EXPERIMENT 4

Acids and Bases

Introduction

Many common substances are either acids or bases. Some acids, like stomach acid are necessary for our health, while others, like sulfuric acid are dangerous and can cause burns and other injuries. Baking soda is a common, weak base used in our homes, while sodium hydroxide, a strong base, is hazardous to skin and eyes. Traditionally, solutions were labeled as being acidic or basic based on their taste and texture. Those that tasted sour were said to be acidic and solutions that tasted bitter and were slippery to the touch were said to be basic. Thus, substances such as lemon juice and vinegar were identified as acids, and solutions of lye and caustic soda as bases. However, in chemistry, we need to have a more definite concept of what makes an acid "acidic" and a base "basic". We also need a means of measuring the relative acidity of different materials. The acidity scale developed is measurement of acidity based on the logarithmic concentration of the hydronium ion, H_3O^+ and assigns a mathematical value to the degree of acidity of different compounds.

Strong Acids and Bases

In addition to the observed properties of tasting sour and producing a burning sensation when touched, acids also have the following chemical properties. First, they do not conduct electricity when pure, but are conductive when dissolved in water. Second, they react with metal and metallic compounds. Both of these observations are the result of an acid dissociating in an aqueous solution to produce hydrogen ions, H^+, which in water forms a hydronium ion H_3O^+. Any compound that donates a hydrogen ion is said to be an **acid**. The more **hydronium ion** that is formed, the more **acidic** the solution.

$$\mathbf{HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)}$$

Bases undergo a similar dissociation in water. When in an aqueous solution, these compounds dissociate to produce hydroxide ions, OH^-. Any compound that dissociates in water to produce **hydroxide ions** is said to be a **base**.

$$\mathbf{KOH\ (aq) \rightarrow K^+\ (aq) + OH^-\ (aq)}$$

The degree of dissociation determines the strength of the acid or base. Acids or bases that dissolve completely in water are designated as "**strong"**, while, those that do not completely dissociate from their anions are designated as "**weak"**. Strong acids produce more hydronium ions for a given concentration of acid than weak acids. **The higher the concentration of hydronium ions produced when adding an acid to water, the stronger the acid solution.** Like weak acids, weak bases produce smaller amounts of ions because of incomplete dissociation. Strong bases such as KOH or NaOH completely dissociate, producing large quantities of OH^- ions while weaker bases do not.

The production of hydroxide ion production affects the acidity of the solution by removing hydronium ion from the solution through an acid-base reaction. Each hydroxide ion reacts with one hydronium ion to form one molecule of water. Thus, for every hydroxide ion in solution, 1 hydronium ion is removed. As hydronium ions are removed from the solution, the hydronium ion concentration, $[H_3O^+]$, decreases and the solution becomes less acidic.

$$\mathbf{H_3O^+(aq) + OH^-\ (aq) \rightarrow 2H_2O\ (l)}$$

When we are discussing the acidity of a solution, we are looking at the formation of hydrogen ion in aqueous solutions only. Water is considered to be **amphoteric**; it can act as either an acid or a base. To understand this, consider the following dissociation of water:

$$H_2O\ (l) + H_2O\ (l) \leftrightarrows H_3O^+(aq) + OH^-\ (aq)$$

When 2 water molecules collide, a proton, H^+, is transferred from one H_2O molecule to the other form 1 hydronium ion, H_3O^+ and 1 hydroxide ion, OH^-. Thus, one water molecule functions as an acid and the other water molecule acts as a base. The equilibrium constant associated with this reaction is called a $\mathbf{K_w}$ **(equilibrium constant of water)** and experimentally was found to be $1.0x10^{-14}$ at 25 °C. Since equal numbers of H_3O^+ ions and $OH^–$ ions are produced in the solution, the concentrations of H_3O^+ and $OH^–$ ions at 25 °C are $1.0 \times 10^{–7}$ M each and the solution is neither acidic or basic, but rather it is "**neutral".**

$$\mathbf{K_w = [H_3O^+] \times [OH^-] = 1.0 \times 10^{-14}\ M}$$

$$\mathbf{[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}\ M}$$

As soon as an acid or a base is added to the water, the concentrations of H_3O^+ and OH^- change. To determine the exact degree of change relative to neutral water, we use a **pH scale**.

