1

A 65 mL sample of HBr gas, measured at 35°C and 722 mmHg, was dissolved in water to yield 275 mL of solution. Calculate the molarity of hydrogen ion, the molarity of hydroxide ion, pH, and pOH in this solution.

\[
n = \frac{722 \text{ atm} \times \frac{65 \text{ mL}}{1000 \text{ L}}}{0.0821 \frac{\text{L atm}}{\text{mol K}} \times (35 + 273) \text{ K}} = 0.00244 \text{ mol HBr}
\]

\[
[HBr]_i = \frac{0.00244 \text{ mol}}{275 \times \frac{1000 \text{ L}}{1000 \text{ mL}}} = 0.0089 \text{ M}
\]

HBr = strong acid:

<table>
<thead>
<tr>
<th></th>
<th>HBr(aq)</th>
<th>→</th>
<th>H⁺(aq)</th>
<th>+</th>
<th>Br⁻(aq)</th>
</tr>
</thead>
<tbody>
<tr>
<td>R</td>
<td>0.0089</td>
<td>0</td>
<td>0</td>
<td></td>
<td></td>
</tr>
<tr>
<td>I</td>
<td>-0.0089</td>
<td>+0.0089</td>
<td>+0.0089</td>
<td></td>
<td></td>
</tr>
<tr>
<td>C</td>
<td>0</td>
<td>0.0089</td>
<td>0.0089</td>
<td></td>
<td></td>
</tr>
<tr>
<td>E</td>
<td>0</td>
<td>0.0089</td>
<td>0.0089</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

\[
[H^+] = 0.0089 \text{ M}
\]

\[
[OH^-] = \frac{1.0 \times 10^{-14}}{0.0089} = 1.1 \times 10^{-12} \text{ M}
\]

\[
pH = -\log(0.0089) = 2.05
\]

\[
pOH = 14.00 - 2.05 = 11.95
\]

2

A 0.086 gram sample of strontium hydroxide was dissolved in water to create 58 mL of solution. Calculate the molarity of hydroxide ion, the molarity of hydrogen ion, pOH, and pH in this solution.

\[
0.086 \text{ g} \left( \frac{1 \text{ mol}}{121.6 \text{ g}} \right) = 7.07 \times 10^{-4} \text{ mol Sr(OH)}_2
\]
\[ [\text{Sr(OH)}_2] = \frac{7.07 \times 10^{-4} \text{ mol}}{\text{L}} = 0.012 \text{ M} \]

\( \text{Sr(OH)}_2 = \) strong base:

<p>| | | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>R</td>
<td>( \text{Sr(OH)}_2(aq) )</td>
<td>( \rightarrow )</td>
</tr>
<tr>
<td>I</td>
<td>0.012</td>
<td>0</td>
</tr>
<tr>
<td>C</td>
<td>-0.012</td>
<td>+0.012</td>
</tr>
<tr>
<td>E</td>
<td>0</td>
<td>0.012</td>
</tr>
</tbody>
</table>

\[ [\text{OH}^-] = 0.024 \text{ M} \]

\[ [\text{H}^+] = \frac{1.0 \times 10^{-14}}{0.024} = 4.2 \times 10^{-13} \text{ M} \]

\[ \text{pOH} = -\log(0.024) = 1.62 \]

\[ \text{pH} = 14.00 - 1.62 = 12.38 \]

3

Write the acid ionization equation and calculate the pH and percent ionization of 0.63 M hypobromous acid, HBrO (\( K_a = 2.5 \times 10^{-9} \)).

<p>| | | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>R</td>
<td>( \text{HBrO(aq)} )</td>
<td>( \rightleftharpoons )</td>
</tr>
<tr>
<td>I</td>
<td>0.63</td>
<td>0</td>
</tr>
<tr>
<td>C</td>
<td>-x</td>
<td>+x</td>
</tr>
<tr>
<td>E</td>
<td>0.63 - x</td>
<td>x</td>
</tr>
</tbody>
</table>

\[ K_a = 2.5 \times 10^{-9} = \frac{x^2}{0.63 - x} \]

\((0 < x < 0.63)\)

\[ x = 4.0 \times 10^{-5} \text{ M} = [\text{H}^+] \]

\[ \text{pH} = -\log(4.0 \times 10^{-5}) = 4.40 \]

\[ \% \text{ionization} = \frac{4.0 \times 10^{-5}}{0.63} \times 100\% = 0.0063\% \]
A 0.85 M lactic acid solution has a pH of 1.97. Write the acid ionization equation and calculate $K_a$ for lactic acid, $\text{HC}_3\text{H}_5\text{O}_3$.

\[
\begin{array}{c}
\text{R)} \\
\text{I)} \\
\text{C)} \\
\text{E)} \\
\end{array}
\begin{array}{c}
\text{HC}_3\text{H}_5\text{O}_3(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{C}_3\text{H}_5\text{O}_3^-(\text{aq}) \\
0.85 \\
-x \\
0.85 - x \\
\end{array}
\begin{array}{c}
0 \\
+x \\
x \\
\end{array}
\begin{array}{c}
0 \\
+x \\
x \\
\end{array}
\begin{array}{c}
0 \\
+x \\
x \\
\end{array}
\]

\[
\begin{align*}
[H^+] &= 10^{-1.97} = 0.011 \text{ M} = x \\
K_a &= \frac{(0.011)^2}{(0.85 - 0.011)} = 1.4 \times 10^{-4}
\end{align*}
\]

A 0.017 M solution of propanoic acid is 2.7% ionized. Write the acid ionization equation and calculate the pH of the solution and $K_a$ for propanoic acid, $\text{HC}_3\text{H}_5\text{O}_2$.

