

Acids, Bases, Salts, and Buffers

Investigation questions Parts 1 and 2

- What is pH and how is it related to the identity and concentration of the substance in a solution?
- What is a salt?

Introduction

I. Brønsted-Lowry acid-base theory

The Brønsted-Lowry concept of acids and bases was introduced by Johannes Brønsted and Thomas Lowry in 1923, and led to an understanding of many *proton transfer reactions* observed to occur in both non-aqueous and aqueous solutions. In this concept, acids are defined as substances that can donate a proton and bases are substances that can accept a proton. Accordingly, the reaction between an acid and base is described as follows:



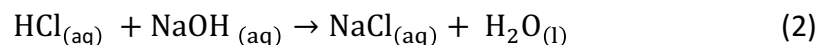
In the reaction above the reaction products are A^- and BH^+ . The ion A^- is called the *conjugate base* of HA because it has donated a proton (H^+) to the base B. Likewise, BH^+ is the *conjugate acid* of B since it has accepted the proton from HA. The substances HA and A^- are called a *conjugate acid-base pair*. Likewise, BH^+ and B are also a conjugate acid-base pair.

II. Acid-Base Reactions

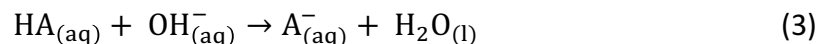
Acids and bases form ions when dissolved in water. These ions are free to move in solution and thus the solutions conduct electricity. Acids or bases that completely ionize produce a large number of ions and are called **strong electrolytes**. Hydrochloric acid is one such

example. **Weak electrolytes** (e.g., acetic acid) only slightly ionize and **non-electrolytes** do not ionize at all when dissolved in water.

Aqueous solutions of acids and bases react to form the corresponding salt and water. When equal moles of monoprotic acid and monobasic base react, it is called “neutralization reaction”. For example:



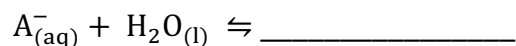
The salt of a weak acid is formed when a weak acid reacts with a base. The net ionic equation for the reaction is:



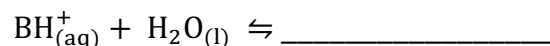
The salt of a weak base is formed when a weak base reacts with an acid. The net ionic equation is:



When a salt of weak acid or a weak base is dissolved in water, the ions formed may undergo reaction with water. The ions that can react with water are the conjugate base of the weak acid and the conjugate acid of the weak base. Refer to equations (3) and (4) and predict the products of the following equilibrium equations:



When the conjugate acid of the weak base reacts with water,



III. Calculations: pH and pOH

The term pH is a mathematical function related to the hydronium ion (H_3O^+) concentration in the solution. This function, which is given below, says that the pH of a substance is the negative base 10 logarithm of the H_3O^+ concentration. A solution that has H_3O^+ concentration of 0.1 M has a pH of 1. The table gives examples of the pH of solutions at various H_3O^+ concentrations.

$$\text{pH} = -\log_{10}[\text{H}_3\text{O}^+] \quad (5)$$

Table 1 pH and Hydronium Ion Concentration

$[\text{H}_3\text{O}^+]$	pH	$[\text{OH}^-]$
0.1 M	1	0.00000000000001 M
0.01 M	2	0.0000000000001 M
0.001 M	3	0.000000000001 M
0.0001 M	4	0.0000000001 M
0.00001 M	5	0.000000001 M
0.000001 M	6	0.00000001 M
0.0000001 M	7	0.0000001 M
0.00000001 M	8	0.000001 M
0.000000001 M	9	0.00001 M
0.0000000001 M	10	0.0001 M
0.00000000001 M	11	0.001 M
0.000000000001 M	12	0.01 M
0.0000000000001 M	13	0.1 M

Table 1 shows how the pH of a solution increases as the H_3O^+ concentration decreases. The neutral point is at pH 7, at which the H_3O^+ concentration is 1×10^{-7} (0.0000001) M. This is an extremely small concentration of H_3O^+ . You might have thought the H_3O^+ concentration would be zero at the neutral point but this is not the case. The number of ions produced in this dissociation is far too small to conduct any electrical current so water is a non-electrolyte.