Acid Concentration and pH

The acidity of a solution is based on its hydrogen ion concentration. Unfortunately, these concentrations are not convenient to work with, as they commonly fall in the range of 0M to $1x10^{-14}$ M. A logarithmic scale, called a pH scale, was developed to more easily determine the relative acidity of solutions and make the numbers more manageable. **Thus, the pH of a solution is defined as the negative logarithm of [H_3O^+] or negative logarithm of [H_3O^+] in the solution**;

$$\mathbf{pH = -\log [H_3O^+]}$$

Therefore, if you know the pH of a solution, you can calculate the hydrogen ion concentration as follows;

$$\mathbf{[H_3O^+] = 10^{-pH}}$$

Similarly, the concentration of hydroxide ions, [OH^-], of a solution is commonly expressed in terms of the **pOH** of the solution, which is defined as the **negative logarithm of [OH^-]**

$$\mathbf{pOH = -\log [OH^-]}$$

The hydroxide ion concentration can be obtained from the pOH of the solution using the equation:

$$\mathbf{[OH\text{-}] = 10^{-pOH}}$$

Additionally, the pH and the pOH of any aqueous solution are related, as are the hydrogen and the hydroxide ion concentrations. The equations for these relationships are:

$$[H_3O^+] \times [OH^-] = 1.0 \times 10^{-14}$$

$$pH + pOH = 14$$

The following are 2 examples to show the relationships that exist between the pH, the pOH, the hydronium ion concentration and the hydroxide concentration.

Example (1)

What is the pH and pOH of a solution that contains 3.5 x 10^{-5} M hydronium ions?

$$pH = -\log [H_3O^+] = -\log (3.5 \times 10^{-5}) = 4.46$$

$$pOH = 14 - pH = 14 - 4.46 = 9.54$$

Example (2)

Calculate the hydronium ion and hydroxide ion concentrations of a solution that has a pOH of 4.40.

$$pH = 14 - pOH = 14 - 4.40 = 9.60$$

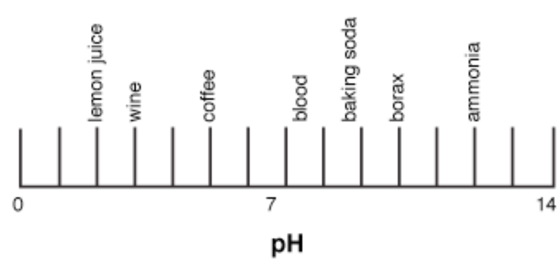
$$[H_3O^+] = 10^{-pH} = 10^{-9.60} = 2.5 \times 10^{-10}\ M$$

$$[OH^-] = 10^{-pOH} = 10^{-4.40} = 4.0 \times 10^{-5}\ M$$

In water, $[H_3O^+]$ is equal to 1.0 x 10^{-7} M, so the pH is 7.00. Because $[H_3O^+] = [OH^-]$ in water, it is neither acidic nor basic.

pH = 7.0 (neutral) **pH < 7.0 (acidic)** **pH > 7.0 (basic)**

The pH scale has a range of 0 (very strongly acidic, 1M HCl) to 14 (very strongly basic, 1M NaOH). Each unit of pH means there is a 10-fold difference in the hydronium ion concentration. For example, a 1.0M solution of HCl has a pH of 0.00 and a diluted solution of 0.10M HCl (1.0x10^{-1}M) has a pH of 1.00. The same relationship is also true for pOH and hydroxide concentration. Concentrations of hydronium ion above 1.0M can result in a pH less than zero and hydroxide concentrations above 1.0M will result in a pH greater than 14.



