Acid-Base Equilibria

By Thomas Cahill, Arizona State University, New College of Interdisciplinary Arts and Sciences.

When a strong acid and a strong base dissolve in water, they dissociate completely into ions, as shown in the following two equations:

\[
\begin{align*}
\text{HCl} (aq) + \text{H}_2\text{O} (l) &\rightarrow \text{H}_3\text{O}^+ (aq) + \text{Cl}^- (aq) \\
\text{NaOH} (s) &\rightarrow \text{Na}^+ (aq) + \text{OH}^- (aq)
\end{align*}
\]

As a result, the concentration of the hydronium or hydroxide ion is equal to the concentration of the strong acid or base respectively. In contrast, weak acids and bases do not dissociate completely in solution, and instead react with water to ionize partly, as shown in the following equation:

\[
\text{HA} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{A}^-
\]

The equilibrium constants for these reactions are known as dissociation or ionization constants, and for a given weak acid or weak base, the ionization constant has a definite value at a specific temperature. The ionization constant expression for the above acid dissociation, using \(\text{H}^+\) in place of \(\text{H}_3\text{O}^+\) for simplicity, is as follows:

**Equation 1:**

\[
K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}
\]

Other useful definitions for solving acid-base problems include:

\[
\begin{align*}
K_w &= [\text{H}_3\text{O}^+][\text{OH}^-] = 1.00 \times 10^{-14} \\
K_w &= K_a \times K_b \\
p\text{H} &= -\log[\text{H}^+] \\
p\text{OH} &= -\log[\text{OH}^-] \\
p\text{H} + p\text{OH} &= 14 \\
pK_a &= -\log K_a \\
pK_b &= -\log K_b \\
pK_a + pK_b &= 14 \quad \text{(for an acid-base conjugate pair)}
\end{align*}
\]

**Acid-Base Properties of Salt Solutions:**
Soluble inorganic ionic compounds (commonly called salts) completely dissociate into their ions when dissolved in water. However, the fate of these ions depends on the acid or base from which they are derived. Here are the guidelines:

1) The anions of strong acids (\(\text{Cl}^-, \text{Br}^-, \Gamma^-, \text{NO}_3^-, \text{ClO}_3^-, \text{ClO}_4^-\)) are not bases. They are simply too weak relative to water to have any effect. The anion of sulfuric acid is \(\text{HSO}_4^-\) and it is an acid since it donates a second hydrogen. However, the ionization of hydrogen sulfate is a weak acid, so the resulting anion (\(\text{SO}_4^{2-}\)) is a weak base.

2) The cations of strong bases (\(\text{Li}^+, \text{Na}^+, \text{K}^+, \text{Rb}^+, \text{Cs}^+, \text{Ca}^{2+}, \text{Sr}^{2+}, \text{Ba}^{2+}\)) are not acids.
3) Some cations with large charges (generally +3 or bigger) can behave as Lewis acids and create rather acidic solutions. Some of the more famous examples of these cations are Fe$^{3+}$, Cr$^{3+}$, and Al$^{3+}$.

4) The anions of weak acids are weak bases. Conversely, the cations of weak bases are weak acids. Therefore, any salt that contains a conjugate ion of either a weak acid or a weak base will affect the pH in a predictable fashion. Conjugate bases of weak acids will react with water and make a basic solution as shown in the following examples:

$$\text{CH}_3\text{COO}^- + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COOH} + \text{OH}^-$$
$$\text{F}^- + \text{H}_2\text{O} \rightleftharpoons \text{HF} + \text{OH}^-$$
$$\text{ClO}^- + \text{H}_2\text{O} \rightleftharpoons \text{HClO} + \text{OH}^-$$

If the solution contains a cation that is the conjugate acid of a weak base, then it will react with water to produce hydronium ion, so the solution will be acidic as in the following example:

$$\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+$$

Finally, a salt solution may contain an anion and a cation that are both capable of reacting with water (e.g. NH$_4$F), and in this case both hydroxide and hydronium ions will be produced. The ion with the larger equilibrium constant ($K_a$ or $K_b$) will have the greater influence on the pH. Therefore, you will need to look up both the $K_a$ of the acidic ion and the $K_b$ of the basic ion. Whichever ion has the larger $K_{eq}$ value will control the pH of the solution. This is also the case for amphoteric ions in salts. For example, the HPO$_4^{2-}$ can act as either an acid or a base as shown below:

