What effect would the addition of $NaOH_{(SOLID)}$ have on the cell potential?

   increases cell potential

Electrochemistry #5 Key.

1. Calculate the electrochemical cell potential for the following reactions and indicate whether the reaction is spontaneous as written.

   (a) $Co + Fe^{3+} \dashrightarrow Co^{2+} + Fe^{2+}$ $E^\circ = 1.05\ V$

   (b) $H_2 + Cl_2 \dashrightarrow 2H^+ + 2Cl^-$ $E^\circ = 1.36\ V$

   (c) $I^- + MnO_4^- \dashrightarrow I_2 + Mn^{2+}$ $E^\circ = 0.97\ V$

2. Write a red-ox reaction for Ag metal in Cl_2 gas and determine whether the reaction is spontaneous by calculating the cell potential. yes, it is

   $2Ag + Cl_2 \rightarrow 2Ag^+ + 2Cl^-$ $E^\circ = +0.56\ V$ spontaneous

3. Will anything happen if an aluminum spoon is used to stir an iron III nitrate solution? Explain, using cell potentials.

   $Al + 3Fe^{3+} \rightarrow Al^{3+} + 3Fe^{2+}$ $E^\circ = +2.43\ V$ it will react

4. Will anything happen if an iron spoon is used to stir an aluminum chloride solution? Explain, using cell potentials.

   $3Fe + 2Al^{3+} \rightarrow 3Fe^{2+} + 2Al$ $E^\circ = -1.21\ V$ nothing happens

5. A common test for silver purity is to add a drop of nitric acid. Explain what this test does, using cell potentials.

   $3Ag + NO_3^- + 4H^+ \rightarrow 3Ag^+ + NO_{(g)} + 2H_2O$ $E^\circ = +0.16\ V$ bubbles of NO form

6. Would the nitric acid test for silver purity work for gold? Explain, using cell potentials.

   $Au + NO_3^- + 4H^+ \rightarrow Au^{3+} + NO_{(g)} + 2H_2O$ $E^\circ = -0.54\ V$ no reaction occurs

7. Good stainless steel is mostly nickel metal. Can copper II sulphate solution be stored in a stainless steel container? Explain, using cell potentials. The Nickel and Cu^{2+} react.

   $Ni + Cu^{2+} \rightarrow Ni^{2+} + Cu$ $E^\circ = +0.60\ V$

8. Will anything happen to the copper plumbing in a house if acidified dichromate solution is poured down the drain? Explain, using cell potentials.

   $Cr_2O_7^{2-} + 14H^+ + 3Cu \rightarrow 2Cr^{3+} + 7H_2O + 3Cu^{2+}$ $E^\circ = +0.89\ V$

   The copper is oxidized by the dichromate solution

7

If the cell voltage for the sum of two half-reactions is positive, the net redox reaction will proceed spontaneously. Use a table of standard reduction potentials to predict whether the reactions below are spontaneous. In each case, show the addition of half-reactions and $E°$ values.

