Week 10: Discussion

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Announcements

• Please follow instructions in email I sent out on how to download your ELN for your own records.

• Quiz 2 (extra credit question included!) will be released Wednesday, 6/4 at 1pm and will close Sunday, 6/8 at 11pm.
Review Problem #5

You conduct a titration of 15.0 mL of a 0.0460 M AgNO$_3$ solution with a 0.0230 M NaI solution within the following cell:

\[
\text{SCE} \ || \ \text{Titration Solution} \ | \ Ag \ (s)
\]

The SCE reduction potential is 0.241 V vs. SHE, \(K_{sp}\) of AgI is \(8.30 \times 10^{-17}\), and the standard reduction potential for the reaction

\[
Ag^+ + e^- \rightarrow Ag \ (s)
\]

is \(E^0 = 0.800\) V vs. SHE.

a) Write out the balanced titration reaction (including states)

b) What is \([I^-]\) and the cell voltage after the addition of the following volume of NaI solution?

i) 2.00 mL

ii) 30.00 mL

iii) 37.00 mL

\[
Ag^+ (aq) + I^- (aq) \rightarrow AgI (s)
\]

\[
V_{I^-,eq} = \frac{15.0 \text{mL} \times 0.0460 \text{M}}{0.0230 \text{M}} \times \frac{1 \text{ mol } I^-}{1 \text{ mol } Ag^+} = 30.0 \text{mL } I^-
\]

To be continued...
Review Problem #5

You conduct a titration of 15.0 mL of a 0.0460 M AgNO₃ solution with a 0.0230 M NaI solution within the following cell:

\[ \text{SCE} \ | \ | \text{Titration Solution} \ | \ Ag(s) \]

The SCE reduction potential is 0.241 V vs. SHE, \( K_{sp} \) of AgI is \( 8.30 \times 10^{-17} \), and the standard reduction potential for the reaction

\[ \text{Ag}^+ + e^- \rightarrow Ag(s) \]

is \( E^0 = 0.800 \) V vs. SHE.

a) Write out the balanced titration reaction (including states)
b) What is \([I^-]\) and the cell voltage after the addition of the following volume of NaI solution?

i) 2.00 mL  
ii) 30.00 mL  
iii) 37.00 mL

i) Before equiv. point

\[ 0.046 \text{ M Ag}^+ \times 15 \text{ mL} = 0.69 \text{ mmol}; \quad 0.023 \text{ M I}^- \times 2 \text{ mL} = 0.046 \text{ mmol} \]

\[ 0.69 \text{ mmol} - 0.046 \text{ mmol} = 0.644 \text{ mmol excess Ag}^+ \]

\[ 0.644 \text{ mmol} / 17 \text{ mL} = 0.0379 \text{ M Ag}^+ \]

\[ [I^-] = K_{sp} / [Ag^+] = (8.30 \times 10^{-17} / 0.0379) = 2.19 \times 10^{-15} \text{ M} \]

\[ E_{cell} = 0.800 - 0.241 - \frac{0.0592}{1} \times \log \frac{1}{0.0379} \]

\[ E_{cell} = 0.479 \text{ V} \]
Review Problem #5

You conduct a titration of 15.0 mL of a 0.0460 M AgNO₃ solution with a 0.0230 M NaI solution within the following cell:

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SCE || Titration Solution | Ag (s)
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The SCE reduction potential is 0.241 V vs. SHE, $K_{sp}$ of AgI is $8.30 \times 10^{-17}$, and the standard reduction potential for the reaction

$$\text{Ag}^+ + e^- \rightarrow \text{Ag} (s)$$

is $E^0 = 0.800$ V vs. SHE.

a) Write out the balanced titration reaction (including states)

b) What is $[I^-]$ and the cell voltage after the addition of the following volume of NaI solution?

i) 2.00 mL

ii) 30.00 mL

iii) 37.00 mL

Ag⁺ (aq) + I⁻ (aq) → AgI (s)

i) Before equiv. point

$$[\text{Ag}^+] = \frac{30.0mL - 2.00mL}{30.0mL} \times \frac{15.0mL}{15mL + 2.00mL} \times 0.0460M = 0.0379M$$

$$[I^-] = \frac{K_{sp}}{[\text{Ag}^+]} = \frac{8.30 \times 10^{-17}}{0.0379} = 2.19 \times 10^{-15} \text{ M}$$

$$E_{cell} = 0.800 - 0.241 - \frac{0.0592}{1} \times \log \frac{1}{0.0379}$$

$$E_{cell} = 0.479 \text{ V}$$
Review Problem #5

You conduct a titration of 15.0 mL of a 0.0460 M AgNO₃ solution with a 0.0230 M NaI solution within the following cell:

\[ \text{SCE} \mid \mid \text{Titration Solution} \mid \text{Ag (s)} \]