Weak Acids, Weak Bases

Unlike strong acids and bases that fully dissociate, many acids and bases do not completely dissociate into cations and anions when in a water solution. The degree of dissociation of an acid is governed by an equilibrium constant we call a $\mathbf{K_a}$. We call these compounds weak acids. Weak bases do not fully react in water either. The equilibrium constant for a base reaction in water is a $\mathbf{K_b}$. The chemical equations and equilibrium constant expressions for both a weak acid and a weak base are given below.

Acetic Acid: $HC_2H_3O_2 + H_2O\ (l) \leftrightarrows H_3O^+(aq) + C_2H_3O_2^-(aq)$ $\quad K_a = 1.8x10^{-5} = \frac{[H_3O^+][\,C_2H_3O_2^-\,]}{[\,HC_2H_3O_2]}$

Acetate ion: $C_2H_3O_2^-(aq) + H_2O\ (l) \leftrightarrows HC_2H_3O_2(aq) + OH^-(aq)$ $\quad K_b = 5.6x10^{-10} = \frac{[OH^-][\,HC_2H_3O_2]}{[\,C_2H_3O_2^-\,]}$

Very little of the original weak acid or base reacts with water. This means that the H_3O^+ and OH^- concentrations are **significantly less** than you would expect based on concentration alone. Thus, the pH of weak acids or bases are found toward the middle of the pH scale, rather than on either end. The pH of a weak acid (or weak base) solution will be based on the degree of dissociation of the weak acid to form H_3O^+ or OH^-. In a weak acid, the larger the degree of dissociation, the more H_3O^+, and the lower the pH. The same argument applies to weak bases, but involves OH^- and pOH.

The relationship between K_a and K_b

The pH of a solution is based on the equilibrium constant expression of water, K_w. The higher the hydronium ion concentration, the lower the hydroxide concentration must be to maintain the value of K_w.

$$K_w = [H_3O^+] \times [OH^-] = 1.0 \times 10^{-14}\ M$$

As we showed for strong acids, if you know the concentration of the hydronium ion, you can find the hydroxide ion using the K_w. You can do the same thing for weak acids and bases as well. The relationship between K_w, K_a and K_b is as follows:

$$K_w = K_a \times K_b$$

To demonstrate how this works, substitute the equilibrium constant expressions for both K_a and K_b:

$$K_w = \frac{[H_3O^+][\,\cancel{C_2H_3O_2^-}\,]}{[\,\cancel{HC_2H_3O_2}]}\ x\ \frac{[OH^-][\,\cancel{HC_2H_3O_2}]}{[\,\cancel{C_2H_3O_2^-}\,]} = 1.0x10^{-14}$$

When simplified everything except the hydronium ion and hydroxide cancel out.

$$K_w = [H_3O^+] \times [OH^-] = 1.0 \times 10^{-14}\ M = 1.8x10^{-5} \times 5.6x10^{-10} = 1.0 \times 10^{-14}\ M$$

For this Activity

In this online activity, you will be using a Phet simulation, created by the University of Colorado at Boulder, to obtain pH values for different concentrations of strong and weak acids and bases. You will then use these values to perform a series of calculations. If you have not yet covered this material in class, remember that you can reach out to your TA via email, or speak with any CHM 114 TA during their webex Office Hours, to ask questions and receive help.

