20. a) \[2\text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{OH}^-(aq)\]

\[K = [\text{H}_3\text{O}^+][\text{OH}^-]\]

Or

\[\text{H}_2\text{O}(l) \rightleftharpoons \text{H}^+(aq) + \text{OH}^-(aq)\]

\[K = [\text{H}^+][\text{OH}^-]\]

b) \[\text{HCN}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{CN}^-(aq)\]

\[K = \frac{[\text{H}_3\text{O}^+][\text{CN}^-]}{[\text{HCN}]}\]

Or

\[\text{HCN}(aq) \rightleftharpoons \text{H}^+(aq) + \text{CN}^-(aq)\]

\[K = \frac{[\text{H}^+][\text{CN}^-]}{[\text{HCN}]}\]

c) \[\text{NH}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^-(aq)\]

\[K = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}\]

For this question you could have picked any weak acid or base or just used a general weak acid (HA) and weak base (B)

24. a) \text{HClO}_4 \quad \text{strong acid}

b) \text{HOCI} \quad \text{weak acid}

c) \text{H}_2\text{SO}_4 \quad \text{strong acid (1\textsuperscript{st} deprotonation only)}

d) \text{H}_2\text{SO}_3 \quad \text{weak acid}

27. The larger the \(K_a\) the stronger the acid (table 7.2).

\text{HClO}_4 > \text{HClO}_2 > \text{NH}_4^+ > \text{H}_2\text{O}

Strongest acid \(\rightarrow\) Weakest acid

Note: \text{HClO}_4 is a strong acid, therefore, will have the highest \(K_a\) value.

\(\text{H}_2\text{O}\) is the conjugate acid of \text{OH}^-. \text{OH}^- is a strong base. Strong bases have the weakest conjugate acids.

\((\text{H}_2\text{O}(l) + \text{OH}^-(aq) \rightarrow \text{OH}^-(aq) + \text{H}_2\text{O}(l))\)

28. The larger the \(K_a\) the stronger the acid (table 7.2), the stronger the acid the weaker the conjugate base, therefore, the strongest base will have the conjugate acid with the smallest \(K_a\).

\text{NH}_3 > \text{ClO}_4^- > \text{H}_2\text{O} > \text{ClO}_4^-

Note: \text{HClO}_4 is a strong acid, therefore, will have the highest \(K_a\) value.

Acids are classified as strong acids if they have a \(K_a\) greater than 1. \text{H}_3\text{O}^+ is one of the weakest strong acids. Therefore, \text{H}_3\text{O}^+ (\text{H}_2\text{O} conjugate base) is one of the weakest strong acid and \text{H}_2\text{O} will be slightly more basic than \text{ClO}_4^-.
29. The larger the $K_a$ the stronger the acid.
   a) $\text{HCl}$ strong acid and $\text{H}_2\text{O}$ is a weak acid $K_w = K_a (1.0 \times 10^{-14})$
   b) $\text{HNO}_2$ The $K_s$ of $\text{HNO}_2 (4.0 \times 10^{-4}) > K_w = K_a (1.0 \times 10^{-14})$
   c) $\text{HCN}$ The $K_s$ $\text{HCN} (6.2 \times 10^{-10}) > K_a$ of $\text{HOC}_6\text{H}_5 (1.6 \times 10^{-10})$

30. The higher the $K_a$ the weaker the conjugate base because $K_p = \frac{K_w}{K_a}$
   a) $\text{Cl}^-$ conjugate base of a strong acid, $\text{HCl}$
   H$_2$O conjugate base of one of the weakest strong acids H$_3$O$^+$
   H$_2$O stronger base
   b) $\text{H}_2\text{O}$ conjugate base of one of the weakest strong acids H$_3$O$^+$
   NO$_2^-$ conjugate base of a weak acid
   NO$_2^-$ stronger base
   c) $\text{CN}^-$ conjugate base of the weak acid HCN ($K_a = 6.2 \times 10^{-10}$)
   OOC$_6$H$_5^-$ conjugate base of the weak acid HOC$_6$H$_5$ ($K_a = 1.6 \times 10^{-10}$)
   OOC$_6$H$_5^-$ stronger base

31. a) $\text{CH}_3\text{CO}_2\text{H}$ acid
   H$_2$O base
   $\text{CH}_3\text{CO}_2^-$ conjugate base
   H$_3$O$^+$ conjugate acid
   Therefore, H$_2$O and $\text{CH}_3\text{CO}_2^-$ are competing for protons.
   b) $\text{CH}_3\text{CO}_2^-$ is the stronger base (H$_3$O$^+$ is a strong acid, therefore, forms a weaker
   conjugate acid than $\text{CH}_3\text{CO}_2\text{H}$)
   c) $\text{CH}_3\text{CO}_2^-$ is a weak base because it is the conjugate base of a weak acid.
   Conjugate bases of weak acid are also weak, therefore, the do no fully dissociate in solution.
   $\text{CH}_3\text{CO}_2^- + \text{H}_2\text{O} (aq) \rightleftharpoons \text{CH}_3\text{CO}_2\text{H} (aq) + \text{OH} (aq)$

33. a) weak acid b) strong acid c) weak base
d) strong base e) weak base f) weak acid
g) weak acid h) strong base i) strong acid (1st deprotonation)

40. The solution is acidic if $[\text{H}^+] > [\text{OH}^-]$, the solution is basic if $[\text{H}^+] < [\text{OH}^-]$, and the solution is neutral if $[\text{H}^+] = [\text{OH}^-]$.
The relationship between $[\text{H}^+]$ and $[\text{OH}^-]$ is: $K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$
   a) $(1.0 \times 10^{-7})[\text{OH}^-] = 1.0 \times 10^{-14}$
      $[\text{OH}^-] = 1.0 \times 10^{-7} M$
      $[\text{H}^+] = [\text{OH}^-]$ neutral solution
   b) $(8.3 \times 10^{-16})[\text{OH}^-] = 1.0 \times 10^{-14}$
      $[\text{OH}^-] = 12. M$
      $[\text{H}^+] < [\text{OH}^-]$ basic solution
   c) $(12)[\text{OH}^-] = 1.0 \times 10^{-14}$
      $[\text{OH}^-] = 8.3 \times 10^{-16} M$
      $[\text{H}^+] > [\text{OH}^-]$ acidic solution
d) $(5.4 \times 10^{-5})[\text{OH}^-] = 1.0 \times 10^{-14}$
      $[\text{OH}^-] = 1.9 \times 10^{-10} M$
      $[\text{H}^+] > [\text{OH}^-]$ acidic solution
41. If the pH < 7 the solution is acidic, if the pH > 7 the solution is basic, and if the pH = 7 the solution is neutral.

If the pOH < 7 the solution is basic, if the pOH > 7 the solution is acidic, and if the pOH = 7 the solution is neutral.

