Name ______________________ Class ______________ Date ______________

Skills Practice

# Titration with an Acid and a Base

Titration is a process in which you determine the concentration of a solution by measuring what volume of that solution is needed to react completely with a standard solution of known volume and concentration. The process consists of the gradual addition of the standard solution to a measured quantity of the solution of unknown concentration until the number of moles of hydronium ion, $H_3O^+$, equals the number of moles of hydroxide ion, $OH^-$. The point at which equal numbers of moles of acid and base are present is known as the equivalence point. An indicator is used to signal when the equivalence point is reached. The chosen indicator must change color at or very near the equivalence point. The point at which an indicator changes color is called the end point of the titration. Phenolphthalein is an appropriate choice for this titration. In acidic solution, phenolphthalein is colorless, and in basic solution, it is pink.

At the equivalence point, the number of moles of acid equals the number of moles of base.

$$(1) \quad \text{moles of } H_3O^+ = \text{moles of } OH^-$$

By definition

$$(2) \quad \textit{molarity}\ (\text{mol/L}) = \frac{\textit{moles}}{\textit{volume}}\ (\text{L})$$

If you rearrange equation **2** in terms of moles, equation **3** is obtained.

$$(3) \quad \textit{moles} = \textit{molarity}\ (\text{mol/L}) \times \textit{volume}\ (\text{L})$$

When equations **1** and **3** are combined, you obtain the relationship that is the basis for this experiment, assuming a one-to-one mole ratio and the units of volume are the same for both the acid and base.

$$(4) \quad \text{molarity of acid} \times \text{volume of acid} = \text{molarity of base} \times \text{volume of base}$$

In this experiment, you will be given a standard hydrochloric acid, HCl, solution and told what its concentration is. You will carefully measure a volume of it and determine how much of the sodium hydroxide, NaOH, solution of unknown molarity is needed to neutralize the acid sample. Using the data you obtain and equation **4,** you can calculate the molarity of the NaOH solution.

## OBJECTIVES

**Use** burets to accurately measure volumes of solution.

**Recognize** the end point of a titration.

**Describe** the procedure for performing an acid-base titration.

**Determine** the molarity of a base.

## MATERIALS

- 0.500 M HCl
- 50 mL burets, 2
- 100 mL beakers, 3
- 125 mL Erlenmeyer flask
- double buret clamp
- NaOH solution of unknown molarity
- phenolphthalein indicator
- ring stand
- wash bottle filled with deionized water

**Always wear safety goggles, gloves, and a lab apron to protect your eyes and clothing.** If you get a chemical in your eyes, immediately flush the chemical out at the eyewash station while calling to your teacher. Know the locations of the emergency lab shower and eyewash station and the procedures for using them.

**Do not touch any chemicals.** If you get a chemical on your skin or clothing, wash the chemical off at the sink while calling to your teacher. Make sure you carefully read the labels and follow the precautions on all containers of chemicals that you use. If there are no precautions stated on the label, ask your teacher what precautions to follow. Do not taste any chemicals or items used in the laboratory. Never return leftovers to their original container; take only small amounts to avoid wasting supplies.

**Call your teacher in the event of a spill.** Spills should be cleaned up promptly, according to your teacher's directions.

**Never put broken glass into a regular waste container.** Broken glass should be disposed of properly.

## Procedure

1. Set up the apparatus as shown in **Figure A.** Label the burets *NaOH* and *HCl.* Label two beakers *NaOH* and *HCl.* Place approximately 80 mL of the appropriate solution into each beaker.
2. Pour 5 mL of NaOH solution from the beaker into the NaOH buret. Rinse the walls of the buret thoroughly with this solution. Allow the solution to drain through the stopcock into another beaker and discard it. Rinse the buret two more times in this manner, using a new 5 mL portion of NaOH solution each time. Discard all rinse solutions.