\[
\begin{array}{c}
\text{R)} \\
\text{I)} \\
\text{C)} \\
\text{E)} \\
\end{array}
\begin{array}{c}
\text{HC}_3\text{H}_5\text{O}_2(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{C}_3\text{H}_5\text{O}_2^-(\text{aq}) \\
0.017 \\
-x \\
0.017 - x \\
\end{array}
\begin{array}{c}
0 \\
+x \\
x \\
\end{array}
\begin{array}{c}
0 \\
+x \\
x \\
\end{array}
\begin{array}{c}
0 \\
+x \\
x \\
\end{array}
\]

\[
\begin{align*}
\frac{x}{0.017} \times 100% &= 2.7% \\
x &= 4.6 \times 10^{-4} \text{ M} = [H^+] \\
pH &= -\log(4.6 \times 10^{-4}) = 3.34 \\
K_a &= \frac{(4.6 \times 10^{-4})^2}{(0.017 - 4.6 \times 10^{-4})} = 1.3 \times 10^{-5}
\end{align*}
\]

A cyanic acid, $\text{HCNO}$, solution has a pH of 2.25. Given that $K_a = 3.5 \times 10^{-4}$ for cyanic acid, write the acid ionization equation and calculate the initial molarity of the cyanic acid solution.

\[
\begin{array}{c}
\text{R)} \\
\text{I)} \\
\text{C)} \\
\text{E)} \\
\end{array}
\begin{array}{c}
\text{HCNO}(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{CNO}^-(\text{aq}) \\
M_i \\
-x \\
M_i - x \\
\end{array}
\begin{array}{c}
0 \\
+x \\
x \\
\end{array}
\begin{array}{c}
0 \\
+x \\
x \\
\end{array}
\begin{array}{c}
0 \\
+x \\
x \\
\end{array}
\]

3
\[
[H^+] = 10^{-2.25} = 0.0056 \text{ M} = x
\]
\[
K_a = 3.5 \times 10^{-4} = \frac{(0.0056)^2}{(M_i - 0.0056)}
\]
\[
M_i = 0.095 \text{ M}
\]

7

Write the stepwise acid ionization equations and calculate the pH of 0.23 M selenous acid, \( \text{H}_2\text{SeO}_3 \), which has the following acid ionization constants:

\[
\begin{align*}
K_{a1} &= 2.3 \times 10^{-3} \\
K_{a2} &= 5.3 \times 10^{-9}
\end{align*}
\]

1st ionization:

\[
\begin{align*}
\text{R)} & \quad \text{H}_2\text{SeO}_3(\text{aq}) & \rightleftharpoons & \quad \text{H}^+(\text{aq}) & + & \quad \text{HSeO}_3^-(\text{aq}) \\
\text{I)} & \quad 0.23 & & 0 & & 0 \\
\text{C)} & \quad -x & & +x & & +x \\
\text{E)} & \quad 0.23 - x & & x & & x
\end{align*}
\]

\[
K_{a1} = 2.3 \times 10^{-3} = \frac{x^2}{0.23 - x}
\]
\[
(0 < x < 0.23)
\]
\[
x = 0.022 \text{ M} = [H^+] \\
pH = -\log(0.022) = 1.66
\]

2nd ionization:

\[
\text{HSeO}_3^- (\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{SeO}_3^{2-} (\text{aq})
\]

8

Identify the Bronsted acids and bases in the forward and reverse directions for the reaction below:

\[
\text{H}_2\text{C}_6\text{H}_5\text{O}_7^- (\text{aq}) + \text{HC}_3\text{H}_2\text{O}_4^- (\text{aq}) \rightleftharpoons \text{HC}_6\text{H}_5\text{O}_7^{2-} (\text{aq}) + \text{H}_2\text{C}_3\text{H}_2\text{O}_4(\text{aq})
\]

forward reaction:
Bronsted acid = \( \text{H}_2\text{C}_6\text{H}_5\text{O}_7^- \) (donates proton), Bronsted base = \( \text{HC}_3\text{H}_2\text{O}_4^- \) (accepts proton)

reverse reaction:
Bronsted acid = \( \text{H}_2\text{C}_3\text{H}_2\text{O}_4 \), Bronsted base = \( \text{HC}_6\text{H}_5\text{O}_7^{2-} \)
9

Write the formula for:

a. the conjugate acid of HC\(_2\)O\(_4^-\)
b. the conjugate base of HAsO\(_4^{2-}\)

a. conjugate acid = add H\(^+\) to formula = H\(_2\)C\(_2\)O\(_4\)
b. conjugate base = remove H\(^+\) from formula = AsO\(_4^{3-}\)

10

Write the base ionization equation and calculate the pH and percent ionization of 0.44 M dimethylamine, (CH\(_3\))\(_2\)NH (K\(_b\) = 5.4 \times 10^{-4}).

R) \( (\text{CH}_3)_2\text{NH}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons (\text{CH}_3)_2\text{NH}_2^+(aq) + \text{OH}^-(aq) \)
I) 0.44 0 0
C) \(-x\) \(+x\) \(+x\)
E) 0.44 - x x x

\[ K_b = 5.4 \times 10^{-4} = \frac{x^2}{0.44 - x} \]
(0 < x < 0.44)
\[ x = 0.015 M = [\text{OH}^-] \]
\[ \text{pOH} = -\log(0.015) = 1.82 \]
\[ \text{pH} = 14.00 - 1.82 = 12.18 \]
\% ionization = \frac{0.015}{0.44} \times 100% = 3.4%

11

A 0.084 M codeine solution has a pH of 10.46. Write the base ionization equation and calculate K\(_b\) for codeine, C\(_{18}\)H\(_{21}\)O\(_3\)N.

