1. The first step to solving selective ion problems is to figure out which solid will precipitate first. We can do this by simply comparing the Ksp values since both compounds ionize in a 1:1 ratio.

\[ \text{BaSO}_4 \quad \text{Ksp} = 2.8 \times 10^{-11} \]
\[ \text{BaSeO}_4 \quad \text{Ksp} = 1.1 \times 10^{-10} \]

BaSO_4 has a lower Ksp value and will precipitate first.

Next, we need to know at what \([\text{Ba}^{2+}]\) will BaSO_4 start precipitating.

\[ 1.1 \times 10^{-10} = [\text{Ba}^{2+}][\text{SO}_4^{2-}] \]

Since \(\text{Na}_2\text{SO}_4\) will dissolve into \(2 \text{Na}^{+}\) and \(\text{SO}_4^{2-}\), we can use the given concentration for \([\text{SO}_4^{2-}]\).

\[ 1.1 \times 10^{-10} = [\text{Ba}^{2+}](1.0 \times 10^{-4} M) \]

\([\text{Ba}^{2+}] = 1.1 \times 10^{-6} M\]

To find out how much \(\text{SeO}_4^{2-}\) we have left in solution, plug into \(\text{BaSeO}_4\)'s Ksp equation.

\[ 2.8 \times 10^{-11} = (1.1 \times 10^{-6} M)[\text{SeO}_4^{2-}] \]

\([\text{SeO}_4^{2-}] = 2.5 \times 10^{-5} M\]

The original \([\text{SeO}_4^{2-}]\) was \(1.0 \times 10^{-4} M\) when \(\text{Na}_2\text{SeO}_4\) was dissociated in solution.

\[ \% \text{SeO}_4^{2-} \text{ Remaining} = \frac{2.5 \times 10^{-5} M}{1.0 \times 10^{-4} M} \times 100\% = 25\% \]

**Meaning 75% of SeO_4^2^- precipitated**

2. Balancing redox reactions can be done with the following steps:
   1) Write out half-reactions
   2) Balance all atoms except O or H.
   3) Balance O by adding H_2O.
4) Balance H by adding \( H^+ \).

5) Balance charges by adding electrons, \( e^- \).

6) Balance the number \( e^- \) by multiplying by lowest common multiple.

7) Add half-reactions together, cancelling \( e^- \) and common terms.

For basic solutions, there is one extra step.

8) Add \( OH^- \) to both sides of the equation, combining all \( H^+ \) and \( OH^- \) to make \( H_2O \).

Balancing:

1) \( Cr(OH)_4^- \rightarrow CrO_4^{2-} \)
   \[ H_2O_2 \rightarrow H_2O \]

2) Already done

3) \( Cr(OH)_4^- \rightarrow CrO_4^{2-} \)
   \[ H_2O_2 \rightarrow 2H_2O \]

5) \( Cr(OH)_4^- \rightarrow CrO_4^{2-} + 4H^+ + 3e^- \)
   \[ H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O \]

6) Multiply top equation by 2 and bottom by 3 (\( LCM = 6 \))

\[ 2Cr(OH)_4^- \rightarrow 2CrO_4^{2-} + 8H^+ + 6e^- \]

\[ 3H_2O_2 + 6H^+ + 6e^- \rightarrow 6H_2O \]

7) \( 2Cr(OH)_4^- + 3H_2O_2 \rightarrow 2CrO_4^{2-} + 6H_2O + 2H^+ \)

8) \( 2Cr(OH)_4^- + 3H_2O_2 + 2OH^- \rightarrow 2CrO_4^{2-} + 6H_2O + 2H^+ + OH^- \)

\[ 2Cr(OH)_4^- + 3H_2O_2 + 2OH^- \rightarrow 2CrO_4^{2-} + 8H_2O \]
\[ 3\text{Cu}(s) \rightarrow 3\text{Cu}^{2+}(aq) + 6e^- \]

* It doesn't matter which half reaction \( H^+ \) and \( H_2O \) belongs to since \( H^+ \) is neither reduced or oxidized.

