Minneapolis Community and Technical College
Chemistry 1020 Laboratory

Note: In today's lab, you will be doing this new experiment as well as complete the aspirin experiment from last week. Therefore, bring both protocols to the lab.

Experiment: Titration and Buffers

Objectives: During this lab section, you will carry out experiments to gain hands-on experience with:

- properties of acids and bases
- determination of acidity (pH) by using an indicator
- ***titration*** of an acid with a base and
- a buffer system—an important application of acid-base neutralization reaction.

Discussion:

Acid and Base Neutralization

As you have learned in Chapter Seven, the Arrhenius theory states that an acid is a substance that can generate hydronium ions (H_3O^+) in aqueous solutions, and a base is a substance that releases hydroxide ions (OH^-) in aqueous solutions. (Remember, the shorthand notation for hydronium ions is: H^+.)

If equal amounts of an acid and a base are added to one another, *neutralization* occurs. When writing out a molecular/formula equation to describe this process, the products of acid-base neutralization are always water and an ionic salt.

$$HA_{(aq)} + MOH_{(aq)} \rightarrow H_2O_{(l)} + MA_{(aq)}$$

Depending on the acid and base involved, the formulas and coefficients may vary from those depicted in this equation—which is only intended to serve as a generic neutralization equation.

Typically, the only true chemical change that takes place when the strong acid and strong base are mixed involves the formation of water. This becomes apparent when the complete and net ionic equations for this process are written:

Net Ionic $$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$$

As most salts generated from a strong acid/strong base reaction are soluble in water, the ions of those salts are simply spectator ions and do not actually participate in the reaction. The neutralization process involves removal of most hydronium and hydroxide ions from solution in the formation of water. The small amounts of these two ions that remain are present **in equal amounts** (both = 1.0×10^{-7} M), and a neutral solution is attained. The concentration of hydronium ion in solution (and therefore acidity of a solution) is often communicated using values known as pH.

Titration: When you have two solutions but the concentration of only one of them is known, you can determine the concentration of the other, by a method called **titration.** In titrations, a known volume of one of the solutions is made to stoichiometrically react with the other. The volume of the second solution required for reaction completion is determined and the concentration of the other solution can then be calculated. In acid-base titrations, the volume of acid (or base) of known concentration required to neutralize known volume of the base (or acid) is determined by titrating them. To determine the neutralization point or end-point of the titration, you can use an indicator. As you learned previously, an acid-base indicator will change its color depending on the pH of the solution.

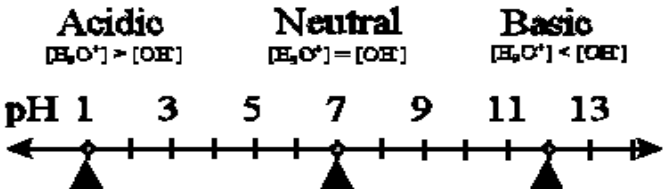
pH & Indicator Refresher

pH indicates the acidity or basicity of a solution. Remember, lower pH values (below 7) imply the solution is acidic, while higher pH values (above 7) imply a solution is basic. Each whole number increase or decrease in pH is equivalent to a 10 fold increase or decrease in acidity.

In acidic solutions, the amount of hydronium ion $[H^+]$ is greater than the amount of hydroxide ion $[OH^{1-}]$, although both are still present.

In basic solutions, the amount of hydroxide ion $[OH^{1-}]$ is greater than the amount of hydronium ion $[H^+]$, and again, both are present.

In neutral solutions, the amount of hydroxide ion $[OH^{1-}]$ equals the amount of hydronium ion $[H^+]$. (Both equal 1.0×10^{-7} M)



Molarity

The indicator you will use in today's lab (called **thymol blue**) **is blue in basic**, **pink in acidic and yellow in neutral, solutions**. If the proper amount of acid is added to a base containing this indicator, the base is neutralized and a yellow color can be observed at neutralization point. If not enough acid is added, the base is still basic – and remains blue. If too much acid is added, the solution becomes acidic and a pink color is observed.

In the previous two sections, you've encountered a unit "M". It stands for **molarity** (or **M)** and is the most commonly used expression of a concentration of solutions in a chemistry lab. It defines the number of moles of solute in one liter of solution:

$$M\ (molarity) = \frac{moles\ of\ solute}{volume\ of\ solution\ in\ liter} = \frac{mol}{L}$$

So if you know, for example, the number of moles of HCl in a given volume of solution, you can determine the molarity of that HCl solution.

