Name ______________________________ **AP Chemistry**

HW 2: Due 3/18/15. Complete both electrochemistry free response questions. Only one will be graded – sorry it keeps me sane and makes you work hard. Do you ever think, "Oh, it's this one" and then get it back and be right and ace that problem. That must feel good. Show all work. Box and clearly label all final free response answers.

1. $X(s) + Cu^{2+}(aq) \rightarrow X^{2+}(aq) + Cu(s)$

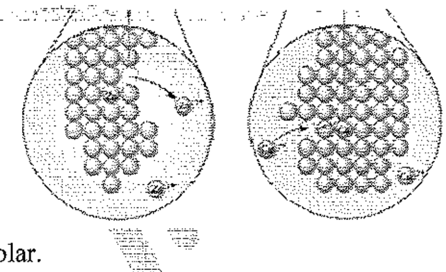
For the reaction above, E° = 0.620 volt at 25 °C.

(a) Determine the standard electrode potential for the reaction half-reaction $X^{2+}(aq) + 2\ e^{-} \rightarrow X(s)$.

(b) What is the chemical symbol for element X?

(c) Diagram the cell and label the cathode, anode, salt bridge and indicate the directional flow of electrons.

(d) Draw a diagram similar to the one on the right showing the atomic view of the reaction that is happening at each electrode. Label the electrodes. **See outline page 3 for help.**



(e) Calculate ΔG° for the cell.

(f) A cell is constructed in which the reaction above occurs. All substances are initially in their standard states, and equal volumes of the solutions are used. The cell is then discharged. Calculate the value of the cell potential, E°, when $[Cu^{2+}]$ has dropped to 0.55 molar.

(g) Find the ratio $[X^{2+}](aq) / [Cu^{2+}](aq)$ when the cell reaction above reaches equilibrium

a. $Cu^{2+} + 2e^{-} \rightarrow Cu$ 0.34

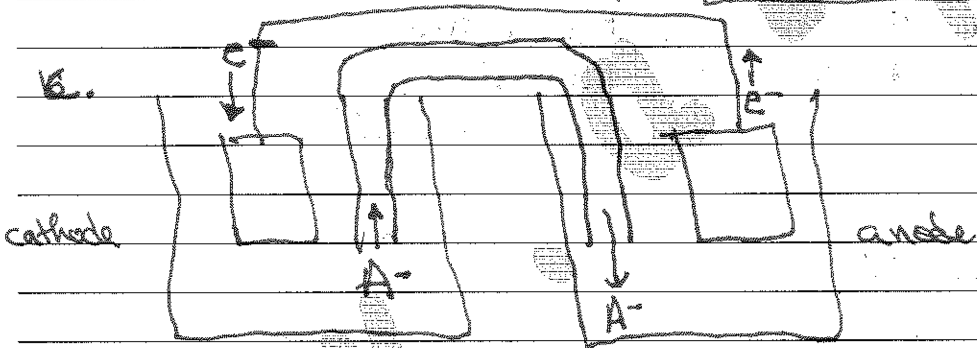
$X \rightarrow X^{2+} + 2e^{-}$ 0.28

0.62V

$X^{2+} + 2e^{-} \rightarrow X$ $\boxed{\mathcal{E} = -0.28V}$

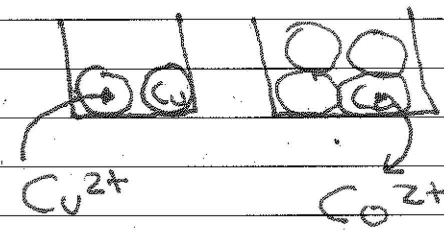
b. $\boxed{\text{Co is X}}$

c.