R) \( \text{C}_{18}\text{H}_{21}\text{O}_3\text{N}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{C}_{18}\text{H}_{21}\text{O}_3\text{N}^+(aq) + \text{OH}^-(aq) \)
I) 0.084 0 0
C) \(-x\) \(+x\) \(+x\)
E) 0.084 - x x x

\[ \text{pOH} = 14.00 - 10.46 = 3.54 \]
\[ [\text{OH}^-] = 10^{-3.54} = 2.9 \times 10^{-4} M = x \]
\[ K_b = \frac{(2.9 \times 10^{-4})^2}{(0.084 - 2.9 \times 10^{-4})} = 1.0 \times 10^{-6} \]
12

A 0.077 M solution of piperidine is 12% ionized. Write the base ionization equation and calculate the pH of the solution and \( K_b \) for piperidine, \( C_5H_{11}N \).

\[
\begin{align*}
\text{R)} & \quad C_5H_{11}N(aq) + H_2O(l) \rightleftharpoons C_5H_{11}NH^+(aq) + OH^-(aq) \\
\text{I)} & \quad 0.077 \quad 0 \quad 0 \\
\text{C)} & \quad -x \quad +x \quad +x \\
\text{E)} & \quad 0.077 - x \quad x \quad x
\end{align*}
\]

\[
\frac{x}{0.077} \times 100\% = 12\%
\]

\[ x = 0.0092 \text{ M} = [OH^-] \]

\[ pOH = -\log(0.0092) = 2.04 \]

\[ pH = 14.00 - 2.04 = 11.96 \]

\[ K_b = \frac{(0.0092)^2}{(0.077 - 0.0092)} = 1.2 \times 10^{-3} \]

13

A quinoline, \( C_9H_7N \), solution has a pH of 9.00. Given that \( K_b = 6.3 \times 10^{-10} \) for quinoline, write the base ionization equation and calculate the initial molarity of the quinoline solution.

\[
\begin{align*}
\text{R)} & \quad C_9H_7N(aq) + H_2O(l) \rightleftharpoons C_9H_7NH^+(aq) + OH^-(aq) \\
\text{I)} & \quad M_i \quad 0 \quad 0 \\
\text{C)} & \quad -x \quad +x \quad +x \\
\text{E)} & \quad M_i - x \quad x \quad x
\end{align*}
\]

\[ pOH = 14.00 - 9.00 = 5.00 \]

\[ [OH^-] = 10^{-5.00} = 1.0 \times 10^{-5} \text{ M} = x \]

\[ K_b = 6.3 \times 10^{-10} = \frac{(1.0 \times 10^{-5})^2}{(M_i - 1.0 \times 10^{-5})} \]

\[ M_i = 0.16 \text{ M} \]
14

Draw Lewis structures for chloric acid, HClO₃, and selenous acid, H₂SeO₃.
Which is the stronger acid? Give two reasons to justify your answer.

![Lewis structures for chloric acid and selenous acid]

Chloric acid is stronger acid because:

1. HClO₃ has more terminal oxygens (3 - 1 = 2) than H₂SeO₃ (3 - 2 = 1).
2. Electronegativity for central atom Cl is higher than for central atom Se.

15

Which of the two acids shown below is the stronger acid? Give two reasons to justify your answer.

![Lewis structures of two acids]

The acid on the right is stronger because:

1. Electronegativity of F is higher than Cl.
2. F is closer to ionizable proton than Cl.

16

Predict whether a solution of each compound below will be acidic, basic, or neutral. For solutions that are not neutral, show all relevant hydrolysis reactions that affect the pH and also calculate the equilibrium constant for each reaction you write using information from the following data tables:

<table>
<thead>
<tr>
<th>Acid</th>
<th>$K_a$</th>
</tr>
</thead>
<tbody>
<tr>
<td>HCN</td>
<td>$4.9 \times 10^{-10}$</td>
</tr>
<tr>
<td>HIO</td>
<td>$2.3 \times 10^{-14}$</td>
</tr>
<tr>
<td>Base</td>
<td>$K_b$</td>
</tr>
<tr>
<td>--------------</td>
<td>-----------</td>
</tr>
<tr>
<td>$(\text{CH}_3)_3\text{N}$</td>
<td>$6.4 \times 10^{-5}$</td>
</tr>
<tr>
<td>$\text{NH}_3$</td>
<td>$1.8 \times 10^{-5}$</td>
</tr>
</tbody>
</table>

a. $(\text{CH}_3)_3\text{NHCl}$ [composed of $(\text{CH}_3)_3\text{NH}^+$ and $\text{Cl}^-$]
b. KCN
c. NaI
d. NH$_4$IO

a. $\text{Cl}^-$ = spectator ion

$(\text{CH}_3)_3\text{NH}^+$ hydrolyzes as weak acid = solution is acidic:

$$(\text{CH}_3)_3\text{NH}^+(aq) \rightleftharpoons \text{H}^+(aq) + (\text{CH}_3)_3\text{N}(aq)$$

$$K_a = \frac{1.0 \times 10^{-14}}{6.4 \times 10^{-5}} = 1.6 \times 10^{-10}$$

b. $\text{K}^+$ = spectator ion

$\text{CN}^-$ hydrolyzes as a weak base = solution is basic:

$$\text{CN}^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HCN}(aq) + \text{OH}^-(aq)$$

$$K_b = \frac{1.0 \times 10^{-14}}{4.9 \times 10^{-10}} = 2.0 \times 10^{-5}$$

c. $\text{Na}^+$ and $\text{I}^-$ = spectator ions = solution is neutral
d. Both ions hydrolyze, so we must compare equilibrium constants:

$$\text{NH}_4^+(aq) \rightleftharpoons \text{H}^+(aq) + \text{NH}_3(aq)$$

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

$$\text{IO}^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HIO}(aq) + \text{OH}^-(aq)$$

$$K_b = \frac{1.0 \times 10^{-14}}{2.3 \times 10^{-11}} = 4.3 \times 10^{-4}$$

$$K_a < K_b = \text{solution is basic}$$
17

For each solution below, show any relevant hydrolysis reactions and calculate the pH.

a. 0.31 M C₆H₅NH₃Br [composed of C₆H₅NH₃⁺ and Br⁻]
b. 1.2 M KC₃H₅O₃

Kₐ = 4.3 × 10⁻¹⁰ for C₆H₅NH₂
Kₐ = 1.4 × 10⁻⁴ for HC₃H₅O₃

a. Br⁻ = spectator ion

C₆H₅NH₃⁺ hydrolyzes as weak acid:

R) C₆H₅NH₃⁺(aq) ⇌ H⁺ + C₆H₅NH₂(aq)
I) 0.31 0 0
C) -x +x +x
E) 0.31 - x x x

Kₐ = \frac{1.0 \times 10^{-14}}{4.3 \times 10^{-10}} = 2.3 \times 10^{-5} = \frac{x^2}{0.31 - x}

