AP Chem Redox Study Sheet!!!

1. Rank from strongest to weakest oxidizing agent: NO_3^{-1} (in acid) ②, Ag^{+1} ③, Ag ④, $Cr_2O_7^{-2}$(in acid) ①

$NO_3^{-1} + 4H^{+1} + 3e^- \rightarrow NO + 2H_2O$ $E^\circ_{red} = 0.96\ V$

$Ag^{+1} + 1e^- \rightarrow Ag$ $E^\circ_{red} = 0.799\ V$

$Cr_2O_7^{-2} + 14H^{+1} + 6e^- \rightarrow 2Cr^{+3} + 7H_2O$ $E^\circ_{red} = 1.33\ V$ ← The strongest oxidizing agent would be the substance most capable of gaining electrons (being reduced).. so highest reduction potential

$Ag + e^- \rightarrow Ag^-$ (?) silver is a metal.... unlikely to gain e-.. can't find on volt table.

2. Rank from strongest to weakest reducing agent: I_2 ③, Fe ①, Fe^{+2} ②

$IO_3^{-1} + 6H^{+1} + 5e^- \rightarrow I_2 + 3H_2O$ $E^\circ_{red} = 1.195V$

$Fe^{+2} + 2e^- \rightarrow Fe_{(s)}$ $E^\circ_{red} = -.440\ V$

$Fe^{+3} + 1e^- \rightarrow Fe^{+2}$ $E^\circ_{red} = +.771\ V$

} look up on reduction table, then reverse rxns to see how favorable it is to oxidize I_2, Fe, Fe^{+2}

$I_2 + 3H_2O \rightarrow 5e^- + \ldots$ $E^\circ_{ox} = -1.195V$

$Fe_{(s)} \rightarrow Fe^{+2} + 2e^-$ $E^\circ_{ox} = +.440V$

$Fe^{+2} \rightarrow Fe^{+3} + 1e^-$ $E^\circ_{ox} = -.771\ V$

The thing with the highest E°_{ox} is the most easily oxidized, and therefore the strongest reducing agent.

$Zn^{+2}_{(aq)} + 2e^- \longrightarrow Zn_{(s)}$	E°red = - 0.76V
$Na^{+1}_{(aq)} + 1e^- \longrightarrow Na_{(s)}$	E°red = - 2.71 V
$F_{2(g)} + 2e^- \longrightarrow 2F^{-1}_{(aq)}$	E°red = 2.87 V
$Br_{2(g)} + 2e^- \longrightarrow 2Br^{-1}_{(aq)}$	E°red = 1.07 V
$I_{2(g)} + 2e^- \longrightarrow 2I^{-1}_{(aq)}$	E°red = 0.53 V
$2H_2O_{(l)} + 2e^- \longrightarrow H_{2(g)} + 2OH^{-1}_{(aq)}$	E°red = - 0.83 V
$O_{2(g)} + 4H^{+1}_{(aq)} + 4e^- \longrightarrow 2H_2O_{(l)}$	E°red = 1.23 V

3. If a voltage is applied to solution of $NaBr_{(aq)}$, and electrolysis occurs, possible reactants are Na^{+1}, Br^{-1}, H_2O

a. Write the half reaction that will occur at the cathode (this is the reduction half-reaction).

$Na^{+1} + 1e^- \not\rightarrow Na$ $E^\circ_{red} = -2.71\ V$

$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ $E^\circ_{red} = -.83\ V$ higher, so → $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ $E^\circ_{red} = -.83\ V$

b. Write the half reaction that will occur at the anode (the oxidation half reaction).

