

NOTE: for this problem set, unless stated otherwise, all standard cell potentials and other standard thermodynamic values are assumed to be reported at 298 K.

1. Consider the demonstration where a piece of copper wire is being immersed in a solution of AgNO_3 .

- a) What is the sign of ΔG° for the demonstration you are observing? Explain.

Negative. Reaction is spontaneous.

- b) Write out a balanced overall reaction for the demonstration, including all phases and stoichiometric coefficients:



- c) In your engineering reference manual you find the following half-reactions:



Determine the standard cell potential associated (E°_{cell}) associated with this reaction.

$$E^\circ_{\text{cell}} = E^\circ_{\text{red}} - E^\circ_{\text{ox}} = 0.80 - 0.34 = 0.46 \text{ V}$$

- d) Calculate the value of ΔG° and K for the demonstration reaction at 298 K.

$$\begin{aligned} \Delta G^\circ &= -n F E^\circ_{\text{cell}} = -2 \times 96485 \text{ C mol}^{-1} \times 0.460 \text{ V} = -88766 \text{ J} \\ &= -89 \text{ kJ} \end{aligned}$$

$$\begin{aligned} \ln K &= -\frac{\Delta G^\circ}{RT} = \frac{88766 \text{ J}}{1 \text{ mol} \times 8.3145 \text{ J mol}^{-1} \text{K}^{-1} \times 298 \text{ K}} = 35.83 \\ K &= e^{35.83} = 3.62 \times 10^{15} \end{aligned}$$

- e) Use your knowledge of thermodynamics to predict what would happen if we were to carry out the opposite experiment – a piece of silver in $\text{Cu}(\text{NO}_3)_2$.

Because dipping Cu in AgNO_3 was a spontaneous process, that means the opposite process should be non-spontaneous.

Note: both this answer and that for part a) make a subtle assumption. Can you see what it is?

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S Ben Porath



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