Titration of Synthesized Aspirin
A continuation of the aspirin synthesis lab

In this lab, you will determine the percent purity of your product from the aspirin synthesis using an acid-base titration. In general, an acid and a base react to produce a salt and water by transferring a proton (H^+):

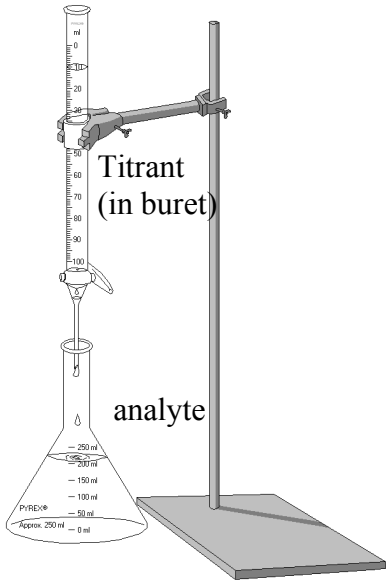
$$\underset{\text{Acid}}{HA\ (aq)} + \underset{\text{base}}{NaOH\ (aq)} \rightarrow H_2O\ (l) + \underset{\text{salt}}{NaA\ (aq)} \qquad (1)$$

The active ingredient in aspirin, and the chemical for which aspirin is the common name, is acetylsalicylic acid. To determine the amount of aspirin (acetylsalicylic acid) in a sample, the precise volume and concentration of the NaOH, and the overall reaction, must be known. The NaOH serves as a **secondary standard**, because its concentration can change over time. To find the precise concentration of the NaOH, it must be titrated against a **primary standard**, an acid that dissolves completely in water, has a high molar mass, that remains pure upon standing, and is not hygroscopic (tending to attract water from the air). Because sodium hydroxide is **hygroscopic**, it draws water from its surroundings. This mean one cannot simply weigh out a sample of sodium hydroxide, dissolve it in water, and determine the number of moles of sodium hydroxide present using the mass recorded, since any sample of sodium hydroxide is likely to be a mixture of sodium hydroxide and water. Thus, the most common way to determine the concentration of any sodium hydroxide solution is by titration. Determining the precise concentration of NaOH using a primary standard is called **standardization**. You will first standardize your NaOH solution, and then use it to analyze your synthesized aspirin for purity.

Think about it: What impurities might be present in your synthesized aspirin?

What is a titration?

A titration is a procedure for determining the concentration of a solution (the analyte) by allowing a carefully measured volume of this solution to react with another solution whose concentration is known (the titrant). The point in the titration where enough of the titrant has been added to react exactly with the analyte is called the equivalence point, and occurs when moles of titrant equals moles of analyte according to the balanced equation. For example, if a monoprotic acid (the analyte) is titrated with a strong base like sodium hydroxide (the titrant), the **equivalence point** occurs when



Number of moles of OH^- = Number of moles of H^+ (2)

The equivalence point is often marked by an **indicator**, a substance that changes color at (or very near) the equivalence point.
The point at which the indicator changes color is called the "**endpoint**". We will normally assume that the endpoint is equal to the equivalence point.

There are many types of titrations. In this lab you will be performing an acid base titration.

Background

Standardizing a Base

Consider the following reaction between oxalic acid, $H_2C_2O_4(aq)$, and potassium hydroxide:

$$H_2C_2O_4(aq) + 2\,KOH(aq) \rightarrow 2\,H_2O(l) + K_2C_2O_4(aq) \quad (3)$$

If 0.4862 g of oxalic acid was dissolved in water and titrated with 17.98 mL of potassium hydroxide solution, the concentration of the potassium hydroxide solution can be calculated. Since 17.98 mL of potassium hydroxide solution is used, that volume is converted to liters and put into the denominator:

$$\text{molarity of KOH} = [KOH] = \frac{\text{mol KOH}}{\mathbf{0.01798\ L\ KOH}} \quad (4)$$

