Dr. Steward's CHM152 Exam #2 Review Spring 2014 (Ch. 16. 17) KEY

1. For each of the following acid / base reactions, determine if each of the following will have more products or reactants ($K_a$ HSO$_4^-$ = $1.3 \times 10^{-2}$, $K_b$ NH$_3$ = $1.8 \times 10^{-5}$, $K_a$ HF = $6.8 \times 10^{-4}$):

- NH$_4^+$ (aq) + F$^-$ (aq) → NH$_3$ (aq) + HF (aq)
  More Products      More reactants

- OH$^-$ (aq) + HNO$_3$ (aq) → H$_2$O (l) + NO$_3^-$ (aq)
  More Products      More reactants

- HSO$_4^-$ (aq) + NH$_3$ (aq) → NH$_4^+$ (aq) + SO$_4^{2-}$ (aq)
  More Products      More reactants

- 2H$_2$O (l) → H$_3$O$^+$ (aq) + OH$^-$ (aq)
  More Products      More reactants

2. For the acid/base reactions determine the conjugate acid/base pairs for each.

- NH$_4^+$ / NH$_3$ (acid / base), F$^-$ / HF (base / acid)
- OH$^-$ / H$_2$O (base / acid), HNO$_3$ / NO$_3^-$ (acid / base)
- HSO$_4^-$ / SO$_4^{2-}$ (acid / base), NH$_3$ / NH$_4^+$ (base / acid)
- H$_2$O / H$_3$O$^+$ (base / acid), H$_2$O / OH$^-$ (acid / base)

3. Explain how NH$_3$ is both a Bronsted/Lowry base and a Lewis base.
   It is a Bronsted/Lowry base because it is able to accept protons. It is a Lewis base because it is able to donate electrons.

4. Determine the pH, pOH, [H$_3$O$^+$], and [OH$^-$] for the following solutions:

- 0.0051 M HBr
  pH = -log(0.0051) = 2.29, pOH = 14.00 − 2.29 = 11.71
  [H$^+$] = $10^{-2.29} = 0.0051$ M, [OH$^-$] = $10^{-11.71} = 1.9 \times 10^{-12}$ M

- 0.010 M NaOH
  pOH = -log(0.010) = 2.00, pH = 14.00 − 2.00 = 12.00
  [OH$^-$] = $10^{-2.00} = 0.10$ M, [H$^+$] = $10^{-12.00} = 1.0 \times 10^{-12}$ M

- 0.010 M Sr(OH)$_2$
  pOH = -log(0.010 x 2) = 1.70, pH = 14.00 − 1.70 = 12.30
  [OH$^-$] = $10^{-1.70} = 0.020$ M, [H$^+$] = $10^{-12.30} = 5.0 \times 10^{-13}$ M

5. Determine whether the following salts are acidic, basic, or neutral:

- LiCl     neutral
- Al(NO$_3$)$_3$ acidic
- BaF$_2$ basic
- NH$_4$I acidic

6. What is the pH of a 0.100 M formic acid solution with a $K_a$ = $1.8 \times 10^{-4}$?
HCOOH + H₂O ⇌ H₃O⁺ + HCOO⁻

\[
\begin{array}{c|c|c|c}
\text{HCOOH} & \text{H₂O} & \text{HCOO}^- & \text{H₃O}^+ \\
0.100M & 0M & 0M & \\
+x & -x & -x & 0.100 + x
\end{array}
\]

\[
pH = -\log[H₃O^+] = -\log(4.24 \times 10^{-3}) = 2.37
\]

\[
x = 4.24 \times 10^{-3}
\]

\[
\frac{4.24 \times 10^{-3}}{0.100} \times 100 = 4.2\% \text{OK assumption}
\]

7. What is the percent dissociation? 4.2%

8. If the concentration of the formic acid was 4.0M instead, explain how the pH, [OH⁻], and % dissociation would be different (higher/lower/same).
   The pH would be lower (more acidic/higher [H₃O⁺]), [OH⁻] would be lower, % dissociation would be lower (high initial concentration = lower % dissociation for same acid or base).