On the other hand, you might already know the molarity of the solution, and would like to find out the number of moles of solute present in a given volume of a solution. To do that,

$$moles\ of\ solute = M\ (molarity) \times volume\ of\ solution\ in\ liters$$

You will be doing both kinds of calculation in this experiment.

Buffer Solutions

A *buffered solution* is resistant to change in its pH, even when strong acid or strong base are added to it. The ability of a solution to resist pH changes in this way is referred to as its *"buffering capacity"*. Buffers are prepared by mixing either a weak acid and its conjugate base or a weak base and its conjugate acid. In today's activity, we will focus on the first combination (weak acid + conjugate base).

Unlike a strong acid, a weak acid does not ionize 100% when dissolved in water. In fact, it is typically only a very small percentage of the acid that produces ions in solution. The majority of the acid molecules remain intact and un-ionized in water.

Dissolution of a Generic Weak Acid (HA)

$$HX_{(aq)} \overset{H_2O}{\rightleftharpoons} H^+_{(aq)} + X^-_{(aq)}$$

Once the HA molecule is dissolved in water, a very small percentage form H^+ and A^- ions in solution—the rest remain as HA molecules.

To learn more about buffers, weak acids, and conjugate bases, it is important to have an understanding of the **Brønsted-Lowry model** of acids and bases. According to this theory, *acids are proton (H^+) donors* and *bases are proton acceptors.* This theory differs from the Arrhenius theory in that a base does not necessarily contribute hydroxide ions to solution. When considering a weak acid, it is relevant to remember that an H^+ from the acid is transferred to water to form the hydronium ion. According to the Brønsted-Lowry definition, water is now acting as a base—as it has *accepted* a proton from the weak acid.

$$\underset{\text{acid}}{HA_{(aq)}} + \underset{\text{base}}{H_2O_{(l)}} \rightleftharpoons H_3O^+_{(aq)} + A^-_{(aq)}$$

Once this reaction takes place, there are really four different species present in the reaction vessel: HA, H_2O, H_3O^+, and A^-. We have identified the first two as the acid and the base in the reaction. The products of such a reaction (a weak acid plus water) can also be labeled. *The species formed when a base accepts a proton* is referred to as a **conjugate acid**, and *the species remaining after an acid donates a proton* is a **conjugate base**.

$$\underset{\text{acid}}{HA_{(aq)}} + \underset{\text{base}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{conjugate acid}}{H_3O^+_{(aq)}} + \underset{\text{conjugate base}}{A^-_{(aq)}}$$

As described earlier, a buffer is prepared by mixing a weak acid with its conjugate base or a weak base with its conjugate acid. These combinations are referred to as **conjugate acid-base pairs**, which consist of *two substances related to each other by the donating and accepting of a single proton*. In the equation listed above, there are two such pairs:

$$HA/A^- \text{ and } H_2O/H_3O^+$$

As you can see, the two components of a conjugate acid-base pair differ from one another only by one proton (H^+).

The buffer we use today will consist of a weak acid and its conjugate base, and is similar to the HA/A^- pair. Two different processes describe how a buffer system maintains a constant pH. One involves when a strong acid is added to the buffer. Normally, addition of a strong acid would lead to the formation of free H^+ ions and the pH of the solution decreases. In a buffer system, any acid added will instead react with the conjugate base. No H^+ ions form and the pH of the solution remains constant. Say, for example, that the strong acid HCl is added:

Addition of a Strong Acid to a Buffer

$HCl_{(aq)} + A^-_{(aq)} \rightarrow Cl^-_{(aq)} + HA_{(aq)}$

strong acid — conjugate base of buffer

The strong acid (HCl) reacts with the conjugate base portion of the buffer, and no H^+ ions are produced.

The second process involves when a strong base is added to the buffer. Normally, addition of a strong base would lead to the formation of free OH^- ions and the pH of the solution increases. In a buffer system, any base added will instead react with the weak acid. No OH^- ions form and the pH of the solution remains constant. Say, for example, that the strong base NaOH is added:

Addition of a Strong Base to a Buffer

$NaOH_{(aq)} + HA_{(aq)} \rightarrow H_2O_{(l)} + NaA_{(aq)}$

strong base — weak acid of buffer

The strong base (NaOH) reacts with the acid portion of the buffer, and no OH^- ions are produced.

Pre-lab exercise

1. In a balanced molecular/formula equation, describe the neutralization reaction that takes place between hydrochloric acid solution and sodium hydroxide solution. Include physical states of all substances.