\[
E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} \\
E^\circ_{\text{cell}} = 0.97V - 0.34V = 0.63V
\]

\[
\Delta G^\circ = -nFE^\circ_{\text{cell}} \\
\Delta G^\circ = -(6)(96500 \frac{J}{e}) (0.63V)
\]

\[
\Delta G^\circ = -364,770 \text{ J} \quad \text{or} \quad -364.8 \text{ kJ}
\]

Ha. Non-standard cell potential can be calculated using the Nernst equation:

\[
E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{RT}{nF} \ln \text{K} \text{eq}
\]

\[
K_{\text{eq}} = \frac{[Zn^{2+}]}{[Cu^{2+}]}
\]

\[
E_{\text{cell}} = 1.10V - \frac{(8.1478 \text{kJ})}{2(96500 \frac{J}{e})} \ln \left( \frac{5.0M}{0.5M} \right)
\]

\[
E_{\text{cell}} = 1.07V
\]

b. For the redox reaction to be nonspontaneous, \( E_{\text{cell}} < 0 \).

\[
1.10V - \frac{(8.1478 \text{kJ})}{2(96500 \frac{J}{e})} \ln \left( \frac{5.0M}{0.5M} \right) < 0
\]

\[
T > 11327K
\]

5a. \( \text{Rate} = K[\text{NO}]^x[\text{O}_2]^y \)

From Exp. 1 to 2, \([\text{NO}]\) remains constant, but \([\text{O}_2]\) doubles. So does the rate. \( \text{O}_2 \) is 1st order → \( y = 1 \)

To solve for \( x \), pick any 2 experiment to divide together. (3 by 2).

\[
\frac{K(4.5 \times 10^{-2} \text{ mol/L})^x (4.5 \times 10^{-2} \text{ mol/L})}{K(9.0 \times 10^{-2} \text{ mol/L})^x (9.0 \times 10^{-2} \text{ mol/L})} = \frac{1.76 \times 10^{-1} \frac{M}{L}}{1.60 \times 10^{-2} \frac{M}{L}}
\]
\[
(2)^2 (2) = 8 \\
2^2 = 4 \\
x = 2
\]

\[ \text{Rate} = \frac{\text{R}[\text{NO}]^2[\text{O}_2]}{\text{dr}/\text{dt}} \]

b. Now we know the reaction orders, we can divide exp. 4 by 3 to solve for the rate.

\[
\frac{R (3.8 \times 10^{-1} \text{m})^2 (4.6 \times 10^{-3} \text{m})}{R (9.0 \times 10^{-2} \text{m})^2 (9.0 \times 10^{-2} \text{m})} = \frac{R}{1.78 \times 10^{-1} \text{m}^2} \\
R = 1.17 \times 10^{-1} \text{m}^2
\]

g. Calculating initial concentrations:

\[
[\text{N}_2]_0 = \frac{1 \text{mol}}{5 \text{L}} = 0.2 \text{ M} \\
[\text{H}_2]_0 = \frac{3 \text{mol}}{5 \text{L}} = 0.6 \text{ M}
\]

\[
[\text{N}_2]_\text{eq} = 0.2 - 0.0783 = 0.16 \text{ M} \\
[\text{H}_2]_\text{eq} = 0.6 - 3(0.0383) = 0.49 \text{ M} \\
[\text{N}_2]_\text{eq} = 2(0.0383) = 0.077 \text{ M}
\]

b. Reaction will shift to the right. An increase in pressure will go towards the side with fewer moles of gas.

c. Reaction will shift to the left. Increasing the volume will decrease pressure. Opposite reason as part b.