$2Fe^{3+} + Co_{(s)} \rightarrow Co^{2+} + 2Fe^{2+}$

1. $2Fe^{3+} + Co(s) \rightarrow Co^{2+} + Fe^{2+}$

Red: $2Fe^{3+} + 2e^- \rightarrow 2Fe^{2+}$ $E° = +0.77$ V

Ox: $Co_{(s)} \rightarrow Co^{2+} + 2e^-$ $E° = +0.28$ V

Full: $2Fe^{3+} + Co_{(s)} \rightarrow Co^{2+} + 2Fe^{2+}$ $E° = +1.05$ V

* spontaneous

$Sr_{(s)} + 2H^+ \rightarrow Sr^{2+} + H_{2(g)}$

2. $Sr(s) + 2H^+ \rightarrow Sr^{2+} + H_2(g)$

Ox: $Sr_{(s)} \rightarrow Sr^{2+} + 2e^-$ $E° = +2.89$ V

Red: $2e^- + 2H^+ \rightarrow H_{2(g)}$ $E° = 0.00$ V

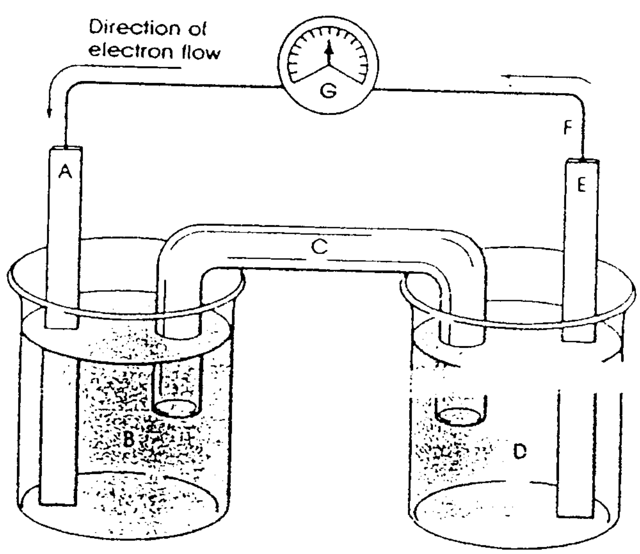
Full: $Sr_{(s)} + 2H^+ \rightarrow Sr^{2+} + H_{2(g)}$ $E° = +2.89$ V

* spontaneous

In an electrochemical cell, an oxidation-reduction reaction produces electrical energy. Identify the labeled parts of the cell that is shown in the figure by writing the corresponding letters into the blanks below. The reaction that is taking place is:

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

Voltmeter	G	Solution of copper(II)	B
Salt bridge	C	Solution of zinc(II)	D
Anode	E	External wire	F
Cathode	A		



Key.

Electrochemistry #6

1. Balance the equation and indicate the
   (a) oxidizing agent. $KMnO_4$
   (b) reducing agent. KNO_2
   (c) the element being reduced. Mn
   (d) the element being oxidized. N

$2\,KMnO_4 + 5KNO_2 + 3H_2SO_4 \longrightarrow 2MnSO_4 + 3H_2O + 5KNO_3 + K_2SO_4$

2. Balance the equation and indicate the
   (a) oxidizing agent. MnO_4^-
   (b) reducing agent. $C_2O_4^{2-}$
   (c) the element being reduced. Mn
   (d) the element being oxidized. C

$16H^+ + 2MnO_4^- + 5C_2O_4^{2-} \longrightarrow 10CO_2 + 2Mn^{2+} + 8H_2O$ (ACID)

3. Write the net ionic electrochemical equation for:

   Oxygen + hydrogen sulfide (ACID)

   $\frac{1}{2}O_2 + H_2S \longrightarrow S + H_2O$

4. Balance the following electrochemical reaction and determine the cell E° value.

   $3Br_2 + 2Cr^{3+} + 7H_2O \longrightarrow Cr_2O_7^{2-} + 14H^+ + 6Br^-$ $E^0_{cell} = -0.14\,V$

5. An iron-nickel electrochemical cell uses a salt bridge to join a half-cell containing a strip of iron in a 1.0 M solution of Fe^{2+} to a half-cell which contains a strip of nickel in a 1.0 M Ni^{2+} solution. A voltmeter connects the two metal strips.

   (a) In which cell does reduction occur? Ni^{2+}/Ni

   (b) Write the two half-cell reactions involved.
   $Ni^{2+} + 2e^- \longrightarrow Ni$ $Fe \longrightarrow Fe^{2+} + 2e^-$

   (c) Which metal is the anode?
   Fe

   (d) In which direction are electrons passing through the voltmeter?
   From Fe through voltmeter to Ni

   (e) What is the expected initial voltmeter reading? 0.19 V

   (f) What would be the effect on the voltmeter reading if the Fe^{2+} concentration only were increased to 2.0 M?
   Voltage would decrease

   (g) What would be the effect if only the $[Ni^{2+}]$ were decreased to 0.50 M?
   Voltage would decrease

   (h) What is the voltmeter reading when the cell reaches equilibrium?

   0.00 V

Key.