Acting as an acid: $\text{HPO}_4^{2-} + \text{H}_2\text{O} = \text{PO}_4^{3-} + \text{H}_3\text{O}^+$  \hspace{1cm} $K_a = 4.2 \times 10^{-13}$
Acting as a base: $\text{HPO}_4^{2-} + \text{H}_2\text{O} = \text{H}_2\text{PO}_4^- + \text{OH}^-$  \hspace{1cm} $K_b = 1.6 \times 10^{-7}$

Since $K_b$ is much larger than $K_a$, the base reaction is the dominant reaction and the solution will be basic.

In Part A of this lab, the pH of various salt solutions will be determined using a pH meter. From the experimental pH, the hydronium [H$_3$O$^+$] ion concentrations will be calculated.

**Buffer Solutions:**

Buffer solutions are those which contain a weak conjugate acid-base pair, and can resist drastic changes in pH upon the addition of small amounts of strong acid or strong base. Human blood is a complex aqueous mixture with a pH buffered at about 7.4. Any deviation from this normal pH range can have extremely disruptive effects on the stability of cell membranes, protein structure, and the activity of enzymes. Death can result if the blood pH falls below 6.8 or rises above 7.8.

Buffers resist changes in pH because they contain both an acid species to neutralize the OH$^-$ ions and a basic species to neutralize the H$^+$ ions. Buffers are often prepared by mixing a weak acid with a salt containing the conjugate base of the acid. For example, you can create a buffer by mixing CH$_3$COOH and NaCH$_3$COO. By choosing appropriate components and adjusting their relative concentrations, buffers can be prepared at any pH.

**Calculating the pH of a buffer:**

Because conjugate acid-base pairs share a common ion, we can use the same procedures to calculate the pH of a buffer that we used to treat the common-ion effect. However, an alternate approach is often taken, known as the **Henderson-Hasselbalch equation**. This equation is derived from
the acid-dissociation constant expression, and by taking the log of both sides and rearranging the result, the Henderson-Hasselbalch equation is as follows:

**Equation 2:**

\[
\text{pH} = pK_a + \log \left( \frac{[A^-]}{[HA]} \right)
\]

In problems that involve the common ion effect, the starting concentrations of the weak acid and its salt are usually given. As long as these concentrations are reasonably high (> 0.1M), we can assume that the starting concentrations are equal to the equilibrium concentrations and can be used in Equation 1 or Equation 2. This assumption is valid due to the fact that the presence of the conjugate base suppresses the ionization of the weak acid and the weak acid suppresses the activity of the conjugate base, so these values are very small.

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**Example 1**

- Let’s calculate the pH of a buffer that is 0.12 M in lactic acid (HC₃H₅O₃) and 0.1 M in sodium lactate using the Henderson-Hasselbalch equation. The \( K_a \) for lactic acid is \( 1.4 \times 10^{-4} \).

\[
\begin{align*}
&\text{HC}_3\text{H}_5\text{O}_3 \leftrightarrow \text{C}_3\text{H}_5\text{O}_3^- + \text{H}^+ \\
pK_a &= -\log K_a = -\log(1.4 \times 10^{-4}) = 3.85 \\
[A^-] &= 0.1 \text{M} \\
[HA] &= 0.12 \text{M}
\end{align*}
\]

\[
\text{pH} = pK_a + \log \left( \frac{[A^-]}{[HA]} \right)
\]

\[
\text{pH} = 3.85 + \log \left( \frac{0.1}{0.12} \right) = 3.77
\]

Two important characteristics of a buffer are its capacity and its effective pH range. Buffer capacity refers to the amount of acid or base the buffer can neutralize before the pH begins to change significantly. A buffer solution containing a higher concentration of acid/conjugate base will have a greater buffering capacity and will be more resistant to a change in pH. For example, the pH range for a 1 L solution that is 1 M in acetic acid and 1 M in sodium acetate will be the same as for a solution that is 0.1 M acetic acid and 0.1 M sodium acetate. However, the solution containing the higher concentrations of acetic acid and sodium acetate will have a greater buffering capacity because it contains more acid and its conjugate base ion.

