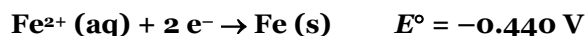


Electrochemistry Quiz

Name: _____ **Key**

May I post your solution? Yes No Yes, but redact my name

Consider a Galvanic/voltaic cell based on the following half reactions under standard conditions.



What will be the cell potential (E_{cell}) when the cathode solution *changes* by 0.827 M?

You may leave your answer in the form of an expression.

To construct a Galvanic/voltaic cell, the overall reaction must be spontaneous: $\Delta G^{\circ} < 0$ and $E^{\circ}_{\text{cell}} > 0 \text{ V}$.

This means that the following reactions must take place:

- Anode oxidation $\text{Fe} (\text{s}) \rightarrow \text{Fe}^{2+} (\text{aq}) + 2 \text{e}^{-}$
- Cathode reduction $\text{Cd}^{2+} (\text{aq}) + 2 \text{e}^{-} \rightarrow \text{Cd} (\text{s})$

The overall (net ionic) equation for the cell is: $\text{Fe} (\text{s}) + \text{Cd}^{2+} (\text{aq}) \rightarrow \text{Fe}^{2+} (\text{aq}) + \text{Cd} (\text{s})$

$$\begin{aligned} E^{\circ}_{\text{cell}} &= E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}} \\ &= -0.403 \text{ V} - (-0.440 \text{ V}) \\ E^{\circ}_{\text{cell}} &= +0.037 \text{ V} \end{aligned}$$

Now we can set up an IC“E” chart to determine how much of each solution will be left after the cell has operated for some time. Note three things though:

- We are starting at standard conditions initially, so $[\text{Fe}^{2+}] = [\text{Cd}^{2+}] = 1.000 \text{ M}$ and $T = 298.15 \text{ K}$.
- The change value is $x = 0.827 \text{ M}$.
- The number of electrons is $n = 2$.

	Fe (s)	+	Cd ²⁺ (aq)	→	Fe ²⁺ (aq)	+	Cd (s)
I	n/a		1.000 M		1.000 M		n/a
C	n/a		- 0.827		+ 0.827		n/a
“E”	n/a		0.173 M		1.827 M		n/a

Now use the Nernst equation to find the new cell potential at these concentrations:

$$\begin{aligned} E_{\text{cell}} &= E^{\circ}_{\text{cell}} - \frac{RT}{nF} \ln \left(\frac{[\text{Fe}^{2+}]}{[\text{Cd}^{2+}]} \right) \\ &= 0.037 \text{ V} - \frac{\left(8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}} \right) (298.15 \text{ K})}{(2 \text{ mol } \text{e}^{-}) \left(96500 \frac{\text{C}}{\text{mol } \text{e}^{-}} \right)} \cdot \ln \left(\frac{1.827}{0.173} \right) \\ E_{\text{cell}} &= 0.007 \text{ V} \end{aligned}$$