(0 < x < 0.31)

x = 0.0027 M = [H⁺]
pH = -\log(0.0027) = 2.57

b. K⁺ = spectator ion

C₃H₅O₃⁻ hydrolyzes as a weak base:

R) C₃H₅O₃⁻(aq) + H₂O(l) ⇌ HC₃H₅O₃(aq) + OH⁻(aq)
I) 1.2 0 0
C) -x +x +x
E) 1.2 - x x x

Kₐ = \frac{1.0 \times 10^{-14}}{1.4 \times 10^{-14}} = 7.1 \times 10^{-11} = \frac{x^2}{1.2 - x}

(0 < x < 1.2)

x = 9.2 \times 10^{-6} M = [OH⁻]
pOH = -\log(9.2 \times 10^{-6}) = 5.04
pH = 14.00 - 5.04 = 8.96
Predict whether a solution of sodium hydrogen arsenate, NaH$_2$AsO$_4$, will be acidic or basic. Show all relevant reactions that affect the pH and also give the value of the equilibrium constant for each reaction you write. For arsenic acid, H$_3$AsO$_4$:

\[
\begin{align*}
K_{a1} &= 5.6 \times 10^{-3} \\
K_{a2} &= 1.0 \times 10^{-7} \\
K_{a3} &= 3.0 \times 10^{-12}
\end{align*}
\]

Na$^+$ = spectator ion

weak acid ionization reaction:

\[H_2AsO_4^- (aq) \rightleftharpoons H^+ (aq) + HAsO_4^{2-}\]

\[K_a \text{ for } H_2AsO_4^- = K_{a2} \text{ for } H_3AsO_4 = 1.0 \times 10^{-7}\]

weak base hydrolysis reaction:

\[H_2AsO_4^- (aq) + H_2O(l) \rightleftharpoons H_3AsO_4(aq) + OH^- (aq)\]

\[K_b \text{ for } H_2AsO_4^- = \frac{1.0 \times 10^{-14}}{K_{a1} \text{ for } H_3AsO_4 = 5.6 \times 10^{-3}} = 1.8 \times 10^{-12}\]

\[K_a > K_b = \text{solution is acidic}\]

Will the reaction of P$_4$O$_{10}$ and water produce H$_3$PO$_3$ or H$_3$PO$_4$? Write a balanced equation for the reaction.

The reaction of P$_4$O$_{10}$ (oxidation number of P = +5) and water will produce H$_3$PO$_4$ (oxidation number of P = +5) rather than H$_3$PO$_3$ (oxidation number of P = +3). The balanced equation will be P$_4$O$_{10}$ + 6H$_2$O \rightarrow 4H$_3$PO$_4$.

a. An unknown monoprotic weak acid was found to be 30.60% carbon, 45.16% chlorine, and 3.85% hydrogen by mass, with the remainder being oxygen. Determine the empirical formula of the acid.

b. In a separate experiment, 3.75 grams of the acid was dissolved in 45 mL of water and then titrated with 0.164 M barium hydroxide. The volume of base required to reach the equivalence point was 72.8 mL. Calculate the molar mass
and determine the molecular formula of the acid.

a. 

$100\% - 30.60\% C - 45.16\% Cl - 3.85\% H = 20.39\% O$

Assume one hundred grams of unknown compound:

$30.60 \text{ g } C \left( \frac{1 \text{ mol}}{12.01 \text{ g}} \right) = 2.548 \text{ mol } C$

$45.16 \text{ g } I \left( \frac{1 \text{ mol}}{35.45 \text{ g}} \right) = 1.274 \text{ mol } Cl$

$3.85 \text{ g } H \left( \frac{1 \text{ mol}}{1.008 \text{ g}} \right) = 3.82 \text{ mol } H$

$20.39 \text{ g } O \left( \frac{1 \text{ mol}}{16.00 \text{ g}} \right) = 1.274 \text{ mol } O$

$\frac{2.548}{1.274} \text{ mol } C : \frac{1.274}{1.274} \text{ mol } Cl : \frac{3.82}{1.274} \text{ mol } H : \frac{1.274}{1.274} \text{ mol } O$

emirical formula $= C_2ClH_3O$

b. $0.164 \text{ M } \text{Ba(OH)}_2 \rightarrow 0.164 \text{ M } \text{Ba}^{2+} \text{ and } 2 \times 0.164 \text{ M } = 0.328 \text{ M } \text{OH}^-$

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

$72.8 \text{ mL } \left( \frac{1 \text{ L}}{1000 \text{ mL}} \right) \left( \frac{0.328 \text{ mol OH}^-}{1 \text{ L}} \right) \left( \frac{1 \text{ mol HA}}{1 \text{ mol OH}^-} \right) = 0.02388 \text{ mol HA}$

$M = \frac{3.75 \text{ g}}{0.02388 \text{ mol}} = 157 \text{ g/mol}$

$\frac{M}{EM} = \frac{157}{78.49} = 2$

molecular formula $= C_2ClH_3O \times 2 = C_4Cl_2H_6O_2$

Since the acid is monoprotic, we can rewrite the molecular formula as $\text{HC}_4\text{Cl}_2\text{H}_5\text{O}_2$. 
21

Calculate the initial molarity of a sodium fluoride, NaF, solution that has a pH of 8.17 given that $K_a = 6.8 \times 10^{-4}$ for hydrofluoric acid, HF.

\[ \text{Na}^+ = \text{spectator ion, F}^- \text{ hydrolyzes as a weak base:} \]

\[
\begin{align*}
\text{R) } & \quad \text{F}^- (aq) + H_2O(l) \rightleftharpoons HF(aq) + \text{OH}^- (aq) \\
\text{I) } & \quad M_i \quad 0 \quad 0 \\
\text{C) } & \quad -x \quad +x \quad +x \\
\text{E) } & \quad M_i - x \quad x \quad x \\
\end{align*}
\]

\[ \text{pOH} = 14.00 - 8.17 = 5.83 \]
\[ [\text{OH}^-] = 10^{-5.83} = 1.5 \times 10^{-6} \text{ M} = x \]
\[ K_b \text{ for F}^- = \frac{1.0 \times 10^{-14}}{6.8 \times 10^{-4}} = 1.5 \times 10^{-11} = \frac{(1.5 \times 10^{-6})^2}{(M_i - 1.5 \times 10^{-6})} \]
\[ M_i = 0.15 \text{ M} \]