The pH of the buffer depends on the \( pK_a \) for the acid and on the relative concentrations of the acid and conjugate base that comprise the buffer. When the concentrations of weak acid and conjugate base are about the same, the buffer most effectively resists a change in pH in either direction, and when the concentrations are equal, \( \text{pH} = pK_a \). Consequently, a buffer is usually prepared using an acid-base system whose \( pK_a \) is close to the desired pH, and the ratio of \([A^-]/[HA]\) is close to one.

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**Example 2**

- Suppose we need to prepare a buffer at pH 4.2.
We would need to choose a weak acid that has a $pK_a$ close to 4.2 and a salt containing the conjugate base of this weak acid. Looking at some acid ionization constants (in the text book or on the class handouts), we see that acrylic acid ($\text{CH}_2=\text{CH}−\text{COOH}$) has a $pK_a$ of 4.25. Next, we substitute the pH and $pK_a$ values into the Henderson-Hasselbalch equation to obtain the ratio $[\text{conjugate base}] / [\text{acid}]$.

$$\text{CH}_2=\text{CH}−\text{COOH} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{CH}_2=\text{CH}−\text{COO}^- \qquad pK_a = 4.25$$

$$\text{pH} = pK_a + \log[\text{A}^-]/[\text{HA}]$$

$$4.2 = 4.25 + \log[\text{A}^-]/[\text{HA}]$$

$$-0.05 = \log[\text{A}^-]/[\text{HA}]$$

$$10^{-0.05} = [\text{A}^-]/[\text{HA}]$$

$$[\text{A}^-]/[\text{HA}] = 0.89$$

With a ratio of 0.89:1 we would dissolve 0.89 moles of sodium acrylate ($\text{NaCH}_2=\text{CH}−\text{COO}$) into 1 mole of acrylic acid and dilute to 1 L solution.

A buffer solution has the ability to neutralize either added acid or added base. For example, in the acetic acid/acetate buffer system, if acid is added, the excess $\text{H}^+$ ions are consumed by the conjugate base according to the equation:

$$\text{CH}_3\text{COO}^-(\text{aq}) + \text{H}^+(\text{aq}) \rightleftharpoons \text{CH}_3\text{COOH}(\text{aq})$$

If a base is added, the $\text{OH}^-$ ions are consumed by the acid in the buffer according to the equation:

$$\text{CH}_3\text{COOH}(\text{aq}) + \text{OH}^-(\text{aq}) \rightleftharpoons \text{CH}_3\text{COO}^-(\text{aq}) + \text{H}_2\text{O}(\text{l})$$

Example 3 illustrates the effects on the pH of a buffer system after the addition of HCl.

**Example 3**

- Determine the pH of the buffer system in Example 1 after the addition of 1.0 mL of 1 M HCl to 50 mL of buffer. Original Buffer: 0.12 M in lactic acid ($\text{HC}_3\text{H}_5\text{O}_3$) and 0.1 M in sodium lactate. The $K_a$ for lactic acid is $1.4 \times 10^{-4}$

Step 1. Write the equation for the addition of HCl to the acid/conjugate base buffer

$$\text{C}_3\text{H}_5\text{O}_3^-(\text{aq}) + \text{H}^+(\text{aq}) \rightleftharpoons \text{HC}_3\text{H}_5\text{O}_3(\text{aq})$$

Step 2. Calculate the initial **moles** of the acid form and the base form of the conjugate acid-base pair.

<table>
<thead>
<tr>
<th>Acid Form</th>
<th>Base Form</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{C}_3\text{H}_5\text{O}_3^-$ (aq)</td>
<td>$\text{HC}_3\text{H}_5\text{O}_3$ (aq)</td>
</tr>
<tr>
<td>Initial</td>
<td>0.005mol</td>
</tr>
</tbody>
</table>

Step 3. Next, calculate the moles of $\text{H}^+$ added to the system by a strong acid. The strong acid is stronger than weak acids, so it can protonate the conjugate base of a weak acid and turns it into the acid form. Therefore, for each mole of $\text{H}^+$ added to the system, the moles of the anion decrease by the same amount and the moles of the acid form increases by the same amount. In simpler words, the moles of $\text{H}^+$ added converts an equal number of moles of the anion to the
acid. In this example, 0.001 mol of H\textsuperscript{+} was added, so the moles of the anion form decreases by 0.001 moles and the acid form increases by 0.001 moles.