The mass of oxalic acid, $H_2C_2O_4(aq)$, and Equation (4) are then used to determine the number of moles of potassium hydroxide present:

$$0.486\text{ g } H_2C_2O_4 \times \frac{\text{mol } H_2C_2O_4}{90.036\text{ g } H_2C_2O_4} \times \frac{2\text{ mol KOH}}{1\text{ mol } H_2C_2O_4} = \mathbf{0.0108\ mol\ KOH} \quad (5)$$

Finally, the molarity for potassium hydroxide is calculated as follows:

$$\text{molarity of KOH} = [KOH] = \frac{\mathbf{0.01080\ mol\ KOH}}{\mathbf{0.01798\ L\ KOH}} = \mathbf{0.600\underline{7}\ M\ KOH} \quad (6)$$

Determining the Number of Moles of an Acid

Once a solution has been standardized (such as a KOH solution above whose molarity has been determined), this solution can be used to determine the molar concentration of another solution, or simply the number of moles of analyte in the flask. Combined with the precise molar concentration of the titrant, the precise volume of titrant delivered yields the number of moles used to react with the analyte. Consider the following reaction between a monoprotic acid and potassium hydroxide:

$$HA(s) + KOH(aq) \rightarrow H_2O(l) + KA(aq) \quad (7)$$

Suppose 0.375 grams of acid, HA, require 3.47 mL KOH for neutralization. The volume (in liters) and the molarity of potassium hydroxide and Equation (8) can be used to determine the number of moles of acid present:

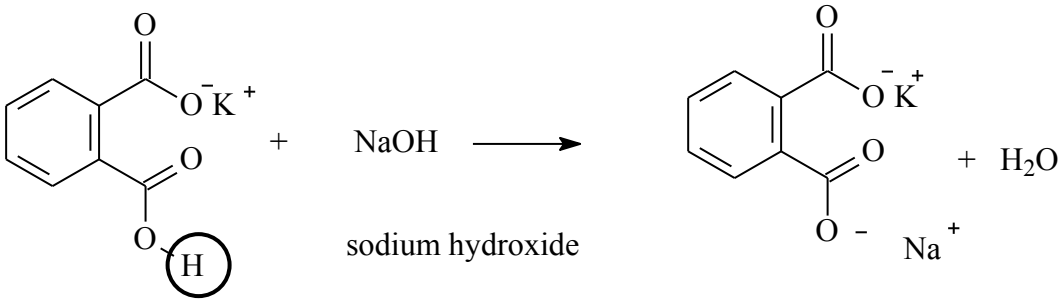
$$0.00347\text{ L} \times \frac{0.600\underline{7}\text{ mol KOH}}{\text{L}} \times \frac{1\text{ mol HA}}{1\text{ mol KOH}} = \mathbf{0.002081\ mol\ HA} \quad (8)$$

Part I: Standardization of NaOH (Titration of NaOH with KHP)

In this experiment the primary standard used will be **potassium hydrogen phthalate, $KHC_8H_4O_4$**, which is abbreviated as **KHP**. (Note that the "P" in KHP is *phthalate*, not phosphorus!) KHP is used to standardize the sodium hydroxide solution. The balanced chemical equation for the reaction is shown below:

Molecular Equation: $KHC_8H_4O_4(aq) + NaOH(aq) \rightarrow H_2O(l) + NaKC_8H_4O_4(aq)$ (9)

(The reacting hydrogen is circled in the equation below.)



Potassium hydrogen phthalate (KHP)
($KHC_8H_4O_4$)

Ionic Equation:

$K^+(aq) + HC_8H_4O_4^{1-}(aq) + Na^+(aq) + OH^-(aq) \rightarrow H_2O(l) + K^+(aq) + C_8H_4O_4^{2-}(aq) + Na^+(aq)$ (10)

Net Ionic Equation: $HC_8H_4O_4^{1-}(aq) + OH^-(aq) \rightarrow H_2O(l) + C_8H_4O_4^{2-}(aq)$ (11)

In this experiment, phenolphthalein will be used as the indicator because it gives a sharp endpoint, meaning that when a titration is being carried out, the phenolphthalein is so sensitive that within a fraction of a drop of titrant (NaOH), phenolphthalein will make the solution turn completely pink. Note that phenolphthalein does not turn pink until the solution is *basic*, so at the endpoint, when the color change occurs, there is a slight excess of hydroxide ions in solution, making the solution basic. The fainter the pink color, the closer one is to the **equivalence point (also called the stoichiometric point**), the theoretical point in the titration when enough titrant has been added to react completely with the analyte, KHP.

