**Lab Activity:**  Strong Acids & Strong Bases, Weak Acids & Weak Bases, and the Hydrolysis of Salts

Strong acids and strong bases completely ionize in water to make hydronium or hydroxide ions. Because the reactions essentially result in 100% ionization, the $[\text{H}_3\text{O}^+]$ and the $[\text{OH}^-]$ can be calculated directly from the initial concentrations of the acid or base. Consider nitric acid, $\text{HNO}_3$, and potassium hydroxide, $\text{KOH}$.

$$\text{HNO}_3 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{NO}_3^- \quad \text{KOH} \rightarrow \text{K}^+ + \text{OH}^-$$

$$[\text{H}_3\text{O}^+]_{\text{soln}} = [\text{HNO}_3]_0 \quad [\text{OH}^-]_{\text{soln}} = [\text{KOH}]_0$$

On the other hand, weak acids and bases do not fully dissociate or ionize when they are in water. The pH of a solution of a weak acid or weak base must be predicted based on their dissociation constants: $K_a$ or $K_b$. For an acid dissociation, consider hypochlorous acid, $\text{HClO}$.

$$\text{HClO} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{ClO}_2^-$$

This reaction does not go to completion, therefore $[\text{H}_3\text{O}^+]_{\text{soln}} \neq [\text{HClO}]_0$. Therefore, we must apply the equilibrium condition and use $K_a$ to calculate the $[\text{H}_3\text{O}^+]$ of the solution. The pH can then be calculated as the $-\log[\text{H}_3\text{O}^+]$.

A weak basic solution, like methylamine will not fully ionize water into hydroxide ions, but instead, an equilibrium is established:

$$\text{CH}_3\text{NH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{NH}_3^+ + \text{OH}^-$$

The $[\text{OH}^-]_{\text{soln}} \neq [\text{CH}_3\text{NH}_2]_0$. The hydroxide ion concentration must be calculated using $K_b$ and the application of the equilibrium condition. The hydronium ion concentration and the pH can then be calculated.

Some salts exhibit acidic or basic properties based on the identities of their cations or anions. Other salts are completely neutral when dissolved in water. Consider $\text{KNO}_3$. When $\text{KNO}_3$ is dissolved in pure water, the pH of the solution will be neutral (7). This is because $\text{K}^+$ is the conjugate acid of a very strong base, $\text{KOH}$. Therefore, the following reaction simply does not occur:

$$\text{K}^+ + \text{H}_2\text{O} \rightarrow \text{KOH} + \text{OH}^-$$

Because $\text{NO}_3^-$ is the conjugate base of a very strong acid, it also does not react with water:

$$\text{NO}_3^- + \text{H}_2\text{O} \rightarrow \text{HNO}_3 + \text{OH}^-$$

However, when the anion in a salt is the conjugate base of a weak acid, then the ion is itself a weak base, and will ionize slightly in water to make the solution basic. Consider potassium hypochlorite, $\text{KClO}$. As we have already discussed, potassium ions do not affect the pH of a solution. However, chlorite is the conjugate base of hypochlorous acid. Therefore, hypochlorite should be a weak base:

$$\text{ClO}^- + \text{H}_2\text{O} \rightleftharpoons \text{HClO} + \text{H}_2\text{O}$$

$K_b$ for the above reaction can be calculated from the $K_a$ for chlorous acid:

$$K_a \cdot K_s = K_w \quad K_s = K_w / K_a = 1 \times 10^{-14} / 3.5 \times 10^{-8} = 2.86 \times 10^{-7}$$
When the cation in a salt is the conjugate acid of a weak base, then the ion is a weak acid, and will dissociate slightly in water to make the solution acidic. Consider the salt methylammonium nitrate (CH$_3$NH$_3$NO$_3$). As already discussed, nitrate ions do not affect the pH of a solution. However, methylammonium ions are the conjugate acid of the weak base, methylamine. Methylammonium is a weak acid:

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

$$K_a = K_w / K_b = 1 \times 10^{-14} / 4.6 \times 10^{-4} = 2.2 \times 10^{-11}$$

When a salt is composed of a weak acid and a weak base, then you must compare the $K_a$ and the $K_b$ of the two ions to determine which will dominate. In the case of the salt methylammonium hypochlorite, CH$_3$NH$_3$ClO, $K_b(\text{ClO}^-) > K_a(\text{CH}_3\text{NH}_3^+)$. Therefore, a solution of this salt should be basic.