6. A fuel cell converts about 75% of the available chemical energy into usable electric energy. One type of fuel cell is based on the combustion of hydrogen forming water:

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

(a) Write the anode reaction occurring in basic solution.

$2OH^- + H_2 \longrightarrow 2H_2O + 2e^-$

(b) Write the cathode reaction also occurring in basic solution.

$H_2O + \frac{1}{2}O_2 + 2e^- \longrightarrow 2OH^-$

(c) Would a low or high gas pressure give the better E° value? Explain.

high gas pressure, gases are reactants so eq^n shifts to products

7. Corrosion of iron involves the oxidation of iron into Fe^{2+} when impurities such as copper serve as the cathode half cell where the reduction of oxygen occurs.

$$Fe_{(s)} + \frac{1}{2}O_{2(g)} + H_2O_{(l)} \longleftrightarrow Fe(OH)_{2(s)}$$

Further oxidation of $Fe(OH)_2$ by oxygen yields $Fe(OH)_3$.

(a) What two factors are involved in corrosion of iron?

oxygen gas + moisture

(b) Write the anode reaction.

$Fe \longrightarrow Fe^{2+} + 2e^-$.

(c) Write the cathode reaction.

$\frac{1}{2}O_2 + H_2O + 2e^- \longrightarrow 2OH^-$

8. (a) Draw a diagram for an experimental setup that demonstrates how a copper-plated spoon could be plated with silver.

See below

(b) Identify the cathode and anode materials and the electrolytic solution. Cathode = copper spoon Solution = silver nitrate
anode = silver metal

9. Predict the principal product discharged at each electrode during the electrolysis of these 1 M solutions. Assume platinum electrodes are used.

(a) KI Anode: I_2
Cathode: $H_2 + OH^-$

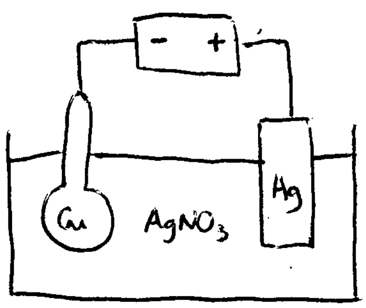
(b) H_2SO_4 Anode: $O_2 + H^+$
Cathode: H_2SO_3 ; H_2O H_2

(c) HCl Anode: Cl_2
Cathode: $\cancel{H_2SO_3\ H_2O}$ H_2

10. Explain why chromium, aluminum or magnesium metal does not corrode as rapidly as a less active metal such as iron.

Iron oxide flakes off, allowing moist air to come into contact with fresh metal. The other metal oxides are more likely to form an impermeable layer on the metal.

8(a)



Chemistry 12: Electrochemistry 1
Review Worksheet

Name Key
Date ________ Block __

1. When a substance is reduced, it

   a. loses electrons.
   b. causes oxidation.
   c. undergoes oxidation.
   d. increases in oxidation number.

2. Identify the oxidizing agent in the following equation:
   $2H^+_{(aq)} + Pb_{(s)} \rightarrow H_{2(g)} + Pb^{2+}_{(aq)}$

   a. $H_{2(g)}$
   b. $H^+_{(aq)}$
   c. $Pb_{(s)}$
   d. $Pb^{2+}_{(aq)}$

3. An example of reduction is

   a. $Mn_{(s)} \rightarrow Mn^{2+}_{(aq)}$
   b. $H^+_{(aq)} + MnO_4^-{}_{(aq)} + K^+_{(aq)} + OH^-_{(aq)} \rightarrow K^+_{(aq)} + MnO_4^-{}_{(aq)} + H_2O_{(l)}$
   c. $Mn^{2+}_{(aq)} + S^{2-}_{(aq)} \rightarrow MnS_{(s)}$
   d. $MnO_{2(s)} + 4H^+_{(aq)} + 2e^- \rightarrow Mn^{2+}_{(aq)} + 2H_2O_{(l)}$