The relationship between pH and [H$^+$] is: \( p\text{H} = -\log[H^+] \)

The relationship between [H$^+$] and [OH$^-$] is: \( K_w = [H^+][OH^-] = 1.0 \times 10^{-14} \)

a) Basic
\[
[H^+] = 10^{-pH} = 10^{-7.40} = 4.0 \times 10^{-8} \text{ M}
\]
\( (4.0 \times 10^{-8})[OH^-] = 1.0 \times 10^{-14} \)
\( [OH^-] = 2.5 \times 10^{-7} \text{ M} \)

b) Basic
\[
[H^+] = 10^{-pH} = 10^{-15.3} = 5 \times 10^{-16} \text{ M}
\]
\( (5 \times 10^{-16})[OH^-] = 1.0 \times 10^{-14} \)
\( [OH^-] = 20 \text{ M} \)

c) Acidic
\[
[H^+] = 10^{-pH} = 10^{-1.0} = 10 \text{ M}
\]
\( (10)[OH^-] = 1.0 \times 10^{-14} \)
\( [OH^-] = 1 \times 10^{-15} \text{ M} \)

d) Acidic
\[
[H^+] = 10^{-pH} = 10^{-3.20} = 6.3 \times 10^{-4} \text{ M}
\]
\( (6.3 \times 10^{-4})[OH^-] = 1.0 \times 10^{-14} \)
\( [OH^-] = 1.6 \times 10^{-11} \text{ M} \)

e) Basic
\[
[OH^-] = 10^{-pOH} = 10^{-5.0} = 1 \times 10^{-5} \text{ M}
\]
\( [H^+](1 \times 10^{-5}) = 1.0 \times 10^{-14} \)
\( [H^+] = 1 \times 10^{-9} \text{ M} \)

f) Acidic
\[
[OH^-] = 10^{-pOH} = 10^{-9.60} = 2.5 \times 10^{-10} \text{ M}
\]
\( [H^+](1.5 \times 10^{-10}) = 1.0 \times 10^{-14} \)
\( [H^+] = 4.0 \times 10^{-5} \text{ M} \)

47. a) HNO$_2$ is a weak acid

The major species in solution are HNO$_2$

The following equilibrium will occur in solution

HNO$_2$(aq) $\Rightarrow$ NO$_2^-$ (aq) + H$^+$(aq) \quad \text{K}_a = 4.0 \times 10^{-4} \text{ (Table 7.2)}

Calculate the concentration of H$^+$ ions at equilibrium

<table>
<thead>
<tr>
<th></th>
<th>HNO$_2$</th>
<th>NO$_2^-$</th>
<th>H$^+$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.250 M</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-x</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.250-x</td>
<td>X</td>
<td>X</td>
</tr>
</tbody>
</table>

\[
\text{K}_a = \frac{[NO_2^-][H^+]}{[HNO_2]} = \frac{xx}{(0.250-x)} = 4.0 \times 10^{-4}
\]

Since K$_a$ is small assume 0.250 $-x = 0.250$
\[
\frac{xx}{(0.250)} = 4.0 \times 10^{-4}
\]
\[
x = 0.010 \text{ M}
\]
Check assumption (5% rule)
\[
\frac{\text{Amount Gained/Lost}}{\text{Original Concentration}} \times 100\% = \frac{0.010}{0.250} \times 100\% = 4\% \quad \text{Good}
\]
Concentration of H⁺
\[
[H^+] = x = 0.010 \, M
\]
Calculate pH
\[
pH = -\log[H^+] = -\log(0.010) = 2.00
\]
b)
HC₃H₅O₂ is a weak acid
The major species in solution HC₃H₅O₂
The following equilibrium will occur in solution
HC₃H₅O₂(aq) ⇌ C₂H₃O₂⁻(aq) + H⁺(aq) \quad K_a = 1.8 \times 10^{-5} \, \text{(Table 7.2)}

Calculate the concentration of H⁺ ions at equilibrium

<table>
<thead>
<tr>
<th></th>
<th>HC₃H₅O₂</th>
<th>C₂H₃O₂⁻</th>
<th>H⁺</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.250 M</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-x</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.250-x</td>
<td>x</td>
<td>x</td>
</tr>
</tbody>
</table>

\[
K_a = \frac{[C_2H_3O_2^-][H^+]}{[HC_3H_5O_2]} = \frac{xx}{(0.250-x)} = 1.8 \times 10^{-5}
\]
Since \( K_a \) is small assume 0.250-x = 0.250
\[
1.8 \times 10^{-5} = \frac{xx}{(0.250)}
\]
\[
x = 0.0021
\]
Check assumption (5% rule)
\[
\frac{\text{Amount Gained/Lost}}{\text{Original Concentration}} \times 100\% = \frac{0.0021}{0.25} \times 100\% = 0.84\% \quad \text{Good}
\]
Concentration of H⁺
\[
[H^+] = x = 0.0021 \, M
\]
Calculate pH
\[
pH = -\log[H^+] = -\log(0.0021) = 2.68
\]

48. a)
HOC₆H₅ is a weak acid therefore
The major species in solution are HOC₆H₅
The following equilibrium will occur in solution
HOC₆H₅(aq) ⇌ OC₆H₅⁻(aq) + H⁺(aq) \quad K_a = 1.6 \times 10^{-10}

Calculate the concentrations of H⁺ at equilibrium

<table>
<thead>
<tr>
<th></th>
<th>HOC₆H₅</th>
<th>OC₆H₅⁻</th>
<th>H⁺</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.250 M</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-x</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.250-x</td>
<td>x</td>
<td>x</td>
</tr>
</tbody>
</table>

\[
K_a = \frac{[OC_6H_5^-][H^+]}{[HOC_6H_5]} = \frac{xx}{(0.250-x)} = 1.6 \times 10^{-10}
\]
Since \( K_a \) is small assume 0.250-x = 0.250
\[
1.6 \times 10^{-10} = \frac{xx}{(0.250)}
\]
\[
x = 6.3 \times 10^{-6}
\]
Check assumption (5% rule)
\[
\frac{\text{Amount Gained/Lost}}{\text{Original Concentration}} \times 100\% = \frac{6.3 \times 10^{-6}}{0.25} \times 100\% = 0.0025\% \quad \text{Good}
\]
Concentration H⁺

\[ [H^+] = x = 6.3 \times 10^{-6} \, M \]

Calculate pH

\[ pH = -\log[H^+] = -\log(6.3 \times 10^{-6}) = 5.20 \]

b) HCN is a weak acid

The major species in solution are HCN

The following equilibrium will occur in solution

\[ HCN(aq) \rightleftharpoons CN^-(aq) + H^+(aq) \quad K_a = 6.2 \times 10^{-10} \quad \text{(Table 7.2)} \]

Calculate the concentration of H⁺ ion at equilibrium

\[
\begin{array}{c|ccc}
 & HCN & CN^- & H^+ \\
\hline
\text{Initial} & 0.250 \text{ M} & 0 & 0 \\
\text{Change} & -x & +x & +x \\
\text{Equilibrium} & 0.250-x & x & x \\
\end{array}
\]

\[ K_a = \frac{[CN^-][H^+]}{[HCN]} = \frac{xx}{(0.250 - x)} = 6.2 \times 10^{-10} \]

Since \( K_a \) is small assume \( 0.250 - x = 0.250 \)

\[ 6.2 \times 10^{-10} = \frac{xx}{(0.250)} \]

\[ x = 1.2 \times 10^{-5} \]

Check assumption (5% rule)

\[ \frac{\text{Amount Gained/Lost}}{\text{Original Concentration}} \times 100\% = \frac{0.25 \times 10^{-5}}{0.25} = 0.0048\% \]

Concentration H⁺

\[ [H^+] = x = 1.2 \times 10^{-5} \, M \]

Calculate pH

\[ pH = -\log[H^+] = -\log(1.2 \times 10^{-5}) = 4.92 \]