\[
\begin{array}{c|c|c}
\text{ } & C_3H_5O_3\textsuperscript{(aq)} & HC_3H_5O_3\textsuperscript{(aq)} \\
\hline
\text{Initial} & 0.005\text{mol} & 0.006\text{mol} \\
\text{Change} & -0.001 & +0.001 \\
\text{Equilibrium} & 0.004\text{mol} & 0.007\text{mol} \\
\end{array}
\]

Step 4. Calculate the equilibrium concentration of the acid/conjugate base in the final volume of 51 mL and determine the pH using the Henderson-Hasselbalch equation

\[
\text{pH} = pK_a + \log[A^-]/[HA]
\]

\[
\text{pH} = (-\log 1.4 \times 10^{-4}) + \log \frac{0.0784}{0.137} = 3.61
\]

Check to make sure the answer makes sense! The original pH of the buffer was calculated in Example 1 to be 3.77 and after the addition of 1.0 mL of 1 M HCl the pH was 3.61. The pH should decrease with the addition of acid, so this answer makes sense.

In Part B of the lab, you will prepare a buffer and study the effects of adding small amounts of strong acids and bases in order to determine how effectively the system resists large changes in pH.
Experimental Procedure:

PART A – Acid/Base Characteristics of Salts

In Part A of this experiment, you will measure the pH of several aqueous salt solutions to determine whether these solutions are acidic, basic, or neutral.

Skip the grayed procedure for the online labs

Go to https://www.public.asu.edu/~jhwang43/
Click on “General Chemistry II Laboratory Exercises, By Thomas Cahill”
Click on “Lab5: Acid-Base Equilibria”
Click on “Part A: Acid/Base Characteristics of Salts”

1. Place 11 test tubes in a test tube rack and label each one for the following solutions that you will test:
   a. Sodium acetate NaCH₃COO
   b. Sodium nitrate NaNO₃
   c. Ammonium chloride NH₄Cl
   d. Sodium carbonate Na₂CO₃
   e. Sodium oxalate Na₂C₂O₄
   f. Potassium iodide KI
   g. Potassium phosphate K₃PO₄
   h. Ammonium nitrate NH₄NO₃
   i. Sodium hypochlorite NaClO
   j. Iron(III) chloride FeCl₃
   k. Ammonium sulfate (NH₄)₂SO₄

2. The reagent stations are spaced around the lab. At each station add 2 ml of each salt solution into the labeled test tubes, by clicking “Click to Add ~2mL 0.1M NaCH₃COO(aq) to Test Tube Below”. The concentration of the salt solutions are generally 0.1 M.

3. Based on the ions in each solution, predict if the pH should be:
   a. Neutral (pH between about 6.5 and 7.5)
   b. Basic (pH > 7.5)
   c. Acidic (pH < 6.5)

4. Use the pre-calibrated pH meter at your station to measure the pH of each sample, by clicking “Click to Measure pH” underneath the test tube. Your instructor will demonstrate how to correctly use the pH meter. It is very important to rinse the pH probe well with DI before and after each solution measurement to prevent contaminating the solutions.

5. Record the pH values on the report sheet for the solution. Repeat step 2, 3, and 4 for the remaining test tubes.

6. When you finish measuring the pH of each solution, put the pH probe back into the storage solution.
7. Dispose of all chemical waste in the designated container in the fume hood.
8. Wash your test tubes well for the next student and put them upside down in the test tube rack to dry. Putting the test tubes upside down lets any remaining water to drain down out of the test tube.
9. Using the pH values, calculate the hydronium (H₃O⁺ or H⁺) ion concentrations for each solution.
10. Indicate the ion responsible for the acid/base characteristics in each solution.
11. Write the net ionic equation for the reaction that generated the hydronium or hydroxide ion.
a. In all cases, the ion will react with water, so water will always be present as a reactant in the equations. The reaction between the ions and water is often called “hydrolysis” (“hydro” meaning “water” and “-lysis” meaning “break apart”), but this is not always correct since acidic ions (e.g. NH$_4^+$) simply transfer a hydrogen to water to form H$_3$O$^+$. 

b. If the solution is neither acidic or basic, then write “no reaction” for the reaction.

c. If the ion present is the conjugate base of a polyprotic acid (e.g. CO$_3^{2-}$), then the reaction with water transfers only one hydrogen to the ion, so it will not take the conjugated base of a polyprotic acid to the fully protonated state.