4. A strip of Zn metal is placed into 0.1M $Ga(NO_3)_3$ and its surface darkens. From this observation it may be concluded that Ga^{3+} is a

   a. weaker reducing agent than Zn^{2+}
   b. weaker oxidizing agent than Zn^{2+}
   c. stronger reducing agent than Zn^{2+}
   d. stronger oxidizing agent than Zn^{2+}

5. Which of the following oxidizing agents will react spontaneously with Br^- at standard conditions?

   a. H^+
   b. Li^+
   c. NO_3^- in acid
   d. $Cr_2O_7^{2-}$ in acid

6. Which of the following most readily loses electrons?

   a. Ag
   b. Cl^-
   c. Sr
   d. Mg^{2+}

7. Which of the following could be a product of a reaction in which SO_3^{2-} acts as a reducing agent?
   a. SO_4^{2-}
   b. SO_2
   c. S_2O
   d. $S_2O_8^{2-}$

8. Given the half-cell reaction $S_2O_8^{2-} + 2H^+ \rightarrow 2HSO_4^-$, which of the following will balance electric charges?
   a. Add $2e^-$ to the left side
   b. Add $2e^-$ to the right side
   c. Add $3e^-$ to the left side
   d. Add $3e^-$ to the right side

Use the following information to answer question 9.

Cl_2 is pale yellow in CCl_4
Cl^- is colorless in water
Br_2 is red in CCl_4
Br^- is colorless in water

9. Aqueous Cl_2 and aqueous KBr are shaken with CCl_4 in a test tube. The CCl_4 layer is red and the water layer is colorless. What is the best conclusion?
   a. Br^- is oxidized
   b. No reaction occurred.
   c. Cl_2 was oxidized
   d. CCl_4 was oxidized

10. What is the oxidation number of Cr in CrO_4^{2-}?
   a. –2
   b. +6
   c. +8
   d. +10

11. The oxidation number for a sulphur atom in $Na_2S_2O_5$ is
   a. –2
   b. +1
   c. +4
   d. +8

12. In which of the following compounds does carbon have an oxidation number of –2?
   a. CO
   b. CO_2
   c. CH_2O
   d. CH_3OH

13. Consider the following reaction:

$$3I_2 + 3H_2O \rightarrow 6H^+ + 5I^- + IO_3^-$$

In this reaction atoms in I_2 undergo

a. oxidation only
b. reduction only
c. neither oxidation nor reduction
d. both oxidation and reduction

14. Which one of the following half-reactions is balanced?

a. $IO_3^-{}_{(aq)} + 6H^+{}_{(aq)} + 5e^- \rightarrow I_{2\,(s)} + 2H_2O_{(l)}$
b. $ClO^-{}_{(aq)} + H_2O_{(l)} + 2e^- \rightarrow Cl^-{}_{(aq)} + 2OH^-{}_{(aq)}$
c. $SO_4^{2-}{}_{(aq)} + 8H^+{}_{(aq)} + 6e^- \rightarrow H_2S_{(g)} + 4H_2O_{(l)}$
d. $NO_2^-{}_{(aq)} + H_2O_{(l)} + 2e^- \rightarrow 2H^+{}_{(aq)} + NO_3^-{}_{(aq)}$

15. Of the following metals, which would be the best one to use to make a container in which to store an aqueous copper(II) sulfate solution?

a. $Ag_{(s)}$
b. $Fe_{(s)}$
c. $Ni_{(s)}$
d. $Pb_{(s)}$

16. The correctly balanced half-reaction for $ClO^-{}_{(aq)} \rightarrow Cl^-{}_{(aq)}$ in a basic solution is

a. $2H^+{}_{(aq)} + ClO^-{}_{(aq)} + 2e^- \rightarrow Cl^-{}_{(aq)} + H_2O_{(l)}$
b. $H_2O_{(l)} + ClO^-{}_{(aq)} \rightarrow Cl^-{}_{(aq)} + 2OH^-{}_{(aq)} + 2e^-$
c. $H_2O_{(l)} + ClO^-{}_{(aq)} + 2e^- \rightarrow Cl^-{}_{(aq)} + 2OH^-{}_{(aq)}$
d. $2H^+{}_{(aq)} + ClO^-{}_{(aq)} \rightarrow Cl^-{}_{(aq)} + H_2O_{(l)} + 2e^-$

17. Experiments were performed with three metal strips, X, Y, and Z, and their corresponding 1.0M nitrate solutions, $X(NO_3)_2$, $Y(NO_3)_2$ and $Z(NO_3)_3$.