9. Calculate the pH if 2.56 grams of sodium fluoride is dissolved in 4.50 liters of water. Kₐ for hydrofluoric acid is 7.1 x 10⁻⁴

\[
\begin{align*}
2.56g \text{NaF} & \quad \frac{1\text{mol} \text{NaF}}{41.99g \text{NaF}} = 6.097 \times 10^{-2}\text{mol} \text{NaF} \\
6.097 \times 10^{-2}\text{mol} \text{NaF} & \quad \frac{6.097 \times 10^{-2}\text{mol} \text{NaF}}{4.50L} = 1.353 \times 10^{-2}\text{M} \text{NaF}
\end{align*}
\]

\[
\begin{array}{c|c|c|c}
\text{F}^- & \text{H₂O} & \text{OH}^- & \text{HF} \\
1.353 \times 10^{-2}\text{M} & 0M & 0M & \\
-x & +x & +x & 1.353 \times 10^{-2} - x
\end{array}
\]

\[
\begin{align*}
K_a \cdot K_b = K_w &= 1.0 \times 10^{-14} \\
\frac{1.0 \times 10^{-14}}{7.1 \times 10^{-4}} &= 1.41 \times 10^{-11} \\
x^2 = 1.41 \times 10^{-11} & \quad \sqrt{x^2} = \sqrt{1.41 \times 10^{-11}} \quad x = 4.37 \times 10^{-7}
\end{align*}
\]

10. What is the initial concentration of benzoic acid (C₆H₅COOH) in a solution with pH = 2.59 (assume the benzoic acid was added to the water with no other ions initially)? The acid dissociation constant for this monoprotic acid is 6.5 x 10⁻⁵.

\[
\begin{align*}
\text{HA} & \quad \text{H₂O} \\
? (y) & \quad 0M & 0M \\
-x & \quad +x & +x \\
y - 2.57 \times 10^{-3} & \quad 2.57 \times 10^{-3} & 2.57 \times 10^{-3}
\end{align*}
\]

\[
\begin{align*}
[H_3O^+] &= 10^{-pH} = 10^{-2.59} = 2.57 \times 10^{-3} \\
\frac{(2.57 \times 10^{-3})^2}{y - 2.57 \times 10^{-3}} &= 6.5 \times 10^{-5} \\
y &= 0.104M
\end{align*}
\]

0.104M - 2.57 \times 10^{-3}M = 0.101M or 0.10M
11. Explain the common-ion effect. Weak acid or weak base’s % dissociation will decrease if they are placed in solutions with one of the products of dissociation. For example, The % dissociation of HCN will be decrease if additional CN⁻ is added to the solution (as NaCN, for example).

$$HCN(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CN(aq)$$

12. Calculate the percent ionization of a 1.00M solution of HF (pKₐ = 3.17).

$$HF + H_2O \rightleftharpoons H_3O^+ + F^-$$

\[ \frac{(x)(x)}{1.00} = 6.76 \times 10^{-4} \]

\[ x = \sqrt{6.76 \times 10^{-4}} = 2.6 \times 10^{-2} \]

\[ \frac{2.6 \times 10^{-2}}{1.00} \times 100 = 2.6\% \]

13. Calculate the percent ionization if there is also a 0.50M concentration of NaF in solution.

$$HF + H_2O \rightleftharpoons H_3O^+ + F^-$$

\[ \frac{(x)(0.500)}{1.00} = 6.76 \times 10^{-4} \]

\[ x = \frac{1.35 \times 10^{-3}}{1.00} \times 100 = 0.14\% \]

14. Calculate the percent ionization if the pH of the solution before the HF was added was 0.30.

$$HF + H_2O \rightleftharpoons H_3O^+ + F^-$$

\[ [H_3O^+] = 10^{-0.30} = 0.50M \]