Water, although it is a non-electrolyte, dissociates into hydronium ions and hydroxide ions. An equation can be written for this dissociation:



The equation is written with arrows in both directions indicating that the reaction is reversible. The equilibrium constant expression for this reversible reaction is written as follows.

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] \quad (7)$$

The extent of dissociation is extremely small. At room temperature, the value for the dissociation constant for water is 1.0×10^{-14} . The H_3O^+ concentration and the hydroxide ion (OH^-) concentration are both the same because of the stoichiometry of the dissociation reaction. The concentration of the H_3O^+ in water at room temperature is as follows.

$$[\text{H}_3\text{O}^+] = \sqrt{K_w} \quad (8)$$

The solution becomes basic as you move past the neutral point to pH values greater than 7. At pH 12, the solution is extremely basic yet the H_3O^+ concentration is not zero. It is extremely small though, 1×10^{-12} M. At room temperature, the product of the H_3O^+ concentration and the OH^- concentration in the solution will always equal 1.0×10^{-14} . We can also use a logarithmic function to represent the H_3O^+ concentration in the solution, as it will also vary over this wide range of values. The pOH of a solution is the negative base 10 logarithm of the OH^- concentration.

$$\text{pOH} = -\log_{10}[\text{OH}^-] \quad (9)$$

The equilibrium constant for the dissociation of water can also be expressed logarithmically:

$$\text{p}K_w = -\log_{10} K_w \quad (10)$$

Using the fact that addition of the logarithms of two numbers gives the logarithm of the product of the two numbers, we can write an expression for the equilibrium constant for the dissociation of water in logarithmic form as follows.

$$pK_w = \text{pH} + \text{pOH} \quad (11)$$

The value for pK_w when the water is at room temperature is 14. The expression can be written as follows for water at room temperature.

$$14 = \text{pH} + \text{pOH} \quad (12)$$

IV. Buffers

A buffer is a solution composed of a weak acid (HA) and the salt of that weak acid (i.e., the conjugate base, A^-), or a weak base (B) and the salt of that weak base (i.e., the conjugate acid BH^+). Buffer systems resist large pH changes because added acids or bases are neutralized by the existing HA/ A^- system (equilibrium). When a strong acid is added to a buffer system, the “basic” component of the buffer (A^-) reacts with it in an attempt to “neutralize” it. When a strong base is added, the “acidic” component (HA) of the buffer reacts with it to neutralize it. Similarly, addition of small amounts of base would slightly increase the A^- concentration.

Buffers are used in a variety of ways. Some of the more important ones are biological and biochemical research and medicinal chemistry. In fact, it is often necessary for the researcher to prepare a buffer of certain pH to support the chemical conditions the system needs to maintain.

Buffers are most effective in the pH range of ± 1 unit to either side of the pK_a , or the pOH (i.e., $14 - \text{pH}$) range ± 1 unit to either side of the pK_b of the desired system. Furthermore, the closer the concentrations of the weak acid and its conjugate weak base salt or the weak base and its conjugate weak acid salt are to one another, the better the chance of achieving greater buffering effect.

V. Calculations: pK_a and pK_b

Acidic buffers consist of a weak acid in equilibrium with its salt (conjugate base) in an aqueous environment. The corresponding chemical equation is:



Notice that this is an equilibrium and, as such, an equilibrium expression can be written:

$$K_{eq} = \frac{[A^-][H_3O^+]}{[HA][H_2O]} \quad (14)$$

Since the reaction occurs in water, water is in such excess that its concentration is considered constant and is incorporated into the constant, which is renamed K_a . K_a is called the “acid dissociation constant”.

$$K_a = \frac{[A^-][H_3O^+]}{[HA]} \quad (15)$$

Often, equimolar amounts of weak acid and its conjugate base are used to make a buffer. In this case, $[A^-]$ equals $[HA]$ and equation (15) becomes:

$$K_a = [H_3O^+] \quad (16)$$

Remembering the definition of pH allows us to write:

$$pK_a = \text{pH} \quad (17)$$

where pK_a is equal to $-\log K_a$. This equation shows us that we can figure out what buffer system to use for a desired pH if we know the acid dissociation constant for the weak acid in each system.