12. Write the equilibrium constant expression for all salts that caused a reaction. Be sure to write the equilibrium expression for the net ionic equation you wrote in the above step.
PART B – BUFFERS

Skip the grayed procedure for the online labs

Go to https://www.public.asu.edu/~jhwang43/
Click on “General Chemistry II Laboratory Exercises, By Thomas Cahill”
Click on “Lab5: Acid-Base Equilibria”
Click on “Part B: BUFFERS”

In this part of the lab, you will prepare a sodium acetate-acetic acid buffer system and observe the effects of the addition of strong acids or bases on the pH.

\[ \text{CH}_3\text{COOH} + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}_3\text{O}^+ \qquad K_a = 1.8 \times 10^{-5} \]

(acetic acid) \hspace{1cm} (acetate ion)

1. Weigh between 3.20 - 3.30 g of NaCH$_3$COO·3H$_2$O to the nearest 0.01 g and record the exact weight on the data sheet, by clicking “Click to get NaCH$_3$COO.3H$_3$O(s) and its grams. Weigh Paper Mass Tared to 0 g”. Carefully add this to a clean 100 mL volumetric flask, by clicking “Click to Add the weighed NaCH$_3$COO.3H$_3$O(s) to the 100 mL volumetric flask Below”.

2. Using a 10 mL graduated cylinder and a transfer pipet, measure exactly 8 mL of 3.0 M acetic acid and add it to the volumetric flask, by clicking “Click to Add exactly 8 mL of 3.0 M acetic acid to the 100 mL volumetric flask Below”.

3. Dilute to the 100 mL mark with DI water by clicking “Dilute to the 100 mL mark with DI water and mix thoroughly” and mix thoroughly, ensuring that the solid has dissolved.

4. Using a 100 mL graduated cylinder, pour exactly 50 mL of the prepared buffer solution into two clean 100 mL beakers, labeled beakers #1 & #2, by clicking “Pour exactly 50 mL of the prepared buffer to Beaker #1 and #2”.

5. Using a pH meter measure the pH of the buffer solution in beaker #1, by clicking “Click to Measure the pH of Beaker #1 BEFORE Adding HCL(aq)”, and record the pH on the report sheet. Make sure that you rinse the probe with DI water before you take your measurements and after you have finished.

6. Using an autopipetter at the designated HCl reagent station, add exactly 1 mL of 1.0M HCl into beaker #1 and swirl to mix, by clicking “Click to Add exactly 1 mL of 1.0M HCl into beaker #1 and swirl to mix”. Measure the pH of the solution by clicking “Click to Measure the pH of Beaker #1 AFTER Adding HCL(aq)”, and record it on the report sheet.

7. Remember to rinse the pH probe with DI water when you are finished.

8. Using an autopipetter at the designated NaOH reagent station, add exactly 1 mL of 1.0M NaOH into beaker #2 and swirl to mix, by clicking “Click to Add exactly 1 mL of 1.0M NaOH into beaker #2 and swirl to mix”. Measure the pH of the solution by clicking “Click to Measure the pH of Beaker #2 AFTER Adding NaOH(aq)”, and record it on the report sheet.

9. Dispose of the solutions in the designated waste container and clean the beakers and pH probe thoroughly with DI water.

Click “Go to Lab-5 Part B Test Water’s Buffer Capacity”
10. Add 50 mL of DI water to a clean 100 mL beaker #1 and beaker #2, by clicking “Click to Pour 50 mL of DI Water to Beaker #1 and #2”. and measure the pH.

11. Add exactly 1 mL of 1.0M HCl to beaker #1 by clicking “Click to Add exactly 1 mL of 1.0M HCl into beaker #1 and swirl to mix”, and measure the pH by clicking “Click to Measure the pH of Water to Beaker #1 AFTER Adding HCl(aq)”. Record the pH in the data table. Discard the solution in the designated waste container and clean the beaker thoroughly.