49. HF is a weak acid therefore, the reactions that are occurring in solution are:

\[ HF(aq) \rightleftharpoons F^-(aq) + H^+(aq) \quad K_a = 7.2 \times 10^{-4} \quad \text{(Table 7.2)} \]

Calculate the concentrations of HF, F⁻, and H⁺ at equilibrium (you will also need to calculate the concentration of OH⁻ which will be in the solution from the following equilibrium \( H_2O(l) \rightleftharpoons OH^-(aq) + H^+(aq) \))

\[
\begin{array}{c|ccc}
 & HF & F^- & H^+ \\
\hline
\text{Initial} & 0.020 \text{ M} & 0 & 0 \\
\text{Change} & -x & +x & +x \\
\text{Equilibrium} & 0.020-x & x & x \\
\end{array}
\]

\[ K_a = \frac{[F^-][H^+]}{[HF]} = \frac{xx}{(0.250 - x)} = 7.2 \times 10^{-4} \]

\[ 1.44 \times 10^{-5} - 7.2 \times 10^{-4}x = x^2 \]

\[ x^2 + 7.2 \times 10^{-4}x - 1.44 \times 10^{-5} = 0 \]

\[ x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} = \frac{-7.2 \times 10^{-4} \pm \sqrt{(7.2 \times 10^{-4})^2 - 5.76 \times 10^{-5}}}{2} \]

\[ x = 0.0035 \, \text{and} \, -0.0042 \]

x is 0.035 M because concentration cannot be negative.

Find the final concentrations

\[ [HF] = 0.20 - x = 0.20 - 0.0035 = 0.17 \, M \]

\[ [F^-] = [H^+] = x = 0.0035 \, M \]
51. a) HA is a weak acid. There are more HA molecules (green and white atoms together) than H\(^+\) ions (single white atom) and A\(^-\) ions (single green atoms). This shows that the acid did not completely dissociate making it a weak acid.

b) Calculate the percentage dissociated

\[
\frac{\text{particles dissociated}}{\text{total particles}} \times 100\% = \frac{1}{1 + 9} \times 100\% = 10\%
\]

Calculate the equilibrium constant

The reaction of interest is HA(aq) \(\rightleftharpoons\) H\(^+\)(aq) + A\(^-\)(aq)

<table>
<thead>
<tr>
<th></th>
<th>HA</th>
<th>H(^+)</th>
<th>A(^-)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.20 M</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-x</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>(0.20 - x = \frac{9}{N_aV})</td>
<td>(x = \frac{1}{N_aV})</td>
<td>(x = \frac{1}{N_aV})</td>
</tr>
</tbody>
</table>

Note from the picture at equilibrium the concentration of HA is \(\frac{9}{N_aV}\) = \(\frac{9}{N_aV}\) \(N_a\) is Avogadro's number and is used to convert the 9 particles to number of moles of HA. The concentration of H\(^+\) and A\(^-\) are \(\frac{1}{N_aV}\)

Solve for V

\[
\begin{align*}
0.20 - x &= \frac{9}{N_aV} \\
0.20 - \frac{1}{N_aV} &= \frac{9}{N_aV} \\
V &= \frac{N_a}{\frac{9}{N_aV}} = \frac{10}{50}N_a \\
V &= \frac{N_a0.20}{N_a} = \frac{9}{N_a}
\end{align*}
\]

Solve for K

\[
K = \frac{[H^+][A^-]}{[HA]} = \frac{\frac{1}{N_a(\frac{10}{N_a})}}{\frac{9}{N_a(\frac{50}{N_a})}} = \frac{1}{(50)(9)} = 0.0022
\]

52. \(\text{HC}_2\text{H}_2\text{ClO}_2\) is a weak acid, therefore, the equilibrium established is:

\[
\text{HC}_2\text{H}_2\text{ClO}_2(aq) \rightleftharpoons \text{C}_2\text{H}_2\text{ClO}_2^-(aq) + \text{H}^+(aq) \\
K_a = 1.35 \times 10^{-3}
\]

Find the concentration of H\(^+\) at equilibrium

<table>
<thead>
<tr>
<th></th>
<th>HC(_2)H(_2)ClO(_2)</th>
<th>C(_2)H(_2)ClO(_2)(^-)</th>
<th>H(^+)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.10 M</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-x</td>
<td>x</td>
<td>x</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.10-x</td>
<td>x</td>
<td>x</td>
</tr>
</tbody>
</table>
\[ K_a = \frac{[C_2H_2ClO_2^-][H^+]}{[HClO_2]} = \frac{xx}{(0.10 - x)} = 1.35 \times 10^{-3} \]
\[ 1.4 \times 10^{-4} - 1.35 \times 10^{-3}x = x^2 \]
\[ x^2 + 1.35 \times 10^{-3}x - 1.4 \times 10^{-4} = 0 \]
\[ x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} = \frac{-1.35 \times 10^{-3} \pm \sqrt{(1.35 \times 10^{-3})^2 - 4 \times 10^{-4}}}{2} \]
\[ = 0.011 \text{ and } 0.013 \]
\[ x \text{ is 0.011 M because concentration cannot be negative.} \]
\[ [H^+] = x = 0.011 \text{ M} \]

Calculate pH
\[ pH = -\log[H^+] = -\log(0.011) = 1.96 \]

53. HIO$_3$ is a weak acid, therefore, the equilibrium established is:
\[ \text{HIO}_3(\text{aq}) = \text{IO}_3(\text{aq}) + \text{H}^+(\text{aq}) \quad K_a = 0.17 \]

Find the concentration of H$^+$ at equilibrium

<table>
<thead>
<tr>
<th></th>
<th>HIO$_3$</th>
<th>IO$_3$</th>
<th>H$^+$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.10 M</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-x</td>
<td>X</td>
<td>x</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.010-x</td>
<td>x</td>
<td>x</td>
</tr>
</tbody>
</table>

\[ K_a = \frac{[IO_3^-][H^+]}{[HIO_3]} = \frac{xx}{(0.10 - x)} = 0.17 \]
\[ 0.0017 - 0.17x = x^2 \]
\[ x^2 + 0.17x - 0.0017 = 0 \]
\[ x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} = \frac{-0.17 \pm \sqrt{(0.17)^2 - 0.0068}}{2} = 0.0094 \text{ and } 0.018 \]
\[ x \text{ is 0.0094 M because concentration cannot be negative.} \]
\[ [H^+] = x = 0.0094 \text{ M} \]

Calculate pH
\[ pH = -\log[H^+] = -\log(0.0094) = 2.03 \]

55. 0.56 g C$_6$H$_5$CO$_2$H \( \left( \frac{1 \text{ mol } C_6H_5CO_2H}{122.13 \text{ g } C_6H_5CO_2H} \right) = 0.0046 \text{ mol } C_6H_5CO_2H \)
\[ M = \frac{n}{V} = \frac{0.0046 \text{ mol}}{1.0 \text{ L}} = 0.0046 \text{ M} \]

C$_6$H$_5$CO$_2$H is a weak acid, therefore, the equilibrium established is:
\[ C_6H_5CO_2H(\text{aq}) = C_6H_5CO_2^- (\text{aq}) + \text{H}^+(\text{aq}) \quad K_a = 6.4 \times 10^{-5} \]

Calculate the equilibrium concentrations

<table>
<thead>
<tr>
<th></th>
<th>C$_6$H$_5$CO$_2$H</th>
<th>C$_6$H$_5$CO$_2^-$</th>
<th>H$^+$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.0046 M</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-x</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.0046-x</td>
<td>X</td>
<td>X</td>
</tr>
</tbody>
</table>

\[ K_a = \frac{[C_6H_5CO_2^-][H^+]}{[C_6H_5CO_2]} = \frac{xx}{(0.0046 - x)} = 6.4 \times 10^{-5} \]
\[ 2.9 \times 10^{-7} - 6.4 \times 10^{-5}x = x^2 \]
\[ x^2 + 6.4 \times 10^{-5}x - 2.9 \times 10^{-7} = 0 \]
\[ x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} = \frac{-6.4 \times 10^{-5} \pm \sqrt{(6.4 \times 10^{-5})^2 - 1.2 \times 10^{-6}}}{5.2 \times 10^{-4} \text{ and } - 5.8 \times 10^{-4}} \]
x is 5.2x10^{-4} M because concentration cannot be negative.