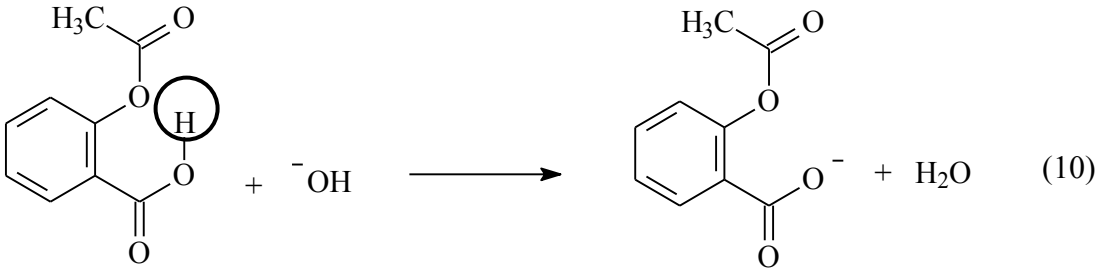
Part II: Determination of the purity of synthesized aspirin (Titration of NaOH with synthesized aspirin)

Second, you will titrate a sample of your aspirin (**acetylsalicylic acid, $C_9H_8O_4$,** MW 180.2 g/mol) with the standardized NaOH to determine the **actual moles** of aspirin(a monoprotic acid) in a given weight of aspirin. The **theoretical moles** of aspirin are calculated from the weighed sample of aspirin. The percent purity will then be calculated as follows.

$$\frac{\text{actual moles of asprin}}{\text{theoretical moles of asprin}} \times 100\% = \%\,\text{purity}$$

Assuming the aspirin is not contaminated with other acids, the titration allows you to quantitatively determine the purity of your aspirin. (The reacting hydrogen is circled in the equation below.)

The **Net Ionic Equation** for the titration in this experiment:



acetylsalycilic acid (aspirin) + hydroxide
($C_9H_8O_4$)
(analyte) (titrant)

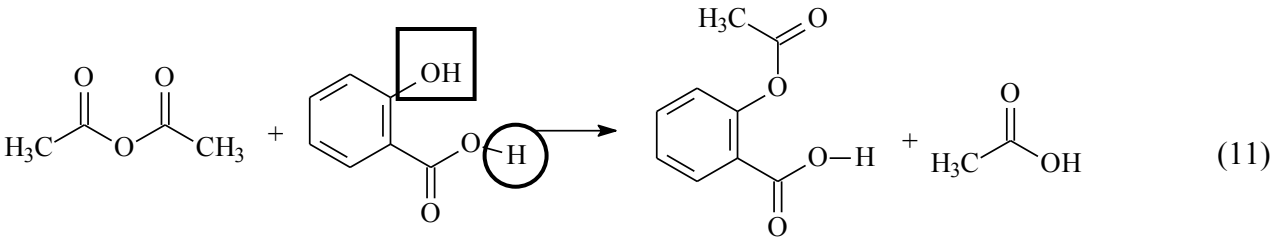
*** The titrant is standardized in Part I of this experiment***

Part III: Detection of Residual Salicylic Acid with $FeCl_3$

Titration of your sample with sodium hydroxide will allow you to determine the total amount of acid present in the sample. If you use this value you are assuming that the only acid present is acetyl salicylic acid (aspirin). When you synthesized aspirin you began with salicylic acid that will also react with sodium hydroxide. How do you know that your results are not from unreacted salicylic acid in your sample?