- metal Y reacted with X^{2+} but not with Z^{3+}.
- metal X did not react with any of the solutions

Which of the following gives the metals in order of <u>decreasing</u> strength as reducing agent (strongest reducing agent first)?

a. Z, Y, X
b. X, Y, Z
c. Y, Z, X
d. X, Z, Y

18. Which of the following sets of coefficients balances the equation

$\_\_HNO_{3\,(aq)} + \_\_H_2S_{(aq)} \rightarrow \_\_NO_{(g)} + \_\_S_{(s)} + 4H_2O_{(l)}$

a. 4, 2, 4, 1
b. 4, 1, 4, 1
c. 2, 3, 2, 3
d. 2, 1, 2, 1

19. Which of the following agents would reduce $Sn^{4+}_{(aq)}$ to $Sn^{2+}_{(aq)}$?

a. $Fe_{(s)}$
b. $I^-_{(aq)}$
c. $Fe^{2+}_{(aq)}$
d. $Ag_{(s)}$

20. In a particular redox reaction, the oxidation number of phosphorus changed from –3 to 0. From this it may be concluded that phosphorus

a. lost 3 electrons and was reduced.
b. lost 3 electrons and was oxidized.
c. gain 3 electrons and was reduced.
d. gain 3 electrons and was oxidized.

SHORT ANSWER QUESTIONS

21. Balance the following half-reaction occurring in acid solution.

$TiO_4^{2-} \rightarrow Ti$

$TiO_4^{2-} + 8H^+ + 6e^- \rightarrow Ti + 4H_2O$

Electrochemistry 2 Review Worksheet

1. In a redox reaction, the species which loses electrons

   A. is oxidized.
   B. is called the cathode.
   C. gains mass at the electrode.
   D. decreases in oxidation number.

Consider the following redox equation:

$$12\,H^{+}_{(aq)} + 2\,IO_3{}^{-}_{(aq)} + 10\,Fe^{2+}_{(aq)} \rightarrow 10\,Fe^{3+}_{(aq)} + I_{2(s)} + 6\,H_2O_{(\ell)}$$

2. The reducing agent is

   A. I_2
   B. H^+
   C. Fe^{2+}
   D. IO_3^-

3. Which of the following is the strongest oxidizing agent?

   A. Cu^{2+}
   B. Pb^{2+}
   C. Ni^{2+}
   D. Sn^{2+}

4. Metallic platinum reacts spontaneously with $Au^{3+}_{(aq)}$ but does not react with $Ag^{+}_{(aq)}$. The metals, in order of increasing strength as reducing agents, are

   A. Ag, Pt, Au
   B. Pt, Au, Ag
   C. Au, Ag, Pt
   D. Au, Pt, Ag

5. The electrolysis of a molten salt involves the migration of

   A. anions only.
   B. cations only.
   C. electrons only.
   D. both cations and anions.

6. When electroplating an iron spoon with silver, the half-reaction taking place on the spoon is

A. $Ag^{+}_{(aq)} + e^{-} \rightarrow Ag_{(s)}$
B. $Ag_{(s)} \rightarrow Ag^{+}_{(aq)} + e^{-}$
C. $Fe^{2+}_{(aq)} + 2e^{-} \rightarrow Fe_{(s)}$
D. $Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^{-}$