$2Br^- \rightarrow Br_2 + 2e^-$ $E^\circ_{ox} = -1.07V$ → $2Br^{-1} \rightarrow Br_2 + 2e^-$ $E^\circ_{ox} = -1.07\ V$

$2H_2O \not\rightarrow O_2 + 4H^{+1} + 4e^-$ $E^\circ_{ox} = -1.23\ V$

c. Write the overall redox reaction that occurs, and calculate E°rxn (E°cell)

$2H_2O + 2Br^{-1} \rightarrow H_2 + Br_2 + 2OH^{-1}$ $E^\circ_{rxn} = -1.90\ V$

4. If a voltage is applied to solution of $ZnF_{2(aq)}$, and electrolysis occurs,

a. Write the half reaction that will occur at the cathode.

$Zn^{+2} + 2e^- \rightarrow Zn$ $E^\circ_{red} = -.76\ V$ → $Zn^{+2} + 2e^- \rightarrow Zn_{(s)}$ $E^\circ_{red} = -.76\ V$

$2H_2O + 2e^- \not\rightarrow H_2 + 2OH^-$ $E^\circ_{red} = -.83V$

b. Write the half reaction that will occur at the anode.

$2F^{-1} \not\rightarrow F_2 + 2e^-$ $E^\circ_{ox} = -2.87\ V$

$2H_2O \rightarrow O_2 + 4H^{+1} + 4e^-$ $E^\circ_{ox} = -1.23\ V$ → $2H_2O \rightarrow O_2 + 4H^{+1} + 4e^-$ $E^\circ_{ox} = -1.23\ V$

c. Write the overall redox reaction that occurs, and calculate E°rxn (E°cell)

$2Zn^{+2} + 2H_2O \rightarrow 2Zn + O_2 + 4H^{+1}$ $E^\circ_{rxn} = -1.99\ V$

5. If a voltage is applied to solution of $CuI_{2(aq)}$, and electrolysis occurs,

a. Write the half reaction that will occur at the cathode.

$Cu^{+2} + 2e^- \rightarrow Cu \quad E^\circ_{red} = +.34\,V$

$2H_2O + 2e^- \not\rightarrow H_2 + 2OH^- \quad E^\circ_{red} = -.83\,V$

so, $Cu^{+2} + 2e^- \rightarrow Cu \quad E^\circ_{red} = .34\,V$

b. Write the half reaction that will occur at the anode.

$2I^{-1} \rightarrow I_2 + 2e^- \quad E^\circ_{ox} = -.53\,V$

$2H_2O \not\rightarrow O_2 + 4H^{+1} + 4e^- \quad E^\circ_{ox} = -1.23\,V$

so $2I^{-1} \rightarrow I_2 + 2e^- \quad E^\circ_{ox} = -.53\,V$

c. Write the overall redox reaction that occurs, and calculate E°rxn (E°cell)

$Cu^{+2}_{(aq)} + 2I^{-1}_{(aq)} \rightarrow Cu_{(s)} + I_{2(s)} \quad E^\circ_{rxn} = -0.19\,V$

d. What mass of Cu(s) can form, if a the solution is electrolyzed for 45 minutes, with an average current of 0.62 A?

$$(45\,min)\left(\frac{60\,s}{min}\right)\left(\frac{0.62\,C}{s}\right)\left(\frac{1\,mole\,e^-}{96485\,C}\right)\left(\frac{1\,mole\,Cu}{2\,mole\,e^-}\right)\left(\frac{63.546\,g}{1\,mole}\right) = 0.55\,g\,Cu$$

6. Suppose that 24.42 mL of 0.50 M $K_2Cr_2O_7$ (in acidic solution) were needed to oxidize 7.59 g solid lead.

a. Write a balanced half reaction for the reduction of the dichromate ion to Cr^{+3}, in acid.

$Cr_2O_7^{-2} + 14H^{+1} + 6e^- \rightarrow 2Cr^{+3} + H_2O$

b. Which lead ion formed: Plumbous or plumbic?

$$(.02442\,L)\left(\frac{.50\,mole\,Cr_2O_7^{-2}}{L}\right)\left(\frac{6\,mole\,e^-}{1\,mole\,Cr_2O_7^{-2}}\right) = 0.07326\,moles\,e^-$$

$$(7.59\,g\,Pb)\left(\frac{1\,mole}{207.2\,g}\right) = 0.03663\,moles\,Pb$$

$$\frac{.07326\,moles\,e^-}{.03663\,mol\,Pb} = 1.9\bar{9} \approx 2$$

Pb^{+2} (Plumbous) ion formed.