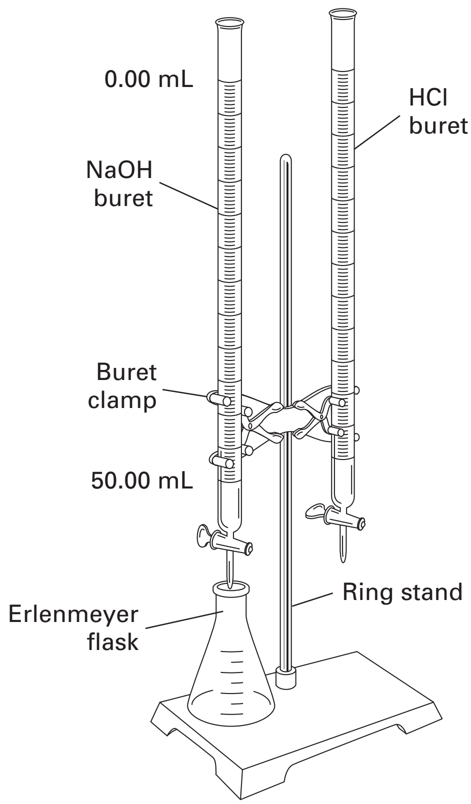


**Figure A**

**3.** Fill the buret with NaOH solution to above the zero mark. Withdraw enough solution to remove any air from the buret tip, and bring the liquid level down within the graduated region of the buret.

**4.** Repeat steps **2** and **3** with the HCl buret, using HCl solution to rinse and fill it.

**5.** For trial 1, record the initial reading of each buret, estimating to the nearest 0.01 mL, in the **Data Table.** For consistent results, have your eyes level with the top of the liquid each time you read the buret. Always read the scale at the bottom of the meniscus.

**6.** Draw off about 10 mL of HCl solution into an Erlenmeyer flask. Add some deionized water to the flask to increase the volume. Add one or two drops of phenolphthalein solution as an indicator.

**7.** Begin the titration by slowly adding NaOH from the buret to the Erlenmeyer flask while mixing the solution by swirling it, as shown in **Figure B.** Stop frequently, and wash down the inside surface of the flask, using your wash bottle.

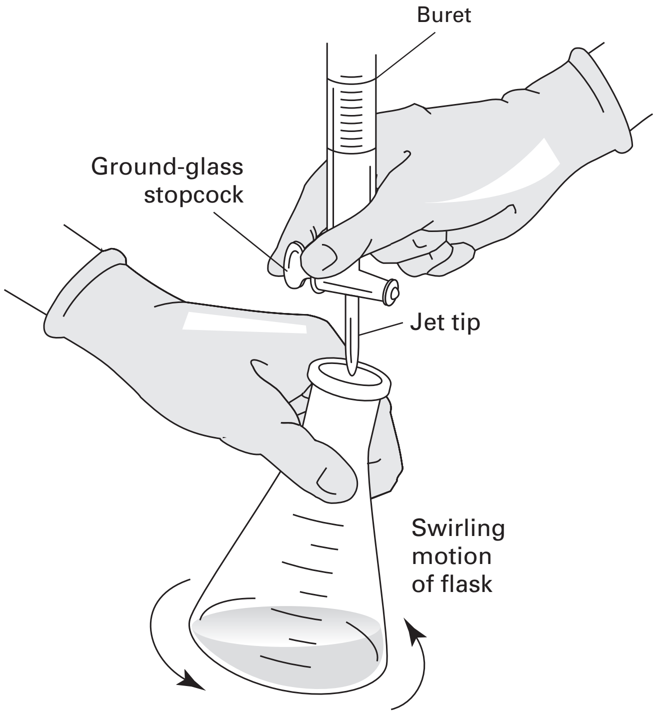


**Figure B**

**8.** When the pink color of the solution begins to appear and linger at the point of contact with the base, add the base drop by drop, swirling the flask gently after each addition. When the last drop added causes the pink color to remain throughout the whole solution and the color does not disappear, stop the titration. A white sheet of paper under the Erlenmeyer flask makes it easier to detect the color change.

**9.** Add HCl solution dropwise just until the pink color disappears. Add NaOH again, dropwise, until the pink color remains. Go back and forth over the end point several times until one drop of the basic solution just brings out a faint pink color. Wash down the inside surface of the flask, and make dropwise addition of NaOH, if necessary, to reestablish the faint pink color. Read the burets to the nearest 0.01 mL, and record these final readings in the **Data Table.**

**10.** Discard the liquid in the flask, rinse the flask thoroughly with deionized water, and run a second and third trial.