A similar series of equations can be used when dealing with basic buffers. In this case, when equimolar amounts of the weak base and its conjugate acid are used, the resulting equations are:

$$K_b = [OH^-] \quad (18)$$

and $pK_b = \text{pOH}$

Keep in mind that, for basic buffers, you are dealing with pOH and will have to use the relationship between pOH and pH to get the desired target pH.

Part 1: Properties of acids and bases

Materials

Conductivity meter
pH meter
30-mL beaker
10-mL graduated cylinder
100-mL graduated cylinder
Graph paper with one logarithmic axis (provided)

Chemicals and Safety Information

Compound	Potential Hazards	Precautions
Hydrochloric acid HCl (aq)	Corrosive. Causes burns and severe respiratory tract irritation	Wear goggles and gloves, use proper ventilation
Acetic acid CH ₃ COOH (aq)	May be harmful if swallowed. Causes eye and skin burns.	Wear goggles and gloves
Formic acid HCOOH (aq)	May be harmful if swallowed. Causes eye and skin burns.	Wear goggles and gloves
Nitric acid HNO ₃ (aq)	Corrosive. May cause severe eye and skin irritation with possible burns	Wear goggles and gloves
Propylamine CH ₃ CH ₂ CH ₂ NH ₂ (aq)	May be harmful if inhaled. Extremely destructive to the tissue of the mucous membranes and upper respiratory tract. Toxic if absorbed through skin. Causes skin and eye burns.	Wear goggles and gloves, use proper ventilation
Potassium hydroxide KOH (aq)	Corrosive. Causes eye, skin, respiratory and digestive tract burns	Wear goggles and gloves, use proper ventilation
Sodium hydroxide NaOH (aq)	Corrosive. Causes eye, skin, respiratory and digestive tract burns	Wear goggles and gloves, use proper ventilation

Ammonia (i.e., Ammonium Hydroxide) NH ₃ (aq) [i.e., NH ₄ OH]	Toxic. Corrosive. Causes eye, skin, respiratory and digestive tract burns	Wear goggles and gloves, use proper ventilation
Ethyl alcohol (Ethanol) CH ₃ CH ₂ OH (l)	Flammable liquid. Causes severe eye irritation. May cause blindness if swallowed. May be fatal if swallowed.	Wear goggles and gloves, use proper ventilation

For more information on safety concerns, go to <http://avogadro.chem.iastate.edu/MSDS/>.

Part 1A. Conductivity Measurements of Acids and Bases

Background

As you may recall, some substances are strong electrolytes, meaning they dissociate into ions that are capable of conducting electrical current in solution. Other substances dissociate to some limited extent producing a few ions in solution but not enough to conduct electrical current well. These substances are generally identified as weak electrolytes. Still other substances do not dissociate to any significant extent and are not capable of conducting electrical current at all. These substances are known as non-electrolytes.

Substances:

Use all nine that are listed above, and distilled water.

Procedure

- Test all the solutions with the conductivity probe. For details on the operation of the conductivity probe, see the Pictorial Guide to the Laboratory, which can be found on Canvas.
- For this test, please insert the probe DIRECTLY in the bottle of the standard solution provided. You will find the standards next to the conductivity meter.

Part 1B. pH and H₃O⁺ concentration

Investigate:

In this part you will predict and measure the actual pH of strong and weak electrolyte solutions using the pH meter. You will then explore the effect of changing concentration on the pH of these solutions.

Substances:

You will use same substances as part 1A, except for the solutions that were identified as non-electrolytes.

Experimental Procedure

Predict the pH of the solutions

- Use the concentration on the bottle labels to predict the pH of each of the solutions identified as strong or weak electrolytes earlier in your investigation. If you are unsure how to do this, begin by reviewing the "Calculations" section of the Introduction.
- Plot the predicted pH vs concentration on 3-cycle semi-log paper for each solution. Use different marks for each solution as you plot the predicted pH values.

Graphing Help: The numbers for the pH axis should be placed along the linear axis of the paper and should range from 0 to 14. The values for the concentration should be plotted along the logarithmic axis of the paper. Each cycle of the logarithmic axis represents an order of magnitude (a power of 10), thus 3-cycle semi-log paper is used to plot numbers that vary over 3 orders of magnitude. Logical labeling of the logarithmic axis has the numbers increasing along the positive direction of the axis. The first cycle begins at 0.0010 M, the second at 0.010 M, and the third at 0.10 M in this experiment.