2. Which indictor will be used in this lab activity?

3. What colors would one expect to see if the indicator specified in Q#2 were present in the following solutions:

    a. A basic solution

    b. A neutral solution

    c. An acidic solution

4. When reading a volume using a burette, how many decimal places should be used?

5. As sodium hydroxide is added to hydrochloric acid, what would one expect to happen to the pH of the mixture? Will it increase, decrease, or remain unchanged?

6. What is molarity? How is it determined?

7. Provide a definition for a buffered solution.

8. What should be done with the wastes from this lab?

Procedure

Part I. Buffer solution

Your instructor will demonstrate how a buffered versus an unbuffered solution responds to added acid and base. Pay attention to this demonstration and be sure to record the results on your data sheet.

Part II. Acid-Base Neutralization Reactions and Titration:

The HCl solution and NaOH solution are provided at your work-stations. **You have to make sure you take the appropriate solutions for the appropriate steps!**

Step-1: Transferring the NaOH solution to the Erlenmeyer (conical) flask:

a) Make note of the concentration (in M, or molarity) of the NaOH from the label on its bottle. **Record this molarity on your data sheet in the appropriate column**!

b) Pour just about 30 mL of the NaOH solution into a small, dry beaker and keep this at your work-station to use for the rest of the experiment. **Label this beaker as "base"**.

c) From the beaker, transfer 10 mL of the NaOH solution into a 10 mL graduated cylinder (you may have to transfer about 9 mL and then use an eye-dropper to add the last few drops to adjust to 10 mL).

d) Transfer this 10.00 mL solution from the graduated cylinder, into the Erlenmeyer (conical) flask.

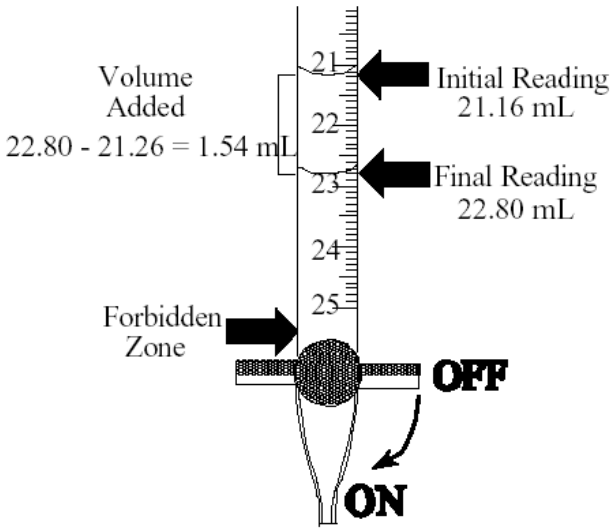
e) Add 2-3 drops of the thymol blue indicator (if it is not already added). **With the indicator, the basic NaOH solution should become blue at this point**. ***Set this flask aside until you are ready for Step-3.***

f) Before proceeding to the procedure on the next page, make sure you have recorded the molarity and volume of NaOH added to your Erlenmeyer flask in the appropriate table on the p.9 data sheet.

Step-2: Set up the burette:

Reading a Burette

A burette *(diagram at right)* may be used to measure the volume of any solution quite precisely. The burette dispenses solution when the burette valve is turned "on" (vertical position). Because the level of the liquid drops as it is dispensed, the burette is read from the TOP to the BOTTOM. The amount of solution dispensed from the burette is obtained by subtracting the initial reading from the final reading.



All burette readings should contain 2 decimal places. Never let the liquid drop into the forbidden zone since that area contains no markings.

a) You are provided with a burette already clamped to the burette-stand. Make sure the stopper at the bottom is shut-off.

b) Label a clean, dry beaker as "**acid**". Pour about 30 mL HCl solution into this beaker.

c) Using a funnel that is provided to you, **carefully pour the HCl solution** from the beaker into the burette to fill the burette somewhere close to the zero reading. There is no need to fill it to exactly zero.

d) **Remove the funnel**. Label another small beaker as **"waste**". Place this waste-beaker under the burette, carefully open the stopper of the burette, to slowly let some solution flow into this waste beaker (until the area below the stopper up to the bottom tip of the burette nozzle is filled without any visible air-bubbles). Close the stopper.

e) Note the volume of HCl in the burette **to 2 decimal places**. Make sure you are reading from the top of the burrette (since the zero level starts at the top). **If you are not confident how to take the reading, check with your instructor.**

f) Record this measurement as "**initial reading**" for Trial 1 on p. 9 data sheet.