7. A voltage was applied to a solution of "Tin sulfate." An electric current was 950 mA (milliamperes) applied over the course of 1.0 hour resulted in the production of 2.10 g solid tin. Was the compound "Tin II sulfate" or "Tin IV sulfate"?

$$(1.0\,hr)\left(\frac{3600\,s}{1\,hr}\right)\left(\frac{0.95\,C}{s}\right)\left(\frac{1\,mole\,e^-}{96485\,C}\right) = 0.035446\,moles\,e^-$$

$$(2.10\,g\,Sn)\left(\frac{1\,mole}{118.71\,g}\right) = 0.01769\,moles\,Sn$$

$$\frac{.035446\,moles\,e^-}{.01769\,moles\,Sn} = 2.0037 \approx 2$$

Tin (II) sulfate or stannous sulfate since it was Sn^{+2}

8. (AP 2002)

The reaction between silver ion and solid zinc is represented by the following equation.

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

(a) A 1.50 g sample of Zn is combined with 250. mL of 0.110 *M* $AgNO_3$ at 25°C.

(i) Identify the limiting reactant. Show calculations to support your answer.

$(0.250\ L)(.110\ \frac{mole}{L}\ AgNO_3)\left(\frac{1\ mole\ Zn^{+2}}{2\ mole\ Ag^{+1}}\right) = 0.01375$ mole Zn^{+2} could form from $AgNO_3$

$(1.50g\ Zn)\left(\frac{1 mole}{65.38g}\right)\left(\frac{1\ mole\ Zn^{+2}}{1 mole\ Zn}\right) = 0.02294$ mole Zn^{+2} could form from Zn

since $AgNO_3$ would produce a lower amount of product, **$AgNO_3$ is the limiting reactant.**

(ii) On the basis of the limiting reactant that you identified in part (i), determine the value of $[Zn^{2+}]$ after the reaction is complete. Assume that volume change is negligible.

Acc to above, .0137(5) moles Zn^{+2} could form.

$[Zn^{+2}] = \frac{0.0137(5)\ moles}{0.250\ L} =$ **0.0550 M**

(b) Determine the value of the standard potential, $E°$, for a galvanic cell based on the reaction between $AgNO_3(aq)$ and solid Zn at 25°C.

← (look up on volt table in AP "greensheets")

$Ag^{+1} + 1e^- \rightarrow Ag \quad E°_{red} = 0.80V$

$Zn^{+2} + 2e^- \rightarrow Zn \quad E°_{red} = -.76\ V$ ← but need to reverse this to get $E°_{ox}$

$E°_{ox}(Zn) = 0.76\ V$

$E°_{red} + E°_{ox} = 0.80\ V + 0.76\ V =$ **1.56 V**

Another galvanic cell is based on the reaction between $Ag^{+}(aq)$ and $Cu(s)$, represented by the equation below. At 25°C, the standard potential, $E°$, for the cell is 0.46 V.

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

(c) Determine the value of the standard free-energy change, $\Delta G°$, for the reaction between $Ag^{+}(aq)$ and $Cu(s)$ at 25°C.

$\Delta G° = -n\mathcal{F}E° = -\left(\frac{2\ mole\ e^-}{1 mole\ rxn}\right)\left(\frac{96485\ C}{Mole\ e^-}\right)\left(\frac{0.46\ J}{C}\right) = -88766$ J/mole ≈ **−89 KJ/mole**

(d) The cell is constructed so that $[Cu^{2+}]$ is 0.045 *M* and $[Ag^{+}]$ is 0.010 *M*. Calculate the value of the potential, *E*, for the cell.