12. Add 50 mL of DI water to a clean 100 mL beaker and measure the pH.

13. Add exactly 1.0 mL of 1.0M NaOH to the beaker by clicking “Click to Add exactly 1 mL of 1.0M NaOH into beaker #2 and swirl to mix”, and measure the pH by clicking “Click to Measure the pH of Water to Beaker #2 AFTER Adding NaOH(aq)”. Record the pH in the data table. Discard the solution in the designated waste container and clean the beaker thoroughly.

14. Calculate the expected pH of the original buffer solution using the Henderson-Hasselbalch equation.
   The pKₐ of acetic acid is 4.74.

15. Calculate the percent error between the expected pH value and the observed pH value of the buffer.

16. Calculate the expected pH of the buffer after the addition of 1.0 mL of 1M HCl.

17. Calculate the expected pH of the buffer after the addition of 1.0 mL of 1M NaOH.

18. Calculate the expected pH of the DI H₂O after the addition of 1.0 mL of 1M HCl

19. Calculate the expected pH of the DI H₂O after the addition of 1.0 mL of 1M NaOH.
Acid-Base Equilibria Experimental Data

PART A – Acid/Base Characteristics of Salts. The first row is given as an example, but you should still test the sodium acetate solution to confirm the results.

<table>
<thead>
<tr>
<th>Solution</th>
<th>Measured pH</th>
<th>[H⁺]</th>
<th>Predicted pH (acidic, basic or neutral)</th>
<th>Acidic or basic ion (if any)</th>
</tr>
</thead>
<tbody>
<tr>
<td>NaCH₃COO</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>NaNO₃</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>NH₄Cl</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Na₂CO₃</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Na₂C₂O₄</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>KI</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>K₃PO₄</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>NH₄NO₃</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>NaClO</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>FeCl₃</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>(NH₄)₂SO₄</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Show one example [H⁺] calculation:

_____ of 3 pts
**DATA – PART A**

Write the net ionic equation for each reaction. If no reaction occurred, then write “No Reaction”. The first row is given as an example. Be sure to include water as a reactant in each equation since the acidic or basic ions are reacting with water to form either hydronium or hydroxide ions. Make sure the reactions are balanced with respect to both atoms and ionic charges (Note that the net charge on each side of the reaction should be the same).

<table>
<thead>
<tr>
<th>Solution (0.1M)</th>
<th>Net Ionic Equation for reactions generating H$_3$O$^+$ or OH$^-$. If no reaction occurs write “No reaction”.</th>
<th>Equilibrium Constant Expression ($K_a$ or $K_b$ or “no reaction”)</th>
</tr>
</thead>
<tbody>
<tr>
<td>NaCH$_3$COO</td>
<td></td>
<td></td>
</tr>
<tr>
<td>NaNO$_3$</td>
<td></td>
<td></td>
</tr>
<tr>
<td>NH$_4$Cl</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Na$_2$CO$_3$</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Na$_2$C$_2$O$_4$</td>
<td></td>
<td></td>
</tr>
<tr>
<td>KI</td>
<td></td>
<td></td>
</tr>
<tr>
<td>K$_3$PO$_4$</td>
<td></td>
<td></td>
</tr>
<tr>
<td>NH$_4$NO$_3$</td>
<td></td>
<td></td>
</tr>
<tr>
<td>NaClO (bleach)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>FeCl$_3$</td>
<td>This one is hard and unusual, so we will give you the equation: $[\text{Fe} \cdot (\text{H}_2\text{O})_6]^{3+} + \text{H}_2\text{O} = [\text{Fe} \cdot (\text{H}_2\text{O})_5(\text{OH})]^{2+}(aq) + \text{H}_3\text{O}^+$</td>
<td></td>
</tr>
<tr>
<td>(NH$_4$)$_2$SO$_4$</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

______ of 5 pts
DATA - PART B

Mass NaCH₃COO·3H₂O: __________
Molar Mass NaCH₃COO·3H₂O: __________
Moles of CH₃COO⁻ __________
Volume 3M CH₃COOH: __________
Moles of CH₃COOH: __________
Final Volume of Buffer Solution: __________
  Molarity of CH₃COO⁻ __________
  Molarity of CH₃COOH: __________

Observed pH Values:

<table>
<thead>
<tr>
<th>pH</th>
<th>pH</th>
<th>pH</th>
<th>pH</th>
<th>pH</th>
</tr>
</thead>
<tbody>
<tr>
<td>Buffer</td>
<td>Buffer + HCl</td>
<td>Buffer + NaOH</td>
<td>DI + HCl</td>
<td>DI + NaOH</td>
</tr>
<tr>
<td></td>
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Calculations: (show all work for full credit)

1. Calculate the expected pH of the original buffer.

2. Calculate the percent error between the expected pH value and the observed pH value using the pH meter.

   \[
   \% \text{ error} = \frac{|\text{expected pH} - \text{observed pH}|}{\text{expected pH}}
   \]

3. What were the greatest sources of error in part B of this experiment?

   _______ of 4 pts
4. Calculate the expected pH of the buffer after the addition of 1.0 mL of 1M HCl. Remember that you are using 50 mL of the buffer, so be sure to calculate the moles of acetic acid and acetate in 50 mL of the buffer.

5. Calculate the expected pH of the buffer after the addition of 1.0 mL of 1M NaOH. Remember that you are using 50 mL of the buffer, so be sure to calculate the moles of acetic acid and acetate in 50 mL of the buffer.

6. Calculate the expected pH of the DI H₂O + 1.0 mL of 1M HCl
*(The final volume after the addition of NaOH is 51 mL)*

7. Calculate the expected pH of the DI H₂O + 1.0 mL of 1M NaOH
*(The final volume after the addition of NaOH is 51 mL)*
POST LAB QUESTIONS

1. The pH of a solution is measured to be 5.4. Calculate the following:
   (a) $[H^+]$
   
   (b) pOH
   
   (c) $[OH^-]$

2. Sodium fluoride (NaF) is a soluble salt.
   (a) From what acid and what base was the salt made?
   
   (b) Which ion (cation or anion) reacts with water in a reaction?
   
   (c) Write the net ionic equation for the reaction
   
   (d) Is the solution acidic, neutral, or basic?
   
   (e) Write the corresponding equilibrium expression (e.g. $K_b = \ldots$)

3. Consider the following weak acids and their $K_a$ values:

<table>
<thead>
<tr>
<th>Acid</th>
<th>$K_a$</th>
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<tbody>
<tr>
<td>Acetic Acid</td>
<td>$1.8 \times 10^{-5}$</td>
</tr>
<tr>
<td>Phosphoric Acid</td>
<td>$7.5 \times 10^{-3}$</td>
</tr>
<tr>
<td>Hypochlorous Acid</td>
<td>$3.5 \times 10^{-8}$</td>
</tr>
</tbody>
</table>

   What weak acid-conjugate base buffer system from the acids listed is the best choice to prepare the following buffers? Explain your reasoning
   
   (a) pH 2.8
   
   (b) pH 4.5
   
   (c) pH 7.5

   ______ of 3 pts
4. What is the resulting pH if 0.005 mol of NaOH is added to 100.0 mL of buffer containing 0.010 mol acetic acid (CH$_3$COOH, pK$_a$ = 4.74) and 0.010 mol sodium acetate (NaCH$_3$COO)? (Assume the final volume is 100 mL)

5. A buffer is prepared by adding 5.0 g of ammonia (NH$_3$) and 20.0 g of ammonium chloride (NH$_4$Cl) to enough water to form a 2.50 L solution. (Ammonia K$_b$ = 1.8 x 10$^{-5}$)
   (a) What is the pH of this buffer?

6. The carbonate buffer system is important in regulating blood pH levels. Carbonic acid is diprotic and therefore has two $K_a$ values:

   $\text{H}_2\text{CO}_3 \rightleftharpoons \text{H}^+ + \text{HCO}_3^-$ $K_{a1} = 4.3 \times 10^{-7}$
   $\text{HCO}_3^- \rightleftharpoons \text{H}^+ + \text{CO}_3^{2-}$ $K_{a2} = 5.6 \times 10^{-11}$

   Since the second dissociation has a $K_a$ value significantly smaller than that of the first dissociation, the second dissociation can be assumed to have no effect on the equilibrium.

   The pH of blood is 7.4. What is the ratio of carbonic acid to bicarbonate ion (HCO$_3^-$) in the blood?

   _____ of 3 pts

   Total _____ of 25 pts