15. You are titrating 10.00mL of 1.00M of HCl with 2.00M NaOH.

- What would be the pH at ½ the equivalence point? NaOH + HCl → H₂O + NaCl

\[ \text{vol of NaOH@eq.} = 0.10000 \text{L HCl} \]

\[ \frac{100 \text{ mol HCl}}{1 \text{ L HCl}} = \frac{1 \text{ mol NaOH}}{1 \text{ mol HCl}} \]

\[ \frac{1 \text{ L NaOH}}{2 \text{ mol NaOH}} = 0.00500 \text{ L NaOH} \]

\[ \text{vol of NaOH@eq.} = 0.00500 \text{ L NaOH} \]

\[ \frac{1}{2} \text{ eq. NaOH} = 0.00250 \text{ L NaOH} \]

\[ \text{mol of NaOH@eq.} = \frac{2.00 \text{ mol}}{1 \text{ L NaOH}} \]

\[ \text{mol of NaOH} = 0.00500 \text{ mol NaOH} \]

\[ \text{HCl} + \text{NaOH} \rightleftharpoons \text{H}_2\text{O} + \text{NaCl} \]

\[ \text{B} \rightarrow 0.0100 \text{ mol HCl} \]

\[ \Delta \rightarrow -0.00500 \text{ mol HCl} \]

\[ \text{A} \rightarrow 0.00500 \text{ mol} \]

\[ \text{pH} = -\log[H^+] = -\log\left(\frac{\text{mol HCl}}{\text{total volume @ \(\frac{1}{2}\) eq}}\right) = -\log\left(\frac{0.00500 \text{ mol}}{0.00250 \text{ L} + 0.0100 \text{ L}}\right) = -\log(0.40) = 0.40 \]

This makes sense since only half of the acid should be left at ½ eq. If you realize this, you do not have to use a change table.
• What would be the pH at the equivalence point?

\[ \text{vol of } \text{NaOH} @ \text{eq.} = \frac{0.01000L \text{ HCl}}{1L \text{ HCl}} \times 1 \text{ mol NaOH} \times \frac{1 \text{ L NaOH}}{2.00 \text{ mol NaOH}} = 0.00500L \text{ NaOH} \]

mol of NaOH @ eq. = (0.00500L) \times \left( \frac{2.00 \text{ mol NaOH}}{1L} \right) = 0.01000mol \text{ NaOH}

\[ \begin{align*}
\text{H}^+ &+ \text{OH}^- \rightleftharpoons \text{H}_2\text{O} \\
\text{B} &\quad 0.0100 \text{ mol} \quad 0.0100 \text{ mol} \\
\Delta &\quad -0.0100 \text{ mol} \quad -0.0100 \text{ mol} \\
\text{A} &\quad 0 \quad 0
\end{align*} \]

There will only be a neutral salt at the equivalence point of a Strong Acid/Strong Base titration.

\[ \text{pH} = 7 \]

• What would be the pH after adding 0.75mL of NaOH after the equivalence point?

Total mol of NaOH = (0.00575L) \times \left( \frac{2.00 \text{ mol NaOH}}{1L} \right) = 0.01150mol \text{ NaOH}

\[ \begin{align*}
\text{H}^+ &+ \text{OH}^- \rightleftharpoons \text{H}_2\text{O} \\
\text{B} &\quad 0.0100 \text{ mol} \quad 0.0115 \text{ mol} \\
\Delta &\quad -0.0100 \text{ mol} \quad -0.0100 \text{ mol} \\
\text{A} &\quad 0 \quad 0.0015 \text{ mol}
\end{align*} \]

After the equivalence point, you are just adding extra OH\(^-\) ions to the solution.

\[ \text{pH} = 14 - \log \left( \frac{0.0015 \text{ mol}}{0.01000L + 0.00500L + 0.00075L} \right) = 12.98 \]

16. You are titrating 10.00mL of 1.00M of benzoic acid (\(K_a\) is 6.4x10\(^{-5}\)) with 2.00M NaOH.

What would be the pH after 2.50 mL has been added?