Step-3: Titration (Your instructor will demonstrate this step)

a) Place the Erlenmeyer flask that has the 10.00 mL NaOH solution and indicator, (from Step-1) under the burette.

b) Open the stopper of the burette slightly to let the HCl solution flow dropwise from the burette, into the Erlenmeyer flask. Keep swirling the flask to mix the contents, as you continue to add the drops.

c) Continue the titration, until you see the color change from blue to yellow (ideally, change to yellow color is the correct neutralization point. A drop in excess beyond the neutralization point will give a pink color).

d) Stop the titration and note the final volume reading in the burette. Record this volume in your data sheet as the "**final reading**" under "Trial 1".

e) Calculate the difference between the initial and final readings of HCl solution. This will give you a good idea of how many mL of HCl solution is needed to titrate about 10 mL of NaOH solution, in order to do a more accurate titration in the second trial.

Repeating the titration

a) Pour the contents of your Erlenmeyer flask after your first titration down the drain and flush with water. Wash the flask with tap water and rinse it with distilled water.

b) Repeat the addition of 10 mL of NaOH, and 2-3 drops of thymol blue indicator in the Erlenmeyer flask, as you did in Trial 1.

c) Check if the HCl in the burette would be enough to carry out the second trial of titration. If not, add some more HCl to the burette, **and note the initial reading** for Trial 2.

d) Carry out the titration as before. Since you know the approximate volume required from the results of the first trial, as you reach within 1 mL of the end-point be very careful and add the HCl slowly dropwise. **Be patient**!

e) When the neutralization is achieved, note the **final reading** in the burette.

f) Calculate the volume of HCl that was required from second trial.

g) Show the results of the two trials to your instructor and get the approval. If the results of the two trials are not in agreement with each other, you will have to carry out a third trial.

Calculations

Follow the steps on p.10 to calculate the concentration of the HCl solution for each titration. Finally, take the average of the concentrations of acid. Report the concentration of the acid solution with the correct number of significant figures.

Waste Disposal

Dispose the contents in the beakers and Erlenmeyer flask into the sink and flush with water. Rinse the glassware with water.

Follow your instructor's directions, regarding contents of the burette.

Report Sheet **Titration and Buffers**

Name _____ Date_____ Lab Section_____

Part One: Buffer Solution (note the results from your instructor's demo)

Number of Drops of HCl added to "water": _____; Color change noticed: _____

Number of Drops of HCl added to "buffer": _____; Color change noticed: _____

Number of Drops of NaOH added to "water": _____: Color change noticed: _____

Number of Drops of NaOH added to "buffer": _____: Color change noticed: _____

Part Two: Titration Results for Acid-Base Neutralization

All measurements must have: quantity & units.

NaOH Solution

	Trial 1	**Trial 2**	**Trial 3 (optional)**
Molarity M_{NaOH}			
Volume taken into flask V_{NaOH} (mL)			

HCl Solution

	Trial 1	**Trial 2**	**Trial 3 (optional)**
Final Reading in Burette (mL)			
Initial Reading in Burette (mL)			
Volume used V_{HCl} (mL)			

Uncertainty

a) Your volume readings using the graduated cylinder are uncertain in the _____ place.

b) Your volume readings using the burette are uncertain in the _____ place.

You are encouraged to work on the calculation on the next page before leaving the lab.

Calculation (Follow the significant figure rules at each step and always show the units)
Your goal is to determine the molarity of HCl that you titrated in the above activity. You first calculate the number of moles of NaOH based on its molarity and volume. In the neutralization reaction between NaOH and HCl, the molar ratio between them is 1:1. Once the number of moles of HCl is determined, its concentration in molarity is calculated based on the moles and volume.

	Trial 1	**Trial 2**	**Trial 3**
Moles of NaOH $(= M_{NaOH} \times V_{NaOH} \times \frac{1L}{1000mL})$			
Moles of HCl (= Moles of NaOH)			
Molarity of HCl $(= \frac{\text{Moles of HCl}}{V_{HCl}} \times \frac{1000\ mL}{1L})$			
Average Molarity of HCl			

Postlab Questions

1. In the buffer portion of the lab, you compared how water and a buffer solution react to the addition of base (NaOH) and acid (HCl). Based on your observations, which seems to have greater "buffering capacity" (ability to maintain a consistent pH): the distilled water or the buffer?

2. The buffer solution consisted of acetic acid and sodium acetate (of which, the acetate ion is the important species). Of this mixture, the weak acid is _____ and its conjugate base is _____.

3. Which component of this buffer system (listed in question #2) reacted with the added sodium hydroxide?

   Write an equation that represents this reaction.

4. Which component of this buffer system (listed in question #2) reacted with the added hydrochloric acid?

   Write an equation that represents this reaction.