Online Activity 2: Experimental Procedures and Data Sheet

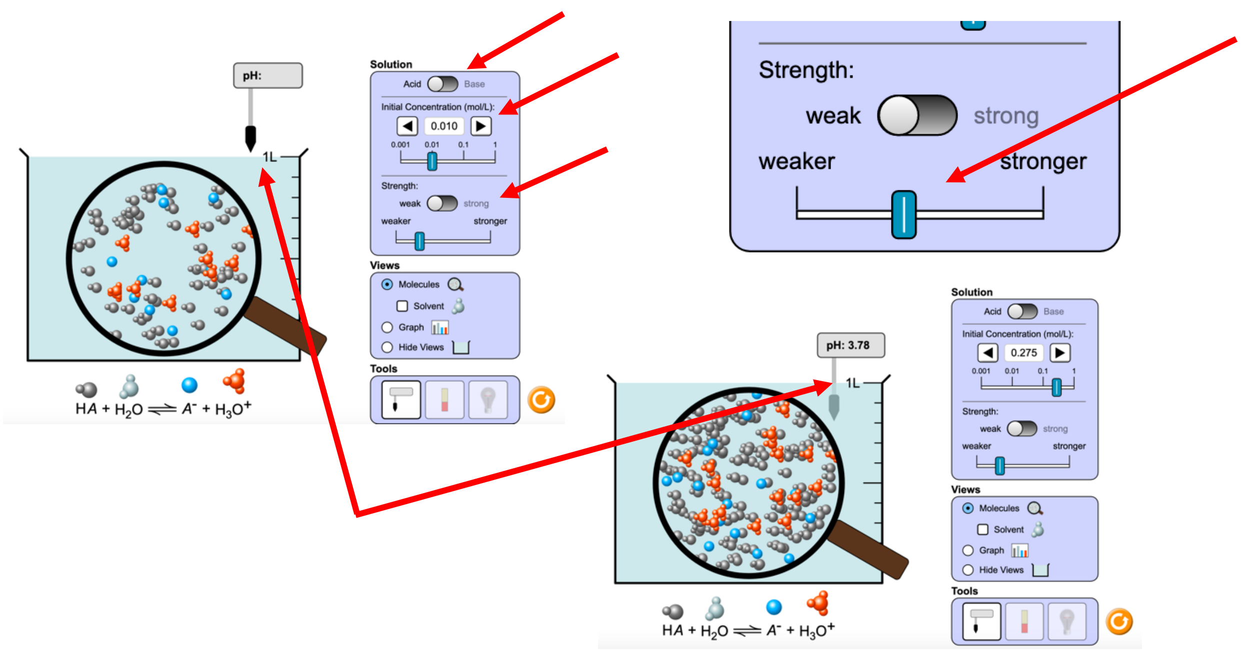
Submit as part of your online activity report

Name: ____ Date: ________ Section: ________

Record all measurements with the correct number of significant figures (all digits provided by simulation) and units.

In order to determine the pH of the strong and weak acid and base solutions that you need for your calculations, you will use the Phet simulation Acid–Base Solutions: https://phet.colorado.edu/en/simulation/acid-base-solutions .

Once you have opened the simulation, click on my solution. Select either acid or base and strong or weak, depending on what you are "testing" then use the black arrows or blue bar to set the concentration. For acetic acid and the acetate ion, when you put the strength selector toggle button on weak, you will also need to move the blue bar under the strength selector so that it is approximately underneath the circle in the strength selector toggle button (see image on top right). When you select strong, the blue adjustment bar will disappear. Lower the pH probe into the solution to get the pH



pH of 0.100M $HC_2H_3O_2$: ________________ pH of 0.100M HCl: ________________

pH of 0.010M $HC_2H_3O_2$: ________________ pH of 0.010M HCl: ________________

pH of 0.100M $NaC_2H_3O_2$: ________________ pH of 0.100M NaOH: ________________

pH of 0.010M $NaC_2H_3O_2$: ________________ pH of 0.010M NaOH: ________________

Grading

Points

Neatness and Clarity of Data	5pts	__________pts
Significant figures and units	5pts	__________pts
All data is present	10pts	__________pts

Deductions (sliding based on TA discretion)

Lab area left unclean	20pts	__________pts
Improper waste disposal	20pts	__________pts
Disruptive behavior	20pts	__________pts
Other: ______________________________________		__________pts

Plagiarism!!! Data are identical to another student 100pts __________pts

Grade for Experimental Procedures and Data ________pts

Online Activity 2: Results Table
Submit as part of your online activity report

Name: ____ Date: ________ Section: ________

TA Signature: _______________________________________ **(after calculations are done!)**

All data must be written in pen at the time it is collected. **Pencil is not allowed!!**
Record all results with the correct number of significant figures and units
No stray marks or notes should be present on this page. Only the tabulated results are allowed

All calculations are to be done during class!!!