Calibrate the pH meter

- Refer to the instructions for your pH meter found under pH Measurement in the Pictorial Guide. Use the instructions for calibration with two buffers.

Measure the pH of each solution

- Take a sample of one of the strong or weak electrolyte solutions (not the non-electrolytes) in a 30-mL beaker. You will need at least 15 mL, but not more than 20 mL. Specify in your ELN which electrolyte you picked!
- Rinse the electrode with distilled water and place it in the solution. If the pH meter has a temperature probe, rinse it and place it in the solution as well. Allow 60 seconds for the electrode and temperature probe to stabilize. Read the pH of the solution.
- Plot this pH on the semi-log paper using the same marks for the different solutions as before. Check the measured pH with the predicted pH.

Make accurate 10-fold dilutions of the solution.

Use the sample in the 30-mL beaker. Measure out 10 mL of that solution in a 10-mL graduated cylinder.

- Pour that 10-mL sample into a 100-mL graduated cylinder and add distilled water to make 100 mL of solution.
 - Calculate the new concentration and predict the pH of this solution.
 - Measure the pH and plot the actual value and the predicted value.
- Make another accurate 10-fold dilution using the solution whose pH you have just measured.
 - Calculate the new concentration and predict the pH of this solution.
 - Measure the pH and plot the actual value and the predicted value.

Waste Disposal

Dispose of the products by pouring them into the proper waste container.

Analysis Questions Part 1

Answer these questions in your ELN before moving on to Part 2 of this experiment. Your TA may choose to do this as a class discussion.

Show data tables and balanced chemical equations for full credit.

1. Which solutions are strong, weak, or non-electrolyte?
2. What is the balanced chemical equation representing each of the solutions you explored?
3. Which ions and/or molecules are present in the greatest concentration in each solution?

In particular:

4. Which of the acids or bases you tested are strong electrolytes?
5. Which of the acids and bases you tested are weak electrolytes? For those solutions:
 - a) Write balanced equations for the dissociation of each acid into a hydronium ion, H_3O^+ , and the appropriate corresponding anion for the acid.
 - b) Write balanced equations for the dissociation of each base into a hydroxide ion, OH^- , and the appropriate corresponding cation for the base.
6. Are these reactions reversible? For those that are, write the equilibrium constant expressions.

Post Discussion Questions Part 1

- For the solutions of acids or bases that are strong electrolytes:
 - What did you observe as you made the measurement of the pH and compared it with the predicted pH of the solution? How well do these numbers agree? Why or why not?
- Are the conductivity and pH observations consistent?
- What do the conductivity and pH results for strong acids and strong bases tell you about the extent of dissociation of these compounds?
- For the solutions of acids or bases that are weak electrolytes:
 - What did you observe as you made the measurement of the pH and compared it with the predicted pH of the solution? How well do these numbers agree? Why or why not?
- Are the conductivity and pH observations consistent? What do the conductivity and pH results for weak acids and weak bases tell you about the extent of dissociation of these compounds?

Part 2: Analyzing acid-base properties of salt solutions

Materials

conductivity meter
pH meter
13x100-mm test tubes
plastic droppers
100-mL beaker
hot plate
microspatula
another small beaker

Chemicals and Safety Information

Compound	Potential Hazards	Precautions
sodium chloride NaCl (aq) and (s)	None	None
potassium nitrate KNO ₃ (aq) and (s)	May cause skin and respiratory tract irritation	Wear goggles and gloves
sodium formate NaHCOO (aq) and (s)	May cause severe burns	Wear goggles and gloves
sodium acetate NaCH ₃ COO (aq) and (s)	May cause mild irritation	Wear goggles and gloves
ammonium nitrate NH ₄ NO ₃ (aq) and (s)	May cause skin irritation. Avoid inhalation.	Wear goggles and gloves, use proper ventilation
propylammonium chloride CH ₃ CH ₂ CH ₂ NH ₃ ·Cl (aq) and (s)	May be harmful if inhaled. Causes respiratory tract irritation. Causes eye and skin irritation.	Wear goggles and gloves, use proper ventilation

Investigate

Some ions react with water to produce H₃O⁺ or OH⁻ while others do not. The reaction of these ions with water, producing H₃O⁺ or OH⁻, is called hydrolysis. The reaction is the result of the acidic or basic nature of the ion involved.

