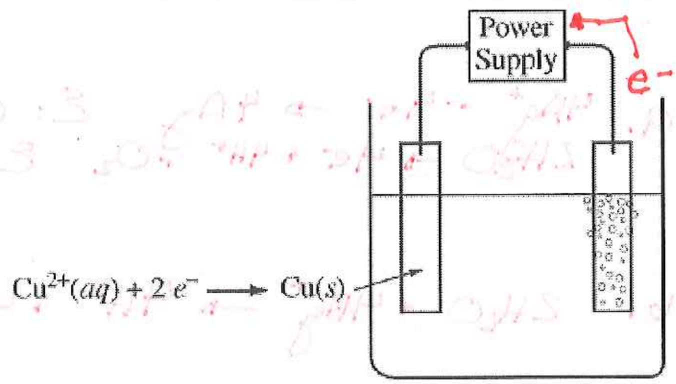
Name ______________________________ **AP Chemistry**

Chapter 16 HW 3: Due 3/15/17. Complete both electrochemistry free response questions. One will be graded and one won't. I usually pick the harder one(for me) to grade. Sometimes I pick the second one because it would be weird if a student did problem 2 but skipped problem 1 on a two problem assignment. Show all work. Box and clearly label all final free response answers.



1. 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)
$O_2(g) + 4\,H^+(aq) + 4\,e^- \rightarrow 2\,H_2O(l)$	+1.23
$Cu^{2+}(aq) + 2\,e^- \rightarrow Cu(s)$	+0.34

(a) On the diagram, indicate the direction of electron flow in the wire.
(b) Write a balanced net ionic equation for the electrolysis reaction that occurs in the cell.
(c) Predict the algebraic sign of ΔG° for the reaction. Justify your prediction.
(d) Calculate the value of ΔG° for the reaction.
An electric current of 1.80 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 1.26 atm, of the $O_2(g)$ that is produced.

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

c. $\Delta G = +$ b/c $\varepsilon = -0.89V$

d. $\Delta G = -nF\varepsilon$
$\Delta G = (-4)(96485)(-0.89)$
$\boxed{\Delta G = +343kJ}$

e. $\frac{40\,min}{} \Big| \frac{60\,sec}{1\,min} \Big| \frac{1.80C}{sec} \Big| \frac{1\,mol\,e^-}{96485\,C} \Big| \frac{1\,mol\,Cu}{2\,mol\,e^-} \Big| \frac{63.55g}{1\,mol\,Cu} = \boxed{1.42g}$

f. $\frac{40\,min}{} \Big| \frac{60\,sec}{1\,min} \Big| \frac{1.80\,C}{sec} \Big| \frac{1\,mole}{96485\,C} \Big| \frac{1\,mol\,O_2}{4\,mole} = 0.0112\,mol$

$V = \frac{nRT}{P} = \frac{(0.0112)(0.08206)(298)}{1.26} = \boxed{0.217\,L}$

2. A 730. mL sample of a 0.60 M silver nitrate solution is electrolyzed. The solution was subjected to current of 2.00 amperes for 55 minutes. The solution became progressively more acidic as the reaction progressed.
(a) Write the two half-reactions that occur and indicate which takes place at the anode and which takes place at the cathode.
(b) Write the overall balanced chemical equation.
(c) What is the minimum voltage required for this electrolysis to occur.
(d) If the gas produced in the electrolysis is collected at 1.16 atm and 40.°C, what volume of oxygen gas was produced?
(e) What mass of solid silver is produced in the electrolysis?

a. $4Ag^+ + 4e^- \rightarrow 4Ag$ $\varepsilon = 0.80V$ cathode
$2H_2O \rightarrow 4e^- + 4H^+ + O_2$ $\varepsilon = -1.23$ anode

b. $2H_2O + 4Ag^+ \rightarrow 4H^+ + 4Ag + O_2$ $\varepsilon = -0.43V$

c. greater than 0.43

d $\frac{55\,min}{} \cdot \frac{60\,sec}{1\,min} \cdot \frac{2.00C}{1\,sec} \cdot \frac{1\,mol}{96485\,C} \cdot \frac{1\,mol\,O_2}{4\,mol\,e^-}$ $\boxed{0.0171\,mol.\ O_2}$

$PV = nRT$

$V = \frac{nRT}{P} = \frac{(0.0171)(0.08206)(313)}{1.16} = \boxed{0.379L}$

e. $\frac{AtM}{Fe^-} = mass$ $\frac{(2.0)(3300)(107.9)}{(96485)(1)} = \boxed{7.38g\,Ag}$