



ELECTROCHEMISTRY

THE NERNST EQUATION: NONSTANDARD CONDITIONS

CHEMISTRY 165 // SPRING 2020

PRACTICE PROBLEM 1



Based on the cell diagram above and standard cell potential, is the constructed cell spontaneous at 298.15 K?

— *answer* —

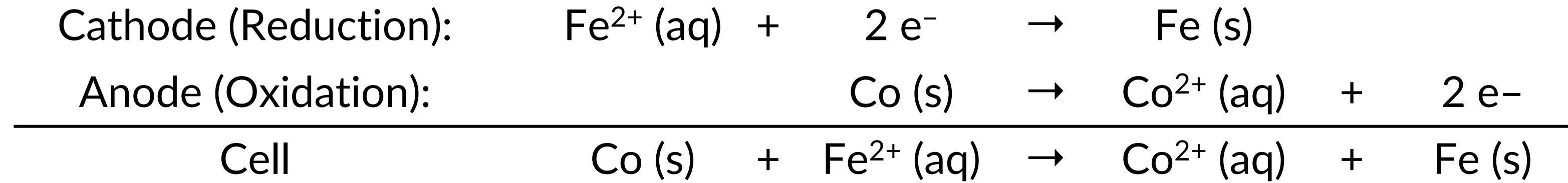
PRACTICE PROBLEM 1



Based on the cell diagram above and standard cell potential, is the constructed cell spontaneous at 298.15 K?

— answer —

First, write out the balanced net ionic equation for the cell reaction.

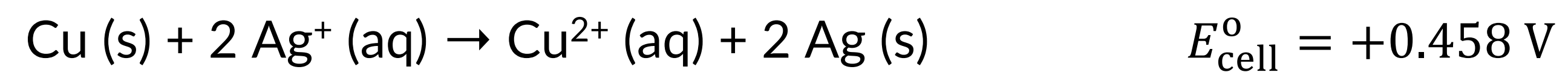


We can use the Nernst equation to determine the nonstandard cell potential:

$$\begin{aligned} E_{\text{cell}} &= E_{\text{cell}}^{\circ} - \frac{RT}{nF} \ln \frac{[\text{Co}^{2+}]}{[\text{Fe}^{2+}]} \\ &= -0.170 \text{ V} - \frac{\left(8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}}\right) (298.15 \text{ K})}{(2 \text{ mol } e^{-}) \left(96485 \frac{\text{C}}{\text{mol}}\right)} \ln \frac{0.15}{1.94} \\ E_{\text{cell}} &= -0.137 \text{ V} \end{aligned}$$

Because the $E_{\text{cell}} < 0$, this reaction/cell is still **nonspontaneous**.

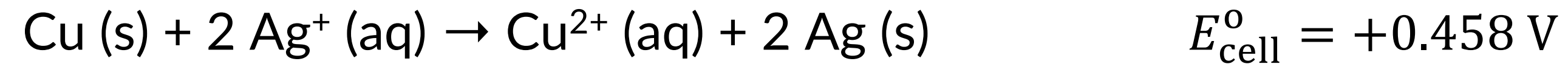
PRACTICE PROBLEM 2



If the value of $E_{\text{cell}} = +0.396 \text{ V}$ when $[\text{Ag}^+] = 2.56 \times 10^{-3} \text{ M}$ at 298 K, what is $[\text{Cu}^{2+}]$?

— *answer* —

PRACTICE PROBLEM 2



If the value of $E_{\text{cell}} = +0.396 \text{ V}$ when $[\text{Ag}^+] = 2.56 \times 10^{-3} \text{ M}$ at 298 K , what is $[\text{Cu}^{2+}]$?

— *answer* —

We can apply the Nernst equation find $[\text{Cu}^{2+}]$.

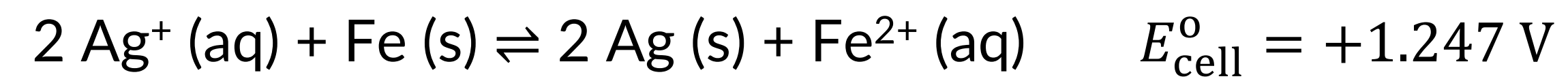
Be careful about the form of Q!

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{RT}{nF} \ln \frac{[\text{Cu}^{2+}]}{[\text{Ag}^+]^2}$$
$$+0.396 \text{ V} = +0.458 \text{ V} - \frac{\left(8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}}\right) (298 \text{ K})}{(2 \text{ mol } e^{-}) \left(96485 \frac{\text{C}}{\text{mol}}\right)} \ln \frac{[\text{Cu}^{2+}]}{(2.56 \times 10^{-3})^2}$$

$$[\text{Cu}^{2+}] = 8.25 \times 10^{-4} \text{ M}$$

PRACTICE PROBLEM 3

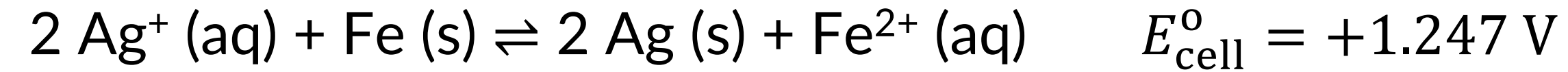
What is the equilibrium constant (K) for the following cell reaction when the reaction reaches equilibrium at 25 °C?



— *answer* —

PRACTICE PROBLEM 3

What is the equilibrium constant (K) for the following cell reaction when the reaction reaches equilibrium at 25 °C?



— answer —

Recall that when a reaction reaches **equilibrium** $\Delta G = 0$ and $E_{\text{cell}} = 0 \text{ V}$ because $\Delta G = -nFE_{\text{cell}}$.

We can apply this in the Nernst equation to find the equilibrium constant (recall $1 \text{ V} = 1 \text{ J/C}$):

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{RT}{nF} \ln K$$

$$0 = E_{\text{cell}}^{\circ} - \frac{RT}{nF} \ln K$$

$$\ln K = \frac{nFE_{\text{cell}}^{\circ}}{RT}$$

$$\ln K = \frac{(2) \left(96485 \frac{\text{C}}{\text{mol}} \right) \left(1.247 \frac{\text{J}}{\text{C}} \right)}{\left(8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}} \right) (298.15 \text{ K})}$$

$$K = 1.4 \times 10^{42}$$

This should make sense since the E_{cell}° value is very positive, so we expect a large K , which is a spontaneous reaction $\Delta G^{\circ} < 0$.