As a class, conduct a series of measurements using the salts listed above to answer these questions. You should use both the conductivity probe and the pH meter, as you learned to do in Part 1.

Part 2A. Conductivity Measurements of Salts

Procedure

Perform these measurements in the same way that you did in Part 1. Solutions are provided for you. DO NOT USE THESE SOLUTIONS FOR PART 2B!

Part 2B. pH Measurements of Salts

Procedure

pH observations for water

Recall the equation for the dissociation of water into hydronium ions and hydroxide. The equilibrium constant for this dissociation reaction is 1.0×10^{-14} .

- Obtain a 1 mL sample of deionized water in a small test tube and observe its pH by adding a drop of Universal Indicator Solution. Compare the color of the solution with the colors given on the posted Universal Indicator chart to determine the pH of your deionized water.
- Fill your 100-mL beaker about half full with deionized water and add a few drops of Universal Indicator. Boil the water to expel dissolved CO₂, until you observe the color that indicates a neutral pH. Cool enough to handle safely.

pH observations for salt solutions

Work quickly so your boiled water does not reabsorb CO_2 . If your results do not make sense, check the pH of your water and re-boil if necessary. **DO NOT USE THE SALT SOLUTIONS PROVIDED FOR CONDUCTIVITY MEASUREMENTS. These solutions have not been boiled and will provide invalid measurements which will cause you to lose points in your ELN report.**

- c) To test a salt, use a plastic dropper to transfer about 1 mL of the boiled, deionized distilled water with indicator to a 13 x 100-mm test tube.
- d) Test all the salts at the same time using a rack with six test tubes.
- e) **Do not put your microspatula into the reagent bottle.** Pour a small amount of the salt onto the bottle lid. Use the lid to transfer the amount you want to use to a small, dry beaker. Pour the salt that remains on the lid back into the reagent bottle and replace the lid. Transfer salt from the beaker to the test tube with your microspatula. Make sure you add more than a few crystals.
- f) Flick the test tube lightly with your finger to mix and dissolve the salt. Observe the color and compare it to the color on the posted Universal Indicator chart.

Analysis Questions Part 2

1. What is the expected pH of water at room temperature?
2. Do the salt solutions available to you in the lab cause hydrolysis?
3. How will you know if this has happened?
4. Is hydrolysis connected to the reactions of acids and bases?
5. For each salt, decide whether it is the salt of a strong acid - strong base reaction, a strong acid - weak base reaction, or a weak acid - strong base reaction.
6. Summarize the conductivity and pH results for the salts that result from the reaction of:
 - i. a strong acid with a strong base
 - ii. a strong acid with a weak base
 - iii. a weak acid with a strong base

Post Discussion Questions Part 2

- For water:
 - Does pure water at room temperature conduct electricity?
 - Why did you boil the water before using it to prepare salts for pH measurement?
 - Are there any ions present in pure water at room temperature? If there are, what are they?
- Answer the following questions for the salts of strong acids and strong bases, the salts of strong acids and weak bases, and the salts of weak acids and strong bases:
 - Are solutions of these salts strong, weak, or non-electrolytes?
 - Identify any ions and/or molecules that are present in each solution.
 - Do these cations hydrolyze? Are these cations acidic, basic, or neutral?
 - Do these anions hydrolyze? Are these anions acidic, basic, or neutral?

Finishing Up (Parts 1 and 2)

Create a heading that says “Reflective Writing” and complete the reflective writing portion of your lab notebook after lab class. See the ELN Rubric on Canvas for details about what to include.

Generate a PDF of your page for today by right clicking on the page in in the Notebook Navigator (two-finger click on the mouse pad for a chrome book). See *Quick Guide for Lab Archives* in the “Student Canvas and LabArchive Tutorials” module on Canvas for details. Also save a PDF of any graphs you have generated.

Upload your PDF files using the assignment link given in the appropriate content area on Canvas by the ELN deadline (see syllabus for details).