Table 1: Equilibrium concentrations from pH	0.100M HCl	0.010M HCl	0.100M $HC_2H_3O_2$	0.010M $HC_2H_3O_2$
$[H_3O^+]$				
pOH				
$[OH^-]$				

	0.100M NaOH	0.010M NaOH	0.100M $NaC_2H_3O_2$	0.010M $NaC_2H_3O_2$
$[H_3O^+]$				
pOH				
$[OH^-]$				

Table 2: Strong acid/strong base calculations	0.100M HCl	0.010M HCl	0.100M NaOH	0.010M NaOH
pH				
pOH				

Table 3: Weak acid calculations	0.100M $HC_2H_3O_2$	0.010M $HC_2H_3O_2$
Theoretical pH based on initial concentration		
Initial concentration based on experimental pH		

Table 4: Weak base calculations	0.100M $NaC_2H_3O_2$	0.010M $NaC_2H_3O_2$
Theoretical pH based on initial concentration		
Initial concentration based on experimental pH		

Grading

<u>Points</u>

Significant figures and units	5pts	__________pts
Table is neat and legible	5pts	__________pts
All results are present	10pts	__________pts

<u>Deductions (sliding based on TA discretion)</u>

Results do not make sense	20pts	__________pts
Results do not match data	20pts	__________pts
Other: ____		__________pts

Plagiarism!!! Results are identical to another student 100pts __________pts

Grade on results table __________pts

Calculations Section: Submit as part of your online activity report

Part 1: Calculations when given the pH of a solution:

These calculations work for both strong and weak acids and bases since the calculations are based on the equilibrium concentrations of hydronium ion and hydroxide ion. Since strong acids and bases fully dissociate, the initial concentration of the given acid or base is the same as the equilibrium concentration.

Calculating the [H_3O^+] from pH

Use the measured pH to calculate the hydronium ion concentration for each of the 8 solutions. For the correct number of significant figures, look at the number of significant figures to the right of the decimal in the pH value. That will be the correct number of significant figures to use when converting to molarity. The number in front of the decimal in a pH value is an estimate of the exponent, not an actual significant figure.

$$[H_3O^+] = 10^{-pH}$$

Example: pH= 5.00 then [H_3O^+] = 1.0x 10^{-5}

[H_3O^+] of 0.100M HCl: ________________

[H_3O^+] of 0.100M $HC_2H_3O_2$: ________________

[H_3O^+] of 0.010M HCl: ________________

[H_3O^+] of 0.010M $HC_2H_3O_2$: ________________

[H_3O^+] of 0.100M NaOH: ________________

[H_3O^+] of 0.100M $NaC_2H_3O_2$: ________________

[H_3O^+] of 0.010M NaOH: ________________

[H_3O^+] of 0.010M $NaC_2H_3O_2$: ________________

pOH of All Solutions

pOH is often used instead of pH when working with basic solutions. Convert measured pH values to pOH.

pOH + pH = 14

(14 is considered to be an exact value so is not used in estimating significant figures)

pOH of 0.100M HCl: ________________

pOH of 0.100M $HC_2H_3O_2$: ________________

pOH of 0.010M HCl: ________________

pOH of 0.010M $HC_2H_3O_2$: ________________

pOH of 0.100M NaOH: ________________

pOH of 0.100M $NaC_2H_3O_2$: ________________

pOH of 0.010M NaOH: ________________

pOH of 0.010M $NaC_2H_3O_2$: ________________

$[OH^-]$ of All Solutions

Use the pOH to calculate the hydroxide ion concentration from the approximate $[OH^-]$ values for each solution in this section. Determine the number of significant figures in the same way as you did the $[H_3O^+]$.

