AP Chemistry
Topic: Electrochemistry
Chapter 20 (20 points)

Name______________________________

Date___________ Period_____

Directions: In the spaces provided to the left, write in the answer that best completes each statement below.(2 points each)

Questions 1 & 2 refer to the following unbalanced equation. Space has been provided for you to assign oxidation numbers to each element in the equation and balance the equation, if you wish to do so.

$\overset{+3\ -1}{}$ $\overset{+1\ -1}{}$ $\overset{-2}{}$ $\overset{-1}{}$

. .3. $SbCl_3$ + .6. HCl + . . . $NaBrO_3$ -----> .3. $SbCl_5$ + . . . NaBr + .3. H_2O

+5 +5

__E__1. Which is the oxidizing agent?
a) Sb b) Br c) HCl d) $SbCl_3$ e) $NaBrO_3$

__C__2. Which is the coefficient of $SbCl_5$ in the balanced equation?
a) 1 b) 2 c) 3 d) 4 e) 5

For questions 3 and 4 consider the following equation:

Zn(s) + Cu^{2+}(aq) ------> Zn^{2+}(aq) + Cu(s) E^o = 1.10 volts

The above equation refers to the Daniel cell, a common electrochemical cell.

(1F = mole e-)
= 96,500 C

__C__3. Which expression gives the value for ΔG^o in kJ/mol for this reaction?

a) -2 X 8.31 X 1.10 X 1,000

b) $\dfrac{-2 \text{ X } 96{,}500 \text{ X } 1.10}{8.31}$

c) $\dfrac{-2 \text{ X } 96{,}500 \text{ X } 1.10}{1{,}000}$

d) $\dfrac{-2 \text{ X } 96{,}500}{1.10 \text{ X } 8.31}$

e) $\dfrac{-2 \text{ X } 8.31 \text{ X } 1.10}{1{,}000}$

$\Delta G^\circ = -nFE^\circ$

$\Delta G^\circ = \dfrac{-2(96{,}500\ \frac{C}{mole})(1.10V)}{1000}$

__A__ 4. Which expression gives the voltage for such a cell at non-standard conditions where [Cu^{2+}] is 1.00 M and [Zn^{2+}] is 0.010 M?

a) 1.10 + 0.0591

b) 1.10 + (0.0591 X 2)

c) 1.10 + $\dfrac{0.0591}{2}$

d) (1.10 + 0.0591) X 2

e) 1.10 X $\dfrac{2}{0.0591}$

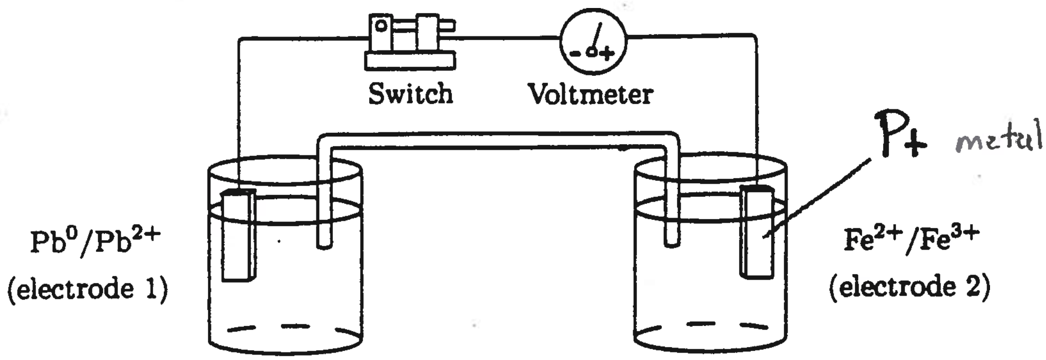
$E = E^\circ - \dfrac{.0592}{n} \log Q$

$= 1.10V - \dfrac{.0592}{n} \log \dfrac{[Zn^{2+}]}{[Cu^{2+}]}$

$= 1.10V - \dfrac{.0592}{2} \log \left(\dfrac{.01M}{1.0M}\right)$

B

5. The diagram below represents a standard Fe^{2+}/Fe^{3+} half cell connected to a standard Pb^0/Pb^{2+} half cell. The electrodes are numbered for the purposes of identification.



$Pb(s) \rightarrow Pb^{2+} + 2e^-$ $Fe^{2+} \rightarrow Fe^{3+}$ e^-

Which describes materials used for the construction of the standard Fe^{2+}/Fe^{3+} half cell?

I. The electrode is made of iron metal.
II. The source of Fe^{3+} ions could be $Fe(OH)_2$.
III. The source of Fe^{3+} ions could be $FeCl_3$.

a) I only b) III only c) II and III only d) I and III only e) I, II and III

II.) Solve the following on separate paper! :) (10 pts)

$$Sr(s) + Mg^{2+} \rightleftharpoons Sr^{2+} + Mg(s)$$

Consider the reaction represented above that occurs at 25°C. All reactants and products are in their standard states. The value of the equilibrium constant, K_{eq}, for the reaction is 4.2×10^{17} at 25°C.

(a) Predict the sign of the standard cell potential, $E°$, for a cell based on the reaction. Explain your prediction.

(b) Identify the oxidizing agent for the spontaneous reaction.

(c) If the reaction were carried out at 60°C instead of 25°C, how would the cell potential change? Justify your answer.

(d) How would the cell potential change if the reaction were carried out at 25°C with a 1.0-molar solution of $Mg(NO_3)_2$ and a 0.10-molar solution of $Sr(NO_3)_2$? Explain.

(e) When the cell reaction in (d) reaches equilibrium, what is the cell potential?

Electrochemistry Review Quiz

$Sr(s) + Mg^{2+} \rightleftharpoons Sr^{2+} + Mg(s)$

1996

(a) • Since the Keq is very large the cell potential will be positive (+1)

One of these For (+1):

• because the reaction is spontaneous

• the calculated E° value is +.52V

• the E°_{ox} for $Sr \rightarrow Sr^{2+} + 2e^-$ is more pos than for oxidizing Mg

$Sr \rightarrow Sr^{2+} + 2e^-$ $E_{ox} = 2.8$

$2e^- + Mg^{2+} \rightarrow Mg$ $= -2.37$

(b) $\boxed{Mg^{2+} \text{ is the oxidizing agent}}$ because Mg^{2+} is being reduced! +1

(c) $E = E^\circ - \frac{RT}{nF} \ln k$; $E = 0$ at equilib so $E^\circ = \frac{RT}{nF} \ln K$ (+1) +1

Since all ions are 1 M so at T increases cell potential increases (+1) !

(OR) If they have...

No change Nernst Eqtn $E_{cell} = E^\circ - \frac{RT}{nF} \ln Q$ (+1) only

$\ln Q = 0$ ∴ $E_{cell} = E^\circ$

(d) $E_{cell} = E^\circ - \frac{.0592}{2} \log Q$ if $Q = \frac{[Sr^{2+}]}{[Mg^{2+}]} = \frac{.10M}{1.0M} = .10$ +1

So ∴ ~~as~~ log Q is negative! so $(-\frac{.0592}{2} \log Q)$ is a + value and added to E° ; $E_{cell} \uparrow$ +2

(e) At Equilibrium, $E_{cell} = 0$ (+1)