\[ \begin{align*}
\text{HA} &+ \text{OH}^- \rightleftharpoons \text{H}_2\text{O} + \text{A}^- \\
\text{B} &\quad 0.0100 \text{ mol} \quad 0.00500 \text{ mol} \quad 0 \\
\Delta &\quad -0.00500 \text{ mol} \quad -0.00500 \text{ mol} \quad +0.00500 \text{ mol} \\
\text{A} &\quad 0.00500 \text{ mol} \quad 0 \quad +0.00500 \text{ mol}
\end{align*} \]

\[ \text{pH} = \text{pK}_a + \log \left( \frac{\text{[A}^-]}{\text{[HA]}} \right) = 4.19 + \log \left( \frac{0.00500 \text{ mol}}{0.00500 \text{ mol}} \right) = 4.19 \]
• What would be the pH after adding a total of 5.00 mL NaOH?

\[
\text{mol of acid (HA) at start} = \text{mol of conjugate base (A\textsuperscript{-}) at eq.}
\]

\[
\text{mol of HA} = (0.01000L \times 1.00mol) / 1L = 0.01000mol HA \text{ at start} = \text{mol of A\textsuperscript{-} at eq}
\]

\[
\text{vol of NaOH at eq.} = 0.01000L \times 6x10^5 = 0.00500L \text{NaOH}
\]

\[
[A] = \frac{\text{mol of A\textsuperscript{-} at eq.}}{\text{total volume at eq.}} = \frac{0.01000mol}{0.01000L + 0.00500L} = 0.6667M
\]

\[
A\textsuperscript{-} + H_2O \rightleftharpoons OH\textsuperscript{-} + HA
\]

\[
K_b = \frac{K_w}{K_a} = \frac{1.56 \times 10^{-10}}{0.6667} = 1.56 \times 10^{-10}
\]

\[
(x)(x) = \sqrt{(1.56 \times 10^{-10})(0.6667)} = 1.0 \times 10^{-5}
\]

\[
\text{pH} = 14 - \log[OH\textsuperscript{-}] = 14 - \log(1.0 \times 10^{-5}) = 9.01
\]

• What would be the pH after adding a total of 5.75 mL NaOH?

\[
\begin{align*}
\text{HA} & \quad + \quad \text{OH}\textsuperscript{-} \rightleftharpoons \text{H}_2\text{O} \quad + \quad \text{A}\textsuperscript{-} \\
\text{B} & \quad 0.01000mol \quad 0.0115mol \quad 0 \\
\Delta & \quad -0.01000mol \quad -0.01000mol \quad +0.01000mol \\
\text{A} & \quad 0 \quad 0.0015mol \quad +0.01000mol \\
\end{align*}
\]

\[
\text{Total mol of NaOH} = (0.00575L \times 2.00mol) / 1L = 0.0115mol NaOH
\]

\[
\text{pH} = 14 - \log\left(\frac{0.0115mol}{0.01000L + 0.00500L + 0.00750L}\right) = 12.98
\]

17. What is the pH of 100.0mL solution containing 0.30 M NH\textsubscript{3} and 0.30 M NH\textsubscript{4}NO\textsubscript{3} (K\textsubscript{b} of NH\textsubscript{3} is 1.8x10\textsuperscript{-5})?

\[
pK_b = -\log K_b = -\log(1.8 \times 10^{-5}) = 4.745
\]

\[
pH = pK_a + \log \frac{[A\textsuperscript{-}]}{[HA]} = 9.255 + \log \frac{0.30M}{0.30M} + 9.26
\]

Since this has the components of a buffer solution, we can use Henderson-Hasselbalch

• What is the pH after adding 1.00mL of 2.00M NaOH

\[
\begin{align*}
\text{NaOH} & \quad + \quad \text{NH}_4\textsuperscript{+} \rightarrow \text{NH}_3 \\
\text{B} & \quad 0.00020mol \quad 0.030mol \quad 0.030mol \\
\Delta & \quad -0.00020mol \quad -0.00200mol \quad +0.00200mol \\
\text{A} & \quad 0 \quad 0.028mol \quad 0.032mol \\
\end{align*}
\]

\[
\text{mol of NaOH added} = 0.00100L \times 2.00mol = 0.00200mol of NaOH
\]

\[
\text{pH} = pK_a + \log \frac{[A\textsuperscript{-}]}{[HA]} = 9.225 + \log \frac{0.032mol}{0.028mol} = 9.28
\]
• What is the pH after adding 2.00mL of 1.00M HCl?