$$[OH^-] = 10^{-pOH}$$

$[OH^-]$ of 0.100M HCl: ________________ $[OH^-]$ of 0.100M $HC_2H_3O_2$: ________________

$[OH^-]$ of 0.010M HCl: ________________ $[OH^-]$ of 0.010M $HC_2H_3O_2$: ________________

$[OH^-]$ of 0.100M NaOH: ________________ $[OH^-]$ of 0.100M $NaC_2H_3O_2$: ________________

$[OH^-]$ of 0.010M NaOH: ________________ $[OH^-]$ of 0.010M $NaC_2H_3O_2$: ________________

Part 2: Calculations when given concentrations of strong acids and bases

pH and pOH from concentration

Since strong acids and bases fully dissolve in solution, you can calculate the hydronium or hydroxide ion concentration directly from the calculation below. You CANNOT do this with weak acids or bases due to incomplete dissociation, so you must know if the acid is strong or weak before calculating pH.

$$pH = -\log [H_3O^+] \text{ and } pOH = -\log [OH^-]$$

$$K_w = [H_3O^+] \times [OH^-] = 1 \times 10^{-14}\ M \text{ so } pK_w = 14$$

so 14= pH + pOH (a "p" before the symbol means -log of that value)

Calculate the pH and pOH from each concentration given for each strong acid and strong base solution. Use the number of significant figures in the concentration as the number of significant figures AFTER the decimal in the pH or pOH.

Ex: $[H_3O^+] = \underline{1.00} \times 10^{-3}$ **so pH = 3.<u>000</u> and pOH = 11.<u>000</u>**

pH of 0.100M HCl: ________________ pOH of 0.100M HCl: ________________

pH of 0.010M HCl: ________________ pOH of 0.010M HCl: ________________

pH of 0.100M NaOH: ________________ pOH of 0.100M NaOH: ________________

pH of 0.010M NaOH: ________________ pOH of 0.010M NaOH: ________________

Part 3: Calculations for weak acids

pH and pOH from the initial concentration of a weak acid

Since weak acids do not fully dissociate, you must use an ICE table and the equilibrium constant to find the hydronium ion concentration of the solution. The equilibrium constant for an acid is called the K_a and for acetic acid it has a value of $1.8x10^{-5}$. Y is the given concentration of the weak acid. Once you know the hydronium ion concentration (x), you can calculate the pH.

	$HC_2H_3O_2 + H_2O\ (l) \leftrightarrows$	$H_3O^+(aq)$	$+\ C_2H_3O_2^-(aq)$
Initial	**Y**	0	0
Change	-**1**x	+**1**x	+**1**x
Equilibrium	(Y-1x)	(0+1x)	(0+1x)

$$K_a = 1.8x10^{-5} = \frac{[H_3O^+][\,C_2H_3O_2^-\,]}{[\,HC_2H_3O_2]} = \frac{[x][\,x\,]}{[\,Y - x]}$$

If the hydronium ion concentration, x, is less than 5% of the value of the acetic acid, Y, you can ignore x in the initial concentration of the acetic acid, and Y- x just becomes Y. If it is more than 5%, then you would have to solve a quadratic equation since x is not negligible. Once you have solved for the hydronium ion concentration, (x), you can solve for pH.

pH of 0.100M $HC_2H_3O_2$: _______________ (Y= 0.100M)

pH of 0.010M $HC_2H_3O_2$: _______________ (Y= 0.010M)

Fact check: For an acetic acid solution with an initial concentration of 0.050M, your pH should be 3.02

Initial concentration of a weak acid from pH

In this case if you know the pH of a solution, you can use the same ICE table you used previously only instead of solving for x, you will solve for Y, the unknown concentration of the acid solution. Calculate the equilibrium H_3O^+ concentration (x) from the pH and solve for Y. Use the pH you measured in lab to determine the initial concentration of each of the weak acid solutions.