The SCE reduction potential is 0.241 V vs. SHE, \( K_{sp} \) of AgI is \( 8.30 \times 10^{-17} \), and the standard reduction potential for the reaction \( \text{Ag}^+ + \text{e}^- \rightarrow \text{Ag (s)} \)

is \( E^0 = 0.800 \) V vs. SHE.

a) Write out the balanced titration reaction (including states)

b) What is [I⁻] and the cell voltage after the addition of the following volume of NaI solution?

i) 2.00 mL

ii) 30.00 mL

iii) 37.00 mL

\[ \text{Ag}^+ (\text{aq}) + \text{I}^- (\text{aq}) \rightarrow \text{AgI (s)} \]

ii) At equiv. point

\[
[\text{Ag}^+] = [\text{I}^-]
K_{sp} = [\text{Ag}^+][\text{I}^-] = [\text{Ag}^+]^2
\]

\[
[\text{I}^-] = [\text{Ag}^+] = \sqrt{K_{sp}} = \sqrt{8.30 \times 10^{-17}} = 9.1 \times 10^{-9} \text{ M}
\]

\[
E_{cell} = 0.800 - 0.241 - \frac{0.0592}{1} \times \log \frac{1}{9.1 \times 10^{-9}}
\]

\[ E_{cell} = 0.0830 \text{ V} \]
Review Problem #5

You conduct a titration of 15.0 mL of a 0.0460 M AgNO₃ solution with a 0.0230 M NaI solution within the following cell:

\[ \text{SCE} \parallel \text{Titration Solution} \parallel \text{Ag (s)} \]

The SCE reduction potential is 0.241 V vs. SHE, \( K_{sp} \) of AgI is 8.30 * 10⁻¹⁷, and the standard reduction potential for the reaction \( \text{Ag}^+ + \text{e}^- \rightarrow \text{Ag (s)} \)

is \( E^0 = 0.800 \) V vs. SHE.

a) Write out the balanced titration reaction (including states)
b) What is \([I^-]\) and the cell voltage after the addition of the following volume of NaI solution?
   i) 2.00 mL
   ii) 30.00 mL
   iii) 37.00 mL

iii) Past equiv. point

\[
0.046 \text{ M Ag}^+ * 15 \text{ mL} = 0.69 \text{ mmol}; \ 0.023 \text{ M I}^- * 37 \text{ mL} = 0.851 \text{ mmol}
\]

\[
0.851 - 0.69 = 0.161 \text{ mmol excess I}^-
\]

\[
0.161 \text{ mmol} / (15+37 \text{ mL}) = 0.00310 \text{ M I}^- 
\]

\[
K_{sp} = [\text{Ag}^+] [\text{I}^-]
\]

\[
[\text{Ag}^+] = K_{sp} / [\text{I}^-] = 8.30 * 10^{-17} / 0.00310 = 2.68 * 10^{-14} \text{ M}
\]

\[
E_{cell} = 0.800 - 0.241 - 0.0592 \times \log \frac{1}{2.68 * 10^{-14}}
\]

\[
E_{cell} = -0.244 \text{ V}
\]

I^- is in excess
Review Problem #5

You conduct a titration of 15.0 mL of a 0.0460 M AgNO3 solution with a 0.0230 M NaI solution within the following cell:

\[ \text{SCE} \parallel \text{Titration Solution} | \text{Ag (s)} \]

The SCE reduction potential is 0.241 V vs. SHE, \(K_{sp}\) of AgI is \(8.30 \times 10^{-17}\), and the standard reduction potential for the reaction

\[ \text{Ag}^+ + \text{e}^- \rightarrow \text{Ag (s)} \]

is \(E^0 = 0.800\) V vs. SHE.

a) Write out the balanced titration reaction (including states)

b) What is \([I^-]\) and the cell voltage after the addition of the following volume of NaI solution?