Think about it: What if your sample is contaminated with other acids besides acetylsalicylic acid? Will they react with NaOH as well?

HINT: Recall the synthesis reaction from the previous lab:



salicylic acid
($C_7H_6O_3$)

acetyl salicylic acid (asprin)
($C_9H_8O_4$)

You can perform a colorimetric test to determine if there is any salicylic acid left in your sample. Addition of $FeCl_3$ to phenols results in the formation of a colored complex. The OH (shown in the box above) attached directly to the benzene ring is called a phenol. This type of OH is different from the OH where the H is circled above or an OH of an alcohol. Addition of $FeCl_3$ to salicylic acid will result in the formation of a deeply purple colored complex.

Laboratory Technique for Burets

Burets are used to deliver a recorded amount of liquid or solution to another container. A buret is marked in milliliters like a graduated cylinder, but buret markings show **0 mL** at the top, and the numbers increase as you go down the buret. The stopcock controls the liquid flow. It is ***open*** when ***parallel*** to the length of the buret and ***closed*** when ***perpendicular*** to the length of the buret.

- **Washing and rinsing the buret:** To clean a buret, wash its interior with soap and tap water. Next, rinse the buret with 5-10 mL portions of DI water. With the buret over the sink and the stopcock open, pour the water into the buret and let it drain out the tip. Use a beaker to pour solutions into the buret—most breakage occurs during washing, and burets do ***NOT*** fit under the faucet.
- **Conditioning the buret:** After the buret is well-drained, close the stopcock and add about 5 mL of the ***titrant*** (the solution to be used into the buret). Tilt the buret sideways and roll the barrel to completely rinse the inner walls of the buret. Drain the solution through the buret tip to insure the tip is also conditioned. Repeat this step at least twice to be sure all interior surfaces are rinsed with titrant.
- **Filling the buret:** Close the stopcock. Use a clean funnel to fill the buret with titrant just above the "0" mark. Place a container under the buret tip, and open the stopcock briefly to fill the buret tip with solution, leaving no air bubbles, and to get the level of meniscus to fall within the markings of the buret. If the tip does not fill with solution when the stopcock is in the open position, there may be an air bubble in the stopcock. Consult your instructor.
  Note: ***The initial level of titrant need not be exactly at 0.00 mL***as the initial level of liquid will be recorded and subtracted from the final volume to determine the volume delivered.
- **Reading the buret:** Always remove the funnel used to fill the buret before taking any measurements. Record the volume of titrant by noting the bottom of the meniscus. On the buret shown below, numbers marked for every 1 mL, and the ten lines between each number represent every 0.1 mL. Thus, the level of titrant in the buret can be estimated to one more decimal place than the markings or to the nearest 0.01 mL.

  Thus, in the figure to the right, the meniscus is about halfway between 25.0 and 25.1 mL, so the level of titrant can be recorded as ***25.04 mL, 25.05 mL, or 25.06 mL*** depending on whether the bottom of the meniscus appears to be just above, just at, or just below halfway, respectively.
- **Cleaning the buret:** Afterwards, empty the buret, disposing of the titrant according to the waste disposal instructions for each experiments. Wash the buret with soap and tap water, then rinse with several portions of tap water, allowing some tap water to run through the tip. Do a final rinse with small portions of DI water, allowing the DI water to run through the tip, then return the buret to the stockroom.

Safety Precautions

NaOH is corrosive. Handle with care. In case of contact with skin, rinse the area with large amounts of water and notify your instructor. Wear goggles at all times in the chemistry laboratory.

Procedure

Part I: Standardization of the NaOH solution (Titration of NaOH with KHP)

Prepare a buret for titration by rinsing it with two small portions of distilled water, followed by two 5-mL portions of the sodium hydroxide solution. Make sure to coat the inside of the buret. Fill the buret and follow the usual procedures for eliminating air bubbles and setting the initial level. Your buret is now ready for the rest of the lab (**NOTE: You should not need more than 50 mLof NaOH for all the titrations**).