Find the final concentrations
\[
\begin{align*}
[HC_6H_5CO_2] &= 0.0046 - x = 0.0046 - 5.2 \times 10^{-4} = 0.0041 M \\
[C_6H_5CO_2^-] &= [H^+] = x = 5.2 \times 10^{-4} M 
\end{align*}
\]
Calculate the concentration of OH⁻ from water equilibrium
\[
K_w = [H^+][OH^-] = 1.0 \times 10^{-14}
\]
\[
(5.2 \times 10^{-4} M)[OH^-] = 1.0 \times 10^{-14}
\]
\[
[OH^-] = 1.9 \times 10^{-11} M 
\]
Calculate pH
\[
pH = -\log[H^+] = -\log(5.2 \times 10^{-4}) = 3.28
\]

56. C₆H₅CO₂H is a solid before it is put into water. Therefore, two equilibriums are established.

\[
\begin{align*}
C_6H_5CO_2H(s) & \rightleftharpoons C_6H_5CO_2H(aq) \\
C_6H_5CO_2H(aq) & \rightleftharpoons C_6H_5CO_2^- + H^+(aq) \\
K_a &= 6.4 \times 10^{-5}
\end{align*}
\]

They want to know how much C₆H₅CO₂H will dissolve. This will be the amount of C₆H₅CO₂H that starts in the second equilibrium which we will denote as s.

<table>
<thead>
<tr>
<th></th>
<th>C₆H₅CO₂H</th>
<th>C₆H₅CO₂⁻</th>
<th>H⁺</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>s</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-x</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>s-x</td>
<td>x</td>
<td>x</td>
</tr>
</tbody>
</table>

The pH of the solution will give us the equilibrium concentration of the H⁺ ions
\[
pH = -\log[H^+] = 10^{-pH} = 10^{-2.80} = 0.0016 M
\]

\[
\begin{align*}
[\text{Initial}] & = C_6H_5CO_2H \\
[\text{Change}] & = -0.0016 +0.0016 +0.0016 \\
[\text{Equilibrium}] & = s-0.0016 0.0016 0.0016 \\
K_a & = \frac{[C_6H_5CO_2^-][H^+]}{[C_6H_5CO_2H]} = \frac{(0.0016)(0.0016)}{(s-0.0016)} = 6.4 \times 10^{-5} \\
S & = 0.042 M \\
\]

Convert to grams per 100 ml
\[
S = \left( \frac{0.042~mol}{L} \right) \left( \frac{1 L}{1000 mL} \right) \left( \frac{122.13 g}{mol} \right) \left( \frac{C_6H_5CO_2H}{1 mol C_6H_5CO_2H} \right) = 0.0051 \frac{g}{mL}
\]

Therefore the water solubility per 100 ml is 0.51 g per 100 ml

61. a) The first solution contains 0.10 M HCl (strong acid) and 0.10 M HOCl (weak acid). Since there is both a strong and weak acid with equal concentrations of each the number of H⁺ ions from the weak acid will be minimal, therefore, the pH will be completely from the strong acid.

\[
pH = -\log[H^+] = -\log(0.10) = 1.00
\]

This can be proven using an ICE table

<table>
<thead>
<tr>
<th></th>
<th>HOCl</th>
<th>OCl⁻</th>
<th>H⁺</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Change</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Equilibrium</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
In 0.10 M  \\
Change  -x  +x  +x  \\
Equilibrium  0.10-x  x  0.10+x

\[
K_a = \frac{[OCl^-][H^+]}{[HOCl]} = \frac{x(0.10 + x)}{(0.10 - x)} = 3.5 \times 10^{-8}
\]

Since \( K_a \) is small assume 0.10-x = 0.10 and 0.10+x=0.10
\[
x(0.10) = 3.5 \times 10^{-8}
\]
\[
x = 3.5 \times 10^{-8}
\]

Check assumption (5% rule)
\[
\frac{\text{Amount Gained/Lost}}{\text{Original Concentration}} \times 100\% = \frac{3.5 \times 10^{-8}}{0.10} \times 100\% = 3.5 \times 10^{-5}\%
\]
Good

Concentration H^+
\[
[H^+] = 0.10 - x = 0.10 - 3.5 \times 10^{-8} = 0.10 M
\]

Calculate pH
\[
pH = -\log[H^+] = -\log(0.10) = 1.00
\]

b) The second solution contains 0.050 M HNO_3 (strong acid) and 0.50 M HC_2H_3O_2 (weak acid). Since there is both a strong and weak acid and even though the concentration of the weak acid is 10x more than the concentration of the strong acid the number of H^+ ions from the weak acid will be minimal, therefore, the pH will be completely from the strong acid.

\[
pH = -\log[H^+] = -\log 0.050 = 1.30
\]

This can be proven using an ICE table
\[
\text{HC}_2\text{H}_3\text{O}_2(aq) \rightleftharpoons \text{H}^+(aq) + \text{C}_2\text{H}_3\text{O}_2^-(aq) \quad K_a = 1.8 \times 10^{-5}
\]

<table>
<thead>
<tr>
<th>Initial</th>
<th>0.50 M</th>
<th>0</th>
<th>0.050</th>
</tr>
</thead>
<tbody>
<tr>
<td>Change</td>
<td>-x</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.50-x</td>
<td>x</td>
<td>0.050+x</td>
</tr>
</tbody>
</table>

\[
K_a = \frac{[C_2H_3O_2^-][H^+]}{[HC_2H_3O_2]} = \frac{x(0.050 + x)}{(0.50 - x)} = 1.8 \times 10^{-5}
\]

Since \( K_a \) is small assume 0.50-x = 0.50 and 0.050+x=0.050
\[
x(0.050) = 1.8 \times 10^{-5}
\]
\[
x = 1.8 \times 10^{-5}
\]

Check assumption (5% rule) Note:0.050<0.50 therefore is the assumption works for 0.050-x then it will work form 0.50-x
\[
\frac{\text{Amount Gained/Lost}}{\text{Original Concentration}} \times 100\% = \frac{1.8 \times 10^{-4}}{0.050} \times 100\% = 0.36\% \quad \text{Good}
\]

Concentration H^+
\[
[H^+] = 0.050 + x = 0.050 + 1.8 \times 10^{-4} = 0.050 M
\]

Calculate pH
\[
pH = -\log[H^+] = -\log(0.050) = 1.30
\]
65. HOCN is a weak acid, therefore, the following equilibrium is established:

\[ \text{HOCN(aq)} \rightleftharpoons \text{OCN}^- (aq) + \text{H}^+ (aq) \]