$E = E° - \frac{RT}{n\mathcal{F}} \ln Q \qquad Q = \frac{[Cu^{+2}]}{[Ag^{+1}]^2} = \frac{(.045\ M)}{(.010\ M)^2} = 450$

$E = 0.46\ V - \frac{(8.314\ \frac{J}{mol \cdot K})(298\ K)}{\left(\frac{2\ mole\ e^-}{mole\ rxn}\right)\left(96485\ \frac{C}{mole\ e^-}\right)} \ln(450)$

$E = 0.46\ V - .012839\ V\ (6.1092) =$ **0.38 V**

(e) Under the conditions specified in part (d), is the reaction in the cell spontaneous? Justify your answer.

Yes. It is spontaneous. You can tell because $E = 0.38\ V$; this is a positive cell potential.

9. (AP 1997)

In an electrolytic cell, a current of 0.250 ampere is passed through a solution of a chloride of iron, producing $Fe(s)$ and $Cl_2(g)$.

(a) Write the equation for the half-reaction that occurs at the anode.

(b) When the cell operates for 2.00 hours, 0.521 gram of iron is deposited at one electrode. Determine the formula of the chloride of iron in the original solution.

(c) Write the balanced equation for the overall reaction that occurs in the cell.

(d) How many liters of $Cl_2(g)$, measured at 25°C and 750 mm Hg, are produced when the cell operates as described in part (b) ?

(e) Calculate the current that would produce chlorine gas from the solution at a rate of 3.00 grams per hour.

(a) anode is where oxidation occurs, so:

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

(also, you know it can't be the Fe rxn because you don't know charge of Fe yet)

(b) $(2.00\,hr)\left(\frac{3600\,s}{hr}\right)\left(\frac{0.250\,C}{s}\right)\left(\frac{1\,mole\ e^-}{96485\,C}\right) = 0.01865(6)$ moles e^-

$(0.521\,g\ Fe)\left(\frac{1\,mole}{55.847\,g}\right) = 0.009329$ moles Fe

$\frac{.018656\ moles\ e^-}{.009329\ moles\ Fe} = 1.999 \approx 2!$ it is Fe^{+2}, so $FeCl_2$ is chloride formula.

(c) iron half rxn is $Fe^{+2}_{(aq)} + 2e^- \rightarrow Fe_{(s)}$

combine w/ rxn in (a)

$$Fe^{+2}_{(aq)} + 2Cl^{-1}_{(aq)} \rightarrow Fe_{(s)} + Cl_{2(g)}$$

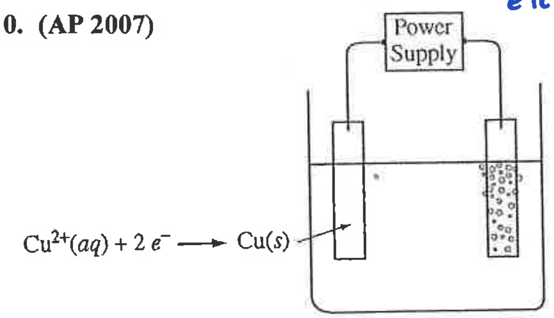
(d) (from (b)) $(.00932(9)\ moles\ Fe)\left(\frac{1\,mole\ Cl_2}{1\,mole\ Fe}\right) = .00932(9)$ moles Cl_2

$$V = \frac{nRT}{P} = \frac{(.00932(9)\ moles)(.0821\ \frac{L\cdot atm}{mol\,K})(298\,K)}{\left(\frac{750}{760}\ atm\right)} = 0.231\ liters$$

(e) $\left(\frac{3.00\,g\ Cl_2}{1\,hr}\right)\left(\frac{1\,hr}{3600\,s}\right)\left(\frac{1\,mole\ Cl_2}{70.906\,g}\right)\left(\frac{2\,mole\ e^-}{1\,mole\ Cl_2}\right)\left(\frac{96485\,C}{1\,mole\ e^-}\right) = 2.27\ \frac{C}{s}$

or 2.27 Amperes!

