

Determination of Aspirin using Back Titration

This experiment is designed to illustrate techniques used in a typical *indirect* or *back* titration. You will use the NaOH you standardized last week to back titrate an aspirin solution and determine the concentration of aspirin in a typical analgesic tablet. You will be graded on your accuracy.

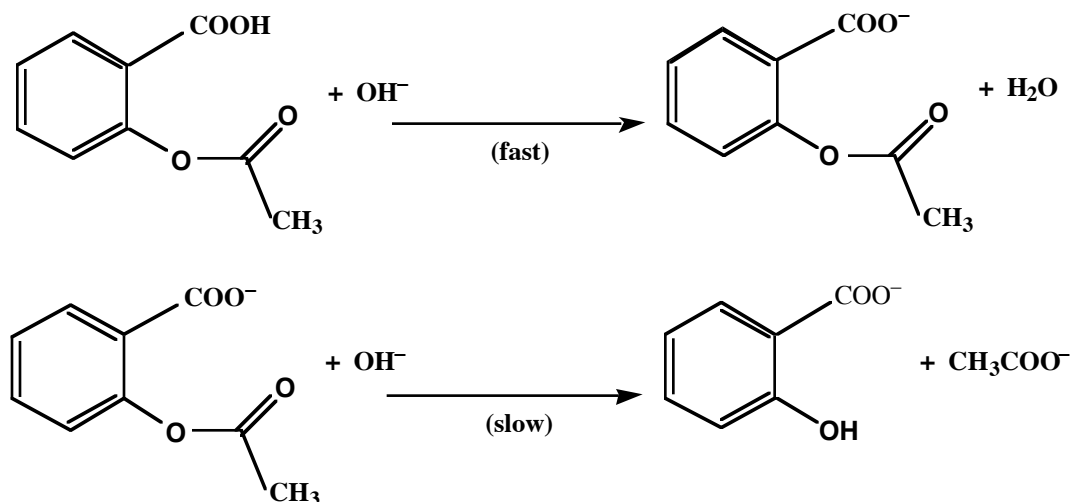
Required Reading

D.C. Harris, Quantitative Chemical Analysis (6th ed., W. H. Freeman, NY, 2003) p. 130, 15.

PreLab Quiz Topics

In addition to being able to explain the purpose of your experiment, the general procedure steps, the use of all chemicals in this experiment and any specific hazards, your prelab quiz may include explanations of any of the following terms: back titration, percent composition, the transition range of an indicator. You should be able to discuss the difference between a direct and a back titration, and know all relevant chemical reactions for this experiment. Describe the chemical and/or physical processes that will occur when you reach the endpoint of your titration. Explain clearly why we use a back titration for this particular analysis; i.e., why can't we just use the endpoint found from the initial NaOH titration?

Many reactions are slow or present unfavorable equilibria for direct titration. Aspirin is a weak acid that also undergoes slow hydrolysis; i.e., each aspirin molecule reacts with two hydroxide ions. To overcome this problem, a known **excess** amount of base is added to the sample solution and an HCl titration is carried out to determine the amount of unreacted base. This is subtracted from the initial amount of base to find the amount of base that actually reacted with the aspirin and hence the quantity of aspirin in the analyte.



Chemicals and their Location

Balance Room

Aspirin tablets (**Do not** bring your own sample)

Above/on lab bench

Ethanol

Phenolphthalein, indicator soln

In the Hood

Hydrochloric acid, conc., 37 wt %

Your Drawer

Sodium hydroxide, standardized soln

Equipment and its Location

Stockroom

Two Burettes
Mortar & Pestle
Weighing bottle

In the Hood

Water bath

Above/on lab bench

Boiling chips

Safety Issues and Chemical Hazard Information

	<i>Physical Hazards</i>	<i>Health Hazards</i>
<i>Aspirin</i>	none	toxic, irritant, sensitizer
<i>Ethanol</i>	flammable	irritant
<i>Hydrochloric acid</i>	water-reactive, corrosive	toxic
<i>Phenolphthalein</i>	none	irritant
<i>Sodium hydroxide</i>	water-reactive, corrosive	toxic, irritant

Concentrated hydrochloric acid is highly corrosive. Be careful. Ethanol is flammable — keep away from heat, sparks, and open flames.

Procedure

There are various components to this procedure. If there is a delay at one step (e.g., all the water baths are in use), perform another section of this procedure while you are waiting. Depending on the size of your lab section, your TA may suggest that half of the class work straight through the procedure, and the other half start at Step 4, returning to Steps 1 – 3 while waiting at Step 8, to reduce congestion in the lab.