Record the initial buret reading on your data table.

Place approximately 0.4 g of KHP (MW = 204.23 g/mol) in a clean 125 mL Erlenmeyer flask. Record the precise mass used. Dissolve the acid in 50 mL of distilled water and add 2 drops of phenolphthalein indicator.

Think about it! Do you have to measure the amount of water precisely? Does your glassware need to be dry?

A pre-lab question asked you to estimate the approximate volume of solution required to react completely with the KHP. Enter this volume in the box. This should be a good approximation of your endpoint. The closer to you approach to the endpoint the more slowly you will need to titrate (dropwise).

mL (approx. volume required)

Run slightly less (1-2 mL less) than this amount into the flask containing the acid, while swirling the flask. Use your wash bottle to clean the walls of the flask of drops of base that may have splattered out the titration mixture. Continue adding the NaOH solution dropwise. As you approach the volume estimated to be required for complete reaction, add the NaOH more slowly, (one drop at a time) while continuing to swirl the flask and wash down the walls of the flask.

Stop the titration when the addition of a single drop of NaOH changes the color of the solution to a light pink, indicating that you have reached the endpoint. The endpoint should persist for 30 seconds without fading.

Record the final buret reading on you data table.

Repeat the titration. Your molarity should reflect the most significant figures you can obtain with the lab equipment, and the molarity of two runs should agree within 5%.

- Calculate the molarity of your NaOH solution.
- Average the molarity of the first two runs.
- Compute the percent difference between the two runs:
  - $\frac{|M_1 - M_2|}{\overline{M}} x100\% = \%$ difference
  - If run 1 and run 2 are within 5% agreement (of molarity), proceed to Part 2.
  - If your runs are not in close agreement, run a third trial. Find the percent difference between the two closest runs, and report this in your Summary Table.
  - If you omit a run explain why. i.e. over titrated sample; passed the endpoint.

Think about it! If your % difference is within 5% what does that mean about your data?

Part II: Titration of Standardized NaOH with the Aspirin Sample

Using a sample of aspirin synthesized in a previous lab experiment:

Weigh out approximately 0.10-0.15g of your aspirin sample into a 125 mL flask and record the precise mass. Add 10 mL of 95% ethanol. Allow the aspirin to dissolve for a few minutes by swirling the mixture. Make sure your aspirin sample is fully dissolved.

Add 3-5 drops of phenolphthalein indicator. Slowly titrate the aspirin with the standardized NaOH solution from part 1 of today's lab. Record the initial and final buret readings to the correct number of significant figures on your data sheets.

Repeat the titration. Do a third trial if necessary (percent difference is >5%). Report the moles of acid and percent difference of the closest two trials. Calculate and report the percent purity of your aspirin sample.

Waste Disposal: All solutions should be emptied into appropriately labeled waste containers in the fume hood.

Part III: Detection of Residual Salicylic Acid with $FeCl_3$

Weigh out about 0.015g (15 mg) of your aspirin sample into a test-tube. Dissolve the sample in about 10 drops of ethanol then about 0.5 mL of water. The sample should be fully dissolved and in solution.

Then add 1-2 drops of 1% aqueous iron (III) chloride solution.

Record your detailed observations in the Report Sheets. Make certain that your notes on this experiment are clear.

Waste Disposal: All solutions should be emptied into appropriately labeled waste containers in the fume hood.

This page intentionally left blank.

Report Sheets

Titration of Synthesized Aspirin

Name_____

Lab partner_____

Section_____

Background

Calculate the molar mass of KHP and enter it here: [ ]

Show the calculation for Pre-lab Exercise #4, estimating the volume of 0.10 M NaOH required to neutralize 0.40 g KHP.