10. (AP 2007)



← electrons flow from right to left through the wire.

An external direct-current power supply is connected to two platinum electrodes immersed in a beaker containing 1.0 *M* $CuSO_4(aq)$ at 25°C, as shown in the diagram above. As the cell operates, copper metal is deposited onto one electrode and $O_2(g)$ is produced at the other electrode. The two reduction half-reactions for the overall reaction that occurs in the cell are shown in the table below.

Half-Reaction	$E°$(V)
$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$	+1.23
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	+0.34

rev: $2H_2O \rightarrow O_2 + 4H^{+1} + 4e^-$ $E°_{ox} = -1.23V$

double $2Cu^{+2} + 4e^- \rightarrow 2Cu(s)$

(a) On the diagram, indicate the direction of electron flow in the wire. (see above)

(b) Write a balanced net ionic equation for the electrolysis reaction that occurs in the cell.

(c) Predict the algebraic sign of $\Delta G°$ for the reaction. Justify your prediction.

(d) Calculate the value of $\Delta G°$ for the reaction.

An electric current of 1.50 amps passes through the cell for 40.0 minutes.

(e) Calculate the mass, in grams, of the $Cu(s)$ that is deposited on the electrode.

(f) Calculate the dry volume, in liters measured at 25°C and 1.16 atm, of the $O_2(g)$ that is produced.

(e) $(40.0\,min)\left(\frac{60\,s}{min}\right)\left(\frac{1.50\,C}{s}\right)\left(\frac{1\,mole\,e^-}{96485\,C}\right)\left(\frac{1\,mole\,Cu}{2\,mole\,e^-}\right)\left(\frac{63.546\,g}{mole}\right)$ = **1.19 g Cu**

(b) need to reverse the first rxn so O_2 is produced, and double the Cu^{+2} rxn so that 4 e- are needed, to cancel w/ 4e- in O_2 rxn.

$$2Cu^{+2} + 2H_2O \rightarrow O_2 + 4H^{+1} + 2Cu$$

(c) $E°_{rxn} = E°_{red} + E°_{ox} = 0.34\,V + -1.23\,V = -0.89\,V$

This rxn is not spontaneous, so $\Delta G°_{rxn}$ must be positive. this is consistent with our negative value for $E°_{rxn}$.

(d) see (c); $E°_{rxn} = -0.89\,V$

$\Delta G° = -nFE° = -\left(\frac{4\,mole\,e^-}{mole\,rxn}\right)\left(\frac{96485\,C}{mole\,e^-}\right)\left(\frac{-.89\,J}{C}\right) = 343487\,J/mole$

340 kJ/mole

(f) $(40.0\,min)\left(\frac{60\,s}{min}\right)\left(\frac{1.50\,C}{s}\right)\left(\frac{1\,mole\,e^-}{96485\,C}\right)\left(\frac{1\,mole\,O_2}{4\,mole\,e^-}\right) = .009326\,mole\,O_2$

$$V = \frac{nRT}{P} = \frac{(0.009326\,mole)(.0821\,L\cdot atm/mol\cdot K)(298\,K)}{1.16\,atm} = 0.197\,liters$$

Problem CC.

$Cu^{+2}_{(aq)} + 2\,e^- \text{ ------> } Cu_{(s)}$ $\quad E^\circ_{reduction} = 0.34$ V

$Cu_{(s)} \text{ ------> } Cu^{+2}_{(aq)} + 2\,e^-$ $\quad E^\circ_{oxidation} =$ **−0.34** V

A "concentration cell" is constructed using the above two half reactions.
The left beaker contains a piece of solid copper and a 2.0 M $CuSO_4$ solution. The right beaker contains a piece of solid copper, and a 0.080 M $CuSO_4$ solution. The salt bridge contains potassium nitrate.