Preparation of Approximate Acid Solution (~ 0.1M HCl)

1. Put ~100 mL distilled water into your other large bottle. Before performing this experiment (i.e., in your **prelab**), calculate the volume of concentrated HCl you will need to prepare 250 mL of 0.1M HCl. (Concentrated reagent grade HCl has a density of 1.188 g mL⁻¹ and is 37 wt %. Check your answer with your TA before making this solution.) Measure approximately this volume of concentrated HCl using a graduated cylinder. Gradually add the acid to the water in your bottle, **mixing well** (remember, always add acid to water, not water to acid.). Add more distilled water, **mixing thoroughly** with each addition, until the total solution volume is ~ 250 mL. You do not have to measure the quantities accurately because you are going to standardize this solution in the next steps to determine its actual concentration.

Comparison of HCl to your standardized NaOH

2. Clean, rinse, and fill your burette with your ~0.1 M HCl solution. To each of three clean, labeled Erlenmeyer flasks, add 50 mL distilled water and three drops of phenolphthalein indicator. Use your HCl burette to add approximately 35, 40 and 45 mL of acid to flasks #1, #2 and #3, respectively, recording **exactly** how much acid is added to each flask within 0.01 mL. Swirl gently to mix.

3. Clean, rinse, and fill your second burette with the ~ 0.1 M NaOH solution that you standardized last week. Titrate the three HCl flasks with the NaOH to the phenolphthalein endpoint.

Sample Preparation

4. Accurately record the weight of a group of three aspirin tablets so that you can determine an average tablet weight. Use a mortar and pestle to crush enough tablets to produce ~ 1 g tablet powder.
5. Using a clean dry weighing bottle, weigh accurately, **by difference**, triplicate ~ 0.3 g samples of tablet, into labeled 250 mL Erlenmeyer flasks. To each flask, add 20 mL of ethanol (measure by graduated cylinder) and three drops of phenolphthalein indicator. Swirl gently to dissolve. (Aspirin is not very soluble in water — the ethanol helps the aspirin dissolve. Note that an aspirin tablet contains other compounds in addition to aspirin. Some of these are not very soluble. Your solution will be cloudy due to insoluble components of the tablet.)

Aspirin Titration with base

6. (If your TA has told you to perform this step before Steps 1 – 3, clean, rinse, and fill a burette with the ~ 0.1 M NaOH solution that you standardized last week, before continuing.) Titrate the first aspirin sample with NaOH to the first permanent cloudy pink colour.
7. The aspirin/NaOH acid-base reaction consumes one mole of hydroxide per mole of aspirin. The slow aspirin/NaOH hydrolysis reaction also consumes one mole of hydroxide per mole of aspirin, and so for a complete titration **we will need to use a total of twice the amount of NaOH that you have already used, plus we will add some excess NaOH** to ensure that we really have reacted with all of the aspirin in your sample (adding excess reactant drives the equilibrium towards products — Le Chatelier's principle).

Calculate how much extra NaOH you will need to add, following this reasoning: The volume of base to add for the hydrolysis reaction is equal to the volume of base you have already used to titrate to the acid-base endpoint in Step 6 plus an additional 10 mL of excess base. (For example: if you used 26 mL of base in the previous step, the volume of base you would add now would be $26 + 10 = 36$ mL. Thus, you would have added a total of $26 + 26 + 10 = 62$ mL of base.)

Use your burette (not a graduated cylinder) to add the appropriate amount of extra NaOH to each of your three sample flasks. (Do not to allow the level of base in the burette to fall below the graduated markings; if necessary, record an intermediate volume and refill the burette to continue with your additions.) Record the total volume of NaOH added to each flask within 0.02 mL.

Heating the reaction to completion

8. Add two or three boiling chips to each flask and heat in a water bath to speed up the hydrolysis reaction. Avoid boiling, because the sample may decompose. While heating, swirl the flasks occasionally. After 15 minutes, remove samples from the water bath and cool for 5 minutes.

9. If the solution is colourless, add a few more drops of phenolphthalein. If it remains colourless, add 10 mL more of the base and reheat. (**Don't forget** to add this additional volume of base to the previously recorded total volume.)

Back titration with acid

10. The only base remaining in each flask will be excess base that has not reacted with the aspirin. Using your burette with your ~ 0.1 M HCl solution, titrate the excess base in each flask with HCl until the pink colour just disappears. The endpoint is best described as “cloudy white”.

Shutdown procedures

IMPORTANT NOTE: Your NaOH solution **must be saved** upon completion of the experiment; you will use it again in future experiments. Save your HCl solution in case you find that something is wrong with your results from today's experiment.

12. Put all waste in the appropriate waste containers. Clean and dry all glassware. Clear up your work area. Return items to the Stockroom.
13. Make sure you have recorded information on the sample you studied, including the manufacturer and brand name and the amount of aspirin reportedly in each tablet from the aspirin packet label, which can be found posted on the Stockroom door.

