

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 ΔG° ?

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

$$\Delta G^\circ_{\text{rxn}} = \sum G_{\text{fP}} - \sum G_{\text{fR}}$$



(4) poll

From the given computed value of ΔG° , what is the value of K?

$\boxed{\text{ENORMOUS}}$

* in calculator, it's error *

$$\Delta G^\circ_{\text{rxn}} = -RT \ln K \quad * \Delta G = -1200 \text{ kJ/mol} *$$

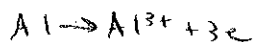
(5) poll

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

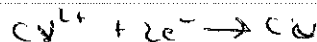
$\boxed{+2V}$

- spontaneous rxn has positive potential

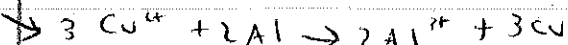
Anode



Cathode



$\frac{1 \text{ mol}}{13}$



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

(c) poll

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

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

$$F = 96485$$

$$E = 2$$

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

$$n = ? = 2 \text{ electrons}$$

NERNST EQUATION

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

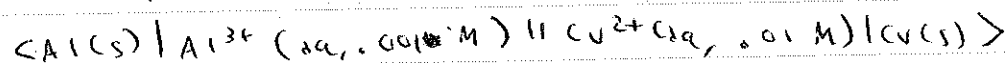
$$-nFE = -nFE^\circ + RT \ln Q$$

$$E = E^\circ - \frac{0.0591}{n} \log Q$$

MEMORIZE!

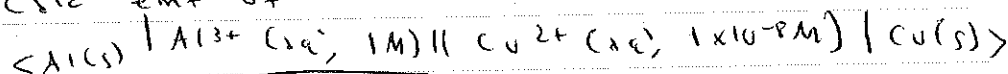
$$Q = \frac{[Al^{3+}]^2}{[Cu^{2+}]^2} = \frac{(0.031)^2}{(0.21)^2} = 1 \quad Q = 1$$
$$n = 6$$

Calc the emf of the cell:



$$\boxed{2V}$$

Calc emf of the cell:



$$\boxed{1.76 V}$$

$$Q = 10^{24}$$

$$n = 6$$