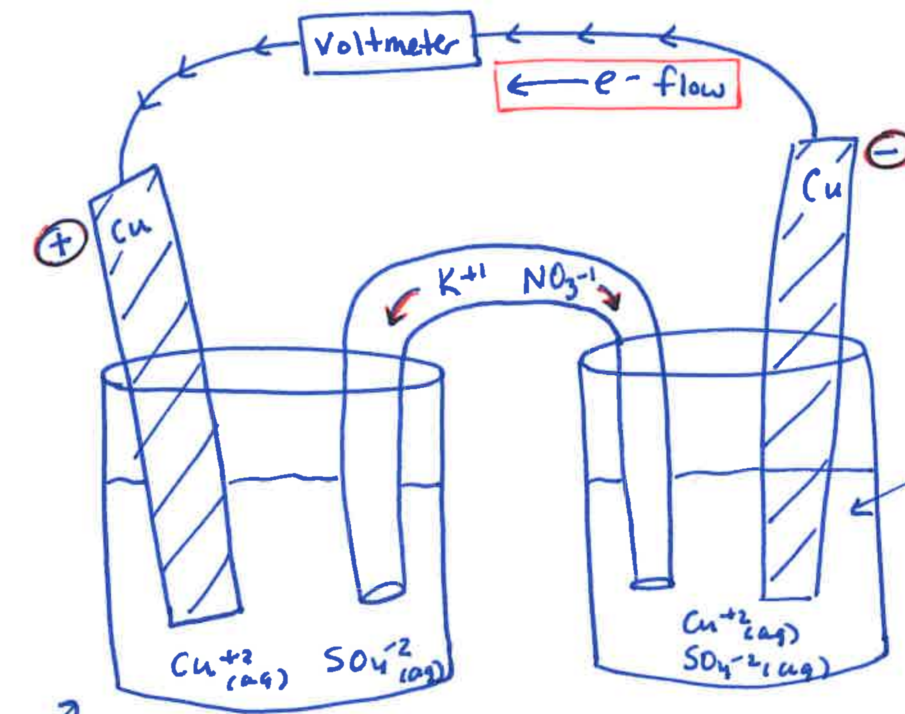
a. Sketch a picture of the voltaic cell.
b. Fill in the oxidation half cell voltage, above, and determine E°_{cell} .

$$E^\circ_{cell} = E^\circ_{red} + E^\circ_{ox} = 0.34\,V + (-.34\,V) = \emptyset \text{ Volts!}$$

c. On your sketch, show the direction of electron flow through the external circuit, and the direction of anion and cation flow in the salt bridge.
d. Label the anode and cathode of the cell, and label each electrode as + or –
e. Which electrode will gain mass during the reaction? which will lose mass?

The cathode will gain mass as $Cu_{(s)}$ forms and plates out.
The anode piece of Cu will lose mass since $Cu_{(s)}$ is a reactant on the anode (right) side.

f. How will the appearance and concentration (of Cu^{+2}) of the solutions change as the cell operates?



$[Cu^{+2}] = 2.0$ M

<u>reduction at cathode</u>:

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

$[Cu^{+2}] = 0.080$ M

<u>oxidation at anode</u>:

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

(f) the solution on the anode side will start out light blue, but the blue color will become darker blue as the rxn proceeds, since $Cu^{+2}_{(aq)}$ is produced; $[Cu^{+2}]$ will increase.