Remember to take your post-lab quiz for this lab and the pre-lab quiz for the next lab by the appropriate deadlines (see the syllabus for details)

Investigation questions Part 3

- What is a buffer?

Part 3: Analyzing Buffer Systems

Materials

13x100-mm test tubes
10-mL graduated cylinder
100-mL graduated cylinder
pH paper

Chemicals and Safety Information

Compound	Potential Hazards	Precautions
Hydrochloric acid 0.1 M HCl (aq)	Corrosive. May cause severe burns. Avoid inhalation	Wear gloves and goggles, use in a well-ventilated area.
Sodium hydroxide 0.1 M NaOH (aq)	May cause severe burns and respiratory damage.	Wear goggles and gloves
Phenolphthalein indicator	May cause mild irritation. Avoid contact with skin or clothing	Wear goggles and gloves
Methyl red indicator	May cause mild irritation. Avoid contact with skin or clothing	Wear goggles and gloves
Acetic acid CH ₃ COOH (aq)	May be harmful if swallowed. Causes eye and skin burns.	Wear goggles and gloves
Formic acid HCOOH (aq)	May be harmful if swallowed. Causes eye and skin burns.	Wear goggles and gloves

Propylamine CH ₃ CH ₂ CH ₂ NH ₂ (aq)	May be harmful if inhaled. Extremely destructive to the tissue of the mucous membranes and upper respiratory tract. Toxic if absorbed through skin. Causes skin and eye burns.	Wear goggles and gloves, use proper ventilation
Ammonia (i.e., Ammonium Hydroxide) NH ₃ (aq) [i.e., NH ₄ OH]	Toxic. Corrosive. Causes eye, skin, respiratory and digestive tract burns	Wear goggles and gloves, use proper ventilation
sodium acetate NaCH ₃ COO (aq)	May cause mild irritation	Wear goggles and gloves
ammonium nitrate NH ₄ NO ₃ (aq)	May cause skin irritation. Avoid inhalation.	Wear goggles and gloves, use proper ventilation
propylammonium chloride CH ₃ CH ₂ CH ₂ NH ₃ ⁺ Cl ⁻ (aq)	May be harmful if inhaled. Causes respiratory tract irritation. Causes eye and skin irritation.	Wear goggles and gloves, use proper ventilation

Investigation

A buffer is a solution composed of a weak acid and the salt of that weak acid or a weak base and the salt of that weak base. As a class, plan an experiment to test a variety of potential buffer solutions to determine how the composition of the buffer affects the pH and the resistance of the pH to change.

Part 3A. Preparation and testing of potential buffer solutions

General Procedure

- a) Select any weak acid or base and a salt from solutions you will find on the reagent bench. Measure and mix together equal volumes of the two solutions in a test tube to make one solution containing both the weak acid or base and a salt. Measure the pH using pH paper.

Investigate

Predict how changing the concentration of one of the two components will change the pH, then design an experiment to see if the prediction was correct. To change the concentration of one of the components of the buffer, you can use the dilution technique from Part 1 to make a less-concentrated solution of one component; then combine it with the other component of the buffer as described above.

NOTE: Save all your solutions for Part 3C

Part 3B. Testing the pH of a boiled and cooled deionized water when a strong acid or strong base is added

Procedure

- b) Take a sample of boiled, deionized water and divide between two test tubes. Observe the pH of the sample. Add 5 drops of 0.10 M HCl to one test tube and measure the new pH.
- c) Add 5 drops of 0.10 M NaOH to the second test tube and measure the new pH.

Part 3C. Testing the buffer solution when a strong acid or strong base is added

For this part, use the buffer solutions (original and diluted buffers) you prepared in Part 3A.

Procedure

- d) Divide the solutions between two test tubes each. To the first part, add 5 drops of 0.10 M HCl and measure the new pH with pH paper. To the second part, add 5 drops of 0.10 M NaOH and measure the new pH.

Analysis Questions Part 3A and B

Create a data table for your observations and measurements for full credit.

1. Double-check that the class has prepared all of the solutions that you planned, and that there is a reference solution for comparisons.
2. Find the buffers for which the compounds were the same, but the amounts were changed. Create an additional table for them.