**11.** Record the known concentration of the standard HCl solution in the **Data Table.**

### DISPOSAL

**12.** Clean all apparatus and your lab station. Return equipment to its proper place. Dispose of chemicals and solutions in the containers designated by your teacher. Do not pour any chemicals down the drain or in the trash unless your teacher directs you to do so. Wash your hands thoroughly before you leave the lab and after all work is finished.

**Data Table**

| Buret readings (ml) | | | | |
|---|---|---|---|---|
| | HCl | | NaOH | |
| **Trial** | **Initial** | **Final** | **Initial** | **Final** |
| 1 | | | | |
| 2 | | | | |
| 3 | | | | |

**Molarity of HCl** __________

## Analysis

**1. Organizing Data** Calculate the volumes of acid used in the three trials. Show your calculations and record your results below.

Trial 1: Volume of HCl = ______________________________

Trial 2: Volume of HCl = ______________________________

Trial 3: Volume of HCl = ______________________________

2. **Organizing Data** Calculate the volumes of base used in the three trials. Show your calculations and record your results below.

   Trial 1: Volume of NaOH = ______________________

   Trial 2: Volume of NaOH = ______________________

   Trial 3: Volume of NaOH = ______________________

3. **Organizing Data** Use equation **3** in the introduction to determine the number of moles of acid used in each of the three trials. Show your calculations and record your results below.

   Trial 1: Moles of acid = ______________________

   Trial 2: Moles of acid = ______________________

   Trial 3: Moles of acid = ______________________

4. **Relating Ideas** Write the balanced equation for the reaction between HCl and NaOH.

   ______________________________________________

5. **Organizing Ideas** Use the mole ratio in the balanced equation and the moles of acid from **Analysis** item **3** to determine the number of moles of base neutralized in each trial. Show your calculations and record your results below.

   Trial 1: Moles of acid = ______________________

   Trial 2: Moles of acid = ______________________

   Trial 3: Moles of acid = ______________________

6. **Organizing Data** Use equation **2** in the introduction and the results of **Analysis** item **2** and **5** to calculate the molarity of the base for each trial. Show your calculations and record your results below.

   Trial 1: Molarity of NaOH ______________________

   Trial 2: Molarity of NaOH ______________________

   Trial 3: Molarity of NaOH ______________________

7. **Organizing Conclusions** Calculate the average molarity of the base. Show your calculations and record your result below.

## Conclusions

1. **Analyzing Methods** In step **6,** you added deionized water to the HCl solution in the Erlenmeyer flask before titrating. Why did the addition of the water not affect the results?

2. **Analyzing Methods** What characteristic of phenolphthalein made it appropriate to use in this titration? Could you have done the experiment without it? How does phenolphthalein's end point relate to the equivalence point of the reaction?

Skills Practice

# Titration with an Acid and a Base

## Teacher Notes

**TIME REQUIRED** 60 min

**SKILLS ACQUIRED**

Collecting data
Communicating
Experimenting
Identifying patterns
Inferring
Interpreting
Organizing and analyzing data

**RATING** Easy ← 1 2 3 4 → Hard

Teacher Prep–3
Student Set-Up–3
Concept Level–3
Clean Up–2

**THE SCIENTIFIC METHOD**

**Make Observations** Students observe the color change at the end point of a titration.

**Analyze the Results** Analysis questions 1 to 9

**Draw Conclusions** Analysis question 7

**Communicate the Results** Analysis questions 1 to 7

**MATERIALS**

To prepare 1.0 L of 0.500 M hydrochloric acid, use a fresh bottle of reagent grade concentrated HCl, preferably one that shows the actual assay of HCl rather than an average assay. Observe the required safety precautions. Assuming that the concentrated HCl is 12 M, slowly and with stirring, add 41.65 mL to enough distilled water to make 1.00 L of solution. If a large number of students are to be provided with solutions, it's best to make up 10.0 L in a large dispenser to ensure the same concentration for all lab groups. However, it will be impossible to make the HCl exactly 0.500 M without the accuracy of a volumetric flask. Some 5 gal carboys with spigots make ideal dispensers for the acid and base solutions.

To prepare 1.0 L of 0.6 M sodium hydroxide solution, observe the required safety precautions. Stir while adding 24 g of NaOH to enough distilled water to make 1.0 L of solution.

To prepare 1.0 L of phenolphthalein solution, dissolve 10 g of phenolphthalein in 500 mL of denatured alcohol and add 500 mL of water.