22

Calculate the initial molarity of a C$_5$H$_5$NHNO$_3$ [composed of C$_5$H$_5$NH$^+$ and NO$_3$ $^-$] solution that has a pH of 2.83 given that $K_b = 1.7 \times 10^{-9}$ for C$_5$H$_5$N.

\[ \text{NO}_3^- = \text{spectator ion, C}_5\text{H}_5\text{NH}^+ \text{ hydrolyzes as a weak acid:} \]

\[
\begin{align*}
\text{R) } & \quad \text{C}_5\text{H}_5\text{NH}^+ (aq) \rightleftharpoons \text{H}^+ (aq) + \text{C}_5\text{H}_5\text{N}(aq) \\
\text{I) } & \quad M_i \quad 0 \quad 0 \\
\text{C) } & \quad -x \quad +x \quad +x \\
\text{E) } & \quad M_i - x \quad x \quad x \\
\end{align*}
\]

\[ [\text{H}^+] = 10^{-2.83} = 0.0015 \text{ M} = x \]
\[ K_a \text{ for C}_5\text{H}_5\text{NH}^+ = \frac{1.0 \times 10^{-14}}{1.7 \times 10^{-9}} = 5.9 \times 10^{-6} = \frac{(0.0015)^2}{(M_i - 0.0015)} \]
\[ M_i = 0.38 \text{ M} \]
Given that a 0.72 M KIO solution has a pH of 12.24, calculate $K_a$ and $pK_a$ for HIO.

$K^+$ = spectator ion, IO$^-$ hydrolyzes as a weak base:

\[
\begin{align*}
\text{R)} & \quad \text{IO}^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HIO}(aq) + \text{OH}^-(aq) \\
\text{I)} & \quad 0.72 \\
\text{C)} & \quad -x \\
\text{E)} & \quad 0.72 - x
\end{align*}
\]

\[
\begin{align*}
\text{pOH} &= 14.00 - 12.24 = 1.76 \\
[\text{OH}^-] &= 10^{-1.76} = 0.017 \text{M} = x \\
K_b \text{ for IO}^- &= \frac{(0.017)^2}{(0.72 - 0.017)} = 4.1 \times 10^{-4} \\
K_a \text{ for HIO} &= \frac{1.0 \times 10^{-14}}{4.1 \times 10^{-4}} = 2.4 \times 10^{-11} \\
pK_a \text{ for HIO} &= -\log(2.4 \times 10^{-11}) = 10.62
\end{align*}
\]

Given that a 0.066 M C$_6$H$_{15}$O$_3$NHCl [composed of C$_6$H$_{15}$O$_3$NH$^+$ and Cl$^-$] solution has a pH of 4.48, calculate $K_b$ and $pK_b$ for C$_6$H$_{15}$O$_3$N.

Cl$^-$ = spectator ion, C$_6$H$_{15}$O$_3$NH$^+$ hydrolyzes as a weak acid:

\[
\begin{align*}
\text{R)} & \quad \text{C}_6\text{H}_{15}\text{O}_3\text{NH}^+(aq) \rightleftharpoons \text{H}^+(aq) + \text{C}_6\text{H}_{15}\text{O}_3\text{N}(aq) \\
\text{I)} & \quad 0.066 \\
\text{C)} & \quad -x \\
\text{E)} & \quad 0.066 - x
\end{align*}
\]

\[
[\text{H}^+] = 10^{-4.48} = 3.3 \times 10^{-5} \text{M} = x
\]
\[ K_a \text{ for } C_6H_{15}O_3\text{NH}^+ = \frac{(3.3 \times 10^{-5})^2}{(0.066 - 3.3 \times 10^{-5})} = 1.7 \times 10^{-8} \]

\[ K_b \text{ for } C_6H_{15}O_3\text{N} = \frac{1.0 \times 10^{-14}}{1.7 \times 10^{-8}} = 5.9 \times 10^{-7} \]

\[ \text{pK}_b \text{ for } C_6H_{15}O_3\text{N} = -\log(5.9 \times 10^{-7}) = 6.23 \]

25

Calculate the pH of a solution containing 4.2 grams of NaC_7H_5O_2 in 75 mL of 0.27 M HC_7H_5O_2 (K_a = 6.3 \times 10^{-5}).

Na^+ = \text{spectator ion, } HC_7H_5O_2/C_7H_5O_2^- = \text{buffer solution:}

\[ 4.2 \text{ g NaC}_7\text{H}_5\text{O}_2 \left( \frac{1 \text{ mol}}{144.1 \text{ g}} \right) = 0.029 \text{ mol NaC}_7\text{H}_5\text{O}_2 \rightarrow 0.029 \text{ mol C}_7\text{H}_5\text{O}_2^- \]

\[ 75 \text{ mL} \left( \frac{1 \text{ L}}{1000 \text{ mL}} \right) \left( \frac{0.27 \text{ mol HC}_7\text{H}_5\text{O}_2}{1 \text{ L}} \right) = 0.020 \text{ mol HC}_7\text{H}_5\text{O}_2 \]

\[ \text{pK}_a \text{ for } HC_7H_5O_2 = -\log(6.3 \times 10^{-5}) = 4.20 \]

\[ \text{pH} = 4.20 + \log \frac{0.029 \text{ mol C}_7\text{H}_5\text{O}_2^-}{0.020 \text{ mol HC}_7\text{H}_5\text{O}_2} = 4.36 \]

26

Calculate the pH of a solution containing 2.4 grams of (CH_3)_2NH_2I [composed of (CH_3)_2NH^+ and I^-] in 84 mL of 0.18 M (CH_3)_2NH (K_b = 5.4 \times 10^{-4}).

