

Date: March 9, 2018

Objectives:

The objective of this experiment was to determine how much aspirin was in a sample using acid base Titration. In this experiment we used aspirin, (acetylsalicylic acid), the acid and sodium hydroxide (NaOH), the base to neutralize the reaction. Phenolphthalein was used as the indicator to signify the end point of the reaction. The solution turns slightly pink and persists for 30 seconds when the end point has been reached. Adding too much of the base to the acid will turn the color bright pink and indicate the reaction was taken too far.

Data/Observations:

A: Preparation for Titration

A ring stand was set up with a 50 mL buret and clamp. The buret was rinsed with deionized water.

The buret was filled with 50 mL of .0949 M of NaOH, (sodium hydroxide). The level of solution was at approximately 0.00 mL at the top.

Four samples of aspirin were obtained. 2 tablets of Bayer 325 mg, and 2 tablets of Bayer 81 mg.

Each sample of aspirin was placed in a 250 mL beaker with 25 mL ethanol and 25 mL deionized water. The mixture was stirred with a rubber policeman. Four drops of the indicator Phenolphthalein were added to each beaker.

First trial – 325 mg Bayer aspirin. 14 mL of NaOH were added to the beaker containing the aspirin solution. Initial reading of NaOH was 0 mL at top, end reading was 14 mL.

Second trial – 325 mg Bayer aspirin. 14 mL of NaOH were added to the beaker containing the aspirin solution, this trial went past the end point and ended up bright pink. Initial reading was 0 mL at top, end reading was 14 mL.

Third trial – 325 mg Bayer aspirin. 15.4 mL of NaOH were added to the beaker containing the aspirin solution. Initial reading was 0 mL at top, end reading was 15.4 mL.

Data was used from the first and third trial to an average of 14.7 mL NaOH to neutralize the aspirin solution. $14\text{ mL} + 15.4\text{ mL} / 2 = 14.7\text{ mL NaOH}$.

Fourth trial – 81 mg Bayer aspirin. 4 mL of NaOH were added to the beaker containing the aspirin solution. Initial reading was 0 mL at top, end reading was 4 mL.

Fifth trial – 81 mg Bayer aspirin. 3.5 mL of NaOH were added to the beaker containing the aspirin solution. Initial reading was 0 mL at top, end reading was 4 mL.

Average NaOH needed to neutralize solution $4 \text{ mL} + 3.5 \text{ mL} / 2 = 3.75 \text{ mL}$.

Trial	Aspirin mg	Initial reading	End reading	Notes
1	325 mg Bayer	0 mL	14 mL	To end point
2	325 mg Bayer	0 mL	14 mL	Past end point
3	325 mg Bayer	0 mL	15.4 mL	To end point
Average	325 mg Bayer		14.7 mL	Data from 1 & 3
4	81 mg Bayer	0 mL	4 mL	To end point
5	81 mg Bayer	0 mL	3.5 mL	To end point
Average	81 mg Bayer		3.75 mL	Data from 4 & 5

Conclusions & Study Questions

Conclusion:

Proper operation of the buret and releasing NaOH into the aspirin mixture is important. It is easy to go too far when adding the base and pass the endpoint. This was an experiment that was easy to conceptualize; adding enough base to the acid to neutralize the solution. It was interesting to see the actual weight of the aspirin in the solution in comparison to the weight of the tablet of aspirin. I also realized after the experiment that, in my previous experience with titration we didn't refill the buret to 0 mL at the top each time. Our end reading from the first trial became the beginning reading for the second trial.

Study Questions:

1. For each titration that you completed determine the molecular weight of your aspirin sample.

Trial 1 – 14 mL NaOH - 0.0949 M.

$$.014 \text{ L NaOH} \times 0.0949 \text{ mol/L} = 0.0013286 = 0.0013 \text{ moles NaOH}$$

$$0.0013286 \text{ moles} \times 180.16 \text{ g/mol} = \mathbf{0.239360576} = \mathbf{239.36 \text{ mg aspirin.}}$$

Trial 3 – 15.4 mL NaOH – 0.0949 M.

$$0.0154 \text{ L NaOH} \times 0.0949 \text{ mol/L} = 0.00146146 \text{ moles NaOH}$$

$0.00146146 \text{ moles} \times 180.16 \text{ g/mol} = \mathbf{0.263296634 \text{ g} = 263.30 \text{ mg aspirin.}}$

Trial 4 – 4 mL NaOH – 0.0949M.

$0.004 \text{ L NaOH} \times 0.0949 \text{ mol/L} = 0.0003796 \text{ moles NaOH}$

$0.0003796 \text{ moles} \times 180.16 \text{ g/mol} = \mathbf{0.068388736 \text{ g} = 68.39 \text{ mg aspirin.}}$

Trial 5 – 3.5 mL NaOH – 0.09494M.

$0.0035 \text{ L NaOH} \times 0.0949 \text{ mol/L} = 0.00033215 \text{ moles NaOH}$

$0.00033215 \text{ moles NaOH} \times 180.16 \text{ g/mol} = \mathbf{0.059840144 \text{ g} = 59.84 \text{ mg aspirin.}}$

2. **Using the two (or three) molecular weights that you determined in question 1, determine the average molecular weight of your aspirin sample.**

325 mg Aspirin sample average molecular weight of two trials:

$239.61 \text{ mg} + 263.30 \text{ mg} / 2 = 251.46 \text{ mg.}$

$251.46 \text{ mg} / 325 \text{ mg} = 77\% \text{ aspirin}$

81 mg Aspirin sample average molecular weight of two trials:

$68.389 \text{ mg} + 59.84 \text{ mg} / 2 = 64.11 \text{ mg}$

$64.11 \text{ mg} / 81 \text{ mg} = 79\% \text{ aspirin}$