(f) The solution on the cathode side will start out deep blue due to Cu^{+2}, but will become lighter blue over time; $[Cu^{+2}]$ will decrease.

problem CC, cont'd

g. Write an overall reaction for the cell, and an expression for Q. (since the Cu^{+2} ions will be on each side, indicate whether you mean the $Cu^{+2}_{(aq)}$ in the concentrated solution or in the dilute solution.)

cathode: $Cu^{+2}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$ concentrated side

anode: $Cu_{(s)} \rightarrow 2e^- + Cu^{+2}_{(aq)}$ dilute side

overall: $Cu^{+2}_{(aq)}$ concentrated $\rightarrow$ $Cu^{+2}_{(aq)}$ dilute

$$Q = \frac{[Cu^{+2}_{(aq)}]_{dilute}}{[Cu^{+2}_{(aq)}]_{concentrated}}$$

h. Calculate the cell voltage, E_{cell}. at 298 K.

$$Q = \frac{[Cu^{+2}]_{dilute}}{[Cu^{+2}]_{conc'd}} = \frac{0.080\ M}{2.0\ M} = 0.040$$

$$E = E^\circ - \frac{RT}{n\mathcal{F}}\ln Q = \emptyset - \frac{(8.314\frac{J}{mol \cdot K})(298\ K)}{\left(\frac{2\ mole\ e^-}{mole\ rxn}\right)\left(\frac{96485\ C}{mole\ e^-}\right)}\ln(.040) = 0.041\ V$$

i. Calculate the cell voltage after it has been operating for a while, so that the left beaker has a cupric ion concentration of 1.2 M, and the right beaker has a cupric ion concentration of 0.88 M.

$$Q = \frac{[Cu^{+2}]_{dilute}}{[Cu^{+2}]_{conc'd}} = \frac{0.88\ M}{1.2\ M} = 0.73(33)\ V$$

$$E = E^\circ - \frac{RT}{n\mathcal{F}}\ln Q = \emptyset - \frac{(8.314)(298)}{(2)(96485)}V(\ln(.7333)) = 0.00398 \rightarrow 0.0040\ V$$

we are getting fewer volts than in part (h) since we are now closer to equilibrium.

j. If this voltaic cell is allowed to run until equilibrium is reached,

i. what will happen to the molarities?
ii. What will be the value of Q?
iii. what will happen to the cell emf?

(i) The molarity of Cu^{+2} will decrease on the cathode side and increase on the anode side until the molarities of Cu^{+2} become equal! In this case, if volumes of solution are the same, both sides would reach equilibrium when $[Cu^{+2}] = \left(\frac{2.0 + .080}{2}\right)M = 1.04\ M$

ii) once the molarities are equal, $Q = 1$

iii) the cell emf will drop to zero!

$$E = E^\circ - \frac{RT}{n\mathcal{F}}\ln(1) = E^\circ = \emptyset$$

AP Multiple Choice (Answers are on the previous page.)
This part of the packet is optional, but good practice. I'll put a few MC on your quiz/test.

19. In which of the following species does sulfur have the same oxidation number as it does in H_2SO_4 ? it is (+6) in H_2SO_4

(A) H_2SO_3 S= +4

(B) $S_2O_3^{2-}$ S = +2

(C) S^{2-} S = -2

(D) S_8 S = 0

(E) SO_2Cl_2 S = +6

Questions 34-35 refer to an electrolytic cell that involves the following half-reaction.

$$AlF_6^{3-} + 3\ e^- \rightarrow Al + 6\ F^-$$

34. Which of the following occurs in the reaction?

(A) AlF_6^{3-} is reduced at the cathode. yep. AlF_6^{-3} gains e-.. and reduction does occur @ cathode!
(B) Al is oxidized at the anode. no.. Al is not a reactant
(C) Aluminum is converted from the −3 oxidation state to the 0 oxidation state. no, Al starts out as +3, not −3
(D) F^- acts as a reducing agent. no. F^{-1} is not a reactant.
(E) F^- is reduced at the cathode. no, F^- is not a reactant

35. A steady current of 10 amperes is passed through an aluminum-production cell for 15 minutes. Which of the following is the correct expression for calculating the number of grams of aluminum produced? (1 faraday = 96,500 coulombs)

(A) $\frac{(10)(15)(96,500)}{(27)(60)}$ g

$\left(15\ min\right)\left(\frac{60\ s}{min}\right)\left(\frac{10\ C}{s}\right)\left(\frac{1\ mole\ e^-}{96500\ C}\right)\left(\frac{1\ mole\ Al}{3\ mole\ e^-}\right)\left(\frac{27.0\ g}{mole}\right)$