7. An oxidizing agent is a substance which

A. accepts protons.
B. donates protons.
C. accepts electrons.
D. donates electrons.

8. Consider the following oxidation-reduction reaction:

$$Mn^{2+} + IO_3^{-} + H_2O \rightarrow MnO_4^{-} + \tfrac{1}{2}I_2 + 2H^{+}$$

The reducing agent is

A. I_2
B. IO_3^{-}
C. H_2O
D. Mn^{2+}

9. A substance that is most likely to gain electrons during a spontaneous redox reaction is

A. I_2
B. Li
C. Au
D. Hg

10. The equation for the spontaneous reaction between Sn and 1.0 M HCl is

A. $Sn + H_2 \rightarrow Sn^{2+} + 2H^{+}$
B. $Sn + 2H^{+} \rightarrow Sn^{2+} + H_2$
C. $Sn + Cl_2 \rightarrow Sn^{2+} + 2Cl^{-}$
D. $Sn + 2Cl^{-} \rightarrow Sn^{2+} + Cl_2$

11. A solution of lead(II) nitrate could be safely stored in a container made of

A. Cu
B. Ni
C. Fe
D. Zn

12. Consider this redox equation:

$$C_2O_4^{2-} + MnO_2 \rightarrow Mn^{2+} + 2CO_2$$

As a result of this reaction the oxidation number of each C atom has

A. increased by 1.
B. increased by 2.
C. decreased by 2.
D. decreased by 4.

13. Consider the following half-reaction:

$$HO_2^- \rightarrow O_2 \quad \text{(basic)}$$

The balanced equation is

A. $HO_2^- \rightarrow O_2 + H^+ + 2e^-$
B. $2HO_2^- + 2e^- \rightarrow O_2 + 2OH^-$
C. $2HO_2^- + 2H^+ \rightarrow 2H_2O_2 + O_2$
D. $HO_2^- + OH^- \rightarrow O_2 + H_2O + 2e^-$

14. To determine the $[Fe^{2+}]$ in a redox titration, a suitable oxidizing agent is

A. SO_4^{2-} in acid.
B. H_3PO_4 in acid.
C. MnO_4^- in acid.
D. MnO_4^- in base.

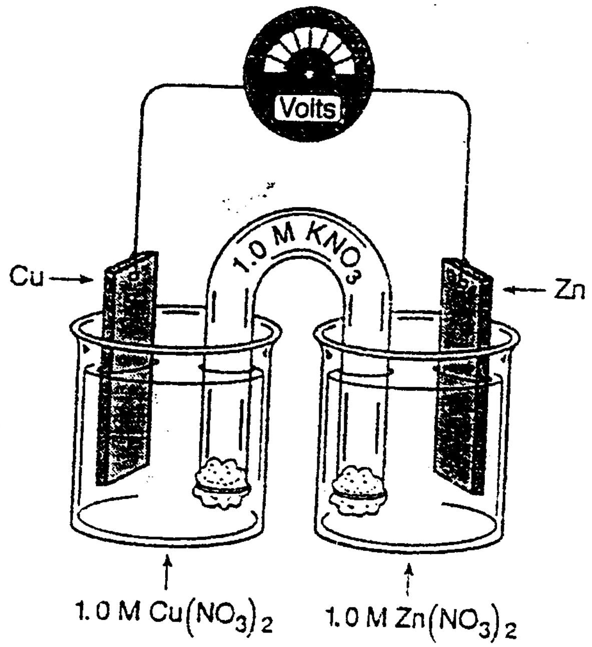
15. Which of the following pairs of ions will react spontaneously in solution?

A. Cu^{2+} and Fe^{2+}
B. Pb^{2+} and Sn^{2+}
C. Co^{2+} and Cr^{2+}
D. Mn^{2+} and Cr^{2+}

16. When NO_2 reacts to form N_2O_4 the oxidation number of nitrogen

A. increases by 2.
B. increases by 4.
C. increases by 8.
D. does not change.

Use the following electrochemical cell diagram to answer questions 17. and 18.



17. As the above cell operates,

A. copper ions migrate into the salt bridge.
B. cations migrate towards the zinc electrode.
C. the mass of the copper electrode increases.
D. anions migrate towards the copper electrode.