Waste Disposal

Waste Chemicals	Waste Container	Location
Excess HCl, dilute or concentrated	Acid	Waste cabinet
Excess NaOH, dilute or concentrated	Base	Waste cabinet
Excess ethanol	Ethanol	Waste cabinet
Aspirin solutions after titration	Neutral	Waste cabinet
Aspirin solid waste	Solid waste	Balance room

Data Analysis

- Using your NaOH molarity from last week, determine the molarity of your HCl solution from the volumes of NaOH needed to titrate your flasks of HCl, remembering that you designed it to be ~ 0.1 M. Report your standardized HCl concentration as the average of your three replicates with appropriate statistical analysis.
- Using the total volume of NaOH added to each, and the molarity of your NaOH standard solution, calculate the total number of moles of NaOH added to each sample.
- Using the volume of HCl needed to back-titrate each aspirin flask and the average HCl concentration determined in Step One, calculate the number of moles of excess NaOH left in each flask after the reaction with aspirin.

4. From the total number of moles of NaOH added to the sample, and the number of moles of excess NaOH left after the reaction was complete, calculate the number of moles of base that reacted with aspirin.
5. Hence determine the number of moles of aspirin in each of your samples. (Don't forget the factor of two.)
6. Calculate the mass of aspirin in each of your samples. Using the mass of tablet you weighed into each sample, calculate the weight percent of aspirin in the tablets.
7. From the mass of the three tablets you weighed, calculate the average mass of a tablet. Use this value, and the weight percent of aspirin you have just calculated, to calculate the mass of aspirin in one of the original tablets.

Discussion Questions

1. Why did you use your burette and not a graduated cylinder to add the excess NaOH?
2. What would be the consequence of competing equilibria, such as from other ingredients in the tablet, on your results? What if there were several pain relievers in the same tablet? How could the error from these interferences be minimized?
3. Ethanol was used in the solutions to help dissolve the acetylsalicylic acid. Ethanol is slightly acidic, and will react with NaOH. Describe a blank correction experiment; i.e., what experiment you might do to determine how much NaOH reacts with the ethanol in your solution. Once you have determined this, how would you use the results of a blank correction experiment in the data analysis?

Conclusions

As part of your conclusions for this experiment, compare the aspirin weight per tablet that you determined to the weight per tablet claimed by the manufacturer on the label.

Spring 2005 CH455 Lab Report Title Page

Student Use			
Name:		Class TA:	
Lab Partners (if any):		Makeup TA (if any):	
Class TA Use			
Prelab Quiz Score (out of 3):		Lab finished:	
(TA signature required before student leaves lab)			
Experiment Date:		Report Due Date:	Due 9 / 9:30 am; 2 pm; Time: 5 pm; 7 pm
Report Received			
Date:	Time:	Initials:	
Grading			
Report Grade (/15):	Late penalty (/15):	Report inc. penalty (/15):	

Preparation of an NaOH Standard Solution using Direct Titration

Your report should follow the Lab Report format listed earlier in this manual. Show samples of all calculations, identify any unknowns, report the average with appropriate significant figures, standard deviation, %RSD and 90% confidence interval and answer the discussion questions. Don't forget to include your conclusion section.

Concentration of analyte in unknown sample	Average	±	Std Dev	%RSD	Average	±	90% confidence interval
Standardized NaOH concentration		±				±	

Identify your NaOH starting solution

Unknown HCl concentration		±				±
---------------------------	--	---	--	--	--	---

Identify your HCl unknown

Preparation of an NaOH Standard Solution using Direct Titration

Report Grading Breakdown

	Your Score	/	Out of a possible
Pre-lab Quiz		/	3
Calculations/Data Analysis		/	2
Results/Accuracy		/	4
Discussion & Conclusions		/	3
Presentation/Impression		/	3
Subtractions			
TOTAL		/	15

Spring 2005 CH455 Lab Report Title Page

Student Use		
Name:	Class TA:	
Lab Partners (if any):	Makeup TA (if any):	
Class TA Use		
Prelab Quiz Score (out of 3):	Lab finished:	
(TA signature required before student leaves lab)		
Experiment Date:	Report Due Date:	Due 9 / 9:30 am; 2 pm; Time: 5 pm; 7 pm
Report Received		
Date:	Time:	Initials:
Grading		
Report Grade (/15):	Late penalty (/15):	Report inc. penalty (/15):

Determination of Aspirin using Back titration

Your report should follow the Lab Report format listed earlier in this manual. Show samples of all calculations, identify any unknowns, report the average with appropriate significant figures, standard deviation, %RSD and 90% confidence interval and answer the discussion questions. Don't forget to include your conclusion section.

	Average ± Std Dev	%RSD	Average ± 90% confidence interval
Your HCl concentration	±		±
Measured aspirin mass per tablet	±		±

Manufacturer's stated aspirin mass per tablet _____

Identify your aspirin sample _____

Determination of Aspirin using Back titration

Report Grading Breakdown

	Your Score	/	Out of a possible
Pre-lab Quiz		/	3
Calculations/Data Analysis		/	3
Results/Accuracy		/	3
Discussion & Conclusions		/	3
Presentation/Impression		/	3
Subtractions			
TOTAL		/	15