(B) $\frac{(10)(15)(27)}{(60)(96,500)}$ g

(C) $\frac{(10)(15)(60)(27)}{(96,500)(3)}$ g

or $\frac{(15)(60)(10)(27)}{(96500)(3)}$

(D) $\frac{(96,500)(27)}{(10)(15)(60)(3)}$ g

(E) $\frac{(27)(3)}{(96,500)(10)(15)(60)}$ g

$$3\ Cu(s) + 8\ H^+(aq) + 2\ NO_3^-(aq) \rightarrow 3\ Cu^{2+}(aq) + 2\ NO(g) + 4\ H_2O(l)$$

(+5) (+2)

51. True statements about the reaction represented above include which of the following?

I. $Cu(s)$ acts as an oxidizing agent. no, Cu is oxidized.
II. The oxidation state of nitrogen changes from +5 to +2. True.
III. Hydrogen ions are oxidized to form $H_2O(l)$. no, H^{+1} and H_2O have H @ ox # of (+1)

(A) I only
(B) II only
(C) III only
(D) I and II
(E) II and III

49. $\ldots Cr_2O_7^{2-}(aq) + \ldots H_2S(g) + \ldots H^+(aq) \rightarrow \ldots Cr^{3+}(aq) + \ldots S(s) + \ldots H_2O(l)$

(+6) (−2) (+3) (0)

When the equation above is correctly balanced and all coefficients are reduced to lowest whole-number terms, the coefficient for $H^+(aq)$ is

(A) 2 (B) 4 (C) 6 (D) 8 (E) 14

$Cr_2O_7^{-2} + 14\ H^{+1} + 6e^- \rightarrow 2Cr^{+3} + 7H_2O \rightarrow Cr_2O_7^{-2} + 14\ H^{+1} + 6e^- \rightarrow 2Cr^{+3} + 7H_2O$

$(H_2S \rightarrow S + 2H^{+1} + 2e^{-1}) \times 3 \rightarrow 3H_2S \rightarrow 3S + 6H^{+1} + 6e^-$

$Cr_2O_7^{-2} + 8H^{+1} + 3H_2S \rightarrow 3S + 2Cr^{+3} + 7H_2O$

$M(s) + 3\ Ag^+(aq) \rightarrow 3\ Ag(s) + M^{3+}(aq)$	$E° = +2.46$ V
$Ag^+(aq) + e^- \rightarrow Ag(s)$	$E° = +0.80$ V

57. According to the information above, what is the standard reduction potential for the half-reaction $M^{3+}(aq) + 3\ e^- \rightarrow M(s)$?

(A) −1.66 V
(B) −0.06 V
(C) 0.06 V
(D) 1.66 V
(E) 3.26 V

$M_{(s)} \rightarrow M^{+3} + 3e^-$ $E° = ?$

$3(Ag^{+1} + e^- \rightarrow Ag)$ $E° = 0.80V$

$M + 3Ag^{+1} \rightarrow M^{+3} + 3Ag$ $E° = 2.46V$

$2.46V - 0.80V = 1.66\ V$ for $M \rightarrow M^{+3} + 3e^-$

but they want reverse of this, so −1.66 V

71. In the electroplating of nickel, 0.200 faraday of electrical charge is passed through a solution of $NiSO_4$. What mass of nickel is deposited?

$\mathcal{F} = \frac{96500\ C}{mole\ e^-}$ .. or here they are thinking of it as the charge on 1 mole e−

(A) 2.94 g
(B) 5.87 g
(C) 11.7 g
(D) 58.7 g
(E) 294 g

$(0.200\ mole\ e^-)\left(\frac{1\ mole\ Ni}{2\ mole\ e^-}\right)\left(\frac{58.7g}{mole}\right) = 5.87\ g$