

Lab #10: Weak Acid and Strong Base Titration

Objectives:

1. To accurately determine the mass content of acetylsalicylic acid in an Aspirin tablet – a common analgesic (painkiller) medicine by titrating it with a strong base.
2. To sketch the titration curve of acetylsalicylic acid by titrating it with a strong base and determine its K_a .

Background Information:

History of Aspirin (taken from Wikipedia)

The Bayer Company of Germany coined the brand name Aspirin. In some countries the name is used as a generic term for the drug rather than the manufacturer's trademark. In countries in which Aspirin remains a trademark, the initial ASA (for acetylsalicylic acid) is used as a generic term.

The name "aspirin" is composed of *a-* (from the acetyl group) *-spir-* (from the plant genus *Spiraea*) and *-in* (a common ending for drugs at the time). It has also been stated that the name originated by another means. "As-" referring to AcetylSalicylic and "-pir-" in reference to one of the scientists who was able to isolate it in crystalline form, Raffaele Piria. Finally, "-in", because it was a common ending for drugs at the time.

On March 6, 1899, Bayer registered it as a trademark. However, the German company lost the right to use the trademark in many countries because the Allies seized and resold its foreign assets after World War I. The right to use "Aspirin" in the United States (along with all other Bayer trademarks) was purchased from the U.S. government by Sterling Drug in 1918. Even before the patent for the drug expired in 1917, Bayer had been unable to stop competitors from copying the formula and using the name elsewhere, so, with a flooded market, the public was unable to recognize "Aspirin" as coming from only one manufacturer. Sterling was subsequently unable to prevent "Aspirin" from being ruled a generic trademark in a U.S. federal court in 1921. Bayer ultimately acquired Sterling in 1994, but this did not restore the U.S. trademark. Other countries (such as Canada and many countries in Europe) still consider "Aspirin" a protected trademark.

Hippocrates, a Greek physician, wrote in the 5th century BC about a bitter powder extracted from willow bark that could ease aches and pains and reduce fevers. This remedy is also mentioned in texts from ancient Sumer, Lebanon, and Assyria. The Cherokee and other Native Americans used an infusion of the bark for fever and other medicinal purposes for centuries. The medicinal part of the plant is the inner bark and was used as a pain reliever for a variety of ailments. The Reverend Edward Stone, a vicar from Chipping Norton, Oxfordshire, England, noted in 1763 that the bark of the willow was effective in reducing a fever.

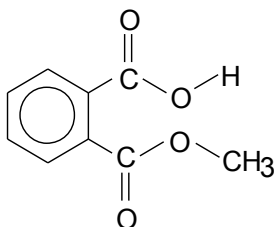
The active extract of the bark, called *salicin*, after the Latin name for the white willow (*Salix alba*), was isolated to its crystalline form in 1828 by Henri Leroux, a French pharmacist, and Raffaele Piria, an Italian chemist. Piria was able to convert the substance into a sugar and a second component, which on oxidation becomes salicylic acid.

Salicylic acid was also isolated from the herb meadowsweet (*Filipendula ulmaria*, formerly classified as *Spiraea ulmaria*) by German researchers in 1839. While their extract was somewhat effective, it also caused digestive problems such as gastric irritation and bleeding and diarrhea, and even death when consumed in high doses. In 1853, a French chemist named Charles Frederic Gerhardt neutralized salicylic acid by buffering it with sodium (sodium salicylate) and acetyl chloride, creating acetosalicylic anhydride. Gerhardt's product worked, but he had no desire to market it and abandoned his discovery. In 1897, researcher Arthur Eichengrün and Felix Hoffmann, a research assistant at Friedrich Bayer & Co. in Germany, derived one of the hydroxyl functional groups in salicylic acid with an acetyl group (forming acetylsalicylic acid, an acetyl ester). This thereby greatly reduced the negative effects caused by the free phenolic group of salicylic acid. When in the body, the ester (aspirin) is hydrolyzed to free the active drug. This was the first synthetic drug, not a copy of something that existed in nature, and the start of the pharmaceuticals industry.