I^- = \text{spectator ion, } (CH_3)_2NH/(CH_3)_2NH_2^+ = \text{buffer solution:}

\[ 2.4 \text{ g (CH}_3\text{)_2NH}_2\text{I} \left( \frac{1 \text{ mol}}{173.0 \text{ g}} \right) = 0.014 \text{ mol (CH}_3\text{)_2NH}_2\text{I} \rightarrow 0.014 \text{ mol (CH}_3\text{)_2NH}_2^+ \]

\[ 84 \text{ mL} \left( \frac{1 \text{ L}}{1000 \text{ mL}} \right) \left( \frac{0.18 \text{ mol (CH}_3\text{)_2NH}}{1 \text{ L}} \right) = 0.015 \text{ mol (CH}_3\text{)_2NH} \]

\[ \text{pK}_b \text{ for } (CH_3)_2NH = -\log(5.4 \times 10^{-4}) = 3.27 \]

\[ \text{pOH} = 3.27 + \log \frac{0.014 \text{ mol (CH}_3\text{)_2NH}_2^+}{0.015 \text{ mol (CH}_3\text{)_2NH}} = 3.24 \]

\[ \text{pH} = 14.00 - 3.24 = 10.76 \]
Write the net ionic equation for the neutralization reaction that occurs and calculate the pH when:

a. 0.002 mol NaOH is added to the solution in Question 25
b. 0.002 mol HBr is added to the solution in Question 25
c. 0.002 mol NaOH is added to the solution in Question 26
d. 0.002 mol HBr is added to the solution in Question 26

a. \( \text{Na}^+ = \text{spectator ion} \)
neutralization reaction: \( \text{HC}_7\text{H}_5\text{O}_2(\text{aq}) + \text{OH}^- (\text{aq}) \rightarrow \text{C}_7\text{H}_5\text{O}_2^- (\text{aq}) + \text{H}_2\text{O}(\text{l}) \)
\[
\text{pH} = 4.20 + \log \frac{(0.029 + 0.002) \text{ mol C}_7\text{H}_5\text{O}_2^-}{(0.020 - 0.002) \text{ mol HC}_7\text{H}_5\text{O}_2} = 4.44
\]
b. \( \text{Br}^- = \text{spectator ion} \)
neutralization reaction: \( \text{C}_7\text{H}_5\text{O}_2^- (\text{aq}) + \text{H}^+ (\text{aq}) \rightarrow \text{HC}_7\text{H}_5\text{O}_2(\text{aq}) \)
\[
\text{pH} = 4.20 + \log \frac{(0.029 - 0.002) \text{ mol C}_7\text{H}_5\text{O}_2^-}{(0.020 + 0.002) \text{ mol HC}_7\text{H}_5\text{O}_2} = 4.29
\]
c. \( \text{Na}^+ = \text{spectator ion} \)
neutralization reaction: \( (\text{CH}_3)_2\text{NH}^+ (\text{aq}) + \text{OH}^- (\text{aq}) \rightarrow (\text{CH}_3)_2\text{NH}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \)
\[
\text{pOH} = 3.27 + \log \frac{(0.014 - 0.002) \text{ mol (CH}_3)_2\text{NH}_2^+}{(0.015 + 0.002) \text{ mol (CH}_3)_2\text{NH}} = 3.12
\]
\[
\text{pH} = 14.00 - 3.12 = 10.88
\]
d. \( \text{Br}^- = \text{spectator ion} \)
neutralization reaction: \( (\text{CH}_3)_2\text{NH}(\text{aq}) + \text{H}^+ (\text{aq}) \rightarrow (\text{CH}_3)_2\text{NH}_2^+ (\text{aq}) \)
\[
\text{pOH} = 3.27 + \log \frac{(0.014 + 0.002) \text{ mol (CH}_3)_2\text{NH}_2^+}{(0.015 - 0.002) \text{ mol (CH}_3)_2\text{NH}} = 3.36
\]
\[
\text{pH} = 14.00 - 3.36 = 10.64
\]

Write the net ionic equation for the neutralization reaction that occurs in each aqueous mixture below. Which one of the four reactions creates a buffer solution? For each mixture, describe the method of calculating the pH of the resulting solution after the neutralization reaction is complete.
a. 0.15 mol HNO₃ + 0.15 mol NaCN
b. 0.25 mol HIO + 0.25 mol KOH
c. 0.15 mol NH₄Br + 0.25 mol NaOH
d. 0.15 mol HI + 0.25 mol C₆H₅NH₂ (a weak base)

a. NO₃⁻ and Na⁺ = spectator ions
   neutralization reaction: CN⁻(aq) + H⁺(aq) → HCN(aq)
   \[ n_{CN⁻} = n_{H⁺} \]
   HCN solution created, pH calculated using RICE chart for ionization of weak acid HCN

b. K⁺ = spectator ion
   neutralization reaction: HIO(aq) + OH⁻(aq) → IO⁻(aq) + H₂O(l)
   \[ n_{HIO} = n_{OH⁻} \]
   IO⁻ solution created, pOH = 14.00 - pH calculated using RICE chart for hydrolysis of weak base IO⁻

c. Br⁻ and Na⁺ = spectator ions
   neutralization reaction: NH₄⁺(aq) + OH⁻(aq) → NH₃(aq) + H₂O(l)
   \[ n_{NH₄⁺} < n_{OH⁻} \]
   OH⁻/NH₃ solution created, pOH = 14.00 - pH = -log [OH⁻]_{excess} (ionization of NH₃ negligible)

d. I⁻ = spectator ion
   neutralization reaction: C₆H₅NH₂(aq) + H⁺(aq) → C₆H₅NH₃⁺(aq)
   \[ n_{C₆H₅NH₂} > n_{H⁺} \]
   C₆H₅NH₂/C₆H₅NH₃⁺ buffer created
   \[ pOH = pK_b + \log \frac{n_{C₆H₅NH₃⁺ \text{produced}}}{n_{C₆H₅NH₂ \text{excess}}} = 14.00 - pH \]