\[
\text{HCl} + \text{NH}_3 \rightarrow \text{NH}_4^+ + \text{H}^+
\]

B 0.00200mol 0.030mol 0.030mol
\Delta -0.00200mol -0.00200mol +0.00200mol
A 0 0.028mol 0.032mol

\[\text{mol of HCl added} = 0.00200 \text{L} \times \frac{1.00 \text{ mol}}{1 \text{ L}} = 0.00200 \text{ mol of HCl}\]

\[\text{pH} = \text{pK}_a + \log \left( \frac{[A^-]}{[HA]} \right) = 9.225 + \log \left( \frac{0.028\text{ mol}}{0.032\text{ mol}} \right) = 9.17\]

18. You have 60.00mL of a buffer solution consisting of 0.900M CH$_3$COOH and 1.10M NaCH$_3$COO. $K_b = 5.6 \times 10^{-10}$

• Determine the pH of the solution after adding 10.11mL of 2.23M HBr.

\[
\text{HBr} + \text{CH}_3\text{COO}^- \rightarrow \text{CH}_3\text{COOH} + \text{Br}^-
\]

B 0.0225mol 0.0660mol 0.0540mol
\Delta -0.0225mol -0.0225mol +0.0225mol
A 0 0.0435mol 0.0860mol

\[K_a = \frac{1.0 \times 10^{-14}}{5.6 \times 10^{-10}} = 1.8 \times 10^{-5}\]

\[\text{pH} = \text{pK}_a + \log \left( \frac{[A^-]}{[HA]} \right) = 4.74 + \log \left( \frac{0.0435}{0.0860} \right) = 4.44\]

• How much total HBr would you have to add until the buffer was used up?

\[
\text{HBr} + \text{CH}_3\text{COO}^- \rightarrow \text{CH}_3\text{COOH} + \text{Br}^-
\]

B 0.0660mol 0.0660mol 0.0540mol
\Delta -0.0660mol -0.0660mol +0.0660mol
A 0 0mol 0.1200mol

\[\text{0.06000L CH}_3\text{COOH} \times \frac{1.10\text{ mol CH}_3\text{COO}^-}{1 \text{ L CH}_3\text{COOH}} \times \frac{1 \text{ mol HBr}}{1 \text{ mol CH}_3\text{COO}^-} \times \frac{1 \text{ L HBr}}{2.23 \text{ mol HBr}} = 0.0296 \text{ L}\]

• What would be the pH when at the point (that you calculated above?)

\[
\text{HBr} + \text{CH}_3\text{COO}^- \rightarrow \text{CH}_3\text{COOH} + \text{Br}^-
\]

B 0.0660mol 0.0660mol 0.0540mol
\Delta -0.0660mol -0.0660mol +0.0660mol
A 0 0.1200mol

\[\frac{0.1200\text{ mol CH}_3\text{COOH}}{(0.06000\text{ L} + 0.0296\text{ L})} = 1.34 \text{ M}\]

\[
\text{CH}_3\text{COOH} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{CH}_3\text{COO}^-
\]

I 1.34M 0 0
C -x +x +x
E 1.34M-x x x \frac{(x)(x)}{1.34} = 1.8 \times 10^{-5}

\[x = \sqrt{(1.8 \times 10^{-5})(1.34)} = 4.9 \times 10^{-3}\]

\[\text{pH} = -\log(4.9 \times 10^{-3}) = 2.31\]
19. If you start over with fresh buffer (same conditions/volumes as original), what would be the pH of the solution after adding 20.00mL of 2.23M KOH?