i) 2.00 mL

ii) 30.00 mL

iii) 37.00 mL

iii) Past equiv. point

\[
[I^-] = \frac{37.0\text{mL} - 30.0\text{mL}}{15.0\text{mL} + 37.0\text{mL}} \times 0.0230M = 0.00310M
\]

\[
K_{sp} = [\text{Ag}^+][I^-]
\]

\[
[\text{Ag}^+] = K_{sp} / [I^-] = 8.30 \times 10^{-17} / 0.00310 = 2.68 \times 10^{-14} M
\]

\[
E_{cell} = 0.800 - 0.241 - 0.0592 \times \log \frac{1}{2.68 \times 10^{-14}}
\]

\[E_{cell} = -0.244\text{ V}\]
Calculating $[\text{Cl}^-]$ from initial $E_{\text{cell}}$ and known $K_{sp}$

You perform a potentiometric titration of silver nitrate into NaCl. Before you begin titrating, you record the potential of your cell setup to be 0.054 V. Given:

$\begin{align*}
\text{Ag}^+_{(aq)} + \text{e}^- &\rightarrow \text{Ag}_{(s)} & E^0(V) = 0.800 \\
\text{Cu}^{2+}_{(aq)} + 2\text{e}^- &\rightarrow \text{Cu}_{(s)} & E^0(V) = 0.338
\end{align*}$

$K_{sp} = 1.8 \times 10^{-10}$, and $[\text{Cu}^{2+}] = 0.100$ M

what is $[\text{Cl}^-]$ in ppm before you begin titrating?

How can we use the information given to get to $[\text{Cl}^-]$?

i.e. What equation do we know that relates $K_{sp}$ and $[\text{Cl}^-]$?

$K_{sp} = [\text{Ag}^+][\text{Cl}^-]$

Now, how do we solve for $[\text{Ag}^+]$?
Calculating $[\text{Cl}^-]$ from initial $E_{\text{cell}}$ and known $K_{sp}$

You perform a potentiometric titration of silver nitrate into NaCl. Before you begin titrating, you record the potential of your cell setup to be 0.054 V. Given:

$$
\begin{align*}
\text{Ag}^+_{(aq)} + e^- &\rightarrow \text{Ag}(s) \quad E^0(V) = 0.800 \\
\text{Cu}^{2+}_{(aq)} + 2e^- &\rightarrow \text{Cu}(s) \quad E^0(V) = 0.338
\end{align*}
\quad \text{, } K_{sp} = 1.8 \times 10^{-10}, \text{ and } [\text{Cu}^{2+}] = 0.100 \text{ M}
$$

what is $[\text{Cl}^-]$ in ppm before you begin titrating?

We can apply the Nernst equation, even before we begin titrating.

$$
E_{\text{cell}} = \text{initial } E_{\text{cell}} = 0.054 \text{ V} = E^0_{\text{cell}} - \frac{0.0592}{2} \log \left[ \frac{[\text{Cu}^{2+}]}{[\text{Ag}^+]^2} \right]
$$

$$
0.054 = (0.800 - 0.338) - \frac{0.0592}{2} \log \left[ \frac{(0.100)}{[\text{Ag}^+]^2} \right]
$$

$$
\frac{0.0592}{2} \log \left[ \frac{(0.100)}{[\text{Ag}^+]^2} \right] = 0.408
$$
Calculating $[\text{Cl}^-]$ from initial $E_{\text{cell}}$ and known $K_{sp}$

You perform a potentiometric titration of silver nitrate into NaCl. Before you begin titrating, you record the potential of your cell setup to be 0.054 V. Given:

$$\text{Ag}^+(aq) + e^- \rightarrow \text{Ag}(s) \quad E^0(V) = 0.800$$
$$\text{Cu}^{2+}(aq) + 2e^- \rightarrow \text{Cu}(s) \quad E^0(V) = 0.338$$

Given: $K_{sp} = 1.8 \times 10^{-10}$, and $[\text{Cu}^{2+}] = 0.100$ M

what is $[\text{Cl}^-]$ in ppm before you begin titrating?

$$\frac{0.0592}{2} \log \left[ \frac{(0.100)}{[\text{Ag}^+]^2} \right] = 0.408$$

$[\text{Ag}^+] = 4.06 \times 10^{-8}$ M

$$[\text{Cl}^-] = K_{sp} / [\text{Ag}^+]$$

$[\text{Cl}^-] = 1.8 \times 10^{-10} / 4.06 \times 10^{-8} = 4.44 \times 10^{-3}$ M

Convert to ppm: $(4.44 \times 10^{-3}$ mol/L $\text{Cl}^-)(35.45$ g/mol $\text{Cl}^-)(1000$ mg/g)

$= 157$ mg/L $= \text{157 ppm}$