d. Reaction will shift to the left. The total volume will have to increase if the pressure is constant. The molar concentrations will decrease. The equilibrium shifts towards the side of more moles of gas to compensate for loss of concentration.
7. When excess Mg(OH)\(_2\) is shaken with a 1.00 L of 1.0 M NH\(_4\)Cl solution, the resulting saturated solution has a pH = 9.00. Calculate the K\(_{sp}\) of Mg(OH)\(_2\): (K\(_{b}\) NH\(_3\)) = 1.76 × 10\(^{-5}\))

\[
\begin{align*}
&\text{[H}_3\text{O}^+]_{eq} = 10^{-\text{pH}} = 10^{-9}\text{M} \\
&\text{[OH}^-]_{eq} = \frac{K_{b}}{\text{[H}_3\text{O}^+]_{eq}} = 10^{-9}\text{M} \\
&K_{a} = \frac{K_{w}}{K_{b}} = 5.68 \times 10^{-10}
\end{align*}
\]

\[
\begin{align*}
&\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{NH}_3 \quad \text{Weak acid} \\
&\text{I} \quad \text{1} \quad 0 \quad 0 \\
&\text{C} \quad -x \quad +x \quad +x \\
&\text{E} \quad 1-x \quad 10^{-9} \quad x
\end{align*}
\]

\[
\begin{align*}
&3.68 \times 10^{-10} = \frac{10^{-9}x}{(1-x)} \\
&x = 3.623
\end{align*}
\]

\[
\begin{align*}
&K_{sp} = [\text{Mg}^{2+}][\text{OH}^-]^2 \\
&K_{sp} = (1.812 \times 10^{-5})^2 \\
&K_{sp} = 1.8 \times 10^{-11}
\end{align*}
\]

8. A 100.0 mL buffer solution is 0.175 M in HClO and 0.150 M in NaClO (K\(_a\)(HClO) = 2.9 × 10\(^{-8}\)).

a) What is the initial pH?

\[
\begin{align*}
\text{pH} &\approx \text{pK}_a + \log \left( \frac{\text{[conjugate base]}}{\text{[acid]}} \right) \\
\text{pH} &\approx -\log (2.9 \times 10^{-8}) + \log (0.15) \\
\text{pH} &\approx 7.47
\end{align*}
\]

b) What is the pH after the addition of 0.0050 mol of HCl?

\[
\begin{align*}
&\text{175 mol HClO} \quad \text{1L} \\
&\text{0.015 mol NaClO} \quad \text{1L} \\
&\text{pH} = -\log \left( 2.9 \times 10^{-8} \right) + \log \left( \frac{0.0175}{0.015} \right) = 7.19
\end{align*}
\]

c) What is the pH after addition of 0.0050 mol of NaOH?

\[
\begin{align*}
&\text{HClO} + \text{OH}^- \rightleftharpoons \text{ClO}^- + \text{H}_2\text{O} \\
&\text{I} \quad 0.0175 \quad 0.015 \\
&\text{C} \quad 0.005 \\
&\text{E} \quad 0.0170 \quad 0.020
\end{align*}
\]

\[
\begin{align*}
&\text{pH} = -\log \left( 2.9 \times 10^{-8} \right) + \log \left( \frac{0.02 \text{mol L}^{-1}}{0.017 \text{mol L}^{-1}} \right) = 7.61
\end{align*}
\]
9. What would each of the solutions do to the solubility of Ca(OH)$_2$ as compared to pure water?

a) A 0.1M solution HCl.

increases, $H^+$ removes OH$^-$ from soln. Le Chatelier shifts to the right.

b) A 0.1M solution NaOH

decreases, rxn shifts left

c) A 0.1M solution of Ca(NO$_3$)$_2$

decreases, rxn shifts left

d) A solution buffer at pH 10.

decrease, pH of 10 is basic $\rightarrow$ [OH$^-$] > [H$^+$], rxn shifts left

e) A solution of NaCl

no change: does not share a common ion.