\[
\begin{align*}
0.06000\text{L} \text{CH}_3\text{COOH} & \quad 0.900\text{mol} \text{CH}_3\text{COOH} \quad \frac{0.900\text{mol} \text{CH}_3\text{COOH}}{1\text{L} \text{CH}_3\text{COOH}} = 0.0540\text{mol} \text{CH}_3\text{COOH} \\
0.06000\text{L} \text{CH}_3\text{COO}^- & \quad 1.10\text{mol} \text{CH}_3\text{COO}^- \quad \frac{1.10\text{mol} \text{CH}_3\text{COO}^-}{1\text{L} \text{CH}_3\text{COO}^-} = 0.0660\text{mol} \text{CH}_3\text{COO}^- \\
0.02000\text{L} \text{NaOH} & \quad 2.23\text{mol NaOH} \quad \frac{2.23\text{mol NaOH}}{1\text{L} \text{NaOH}} = 0.0446\text{mol NaOH} \\
\end{align*}
\]

<table>
<thead>
<tr>
<th>NaOH</th>
<th>CH(_3)COOH</th>
<th>CH(_3)COO(^-)</th>
<th>Na(^+)</th>
</tr>
</thead>
<tbody>
<tr>
<td>B</td>
<td>0.0446mol</td>
<td>0.0660mol</td>
<td>0.0540mol</td>
</tr>
<tr>
<td>(\Delta)</td>
<td>-0.0446mol</td>
<td>-0.0446mol</td>
<td>+0.0446mol</td>
</tr>
<tr>
<td>A</td>
<td>0</td>
<td>0.0214mol</td>
<td>0.0986mol</td>
</tr>
</tbody>
</table>

\[
\text{pH} = \text{pK}_a + \log \frac{[A^-]}{[HA]} = 4.74 + \log \frac{0.0986}{0.0214} = 5.40
\]

- After adding 30.00mL of 2.23 KOH?

\[
\begin{align*}
0.03000\text{L} \text{NaOH} & \quad 2.23\text{mol NaOH} \quad \frac{2.23\text{mol NaOH}}{1\text{L} \text{NaOH}} = 0.0669\text{mol NaOH} \\
\end{align*}
\]

<table>
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<tr>
<td>B</td>
<td>0.0660mol</td>
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<td>0.0540mol</td>
</tr>
<tr>
<td>(\Delta)</td>
<td>-0.0660mol</td>
<td>-0.0660mol</td>
<td>+0.0660mol</td>
</tr>
<tr>
<td>A</td>
<td>0.0009 mol</td>
<td>0 mol</td>
<td>0.1200mol</td>
</tr>
</tbody>
</table>

\[
p\text{OH} = -\log(\text{OH}^-) = -\log(0.01) = 2.0 \\
P\text{H} = 14.0 - 2.0 = 12.0
\]

20. If you wanted to make a buffer with a pH of 3.90 using the same buffer as before with a total concentration of 2.00M, what would the concentration of each part of the buffer need to be?

\[
p\text{H} = \text{pK}_a + \log \frac{[A^-]}{[HA]} = 3.90 = 4.74 + \log \frac{x}{(2.0-x)} \\
-0.84 = \log \frac{x}{(2.0-x)} \\
10^{-0.84} = 10^{\left(\frac{-0.84}{(2.0-x)}\right)}
\]

\[
0.145 = \frac{x}{(2.0-x)} \\
0.29 - 0.15x = x \\
0.29 = 1.15x \\
0.29 = 1.15 \times \frac{0.25}{0.25} [A^-] = x = 1.75 \text{M}
\]

[HA] = 2.00M - 0.25M = 1.75M
21. If 0.0056 moles of AgNO₃ is added to 1.00L of 0.0020M NaCl, will AgCl precipitate form? (K_{sp} AgCl = 1.8x10^{-10}). Explain.

AgCl(s) ⇌ Ag^+(aq) + Cl^-(aq)

Q_{sp} = (0.0056)(0.0020) = 1.1x10^{-6} since Q > K will shift to the left (solid AgCl), so it will precipitate!

What would the minimum concentration of NaCl need to be to form a precipitate?

K_{sp} = [Ag][Cl] = (0.0056)(0.0020) = 1.1x10^{-6} [Cl^-] = \frac{3.2x10^{-7}}{x} M

22. Ni(OH)₂ is an insoluble compound in water. How would the molar solubility be affected in each of the following situations?