$Cu^{2+} + 2e^{-} \rightarrow Cu$ $X \rightarrow X^{2+} + 2e^{-}$

d.



e. $\Delta G = -nFE$

$\Delta G = (-2)(96485)(0.620)$

$\boxed{\Delta G = -120.\ kJ}$

f. $\mathcal{E} = \mathcal{E}° - \frac{0.0591}{2} \log\left(\frac{1.45}{0.55}\right)$

$\mathcal{E} = 0.620 - 0.0125$

$\boxed{\mathcal{E} = 0.607\ V}$

	$Co + Cu^{2+} \rightarrow$	$\cancel{X}Co^{2+} + Cu$
I	1.0	1.0
C	−x	+x
E	0.55	1.45

g. $0 = 0.620 - \frac{0.0591}{2} \log X$

$-0.620 = -0.0296 \log X$

$20.98 = \log X$

$\boxed{X = 9.58 \times 10^{20}}$

2. Hydrogen gas burns in air according to the equation to the right: $2H_2(g) + O_2(g) \rightarrow 2\ H_2O(l)$

(a) Calculate the standard enthalpy change, $\Delta H°$, for the reaction equation above. $\Delta H_f°$, for $H_2O(l)$ is -285.8 kJ mol^{-1}

(b) Calculate the amount of heat, in kJ, that is released when 55.0 g of $H_2(g)$ is burned in air.

(c) Given that the molar enthalpy of vaporization, $\Delta H°_{vap}$, for $H_2O(l)$ is 44.0 kJ mol^{-1} at 298 K, what is the standard enthalpy change, $\Delta H°$, for the reaction $2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$?

A fuel cell is an electrochemical cell that converts the chemical energy stored in a fuel into electrical energy. A cell that uses H_2 as the fuel can be constructed based on the following half-reactions.

Half-reaction	$E°$ (298 K)
$2\ H_2O(l) + O_2(g) + 4\ e^- \rightarrow 4\ OH^-(aq)$	0.40 V
$2\ H_2O(l) + 2\ e^- \rightarrow H_2(g) + 2\ OH^-(aq)$	-0.83 V

(d) Write the equation for the overall cell reaction.

(e) Calculate the standard potential for the cell at 298 K.

(f) Calculate the $\Delta G°$ for the cell at 298 K.

(g) Assume that 0.73 mol of $H_2(g)$ is consumed as the cell operates for 605 seconds.

(i) Calculate the number of moles of electrons that pass through the cell.

(ii) Calculate the average current, in amperes, that passes through the cell.

(h) Some fuel cells use butane gas, C_4H_{10}, rather than hydrogen gas. The overall reaction that occurs in a butane fuel cell is:

$$2\ C_4H_{10}(g) + 13\ O_2(g) \rightarrow 8\ CO_2(g) + 10\ H_2O(l).$$

What is one environmental advantage of using fuel cells that are based on hydrogen rather than on hydrocarbons such as butane?

a. $(2)(-285.8) = \boxed{-571.6\text{ KJ}}$

b. $\frac{4.04}{-571.6} = \frac{55.0}{x}$ $\boxed{x = -7780\text{ KJ}}$

c. $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$ $\Delta H = -571.6$

$2H_2O(l) \rightarrow 2H_2O(g)$ $\Delta H = +88$

$\boxed{-483.6\text{ KJ}}$

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

e. $0.83 + 0.40 = \boxed{1.23\text{ V}}$

f. $\Delta G = -nFE$

$\Delta G = (-4)(96485)(1.23)$

$\boxed{\Delta G = -475\text{ KJ}}$

gi. $\frac{0.73\text{ mol }H_2}{} \Big| \frac{4\text{ mol }e^-}{2\text{ mol }H_2} = \boxed{1.46\text{ mol }e^-}$

gii. $1.46\text{ mole} \times \frac{96485\text{ C}}{1\text{ mol }e^-} = 140868\text{ C} \div 605 = \boxed{233\text{ A}}$

h. Hydrogen fuel cells only produce water. Butane fuel cells also produce CO_2 which is a greenhouse gas