10. Determine whether the following solutions will form a precipitate in a AgBr solution given that $K_{sp} = 5.3 \times 10^{-12}$.

a) a 10 mL solution of $1.00 \times 10^{-5}$ M NaBr and a 10 mL solution of $1.00 \times 10^{-5}$ M Ag(NO$_3$)$_3$ are mixed.

\[
AgBr(s) \rightleftharpoons Ag^+ + Br^- \\
 mol = [Ag^+] L = 10^{-5} \times 0.01 = 10^{-7} \text{mol Ag}^+ \text{L}^{-1}
\]

\[
[Ag^+]_{\text{new}} = \frac{10^{-7} \text{mol}}{0.02 \text{L}} = 5 \times 10^{-6} \text{M}
\]

\[
Q = [Ag^+]_{\text{new}} [Br^-]_{\text{new}} = 2.5 \times 10^{-11}
\]

$Q > K_{sp}$, precipitate forms

b) a 10 mL solution of $1.00 \times 10^{-7}$ M NaBr and a 10 mL solution of $1.00 \times 10^{-7}$ M Ag(NO$_3$)$_3$ are mixed.

\[
 mol = [Ag^+] L = 10^{-7} \times 0.01 = 10^{-9} \text{mol Ag}^+ \\
[Ag^+]_{\text{new}} = \frac{10^{-9} \text{mol}}{0.02 \text{L}} = 5 \times 10^{-8} \text{M}
\]

\[
[Ag^+]_{\text{new}} [Br^-]_{\text{new}} = 2.5 \times 10^{-15}
\]

$Q < K_{sp}$, precipitate does not form
11. The proposed mechanism for the reaction in this question is shown below:

\[
\begin{align*}
\text{Step 1: } & \quad 2\text{NO}(g) \rightarrow \text{N}_2\text{O}_2(g) \quad \text{(fast)} \\
\text{Step 2: } & \quad \text{N}_2\text{O}_2(g) + \text{H}_2 \rightarrow \text{H}_2\text{O}_2 (g) + \text{N}_2(g) \quad \text{(slow)} \\
\text{Step 3: } & \quad \text{H}_2(g) + \text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} \quad \text{(fast)} \\
2\text{NO} & + 2\text{H}_2 \rightarrow \text{N}_2 + 2\text{H}_2\text{O}
\end{align*}
\]

(a) Write a balanced equation for this reaction:

\[
2\text{NO} + 2\text{H}_2 \rightarrow \text{N}_2 + 2\text{H}_2\text{O}
\]

(b) Identify the rate determining step.

(c) Write a rate law that is most consistent with this mechanism.

\[
\text{rate} = k_1[H_2][N_2O_2] \\
\text{rate} = \frac{k_1[H_2]}{k_2[N_2O_2]^2} \\
\text{rate} = k[H_2][N_2O_2]^2
\]

(d) Based on your rate law, determine the effect of doubling the concentration of NO (while maintaining constant amounts of all other reactants) on the reaction rate.

reaction rate increases by 4

12. The hydrolysis of sucrose \((C_{12}H_{22}O_{11})\) into glucose and fructose in acidic water has a rate constant of \(1.8 \times 10^{-4} \text{ s}^{-1}\) at \(25^\circ\text{C}\). Assuming the reaction is first order in sucrose, determine the mass of sucrose that is hydrolyzed when 2.55 L of a 0.150 M sucrose solution is allowed to react for 195 minutes.

\[
\ln[\text{C}_{12}\text{H}_{22}\text{O}_{11}]_t = -kt + \ln[\text{C}_{12}\text{H}_{22}\text{O}_{11}]_0 \\
\ln[\text{C}_{12}\text{H}_{22}\text{O}_{11}]_t = -1.8 \times 10^{-4}(17700s) + \ln[0.156 \text{M}] \\
[\text{C}_{12}\text{H}_{22}\text{O}_{11}]_t = 0.01826 \text{M}
\]

\[
[\text{C}_{12}\text{H}_{22}\text{O}_{11}]_{\text{dissolved}} = 0.15 - 0.01826 = 0.13174 \text{M}
\]

\[
\text{mol C}_{12}\text{H}_{22}\text{O}_{11} = 0.336 \text{ mol C}_{12}\text{H}_{22}\text{O}_{11}
\]

\[
\frac{0.336 \text{ mol C}_{12}\text{H}_{22}\text{O}_{11} \times 342.3 \text{ g}}{1 \text{ mol C}_{12}\text{H}_{22}\text{O}_{11}} = 115 \text{ g C}_{12}\text{H}_{22}\text{O}_{11}
\]