In some cases, a cation or anion may be amphiprotic – it can accept or donate a proton to solution. An example of such an amphiprotic salt is potassium dihydrogen phosphate, KH$_2$PO$_4$. Potassium undergoes no reaction with water. However, the dihydrogen phosphate has two possible reactions with water:

$$\text{H}_2\text{PO}_4^- + \text{H}_2\text{O} \rightleftharpoons \text{HPO}_4^{2-} + \text{H}_3\text{O}^+ \quad K_a = 6.2 \times 10^{-8}$$

$$\text{H}_2\text{PO}_4^- + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{PO}_4 + \text{OH}^- \quad K_b = 1.3 \times 10^{-12}$$

Both reactions occur. However, because the $K_a$ value is much larger than the $K_b$ value, the acid dissociation dominates and the pH of the resulting solution is expected to be acidic. The pH can be predicted on the basis of the first reaction alone (because $K_a >> K_b$).

Some metal ions are slightly acidic or basic when dissolved in water because they form hydrated complexes that may then gain or lose protons from water. Consider the cobalt (II) ion in water. It produces a slightly acidic solution.

$$\text{Co(H}_2\text{O)}_6^{2+} + \text{H}_2\text{O} \rightleftharpoons \text{Co(H}_2\text{O)}_5\text{OH}^+ + \text{H}_3\text{O}^+ \quad K_a = 1.3 \times 10^{-9}$$

**PROCEDURE**

For this activity, you will study a number of solutions to predict and analyze their pH. In some cases, you will prepare the solutions yourself. In others, you will simply take prepared solutions from a bottle. In each case:

1. PRELAB: Prepare a DATA TABLE in your LAB NOTEBOOK to record the pH paper and pH meter readings for each system, 1-12.
2. Predict the pH of the solution by calculation – considering the appropriate reaction with water and the $K_a$ or $K_b$ for the acid or base.
3. Collect or prepare the solution.
4. Measure the pH with a pH paper to get a preliminary pH reading.
5. Calibrate a pH meter for the appropriate acid (4-7) or base (7-10) range.
6. Measure the pH of the solution using the pH meter.
Data and Calculations: Acids, Bases, and Salts

System #1 - Deionized Tap water:

Predicted pH: __________ pH paper: __________ pH meter: __________

Briefly explain why your predicted pH does not match the pH meter reading:

System #2 – Boiled (10 minutes +) Deionized Tap water:

Predicted pH: __________ pH paper: __________ pH meter: __________

Briefly explain why your predicted pH does not match the pH meter reading:

System #3 – 0.10 M HCl solution:

Predicted pH: __________ pH paper: __________ pH meter: __________

System #4 – 0.10 M NaOH solution:

Predicted pH: __________ pH paper: __________ pH meter: __________
System #5: 0.10 \textit{M} acetic acid (HC$_2$H$_3$O$_2$)

Reaction with water:

pH Calculation:

Predicted pH: \_\_\_\_\_\_ pH paper: \_\_\_\_\_ pH meter: \_\_\_\_\_

System #6: 0.10 \textit{M} ammonia (NH$_3$)

Reaction with water:

pH Calculation:

Predicted pH: \_\_\_\_\_\_ pH paper: \_\_\_\_\_ pH meter: \_\_\_\_\_

System #7: 0.10 \textit{M} sodium acetate

Reaction with water:

pH Calculation:

Predicted pH: \_\_\_\_\_\_ pH paper: \_\_\_\_\_ pH meter: \_\_\_\_\_
System #8: 0.10 M ammonium chloride

Reaction with water:

pH Calculation:

Predicted pH: __________ pH paper: __________ pH meter: __________

System #9: 0.10 M ammonium acetate

Reaction with water:

pH Calculation:

Predicted pH: __________ pH paper: __________ pH meter: __________

System #10: 0.10 M sodium carbonate

Reaction with water:

pH Calculation:

Predicted pH: __________ pH paper: __________ pH meter: __________
System #11: 0.10 M sodium bicarbonate

Reaction with water:

pH Calculation:

Predicted pH: _________  pH paper: _________  pH meter: _________

System #12: 0.10 M iron (III) chloride  (source of Fe(H₂O)₆³⁺ complex)

Reaction with water:

pH Calculation:

Predicted pH: _________  pH paper: _________  pH meter: _________
1. Consider the pH values for sodium bicarbonate and sodium carbonate solutions in the experiment. Provide brief explanations for the following questions based on $K_a$ and/or $K_b$ values.

A) Why is the bicarbonate solution basic?

B) Why is the carbonate solution more basic than the bicarbonate solution?

2. Consider your calculation for System #5, what percent of the acetic acid is dissociated?

3. A) What would be the predicted pH of a 0.010 $M$ acetic acid solution? 
   B) What percent of the acid is dissociated?

4. Consider the following solutions of hydrochloric acid: 0.10 $M$ HCl and 0.010 $M$ HCl.

A. What are their predicted pH’s?  
   0.10 $M$ HCl = __________  
   0.010 $M$ HCl = __________

B. Note that the solution in question #3 is a 10-fold decrease in the concentration of the acid versus the 0.10 $M$ acetic acid solution that was studied in the experiment, just as the two HCL solutions in #4A differ by a factor of 10. Briefly explain why is there a more significant change in the pH of the two HCl solutions than in the two acetic acid solutions?