# ***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 $[H_3O_+]$ and the [OH−] can be calculated directly from the initial concentrations of the acid or base. Consider nitric acid, $HNO_3$, and potassium hydroxide, KOH.

$$HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^- \qquad\qquad KOH \rightarrow K^+ + OH^-$$

$$[H_3O^+]_{soln} = [HNO_3]_0 \qquad\qquad [OH^-]_{soln} = [KOH]_0$$

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: Ka or Kb. For an acid dissociation, consider hypochlorous acid, HClO.

$$HClO + H_2O \Leftrightarrow H_3O^+ + ClO_2^-$$

This reaction does not go to completion, therefore $[H_3O^+]_{soln} \neq [HClO]_0$. Therefore, we must apply the equilibrium condition and use Ka to calculate the $[H_3O^+]$ of the solution. The pH can then be calculated as the $-\log[H_3O^+]$

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

$$CH_3NH_2 + H_2O \Leftrightarrow CH_3NH_3^+ + OH^-$$

The $[OH^-]_{soln} \neq [CH_3NH_2]_0$. 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 $KNO_3$. When $KNO_3$ is dissolved in pure water, the pH of the solution will be neutral (7). This is because K+ is the conjugate acid of a very strong base, KOH. Therefore, the following reaction simply does not occur:

$$K^+ + H_2O \not\rightarrow KOH + H^+$$

Because $NO_3^-$ is the conjugate base of a very strong acid, it also does not react with water:

$$NO_3^- + H_2O \not\rightarrow NO_3^- + H_2O$$

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, 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:

$$ClO^- + H_2O \Leftrightarrow HClO + H_2O$$

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

$$K_a \bullet K_b = K_w \qquad\qquad K_b = K_w / K_a = 1 \text{ x } 10^{-14} / 3.5 \text{ x } 10^{-8} = 2.86 \text{ x } 10^{-7}$$

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 ($CH_3NH_3NO_3$). 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:

$$CH_3NH_3^+ + H_2O \Leftrightarrow H_3O^+ + CH_3NH_2$$

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

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, $CH_3NH_3ClO$, $K_{b(ClO-)} > K_{a(CH3NH3+)}$. 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, $KH_2PO_4$. Potassium undergoes no reaction with water. However, the dihydrogen phosphate has two possible reactions with water:

$$H_2PO_4^- + H_2O \Leftrightarrow HPO_4^{2-} + H_3O^+ \qquad K_a = 6.2 \text{ x } 10^{-8}$$

$$H_2PO_4^- + H_2O \Leftrightarrow H_3PO_4 + OH^- \qquad K_b = 1.3 \text{ x } 10^{-12}$$

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 Ka>>Kb).

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.

$$Co(H_2O)_6^{2+} + H_2O \Leftrightarrow Co(H_2O)_5OH^+ + H_3O^+ \qquad K_a = 1.3 \text{ x } 10^{-9}$$

**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. Predict the pH of the solution by calculation – considering the appropriate reaction with water and the Ka or Kb for the acid or base.
2. Collect or prepare the solution.
3. Measure the pH with a pH paper to get a preliminary pH reading.
4. Calibrate a pH meter for the appropriate acid (4-7) or base (7-10) range.
5. Measure the pH of the solution using the pH meter.

# Data and Calculations: Acids, Bases, and Salts

## System #1 - Deionzed Tap water:

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

*Explain why your predicted pH does or does not match the pH meter reading:*

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

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

## 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 M acetic acid ($HC_2H_3O_2$)

Reaction with water:

pH Calculation:

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

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## System #6: 0.10 M ammonia ($NH_3$)

Reaction with water:

pH Calculation:

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

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## System #7: 0.10 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: _________

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## System #9:  0.10 M ammonium acetate

Reaction with water:

pH Calculation:

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

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## 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: _________

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## System #12: 0.10 M iron (III) chloride (source of $Fe(H_2O)_6^{3+}$ complex)

Reaction with water:

pH Calculation:

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

# Post-Lab Questions: Intro to Acids, Bases, & Salts

1. Compare the observed pH values for sodium bicarbonate and sodium carbonate. In terms of $K_a$ and $K_b$, explain why these differ.

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

3. What would be the predicted pH of a 0.010 *M* acetic acid solution? What percent of the acid is dissociated?

4. The solution in question #3 is a 10-fold decrease in the concentration of the acid vs. 0.10 M acetic acid. 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. Why do the HCl solutions have higher pH's?

   C. Why is there a more significant change in the pH of the two HCl solutions than in the two acetic acid solutions?

# Calibration and Measurement

## A. Autocalibration with Two Buffers

1. Attach the pH electrode to the meter.
2. Choose either pH 4.01 and pH 7.00, or pH 7.00 and pH 10.01 buffers; whichever will bracket your expected sample range.
3. Press CAL key to initiate calibration sequence. CAL is displayed for two seconds. Press the YES key to accept the last calibration range (7-4 or 7-10) or select one of these calibration ranges using the SCROLL (▲,▼) keys. Press the YES key to accept. The 7 buffer annunciator will light up. Place electrode in pH 7.00 buffer. Reading will be displayed and updated as calibration continues. When the READY light comes on, indicating electrode stability, press the YES key to accept.
4. The pH 4 (or pH 10) annunciator will light up. Remove electrode from pH 7.00 buffer. Rinse with deionized water and place electrode in either pH 4.01 or 10.01 buffer (depending on calibration range that was selected). When the READY light comes on, press the YES key to accept buffer value. SLP will be displayed while the calculated slope is displayed.
5. The annunciators for the type of calibration performed will remain lit until another calibration is performed.
6. Meter will automatically go into the MEASURE mode. The READY light will come on to indicate electrode stability.
7. If operating in the LogR mode or using an ATC probe, the temperature-corrected pH reading is displayed.
8. Rinse electrodes with deionized water and place into sample. Record pH directly from the main meter display when the READY light is displayed or when electrode signal is stable.

> ***NOTE:*** *When making measurements in LogR mode, you must calibrate for temperature before attempting a pH calibration.* *See page 33 for temperature calibration procedure.*