29

a. Write the net ionic equation for the neutralization reaction that occurs during the titration of 22 mL of 0.40 M HNO₃ with 0.20 M NaOH and calculate the volume of base needed to reach the equivalence point.
b. Calculate the following:

i. the initial pH
ii. the pH after 31 mL of base has been added
iii. the pH at the equivalence point
iv. the pH after 49 mL of base has been added

a. \( \text{NO}_3^- \) and \( \text{Na}^+ \) = spectator ions

\[
\text{neutralization reaction: } \text{H}^+ (\text{aq}) + \text{OH}^- (\text{aq}) \rightarrow \text{H}_2\text{O}(l)
\]

\[
0.022 \text{ L} \left( \frac{0.40 \text{ mol H}^+}{1 \text{ L}} \right) = 0.0088 \text{ mol H}^+ = 0.0088 \text{ mol OH}^- \left( \frac{1 \text{ L}}{0.20 \text{ mol OH}^-} \right)
\]

\[= 0.044 \text{ L} = 44 \text{ mL base needed}\]

b.

i. \[\text{pH} = -\log(0.40) = 0.40\]

ii. \[0.031 \text{ L} \left( \frac{0.20 \text{ mol OH}^-}{1 \text{ L}} \right) = 0.0062 \text{ mol OH}^- \text{ added}\]

\[\text{pH} = -\log \left( \frac{0.0088 - 0.0062}{0.022 + 0.031} \right) = 1.31\]

iii. only \( \text{H}_2\text{O} \) present, pH = 7

iv. \[0.049 \text{ L} \left( \frac{0.20 \text{ mol OH}^-}{1 \text{ L}} \right) = 0.0098 \text{ mol OH}^- \text{ added}\]

\[\text{pOH} = -\log \left( \frac{0.0098 - 0.0088}{0.022 + 0.049} \right) = 1.85\]

\[\text{pH} = 14.00 - 1.85 = 12.15\]

30

a. Write the net ionic equation for the neutralization reaction that occurs during the titration of 24 mL of 0.30 M \( \text{HC}_3\text{H}_5\text{O}_2 \) (\( K_a = 1.3 \times 10^{-5} \)) with 0.20 M KOH and calculate the volume of base needed to reach the equivalence point.

b. Calculate the following:

i. the initial pH
ii. the pH after 23 mL of base has been added
iii. the pH at the equivalence point
iv. the pH after 43 mL of base has been added

a. $K^+ = \text{spectator ion}$

neutralization reaction: $\text{HC}_3\text{H}_5\text{O}_2(\text{aq}) + \text{OH}^- (\text{aq}) \rightarrow \text{C}_3\text{H}_5\text{O}_2^- (\text{aq}) + \text{H}_2\text{O}(\text{l})$

\[
0.024 \text{ L} \left( \frac{0.30 \text{ mol HC}_3\text{H}_5\text{O}_2}{1 \text{ L}} \right) = 0.0072 \text{ mol HC}_3\text{H}_5\text{O}_2 = 0.0072 \text{ mol OH}^- \left( \frac{1 \text{ L}}{0.20 \text{ mol OH}^-} \right) = 0.036 \text{ L} = 36 \text{ mL base needed}
\]

b.

i. 

\[
\begin{align*}
\text{R)} & \quad \text{HC}_3\text{H}_5\text{O}_2(\text{aq}) & \leftrightharpoons & \quad \text{H}^+(\text{aq}) & \quad + & \quad \text{C}_3\text{H}_5\text{O}_2^- (\text{aq}) \\
\text{I)} & \quad 0.30 & \quad 0 & \quad 0 \\
\text{C)} & \quad -x & \quad +x & \quad +x \\
\text{E)} & \quad 0.30 - x & \quad x & \quad x
\end{align*}
\]

\[
K_a \text{ for HC}_3\text{H}_5\text{O}_2 = 1.3 \times 10^{-5} = \frac{x^2}{0.30 - x}
\]

\[
x = 0.0020 \text{ M} = [\text{H}^+]
\]

\[
pH = -\log(0.0020) = 2.70
\]

ii. 

\[
0.023 \text{ L} \left( \frac{0.20 \text{ mol OH}^-}{1 \text{ L}} \right) = 0.0046 \text{ mol OH}^- \text{ added}
\]

\[
= 0.0046 \text{ mol HC}_3\text{H}_5\text{O}_2 \text{ reacted} = 0.0046 \text{ mol C}_3\text{H}_5\text{O}_2^- \text{ produced}
\]

\[
= 0.0072 \text{ mol HC}_3\text{H}_5\text{O}_2 \text{ initial} - 0.0046 \text{ mol HC}_3\text{H}_5\text{O}_2 \text{ reacted} = 0.0026 \text{ mol HC}_3\text{H}_5\text{O}_2 \text{ excess}
\]

\[
pK_a \text{ for HC}_3\text{H}_5\text{O}_2 = -\log(1.3 \times 10^{-5}) = 4.89
\]

\[
pH = 4.89 + \log \frac{0.0046 \text{ mol C}_3\text{H}_5\text{O}_2^-}{0.0026 \text{ mol HC}_3\text{H}_5\text{O}_2} = 5.14
\]

iii. 

\[
\begin{align*}
\text{R)} & \quad \text{C}_3\text{H}_5\text{O}_2^- (\text{aq}) & \quad + & \quad \text{H}_2\text{O}(\text{l}) & \leftrightharpoons & \quad \text{HC}_3\text{H}_5\text{O}_2(\text{aq}) & \quad + & \quad \text{OH}^- (\text{aq}) \\
\text{I)} & \quad M_i & \quad 0 & \quad 0 \\
\text{C)} & \quad -x & \quad +x & \quad +x \\
\text{E)} & \quad M_i - x & \quad x & \quad x
\end{align*}
\]

\[
M_i = \frac{n_{\text{C}_3\text{H}_5\text{O}_2^-}}{L_{\text{total}}} = \frac{n_{\text{OH}^- \text{ added}}}{L_{\text{total}}} = \frac{0.0072 \text{ mol}}{(0.024 + 0.036) \text{ L}} = 0.12 \text{ M}
\]
\[ K_b \text{ for } C_5H_5O_2^- = \frac{K_w}{K_a \text{ for } HC_5H_5O_2} = \frac{1.0 \times 10^{-14}}{1.3 \times 10^{-5}} = 7.7 \times 10^{-10} = \frac{x^2}{0.12 - x} \]

\[ x = 9.6 \times 10^{-6} \text{ M} = [\text{OH}^-] \]

\[ \text{pOH} = -\log(9.6 \times 10^{-6}) = 5.02 \]

\[ \text{pH} = 14.00 - 5.02 = 8.98 \]

iv.