Experimental initial concentration of 0.100M $HC_2H_3O_2$ based on pH: _______________ (solve for Y)

Experimental initial concentration of 0.0100M $HC_2H_3O_2$ based on pH: _______________ (solve for Y)

Fact check: For an acetic acid solution with a pH of 3.02, the initial concentration of the weak acid should be 0.050M.

Part 4: Calculations for weak bases

pH from the initial concentration of a weak base

Since weak bases do not fully dissociate, you must use an ICE table and the equilibrium constant to find the hydronium ion concentration of the solution. The equilibrium constant for a base is called the K_b and for acetic acid it has a value of $5.6x^{-10}$. Once you know the hydroxide ion concentration (x), you can calculate the pH. Note that any sodium salt is always fully soluble so the concentration of the acetate ion is the same as that of the sodium acetate solution.

	$C_2H_3O_2^-$ + H_2O (l) ⇆	OH^- (aq) +	$HC_2H_3O_2$ (aq)
Initial	**Y**	0	0
Change	-**1**x	+**1**x	+**1**x
Equilibrium	(Y-1x)	(0+1x)	(0+1x)

$$K_b = 5.6x10^{-10} = \frac{[x][x]}{[Y-x]}$$

If the hydroxide ion concentration, x, is less than 5% of the value of the acetate ion, Y, you can ignore x in the initial concentration of the acetate ion, and (Y- x) just becomes Y. If it is more than 5%, then you would have to solve a quadratic equation since x is not negligible. Once you have solved for the hydroxide ion concentration, (x), you can solve for pOH and then pH

pOH of 0.100M $NaC_2H_3O_2$: _______________ pH of 0.100M $NaC_2H_3O_2$: _______________

pOH of 0.010M $NaC_2H_3O_2$: _______________ pH of 0.010M $NaC_2H_3O_2$: _______________

Fact check: For a sodium acetate solution with an initial concentration of 0.050M, your pH should be 8.72.

Initial concentration of a weak base from pH

In this case if you know the pH of a solution, you can use the same ICE table you used previously only instead of solving for x, you will solve for Y, the unknown concentration of the base solution. Calculate the equilibrium OH^- concentration (x) from the pH and then use the ICE table to solve for Y.

$$pOH = 14 - pH \text{ and } [OH^-] = 10^{-pOH} = x$$

Use the pH you measured in lab to determine the initial concentration of each of the weak base solutions.

Experimental initial concentration of 0.100M $NaC_2H_3O_2$ based on pH: _______________

Experimental initial concentration of 0.010M $NaC_2H_3O_2$ based on pH: _______________

Fact check: For an sodium acetate solution with a pH of 8.72, the initial concentration of the weak base should be 0.050M.

Part 5: Calculations

Online Activity 2: Additional Questions
Submit as part of your online activity report

Name: ________________________________ Date: ________ Section: ________ Grade: ________

Record all values with the correct number of significant figures and units.
Place all answers on the line next to the question.
Show calculations for any numerical answers.

1) Calculate the concentration of the hydronium ion in a solution with a pH of 2.5. Include units in your answer.

____

2) Write the balanced chemical equation describing the reaction that occurs when acetic acid is dissolved in water to make an aqueous solution. Use the correct arrow notation.

____

3) Write the K_a expression for acetic acid.

____

4) Write the chemical equation for hydrochloric acid ionizing in water. Underline the spectator ion that results when the strong acid ionizes in solution.

____

5) Write the chemical equation for the weak base $NaC_2H_3O_2$ in water. (would be represented by K_b)

____

6) Write the K_b expression for $NaC_2H_3O_2$.

____