Part 3D. Estimating the pK of a weak acid or a weak base

In this part, you will estimate the pK of a weak acid or a weak base by measuring the pH of a solution containing an equal concentration of either the weak acid or base and one of its salts.

Procedure

- a) Measure 25 mL of one of the weak acid solutions into each of two Erlenmeyer flasks. Calculate the number of moles of the weak acid that are in each flask. Remember to specify which solution you picked.
- b) Select one of these flasks and add 5 drops of phenolphthalein indicator. Titrate with 0.10 M NaOH to the endpoint. Do not over-titrate or you will have to start over!
- c) Calculate the number of moles of the sodium salt of the weak acid that is in this flask.

- d) Pour all of the titrated solution into the flask that was not titrated. How many moles of weak acid are now in the flask? How many moles of the sodium salt of the weak acid are now in the flask? Estimate the volume of the solution in the flask. Calculate the concentration of the weak acid in the flask. Calculate the concentration of the sodium salt of the weak acid in the flask. Put the answers to these questions in the "Analysis Questions Part 3C" section of your ELN.
- e) Calibrate a pH meter using two standard buffers and measure the pH of the solution in the flask.
- f) Likewise, measure 25 mL of one of the weak bases into two Erlenmeyer flasks. Calculate the number of moles of the weak base that are in each flask.
- g) Select one of these flasks and add 5 drops of methyl red indicator. Titrate with 0.10 M HCl to the endpoint that will be a yellow to red color change. The solution should appear orange at the endpoint. Do not over titrate or you will have to start over!
- h) Calculate the number of moles of the hydrochloride salt of the weak base that are in this flask.
- i) Pour all of the titrated solution into the flask that was not titrated.
- j) Calibrate a pH meter using two standard buffers and measure the pH of the solution in the flask.

Analysis Questions Part 3D

Show all your calculations for full credit.

At point (d) in the procedure:

1. How many moles of weak acid are now in the flask?
2. How many moles of the sodium salt of the weak acid are now in the flask?
3. What is the estimated volume of the solution in the flask?
4. Calculate the concentration of the weak acid in the flask.
5. Calculate the concentration of the sodium salt of the weak acid in the flask.
6. Use your knowledge of the pH and the concentrations of the weak acid and the salt of the weak acid to find the value of the pK_a for the weak acid. Consult equations 5-12 in the calculations section for assistance. Compare this value with the value in the table. How closely do they agree?

At point (i) in the procedure:

7. How many moles of weak base are now in the flask?
8. How many moles of the hydrochloride salt of the weak base are now in the flask?
9. Estimate the volume of the solution in the flask.
10. Calculate the concentration of the weak base in the flask.
11. Calculate the concentration of the hydrochloride salt of the weak base in the flask.
12. Use your knowledge of the pH and the concentrations of the weak base and the salt of the weak base to find the value of the pK_b for the weak base. Consult the equations in the calculations section if necessary. Compare this value with the value in the table. How closely do they agree?

Post Discussion Questions Part 3

- a. How much of weak acid and its salt or weak base and its salt is needed to make the best buffer?
- b. How can you control the pH of the buffer system?
- c. Which combinations of acid or base and salt were the best at holding a constant pH when strong acid was added? Which were the best when strong base was added?
- d. How did changing the amounts of acid or base and salt change:
 - the pH of the solution?
 - the resistance to pH change when strong acid and base were added?
- e. How was the magnitude of pH change with addition of an acid or base to water compared to when a similar quantity of acid or base was added to a buffer?

Finishing Up

Create a heading that says “Reflective Writing” and complete the reflective writing portion of your lab notebook after lab class. See the ELN Rubric on Canvas for details about what to include.

Generate a PDF of your page for today by right clicking on the page in in the Notebook Navigator (two-finger click on the mouse pad for a chrome book). See *Quick Guide for Lab Archives* in the “Student Canvas and LabArchive Tutorials” module on Canvas for details. Also save a PDF of any graphs you have generated.

Upload your PDF files using the assignment link given in the appropriate content area on Canvas by the ELN deadline (see syllabus for details).

Remember to take your post-lab quiz for this lab and the pre-lab quiz for the next lab by the appropriate deadlines (see the syllabus for details).