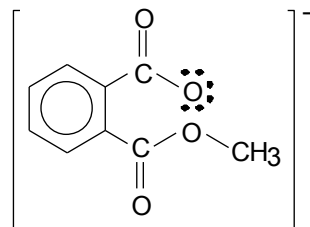
Hoffmann made some of the formula and gave it to his father, who was suffering from the pain of arthritis and could not stand the side effects of salicylic acid. With good results, he then convinced Bayer to market the new wonder drug. 'Aspirin' was patented on March 6, 1899. It was marketed alongside another of Hoffmann's products, an acetylated synthetic of morphine called 'Heroin' that he invented 11 days after aspirin. Heroin was initially the more successful of the two painkillers and it was common belief that it was healthier than aspirin. But, as heroin's shortcoming of addictiveness became more obvious, aspirin stepped to the forefront. Aspirin was originally sold as a powder (still the preferred form in many European countries) and was an instant success; in 1915, Bayer introduced aspirin tablets.

Aspirin as a Weak Acid

Aspirin is a weak acid with a $K_a = 3.27 \times 10^{-4}$. The chemical structures of the acid and its conjugate base are as follow. In this lab, we will use the short form HASA for acetylsalicylic acid and ASA⁻ for the acetylsalicylate.



Aspirin - Acetylsalicylic Acid (HASA)

Acetylsalicylate ion (ASA⁻)*Indicators*

Phenolphthalein and the Universal Indicator Solution will be used in this lab. The Universal Indicator Solution has the following colours at various pH ranges.

pH	Colour	pH	Colour
3	red-orange	7	green
4	orange	8	blue-green
5	yellow	9	blue-gray
6	greenish-yellow	10	violet

Materials:

Ring Stand	Buret Funnel and Buret Clamp	NaOH _(aq) (0.100 M)
pH meter	3 Small / Medium Erlenmeyer Flasks	3 to 4 Aspirin Tablets _(aq) (325 mg each)
Stirring Rod	Mortar and Pestle	Phenolphthalein
Buret	Electronic Balance	Universal Indicator Solution
Graduated Cylinder	1 Small Beaker and 1 Medium Beaker	Ethanol

Hypothesis / Pre-lab Exercise:

1. Write the balanced molecular equation for the neutralization of HASA_(aq) with NaOH_(aq).
2. Given that each Aspirin tablet is 325 mg, calculate the volume of 0.100 M of NaOH_(aq) required to completely neutralize the tablet.
3. Determine the both endpoint colours of phenolphthalein and universal indicator appropriate for this experiment (from reading the procedure below and the background information above).

Procedure:**Part A: Titration of Aspirin with NaOH using Phenolphthalein**

1. Coat the buret with the NaOH_(aq) at least twice and discard the wash fluid.
2. Set up the titration apparatus with the ring stand, buret clamp, buret and buret funnel.
3. Fill the buret with NaOH_(aq) using the buret funnel. Be sure not to pass the 0 mL mark.
4. Record the starting volume of the NaOH_(aq).
5. Measure and record the mass of an aspirin tablet.
6. Label a medium beaker as ethanol. Obtain approximately 100 mL of ethanol from your instructor.
7. Grind the aspirin tablet into powder using mortar and pestle. Pour about 10 mL of ethanol into the mortar to dissolve the powder (not all the powder will dissolve). Using a stirring rod as a guide, transfer the content into an Erlenmeyer flask. Repeat until all the powder from the mortar is transferred into the flask. (Alternatively, place the tablet directly into an Erlenmeyer flask. Add about 20 mL of ethanol into the flask. Using the stirring rod, break the tablet apart and dissolve it as much as possible into the ethanol solvent. Wash the stirring rod with a small amount of ethanol when finished.)
8. Add a few drops of phenolphthalein in the flask.
9. Begin titration of the aspirin. Swirl the Erlenmeyer flask when adding the NaOH_(aq). Depending on the amount of ethanol used as the initial solvent for the aspirin tablet, the equivalence point is around pH 8.0 – 8.4. Stop the titration when the endpoint is reached. Record the final volume of the NaOH_(aq). Calculate the net volume of base added.