18. The initial cell voltage is

A. 0.42 V
B. 0.91 V
C. 1.10 V
D. 1.28 V

19. Consider the following overall reaction:

$$2Rh^{+}_{(aq)} + Pb_{(s)} \rightarrow 2Rh_{(s)} + Pb^{2+}_{(aq)} \qquad E° = +0.73\ V$$

The E° for the half-reaction $Rh^{+}_{(aq)} + e^{-} \rightleftarrows Rh$ is

A. −0.86 V
B. −0.60 V
C. +0.60 V
D. +0.86 V

20. Hydrogen and oxygen react to provide energy in a(n)

A. dry cell.
B. fuel cell.
C. alkaline cell.
D. lead-acid storage cell.

21. The corrosion of iron can be prevented by attaching a piece of zinc to the iron because

A. iron acts as an anode.
B. zinc reduces more readily than iron.
C. electrons flow from the zinc to the iron.
D. iron ions form more readily than zinc ions.

22. **An impure sample of iron was dissolved in acid. The Fe^{2+} in this solution was titrated with 0.0210 M $KMnO_4$. Use the following data table and redox equation to determine the moles of Fe^{2+} in the sample.** **(3 marks)**

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

TRIAL	VOLUME $KMnO_4^-$
1	37.26 mL
2	35.18 mL
3	35.22 mL

Average volume = $\frac{35.18 + 35.22}{2}$

= 35.20 mL

mols MnO_4^- = 35.20 mL × 0.0210 M

= 0.7392 mmol

mols Fe^{2+} = 0.7392 mmol $\left(\frac{5 \text{ mol } Fe^{2+}}{1 \text{ mol } MnO_4^-}\right)$ = 3.70×10^{-3} mol Fe^{2+}

23. **The overall reaction in a fuel cell is:**

$$2H_2 + O_2 \rightarrow 2H_2O$$

a) **Write the equation for the half-reaction at the anode.** **(1 mark)**

$2OH^- + H_2 \longleftrightarrow 2H_2O + 2e^-$

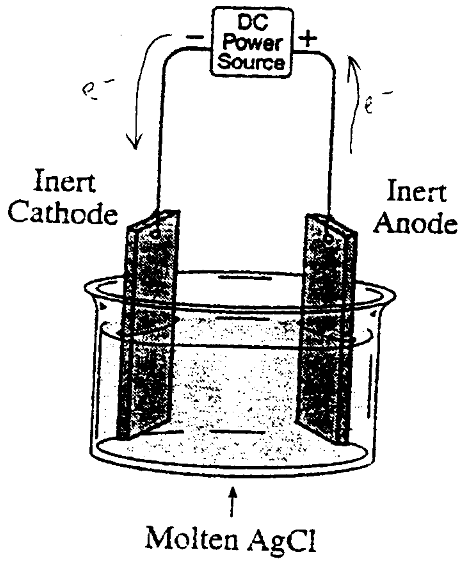
b) **Is the overall reaction spontaneous? Explain.** **(1 mark)**

$\frac{1}{2}O_2 + 2H^+ + 2e^- \longleftrightarrow H_2O \quad E^\circ = 0.82\ V$

$2OH^- + H_2 \longleftrightarrow 2H_2O + 2e^- \quad E^\circ = 0.41\ V$

Overall $2H_2 + O_2 \longleftrightarrow 2H_2O \quad E^\circ_{cell} = 1.23\ V$ as the cell potential is positive the reaction is spontaneous

24. Consider the following electrolytic cell used for the electrolysis of molten AgCl.



a) Clearly indicate on the diagram above, the direction of the electron flow through the wire. **(1 mark)**

See above (from anode to cathode)

b) Write the equation for the half-reaction taking place at the anode. **(1 mark)**

$$2\,Cl^- \longrightarrow Cl_2 + 2e^-$$

c) Write the equation for the half-reaction taking place at the cathode. **(1 mark)**

$$Ag^+ + e^- \longrightarrow Ag$$

d) Write the equation for the overall reaction. **(1 mark)**

$$2\,AgCl \longrightarrow 2\,Ag + Cl_2$$

or

$$2\,Ag^+ + 2\,Cl^- \longrightarrow 2\,Ag + Cl_2$$