<table>
<thead>
<tr>
<th></th>
<th>HOCN</th>
<th>OCN^-</th>
<th>H^+</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>1.00×10^{-2} M</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-x</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>1.00×10^{-2}-x</td>
<td>x</td>
<td>x</td>
</tr>
</tbody>
</table>

The pH of the solution at equilibrium will allow us to calculate the concentration of \( \text{H}^+ \) ions at equilibrium:

\[
pH = -\log[\text{H}^+]\]

\[
[\text{H}^+] = 10^{-pH} = 10^{-2.77} = 0.0017 \text{ M}
\]

Updating the ICE table gives:

<table>
<thead>
<tr>
<th></th>
<th>HOCN</th>
<th>OCN^-</th>
<th>H^+</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>1.00×10^{-2} M</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-0.0017</td>
<td>+0.0017</td>
<td>+0.0017</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.0083</td>
<td>0.0017</td>
<td>0.0017</td>
</tr>
</tbody>
</table>

\[
K_a = \frac{[\text{OCN}^-][\text{H}^+]}{[\text{HOCN}]} = \frac{(0.0017)(0.0017)}{0.0083} = 3.5 \times 10^{-4}
\]

69. The larger the \( K_b \) the stronger the base. Therefore the strongest base will have the largest \( K_b \).

\[
\text{NH}_3 > \text{C}_5\text{H}_5\text{N} > \text{H}_2\text{O} > \text{NO}_3^-
\]

Strongest base → Weakest base

Note: HNO\textsubscript{3} is a strong acid. The stronger the acid the weaker the conjugate base. Therefore, NO\textsubscript{3}\textsuperscript{-} is a very weak base.

Note: H\textsubscript{2}O is the conjugate base of (H\textsubscript{3}O\textsuperscript{+}). H\textsubscript{2}O\textsuperscript{+} is one of the weakest strong acids. Therefore, H\textsubscript{2}O is only slightly more basic than NO\textsubscript{3}\textsuperscript{-}

70. The smaller the \( K_b \) the stronger the conjugate acid.

\[
\text{HNO}_3 > \text{C}_5\text{H}_5\text{NH}^+ > \text{NH}_4^+ > \text{H}_2\text{O}
\]

Strongest acid → Weakest acid

Note: HNO\textsubscript{3} is a strong acid.

Note: C\textsubscript{5}H\textsubscript{5}N and NH\textsubscript{3} are both weak bases, therefore, will have weak conjugate acids. The \( K_b \) of NH\textsubscript{3} > \( K_b \) of C\textsubscript{5}H\textsubscript{5}N resulting in C\textsubscript{5}H\textsubscript{5}NH\textsuperscript{+} being a stronger acid than NH\textsubscript{4}\textsuperscript{+}.

Note: H\textsubscript{2}O is the conjugate acid of OH\textsuperscript{-}. OH\textsuperscript{-} is a strong base. Strong bases have the weakest conjugate acids.

(\text{H}_2\text{O(l)} + \text{OH}^-(aq) \rightarrow \text{OH}^-(aq) + \text{H}_2\text{O(l)})

74. a) Ca(OH)\textsubscript{2} is a strong base (fully dissociates)

\[
[\text{OH}^-] = 0.00040 M \text{ Ca(OH)}_2 \left( \frac{2 \text{ mol OH}^-}{1 \text{ mol Ca(OH)}_2} \right) = 0.00080 M \text{ OH}^-
\]

\[
pOH = -\log[\text{OH}^-] = -\log(0.00080) = 3.10
\]

\[
pH + pOH = 14.00
\]

\[
pH = 14.00 - pOH = 14.00 - 3.10 = 10.90
\]

b) KOH is a strong base (fully dissociates)

\[
25 \text{ g KOH} \left( \frac{1 \text{ mol KOH}}{56.11 \text{ g KOH}} \right) \left( \frac{1 \text{ mol OH}^-}{1 \text{ mol KOH}} \right) = 0.45 \text{ mol OH}^-
\]

\[
[\text{OH}^-] = \frac{n}{V} = \frac{0.45 \text{ mol}}{1.00 \text{ L}} = 0.45 \text{ M}
\]
\[ pOH = -\log[OH^-] = -\log(0.45) = 0.35 \]
\[ pH + pOH = 14.00 \]
\[ pH = 14.00 - pOH = 14.00 - 0.35 = 13.65 \]

(c) NaOH is a strong base (fully dissociates)
\[
150.0 \text{ g NaOH} \left( \frac{1 \text{ mol NaOH}}{40.00 \text{ g NaOH}} \right) \left( \frac{1 \text{ mol } OH^-}{1 \text{ mol NaOH}} \right) = 3.750 \text{ mol } OH^- \\
[OH^-] = \frac{n}{V} = \frac{3.750 \text{ mol}}{1.000 \text{ L}} = 3.750 \text{ M} \\
pOH = -\log[OH^-] = -\log(3.750) = -0.5740 \\
pH + pOH = 14.0000 \\
pH = 14.0000 - pOH = 14.0000 - (-0.5740) = 14.5740
\]

75. Ba(OH)_2 is a strong base (fully dissociates)

Calculate the concentration of OH^-
\[
14.00 = pH + pOH \\
pOH = 14.00 - pH = 14.00 - 10.50 = 3.50 \\
pOH = -\log[OH^-] \\
[OH^-] = 10^{-pOH} = 10^{-3.50} = 3.2 \times 10^{-4}
\]
Calculate the concentration of Ba(OH)_2
\[
3.2 \times 10^{-4} \text{ M} \left( \frac{1 \text{ mol Ba(OH)}_2}{1 \text{ mol } OH^-} \right) = 1.6 \times 10^{-4} \text{ M } \text{Ba(OH)}_2
\]

79. The presence of nitrogen most often results in a basic molecule. Nitrogen has an unpaired electron, therefore, bonds easily to H^+ ions.

101. The larger the pH the more basic a solution is

a) HI strong acid
   HF weak acid
   NaF (weak base)
   \[
   \text{NaF(s)} \rightarrow \text{Na}^+(aq) + \text{F}^-(aq) \\
   \text{Na}^+ + \text{H}_2\text{O} \rightleftharpoons \text{NaOH} + \text{H}^+ \text{ (NaOH is a strong base)} \\
   \text{Na}^+ \text{ is neutral} \\
   \text{F}^- + \text{H}_2\text{O} \rightleftharpoons \text{HF} + \text{OH}^- \\
   \text{F}^- \text{ is basic}
   
   \text{NaI} \text{ (neutral)}
   \[
   \text{NaI(s)} \rightarrow \text{Na}^+(aq) + \text{I}^-(aq) \\
   \text{Na}^+ + \text{H}_2\text{O} \rightleftharpoons \text{NaOH} + \text{H}^+ \text{ (NaOH is a strong base)} \\
   \text{Na}^+ \text{ is neutral} \\
   \text{I}^- + \text{H}_2\text{O} \rightleftharpoons \text{HI} + \text{OH}^- \text{ (HI is a strong acid)} \\
   \text{I}^- \text{ is neutral}
   