- Repeat the titration with the other two Erlenmeyer flasks. Be sure to record the initial and final volume of the buret each time. Try to adjust the buret valve in such a way so NaOH _(aq) is added one drop at a time around the endpoint.

Part B: Titration of Aspirin with NaOH using the Universal Indicator Solution and a pH meter.

- Set up the titration as outlined in Part A, except we will use a small beaker instead of an Erlenmeyer flask and universal indicator solution instead of phenolphthalein.
- Use a pH meter to measure and record the pH of the aspirin-alcohol solution. Record the initial volume of the buret. Add about 1.0 mL of NaOH into the beaker. After mixing thoroughly, measure and record the pH again. Record the final volume of the buret.
- Repeat the previous step until you are 0.5 mL before the volume equivalence point. Add 0.1 mL at a time and record the pH and volumes. (For example: if you find that the equivalence volume is at 22.5 mL from the other trials, you should start measuring the pH at every 0.1 mL interval between 22.0 to 23.0 mL of NaOH added.)
- Continue to add at 1.0 mL increments five more times for five more pH.

Observations:

	Trial 1	Trial 2	Trial 3
Mass of Aspirin Tablet used			
Initial Volume			
Final Volume			
Volume of NaOH added			
Phenolphthalein Colour			

Table 2: Volumes of 0.100 mol/L of NaOH Added to Aspirin Tablet and measured pH

Mass of Aspirin Tablet used: _____

Initial Volume (mL)	Final Volume (mL)	Volume Added (mL)	Total NaOH Volume Added (mL)	pH	Colour of Universal Indicator	Initial Volume (mL)	Final Volume (mL)	Volume Added (mL)	Total NaOH Volume Added (mL)	pH	Colour of Universal Indicator

Analysis:

1. Average the mass of the three Aspirin tablets used.
2. For each trial in Part A, determine the experimental mass of HASA used from the endpoint volumes of Table 1. Average these calculated experimental HASA masses.
3. Find the average mass percentage of acetylsalicylic acid in an average tablet.
4. Using the data in Table 2, graph the pH Curve between the Aspirin tablet and 0.100 M of NaOH_(aq). Indicate the equivalence pH, volume of base added at stoichiometric point, and the buffer region.
5. Using the pH recorded when there is no NaOH added, calculate the theoretical concentration of HASA_(aq) in the small beaker (K_a of HASA is 3.27×10^{-4}).

Evaluation:

1. Compare the equivalence volume from the pH curve with the ones obtained in Part A of this experiment.
2. Given that the stated amount of acetylsalicylic acid in each tablet is 325 mg, calculate the % error for the mass of HASA.
3. Predict and explain what would happen to the calculated mass of HASA calculated when there is/are
 - a. water or ethanol left in the Erlenmeyer flask when the tablet is added to the flask.
 - b. water left in the buret when NaOH_(aq) is added.
 - c. air bubbles in the buret when NaOH_(aq) is added that is dislodged during the titration.
 - d. air bubbles in the buret when NaOH_(aq) is added that is not dislodged during the titration.
4. Besides human errors, what other sources of error could contribute to the % errors calculated above? (*Hint*: think about the nature of the solvent.) How might this error affect your average mass of HASA calculated?
5. Why is alcohol used as the solvent instead of water?
6. Why is it not a good idea to take aspirin with alcoholic beverages?
7. How is aspirin different from Tylenol? (Research on the Internet.)
8. What amounts of aspirin do regular, extra strength, and children aspirin have in them?
9. Why is phenolphthalein a good choice of an indicator for this experiment?
10. In most cases, each tablet of aspirin has more mass than the indicated 325 mg of acetylsalicylic acid. What are the other substances that make up this difference? What is the main role of these “filler” substances? (Research on the Internet.)
11. Why are aspirin substitutes used by many people?

Conclusion:

Summarize what you have learned from this lab.