Data

Table 1: Standardization of NaOH

	Trial 1	Trial 2	Trial 3	Trial 4
Mass KHP (g)				
Moles KHP				
Moles NaOH				
Final volume NaOH (mL)				
Initial volume NaOH (mL)				
Volume of NaOH delivered (mL)				
Molarity of NaOH, mol/L				
Average Molarity				
% Difference				
Observations:				

Don't forget to use units and significant figures.

Table 2: Determination of acetylsalicylic acid in aspirin samples (purity)

	Trial 1	Trial 2	Trial 3
Mass of aspirin (g)			
Theoretical moles of aspirin			
Final volume NaOH (mL)			
Initial volume NaOH (mL)			
Volume of NaOH delivered (mL)			
Moles of OH^- delivered			
Actual moles of aspirin			
Percent purity of aspirin			
Average purity			

D**on't forget to use units and significant figures.**

Calculations

(For all of the calculations that follow, show a sample calculation for **at least one trial**.)

Part I: Standardization of the NaOH solution

a) Calculate the moles of KHP used.

b) Calculate the moles of NaOH used.

c) Calculate the volume of NaOH (in liters) used to neutralize the KHP.

d) Calculate the molarity of the NaOH.

e) Average these values to find the average molarity $\left(\overline{M}\right)$ that will be used in part II.

Average Molarity of standardized NaOH (M):

f) Find the percent difference for the closest two trials, and report this value (also indicate which trials were included) in Table 2. (This should be done in lab, before moving on to Part 2.)

Part II: Titration of the Aspirin Sample

a) Assume that your aspirin sample is pure (not a good assumption) and calculate the moles of aspirin (acetylsalicylic acid, $C_9H_8O_4$, MW 180.2 g/mol) to be titrated in each trial. These moles are the **theoretical** moles of aspirin (in theory, your sample should be pure).

b) Calculate the moles of NaOH used in titration using the average molarity of NaOH determined in Part I.

c) Calculate the moles of aspirin that reacted with the NaOH. These are the **actual** moles of aspirin in your sample (the moles of acid that actually reacted with the NaOH).

d) Calculate the percent purity of aspirin in each trial, and report the average percent purity.

Results

Table 3: Summary of Results for NaOH Standardization

Part I:	Molarity NaOH	Average molarity	Percent difference
Trial # ___			
Trial # ___			

Table 4: Summary of Results for Aspirin Titration

Part II:	mass sample (g)	mL NaOH delivered	moles aspirin	percent purity	percent difference in purity
Trial # ___					
Trial # ___					

Summarize your results using full sentences (in the past tense, third person), and restate the values obtained and the percent differences. What are some sources of error inherent to the techniques used? What potential sources of error could affect a future experimenter?

Post-lab Questions:

1. Give one plausible explanation for why a student might achieve less than 100% purity. (Assume calculations are correct.)

2. Give one plausible explanation for a student who obtains over 100% purity. (Assume calculations are correct.)

Pre-Lab Assignment:

Titration of Synthesized Aspirin

Name_____
Section_____

Read the lab instructions several times. Refer to the sections in your textbook on acid-base titrations.

1) Sodium hydroxide is a hygroscopic solid. It can be purchased as solid pellets. Why can't you weigh out sodium hydroxide on the balance and make a solution of known concentration? (In other words, why does NaOH need to be standardized?)

2) When mixed HCl an acid and NaOH a base will react with each other.
Write the balanced neutralization equation, ionic and net ionic equation for this reaction.

   Molecular Equation:

   Ionic Equation:

   Net ionic equation:

3) A student found that the titration had taken 10.00 mL of 0.1002 M NaOH to titrate 0.132 g of aspirin, a monoprotic acid. Calculate the percent purity of aspirin ($C_9H_8O_4$, MW 180.2 g/mol) sample. Give a possible explanation of what might have affected percent purity of aspirin.

4) The NaOH solution is standardized with KHP in part I of this lab. If 0.400 grams of KHP (MW = 204.23 g/mol) is titrated with approximately 0.100 M NaOH, what is the approximate volume (mL) of NaOH required to reach the endpoint? Show your calculation below. Write this value in the box located in the procedure section for Part I.