   \text{HI} < \text{HF} < \text{NaI} < \text{NaF} \\
   \rightarrow \text{Increasing pH}
   
   b) \text{NH}_4\text{Br} \text{ (weak acid)}
   \[
   \text{NH}_4\text{Br(s)} \rightarrow \text{NH}_4^+(aq) + \text{Br}^-(aq) \\
   \text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_2\text{O}^+ \\
   \text{NH}_4^+ \text{ is acidic} \\
   \text{Br}^- + \text{H}_2\text{O} \rightleftharpoons \text{HBr} + \text{OH}^- \text{ (HBr is a strong acid)} \\
   \text{Br}^- \text{ is neutral}
HBr strong acid
KBr (neutral)
\[
\text{KBr(s) } \rightarrow \text{ K}^+ (aq) + \text{Br}^- (aq)
\]
\[
\text{K}^+ + \text{H}_2\text{O} \quad \text{KOH} + \text{H}^+ (\text{KOH is a strong base})
\]
K\(^+\) is neutral
\[
\text{Br}^- + \text{H}_2\text{O} \quad \text{HBr} + \text{OH}^- (\text{HBr is a strong acid})
\]
Br\(^-\) is neutral

NH\(_3\) weak base
HBr < NH\(_4\)Br < KBr < NH\(_3\)
Increasing pH
c) \[
\text{C}_6\text{N}_5\text{NH}_3\text{NO}_3 \text{(weak acid)}
\]
\[
\text{C}_6\text{N}_5\text{NH}_3\text{NO}_3(s) \rightarrow \text{C}_6\text{N}_5\text{NH}_3^+(aq) + \text{NO}_3^-(aq)
\]
\[
\text{C}_6\text{N}_5\text{NH}_3^+ + \text{H}_2\text{O} \quad \text{C}_6\text{N}_5\text{NH}_2 + \text{H}_3\text{O}^+
\]
\[
\text{C}_6\text{N}_5\text{NH}_3^+ \text{ is acidic}
\]
\[
\text{NO}_3^- + \text{H}_2\text{O} \quad \text{HNO}_3 + \text{OH}^- (\text{HNO}_3 \text{ is a strong acid})
\]
NO\(_3^-\) is neutral

NaNO\(_3\) (neutral)
\[
\text{NaNO}_3(s) \rightarrow \text{Na}^+(aq) + \text{NO}_3^-(aq)
\]
\[
\text{Na}^+ + \text{H}_2\text{O} \quad \text{NaOH} + \text{H}^+ (\text{NaOH is a strong base})
\]
Na\(^+\) is neutral
\[
\text{NO}_3^- + \text{H}_2\text{O} \quad \text{HNO}_3 + \text{OH}^- (\text{HNO}_3 \text{ is a strong acid})
\]
NO\(_3^-\) is neutral

NaOH strong base
HOC\(_6\)H\(_5\) weak acid \((K_a = 1.6\times10^{-10})\)
KOC\(_6\)H\(_5\) (weak base)
\[
\text{KOC}_6\text{H}_5(s) \rightarrow \text{K}^+(aq) + \text{OC}_6\text{H}_5^- (aq)
\]
\[
\text{K}^+ + \text{H}_2\text{O} \quad \text{KOH} + \text{H}^+ (\text{KOH is a strong base})
\]
K\(^+\) is neutral
\[
\text{OC}_6\text{H}_5^- + \text{H}_2\text{O} \quad \text{HOC}_6\text{H}_5 + \text{OH}^-
\]
OC\(_6\)H\(_5^-\) is basic

C\(_6\)H\(_5\)NH\(_2\) weak base \((K_b = 3.8\times10^{-10})\)
HNO\(_3\) Strong acid
There are 2 weak acid \(\text{C}_6\text{N}_5\text{NH}_3\text{NO}_3 \text{(C}_6\text{H}_5\text{NH}_3^+\text{)}\) and HOC\(_6\)H\(_5\)
Compare the \(K_a\) values to determine the weaker acid
HOC\(_6\)H\(_5\) \quad K_a = 1.6\times10^{-10}
The \(K_a\) value of \(\text{C}_6\text{H}_5\text{NH}_3^+\) is not given therefore, it must be calculated
\[
\text{C}_6\text{H}_5\text{NH}_2(aq) \quad K_b = 3.8\times10^{-10}
\]
\[
K_a = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{3.8 \times 10^{-10}} = 2.6 \times 10^{-5}
\]
Therefore, \(\text{C}_6\text{N}_5\text{NH}_3\text{NO}_3\) is a stronger acid than HOC\(_6\)H\(_5\)
There are 2 weak bases \(\text{C}_6\text{N}_5\text{NH}_2\) and KOC\(_6\)H\(_5\) (OC\(_6\)H\(_5^-\))
Compare the \(K_b\) values to determine the weaker base
C\(_6\)N\(_5\)NH\(_2\) \quad K_b = 3.8\times10^{-10}
The \(K_b\) value of \(\text{OC}_6\text{H}_5^-\) is not given therefore, it must be calculated
HOC\(_6\)H\(_5\) \quad K_a = 1.6\times10^{-10}
103. See chart for problem 85.

a) Sr(NO$_3$)$_2$ (salt formed form a strong acid and strong base)

Sr(NO$_3$)$_2$ → Sr$^{2+}$(aq) + 2NO$_3$ (aq)

Sr$^{2+}$ + 2H$_2$O $\rightarrow$ Sr(OH)$_2$ + 2H$^+$ (Sr(OH)$_2$ is a strong base)

NO$_3^-$ + H$_2$O $\rightarrow$ HNO$_3$ + OH$^-$ (HNO$_3$ is a strong acid)

Therefore, Sr(NO$_3$)$_2$ is neutral

b) C$_2$H$_3$NH$_3$CN

C$_2$H$_3$NH$_3$CN(s) → C$_2$H$_3$NH$_3^+$(aq) + CN$^-$ (aq)

C$_2$H$_3$NH$_3^+$ + H$_2$O $\rightarrow$ C$_2$H$_3$NH$_2$ + H$_3$O$^+$

C$_2$H$_3$NH$_3^+$ is acidic

CN$^- + $H$_2$O $\rightarrow$ HCN + OH$^-$

CN$^-$ is basic

Therefore, need to compare K$_a$ and K$_b$ values of reactions

We are not given the K$_a$ of C$_2$H$_3$NH$_3^+$, therefore, we need to calculate it

\[
K_a = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{5.6 \times 10^{-4}} = 1.8 \times 10^{-11}
\]

We are not given the K$_b$ of CN$, therefore, we need to calculate it

\[
K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{6.2 \times 10^{-10}} = 1.6 \times 10^{-5}
\]

Since K$_a < K_b$, C$_2$H$_3$NH$_3$CN is basic

c) C$_2$H$_3$NHF

C$_2$H$_3$NHF(s) → C$_2$H$_3$NH$^+$(aq) + F$^-$ (aq)

C$_2$H$_3$NH$^+$ + H$_2$O $\rightarrow$ C$_2$H$_3$N + H$_3$O$^+$

C$_2$H$_3$NH$^+$ is acidic

F$^- + $H$_2$O $\rightarrow$ HF + OH$^-$

F$^-$ is basic

Therefore, need to compare K$_a$ and K$_b$ values of reactions

We are not given the K$_a$ of C$_2$H$_3$NH$^+$, therefore, we need to calculate it

\[
K_a = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{1.7 \times 10^{-9}} = 5.9 \times 10^{-6}
\]

We are not given the K$_b$ of F$, therefore, we need to calculate it

\[
K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{7.2 \times 10^{-4}} = 1.4 \times 10^{-11}
\]

