#-4

4/23/13

electrical work = charge times voltage

$w = -QV$

neg sign means work is done by system

$\Delta G = -nFE$

E = voltage

n = # electrons/mole reaction

F = Faraday's constant

(3) poll

From the given thermo data, what is the value of $\Delta G°$?

$\boxed{-1159 \text{ kJ/mol}}$

$\Delta G°_{rxn} = \Sigma G_{fP} - \Sigma G_{fR}$

$3Cu^{2+} + 2Al \rightarrow 3Cu + 2Al^{3+}$

(4) poll

From the given computed value of $\Delta G°$, what is the value of K?

$\boxed{\text{ENORMOUS}}$ *in calculator, it's error*

$\Delta G°_{rxn} = -RT\ln K$ *$\Delta G = -1200$ kJ/mol*

(5) poll

From the given standard reduction potential data, what is the value of $E_{cell}°$?

$\boxed{+2V}$ - spontaneous rxn has positive potential

Anode

$Al \rightarrow Al^{3+} + 3e$

Cathode

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

$\frac{1 \text{ mol}}{m}$

$3Cu^{2+} + 2Al \rightarrow 2Al^{3+} + 3Cu$

$\Delta G = -nFE°$

(6) poll

what is the maximum amount of electrical work that can be extracted from running this cell under standard conditions.

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

$F = 96485$

$E = 2$

$n = ?$ $=$ 6 electrons

Answer: $\boxed{1158 \text{ kJ/mol}}$

NERNST EQUATION

$\Delta G = \Delta G^\circ + RT\ln Q$

$-nFE = -nFE^\circ + RT\ln Q$

$$\boxed{E = E^\circ - \frac{.0591}{n}\log Q}$$ ← MEMORIZE!

$Q = \frac{[Al^{3+}]^2}{[Cu^{2+}]^3} = \frac{(.031)^2}{(.21)^3} = 1$

$Q = 1$

$n = 6$

Calc the emf of the cell:

$\langle Al(s) | Al^{3+}(aq, .001 M) || Cu^{2+}(aq, .01 M) | Cu(s) \rangle$

$\boxed{2V}$

$Q = 10^{24}$

$n = 6$

Calc emf of the cell:

$\langle Al(s) | Al^{3+}(aq, 1M) || Cu^{2+}(aq, 1\times10^{-8}M) | Cu(s) \rangle$

$\boxed{1.76 \text{ V}}$