\[ 0.043 \left( \frac{0.20 \text{ mol OH}^-}{1 \text{ L}} \right) = 0.0086 \text{ mol OH}^- \text{ added} \]

\[ \text{pOH} = -\log \left( \frac{0.0086 - 0.0072}{(0.024 + 0.043) \text{ L total}} \right) = 1.68 \]

\[ \text{pH} = 14.00 - 1.68 = 12.32 \]

31

a. Write the net ionic equation for the neutralization reaction that occurs during the titration of 48 mL of 0.10 M C_5H_5N (K_b = 1.7 \times 10^{-9}) with 0.30 M HI and calculate the volume of acid needed to reach the equivalence point.

b. Calculate the following:

i. the initial pH
ii. the pH after 12 mL of acid has been added
iii. the pH at the equivalence point
iv. the pH after 27 mL of acid has been added

a. \( I^- \) = spectator ion

neutralization reaction: \( C_5H_5N(aq) + H^+(aq) \rightarrow C_5H_5NH^+(aq) \)

\[ 0.048 \left( \frac{0.10 \text{ mol C}_5\text{H}_5\text{N}}{1 \text{ L}} \right) = 0.0048 \text{ mol C}_5\text{H}_5\text{N} = 0.0048 \text{ mol H}^+ \left( \frac{1 \text{ L}}{0.30 \text{ mol H}^+} \right) \]

\[ = 0.016 \text{ L} = 16 \text{ mL acid needed} \]

b.

i.

\[
\begin{array}{c|ccc}
\text{R) } & \text{C}_5\text{H}_5\text{N}(aq) & + & \text{H}_2\text{O}(l) & \rightarrow & \text{C}_5\text{H}_5\text{NH}^+(aq) & + & \text{OH}^-(aq) \\
\text{I) } & 0.10 & & 0 & & 0 & & 0 \\
\text{C) } & -x & & +x & & +x \\
\text{E) } & 0.10 - x & & x & & x \\
\end{array}
\]
\[ K_b \text{ for } C_5H_5N = 1.7 \times 10^{-9} = \frac{x^2}{0.10 - x} \]

\[ x = 1.3 \times 10^{-5} \text{ M} = [\text{OH}^-] \]

\[ \text{pOH} = -\log(1.3 \times 10^{-5}) = 4.89 \]

\[ \text{pH} = 14.00 - 4.89 = 9.11 \]

ii.

\[ 0.012 \text{ L} \left( \frac{0.30 \text{ mol H}^+}{1 \text{ L}} \right) = 0.0036 \text{ mol H}^+ \text{ added} \]

\[ = 0.0036 \text{ mol } C_5H_5N \text{ reacted} = 0.0036 \text{ mol } C_5H_5NH^+ \text{ produced} \]

\[ 0.0048 \text{ mol } C_5H_5N \text{ initial} - 0.0036 \text{ mol } C_5H_5N \text{ reacted} = 0.0012 \text{ mol } C_5H_5N \text{ excess} \]

\[ pK_b \text{ for } C_5H_5N = -\log(1.7 \times 10^{-9}) = 8.77 \]

\[ \text{pOH} = 8.77 + \log \frac{0.0036 \text{ mol } C_5H_5NH^+}{0.0012 \text{ mol } C_5H_5N} = 9.25 \]

\[ \text{pH} = 14.00 - 9.25 = 4.75 \]

iii.

R) \( C_5H_5NH^+ (aq) \rightleftharpoons H^+(aq) + C_5H_5N(aq) \)

I) \( M_i \)

\[ \begin{array}{ccc}
-x & +x & +x \\
0 & 0 & 0 \\
\end{array} \]

E) \( M_i - x \)

\[ x \quad x \quad x \]

\[ M_i = \frac{n_{C_5H_5NH^+}}{L \text{ total}} = \frac{\text{mol } H^+ \text{ added}}{L \text{ total}} = \frac{0.0048 \text{ mol}}{0.048 + 0.016} = 0.075 \text{ M} \]

\[ K_a \text{ for } C_5H_5NH^+ = \frac{K_w}{K_b \text{ for } C_5H_5N} = \frac{1.0 \times 10^{-14}}{1.7 \times 10^{-9}} = 5.9 \times 10^{-6} = \frac{x^2}{0.075 - x} \]

\[ x = 6.6 \times 10^{-4} \text{ M} = [H^+] \]

\[ \text{pH} = -\log(6.6 \times 10^{-4}) = 3.18 \]

iv.

\[ 0.027 \text{ L} \left( \frac{0.30 \text{ mol H}^+}{1 \text{ L}} \right) = 0.0081 \text{ mol H}^+ \text{ added} \]

\[ \text{pH} = -\log \frac{(0.0081 - 0.0048) \text{ mol H}^+ \text{ excess}}{(0.048 + 0.027) \text{ L total}} = 1.36 \]
Which indicator, bromphenol blue ($K_a = 1 \times 10^{-4}$) or phenolphthalein ($K_a = 5 \times 10^{-10}$), would be the better choice for the titration in Question 30?

\[
\begin{align*}
\text{pH at equivalence point} & = 8.98 \\
\text{pK}_a \text{ for bromphenol blue} & = -\log(1 \times 10^{-4}) = 4.0 \\
\text{pK}_a \text{ for phenolphthalein} & = -\log(5 \times 10^{-10}) = 9.3 \\
\text{pK}_a \text{ for phenolphthalein close to pH at equivalence point, so phenolphthalein is better choice}
\end{align*}
\]