Since K$_a > K_b$, C$_2$H$_3$NHF is acidic
d) \[ \text{NH}_4\text{C}_2\text{H}_3\text{O}_2 \]

\[ \text{NH}_4\text{C}_2\text{H}_3\text{O}_2(s) \rightarrow \text{NH}_4^+(aq) + \text{C}_2\text{H}_3\text{O}_2^-(aq) \]

\[ \text{NH}_4^+ + \text{H}_2\text{O} \xleftrightarrow{\text{NH}_3 + \text{H}_2\text{O}^+} \]

\( \text{NH}_4^+ \) is acidic

\[ \text{C}_2\text{H}_3\text{O}_2^- + \text{H}_2\text{O} \xleftrightarrow{\text{HC}_2\text{H}_3\text{O}_2 + \text{OH}^-} \]

\( \text{C}_2\text{H}_3\text{O}_2^- \) is basic

Therefore, need to compare \( K_a \) and \( K_b \) values of reactions

\[ \text{NH}_4^+ \] \( K_a = 5.6 \times 10^{-10} \)

We are not given the \( K_b \) of \( \text{C}_2\text{H}_3\text{O}_2^- \); therefore we need to calculate it

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

\( K_a = K_b \) \( \text{NH}_4\text{C}_2\text{H}_3\text{O}_2 \) is neutral

e) \[ \text{NaHCO}_3 \]

\[ \text{NaHCO}_3(s) \rightarrow \text{Na}^+(aq) + \text{HCO}_3^-(aq) \]

\[ \text{Na}^+ + \text{H}_2\text{O} \xleftrightarrow{\text{NaOH} + \text{H}^+ (\text{NaOH is a strong base})} \]

\( \text{Na}^+ \) is neutral

\[ \text{HCO}_3^- + \text{H}_2\text{O} \xleftrightarrow{\text{H}_2\text{CO}_3 + \text{OH}^-} \]

\[ \text{HCO}_3^- + \text{H}_2\text{O} \xleftrightarrow{\text{CO}_2^2- + \text{H}_3\text{O}^+} \]

Since \( \text{H}_2\text{CO}_3 \) is a polyprotic acid must look at \( K_a \) and \( K_b \) of the two possible reactions.

\[ \text{HCO}_3^- \] \( K_a = 4.8 \times 10^{-11} \)

We are not given the \( K_b \) of \( \text{HCO}_3^- \); therefore we need to calculate it

\[ K_a = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{4.3 \times 10^{-7}} = 2.3 \times 10^{-8} \]

Since \( K_a < K_b \) \( \text{HCO}_3^- \) is basic

Therefore, \( \text{NaHCO}_3 \) is basic

105. a) \[ \text{KCl KNO}_2 \]

\[ \text{KCl}(s) \rightarrow \text{K}^+(aq) + \text{Cl}^-(aq) \]

\[ \text{K}^+ + \text{H}_2\text{O} \xleftrightarrow{\text{KOH} + \text{H}^+ (\text{KOH is a strong base})} \]

\( \text{K}^+ \) is neutral

\[ \text{Cl}^- + \text{H}_2\text{O} \xleftrightarrow{\text{HCl} + \text{OH}^- (\text{HCl is a strong acid})} \]

\( \text{Cl}^- \) is neutral

Therefore, \( \text{KCl} \) is neutral

Therefore, the solution will be neutral \( \text{pH}=7.00 \). For neutral solutions the concentration of \( \text{H}^+ \) and \( \text{OH}^- \) equals \( 1.0 \times 10^{-7} \) M.

b) \[ \text{KF} \]

\[ \text{KF}(s) \rightarrow \text{K}^+(aq) + \text{F}^-(aq) \]

\[ \text{K}^+ + \text{H}_2\text{O} \xleftrightarrow{\text{KOH} + \text{H}^+ (\text{KOH is a strong base})} \]

\( \text{K}^+ \) is neutral

\[ \text{F}^- + \text{H}_2\text{O} \xleftrightarrow{\text{HF} + \text{OH}^-} \]

\( \text{F}^- \) is basic

Therefore, \( \text{KF} \) is basic
Interested in the following reaction
\[ F^{-}(aq) + H_2O(l) \rightleftharpoons HF(aq) + OH^{-}(aq) \]
Problem we do not know \( K_b \)
\[ HF(aq) \quad K_a = 7.2 \times 10^{-4} \]
\[ K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{7.2 \times 10^{-4}} = 1.4 \times 10^{-11} \]
Calculate the concentration of OH⁻ ions

<table>
<thead>
<tr>
<th></th>
<th>F⁻</th>
<th>HF</th>
<th>OH⁻</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>1.0 M</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-x</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>1.0-x</td>
<td>x</td>
<td>x</td>
</tr>
</tbody>
</table>

\[ K_b = \frac{[OH^{-}][HF]}{[F^-]} = \frac{xx}{1.0 - x} = 1.4 \times 10^{-11} \]
Since \( K_b \) is small assume that 1.0-x = 1.0
\[ xx = 1.4 \times 10^{-11} \]
\[ x = 3.7 \times 10^{-6} \]
Check assumption (5% rule)
\[ \frac{Amount\ Gained/Lost}{Original\ Concentration} = \frac{3.7 \times 10^{-6}}{1.0} = 3.7 \times 10^{-4}\% \quad Good \]
Calculate OH⁻ concentration
\[ [OH^-] = x = 3.7 \times 10^{-6} \text{ M} \]
Calculate H⁺ concentration
\[ [H^+] = 1.0 \times 10^{-14} \]
\[ [H^+] = 3.7 \times 10^{-6} \]
Calculate pH
\[ pH = -\log[H^+] = -\log(2.7 \times 10^{-9}) = 8.57 \]

109. a) KNO₂
KNO₂(s) \( \rightarrow \) K⁺(aq) + NO₂⁻(aq)
K⁺ + H₂O \( \rightleftharpoons \) KOH + H⁺(KOH is a strong base)
K⁺ is neutral
NO₂⁻ + H₂O \( \rightleftharpoons \) HNO₂ + OH⁻
NO₂⁻ is basic
Therefore, KNO₂ is basic
Interested in the following reaction
\[ NO_2^{-}(aq) + H_2O(l) \rightleftharpoons HNO_2(aq) + OH^{-}(aq) \]
Problem we don’t know \( K_b \)
Know
\[ HNO_2(aq) \quad K_a = 4.0 \times 10^{-4} \]
\[ K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{4.0 \times 10^{-4}} = 2.5 \times 10^{-11} \]
Calculate the concentration of OH⁻ ions

<table>
<thead>
<tr>
<th></th>
<th>NO₂⁻</th>
<th>HNO₂</th>
<th>OH⁻</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.12 M</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-x</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.12-x</td>
<td>x</td>
<td>x</td>
</tr>
</tbody>
</table>
\[
K_b = \frac{[HNO_2][OH^-]}{[NO_2^-]} = \frac{xx}{0.12 - x} = 2.5 \times 10^{-11}
\]

Since \( K_b \) is small assume that \( 0.12 - x = 0.12 \)

\[
\frac{xx}{0.12} = 2.5 \times 10^{-11}
\]

\[
x = 1.7 \times 10^{-6} M
\]

Check assumption (5% rule)

\[
\frac{\text{Amount Gained/Lost}}{\text{Original Concentration}} \times 100\% = \frac{1.7 \times 10^{-6}}{0.12} \times 100\% = 0.0014\%
\]