- NaOH is added – Decrease solubility (common ion)
- NaCl is added – No effect on solubility. Does not react with Ni²⁺ or OH⁻
- HCl is added – Increase solubility. H⁺ reacts with OH⁻.

23. Assume you have a saturated solution of nickel (II) phosphate with some solid in the container (K_{sp} = 4.74 x 10^{-32}). What is its molar solubility? Determine the concentration of the Ni²⁺ and PO₄³⁻ ions in the solution.

\[ \text{Ni₃(PO₄)₂} \rightleftharpoons 3\text{Ni}^{2+} + 2\text{PO}_4^{3-} \]

\[ K_{sp} = [\text{Ni}^{2+}]^3[\text{PO}_4^{3-}]^2 = 4.74 \times 10^{-32} = (3x)^3(2x)^2 = (27x^3)(4x^2) \]

\[ I \quad 0 \quad 0 \]
\[ C \quad +3x \quad +2x \quad 4.74 \times 10^{-32} = 108x^5 \quad x = 2.13 \times 10^{-7} \quad x = \text{molar solubility} = 2.13 \times 10^{-7} \text{M} \]

\[ [\text{Ni}^{2+}] = 3(2.13 \times 10^{-7}) = 6.39 \times 10^{-7} \text{M} \quad [\text{PO}_4^{3-}] = 2(2.13 \times 10^{-7}) = 4.26 \times 10^{-7} \text{M} \]

24. What would it’s molar solubility be if nickel (II) phosphate if 0.100 moles of Na₂PO₄ was added to 1.00L of the solution above? What would be the concentrations of the Ni²⁺ and PO₄³⁻ ions?

\[ \text{Ni₃(PO₄)₂} + \text{H}_2\text{O} \rightleftharpoons 3\text{Ni}^{2+} + 2\text{PO}_4^{3-} \]

\[ K_{sp} = [\text{Ni}^{2+}]^3[\text{PO}_4^{3-}]^2 = 4.74 \times 10^{-32} = (3x)^3(0.100)^2 = \]

\[ I \quad 0 \quad 0 \quad 0.100 \]
\[ C \quad +3x \quad +2x \quad 4.74 \times 10^{-32} = 0.270x^3 \quad x = 5.60 \times 10^{-11} \]
\[ E \quad 3x \quad 0.100+2x \quad x = \text{molar solubility} = 5.60 \times 10^{-11} \text{M} \]

\[ [\text{Ni}^{2+}] = 3(5.60 \times 10^{-11}) = 1.68 \times 10^{-10} \text{M} \quad [\text{PO}_4^{3-}] = 0.100+(2)(5.60 \times 10^{-11}) \approx 0.100 \text{M} \]
25. How much (in g) barium sulfate (BaSO$_4$, MW=233.39) will dissolve in 500. mL of water. For BaSO$_4$: $K_{sp} = 1.1 \times 10^{-10}$

\[
\text{Ba}_2 \text{SO}_4 + \text{H}_2\text{O} \rightleftharpoons \text{Ba}^{2+} + \text{SO}_4^{2-} \\
K_{sp} = [\text{Ba}^{2+}][\text{SO}_4^{2-}] = 1.1 \times 10^{-10} = x^2
\]

<table>
<thead>
<tr>
<th>I</th>
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<tbody>
<tr>
<td>C</td>
<td>+x</td>
<td>x</td>
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<tr>
<td>E</td>
<td>x</td>
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</table>

\[x = 1.04 \times 10^{-5} \text{ molar solubility} = 1.04 \times 10^{-5} \text{M}\]

\[
0.500 \text{L H}_2\text{O} \quad \frac{1.04 \times 10^{-5} \text{mol BaSO}_4}{1 \text{L H}_2\text{O}} \quad \frac{233.39 \text{g BaSO}_4}{1 \text{mol BaSO}_4} = 1.2 \times 10^{-3} \text{ g BaSO}_4
\]