Good

Calculate \( OH^+ \) concentration

\[
[OH^-] = x = 1.7 \times 10^{-6} M
\]

Calculate pH

\[
pOH = -\log[OH^-] = -\log(1.7 \times 10^{-6}) = 10.56
\]

\[
\text{pH} + \text{pOH} = 14.00
\]

\[
\text{pH} = 14.00 - \text{pOH} = 14.00 - 5.77 = 8.23
\]

b) \( \text{NaOCl} \)

\[
\text{NaOCl}(s) \rightarrow \text{Na}^+(aq) + \text{OCl}^- (aq)
\]

\[
\text{Na}^+ + \text{H}_2\text{O} \rightarrow \text{NaOH} + \text{H}^+ \quad \text{(NaOH is a strong base)}
\]

\[
\text{Na}^+ \text{ is neutral}
\]

\[
\text{OCl}^- + \text{H}_2\text{O} \rightarrow \text{HOCl} + \text{OH}^-
\]

\( \text{OCl}^- \) is basic

Therefore, \( \text{NaOCl} \) is basic

Interested in the following reaction

\[
\text{OCl}^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HOCl}(aq) + \text{OH}^-(aq)
\]

Problem we don’t know \( K_b \)

\[
\text{HOCl}(aq) \quad \quad K_a = 3.5 \times 10^{-8}
\]

\[
K_b = \frac{1.0 \times 10^{-14}}{3.5 \times 10^{-8}} = 2.9 \times 10^{-7}
\]

Calculate the concentration of \( OH^- \) ions

<table>
<thead>
<tr>
<th>( \text{OCl}^- )</th>
<th>( \text{HOCl} )</th>
<th>( \text{OH}^- )</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Initial</strong></td>
<td>0.45 M</td>
<td>0</td>
</tr>
<tr>
<td><strong>Change</strong></td>
<td>-x</td>
<td>+x</td>
</tr>
<tr>
<td><strong>Equilibrium</strong></td>
<td>0.45-x</td>
<td>X</td>
</tr>
</tbody>
</table>

\[
K_b = \frac{[\text{HOCl}][\text{OH}^-]}{[\text{OCl}^-]} = \frac{xx}{0.45 - x} = 2.9 \times 10^{-7}
\]

Since \( K_b \) is small assume that \( 0.45 - x = 0.45 \)

\[
\frac{xx}{0.45} = 2.9 \times 10^{-7}
\]

\[
x = 3.6 \times 10^{-4} M
\]

Check assumption (5% rule)

\[
\frac{\text{Amount Gained/Lost}}{\text{Original Concentration}} \times 100\% = \frac{3.6 \times 10^{-4}}{0.45} \times 100\% = 0.080\%
\]

Good

Calculate \( OH^- \) concentration

\[
[OH^-] = x = 3.6 \times 10^{-4} M
\]

Calculate pH

\[
pOH = -\log[OH^-] = -\log(3.6 \times 10^{-4}) = 3.44
\]

\[
\text{pH} + \text{pOH} = 14.00
\]

\[
\text{pH} = 14.00 - \text{pOH} = 14.00 - 3.44 = 10.56
\]
c) \( \text{NH}_4\text{OCl}_4 \) (a salt from a weak base and a strong acid)

\[
\text{NH}_4\text{OCl}_4(s) \rightarrow \text{NH}_4^+(aq) + \text{OCl}_4^-(aq)
\]

\( \text{NH}_4^+ \) is acidic

\[
\text{OCl}_4^- + \text{H}_2\text{O} \rightarrow \text{HOCl}_4^- + \text{OH}^-(\text{HOCl}_4 \text{ strong acid})
\]

\( \text{OCl}_4^- \) is neutral

Therefore, \( \text{NH}_4\text{OCl}_4 \) is acidic

Interested in the following reaction

\[
\text{NH}_4^+(aq) \rightleftharpoons \text{NH}_3(aq) + \text{H}^+(aq) \quad K_a = 5.6 \times 10^{-10}
\]

Calculate the concentration of \( \text{H}^+ \) ions

<table>
<thead>
<tr>
<th>Initial</th>
<th>( \text{NH}_4^+ )</th>
<th>( \text{NH}_3 )</th>
<th>( \text{H}^+ )</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.40 M</td>
<td>0</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-x</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.40-x</td>
<td>x</td>
<td>X</td>
</tr>
</tbody>
</table>

\[
K_a = \frac{[\text{NH}_3][\text{H}^+]}{[\text{NH}_4^+]} = \frac{xx}{0.40 - x} = 5.6 \times 10^{-10}
\]

Since \( K_a \) is small assume that \( 0.40-x = 0.40 \)

\[
\frac{xx}{0.40} = 5.6 \times 10^{-10} \quad \frac{x}{0.40} = 1.5 \times 10^{-5} \quad M
\]

Check assumption (5% rule)

\[
\frac{\text{Amount Gained/Lost}}{\text{Original Concentration}} \times 100\% = \frac{1.5 \times 10^{-5}}{0.40} \times 100\% = 0.0038\% \quad \text{Good}
\]

Calculate \( \text{H}^+ \) concentration

\[
[H^+] = x = 1.5 \times 10^{-5}
\]

Calculate \( pH \)

\[
pH = -\log[H^+] = -\log(1.5 \times 10^{-5}) = 4.82
\]

Because the solution has a \( pH \) of 4.6 we know that the solution is acidic, therefore, \( \text{NaOH}, \text{NH}_3, \) and \( \text{NaCN} \) cannot be the solute.

Because the bulb is bright it means that there are a lot of ions in solution. Therefore, a strong electrolyte is needed. Only \( \text{HCl}, \text{NaOH}, \text{NH}_4\text{Cl}, \text{NaCN} \) are strong electrolytes.

\( \text{NaOH} \) and \( \text{NaCN} \) were already eliminated because they were basic.

\( \text{HCl} \) could light the bulb bright because it is a strong acid but the \( pH \) of a 1.0 M solution of \( \text{HCl} \) is

\[
pH = -\log[H^+] = -\log(1.0) = 0
\]

Confirm the solution is made from \( \text{NH}_4\text{Cl} \) by compare actual to calculated \( pH \) values

Determine the \( K_a \) of the weak acid.

\[
\text{HA}(aq) \rightleftharpoons A^-(aq) + \text{H}^+(aq) \quad K_a = \text{unknown}
\]

<table>
<thead>
<tr>
<th>Initial</th>
<th>HA</th>
<th>A^-</th>
<th>H^+</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.0 M</td>
<td>0</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-x</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>1.0-x</td>
<td>x</td>
<td>x</td>
</tr>
</tbody>
</table>
\[ K_a = \frac{[A^-][H^+]}{[HA]} = \frac{xx}{(1.0 - x)} \]

Determine \( x \)

\[ x = [H^+] \]
\[ pH = -\log[H^+] \]
\[ 4.6 = -\log[H^+] \]
\[ [H^+] = 2.5 \times 10^{-5} \]

\[ K_a = \frac{[A^-][H^+]}{[HA]} = \frac{(2.5 \times 10^{-5})(2.5 \times 10^{-5})}{1.0 - 2.5 \times 10^{-5}} = 6.3 \times 10^{-10} \]

This \( K_a \) of \( \text{NH}_4^+ \) is \( 5.6 \times 10^{-10} \) which is close to the calculated value. Therefore the solution must contain \( \text{NH}_4\text{Cl} \)