A wet cloth mop can be rinsed out a few times and used until it falls apart.

Bromthymol blue may be used as an indicator and is the more appropriate indicator for the titration of a strong acid with a strong base. For the concentrations of acid and base used, phenolphthalein and bromothymol blue give the same molarity results.

## SAFETY CAUTIONS

Read all safety precautions, and discuss them with your students.

- Safety goggles and a lab apron must be worn at all times.
- In case of an acid or base spill, first dilute with water. Then mop up the spill with wet cloths or a wet cloth mop while wearing disposable plastic gloves. Designate separate cloths or mops for acid and base spills.
- Wear safety goggles, a face shield, impermeable gloves, and a lab apron when you prepare the NaOH and HCl solutions. For preparing HCl, work in a hood known to be in operating condition, and have another person stand by to call for help in case of an emergency. Be sure you are within a 30 s walk from a safety shower and an eyewash station known to be in good operating condition.

## TIPS AND TRICKS

Demonstrate all techniques needed for successful titration: cleaning the burets, reading the buret with the eye at the liquid level, reading the buret scale correctly, swirling the flask, manipulating the stopcock, washing down the sides of the flask, and evaluating the color of the indicator.

Discuss the role of the indicator and the meaning of the terms *end point* and *equivalence point*. Indicators change colors at diffferent pH values. It is important to chose an indicator that changes color at a pH which is close to the pH of the equivalence point of the titration. For this titration, phenolphthalein is a good indicator.

It may help students if you work through a set of sample titration data.

## DISPOSAL

Set out three disposal containers for the students: one for unmixed acid solutions, one for unmixed base solutions, and one for partially neutralized substances and the contents of the waste beaker. To neutralize the acid and base, slowly combine the solutions while stirring. Adjust the pH of the final waste liquid with 1.0 M acid or base until the pH is between 5 and 9. Pour the neutralized liquid down the drain.

Name ______________________ Class ______________ Date ____________

Skills Practice

# Titration with an Acid and a Base

Titration is a process in which you determine the concentration of a solution by measuring what volume of that solution is needed to react completely with a standard solution of known volume and concentration. The process consists of the gradual addition of the standard solution to a measured quantity of the solution of unknown concentration until the number of moles of hydronium ion, $H_3O^+$, equals the number of moles of hydroxide ion, $OH^-$. The point at which equal numbers of moles of acid and base are present is known as the equivalence point. An indicator is used to signal when the equivalence point is reached. The chosen indicator must change color at or very near the equivalence point. The point at which an indicator changes color is called the end point of the titration. Phenolphthalein is an appropriate choice for this titration. In acidic solution, phenolphthalein is colorless, and in basic solution, it is pink.

At the equivalence point, the number of moles of acid equals the number of moles of base.

$$(1) \quad \text{moles of } H_3O^+ = \text{moles of } OH^-$$

By definition

$$(2) \quad \textit{molarity}\ (\text{mol/L}) = \frac{\textit{moles}}{\textit{volume}}\ (\text{L})$$

If you rearrange equation **2** in terms of moles, equation **3** is obtained.

$$(3) \quad \textit{moles} = \textit{molarity}\ (\text{mol/L}) \times \textit{volume}\ (\text{L})$$

When equations **1** and **3** are combined, you obtain the relationship that is the basis for this experiment, assuming a one-to-one mole ratio and the units of volume are the same for both the acid and base.

$$(4) \quad \text{molarity of acid} \times \text{volume of acid} = \text{molarity of base} \times \text{volume of base}$$

In this experiment, you will be given a standard hydrochloric acid, HCl, solution and told what its concentration is. You will carefully measure a volume of it and determine how much of the sodium hydroxide, NaOH, solution of unknown molarity is needed to neutralize the acid sample. Using the data you obtain and equation **4,** you can calculate the molarity of the NaOH solution.

## OBJECTIVES

**Use** burets to accurately measure volumes of solution.

**Recognize** the end point of a titration.

**Describe** the procedure for performing an acid-base titration.

**Determine** the molarity of a base.

Name ______________________ Class ______________ Date ______________

## Titration with an Acid and a Base *continued*

### MATERIALS

- 0.500 M HCl
- 50 mL burets, 2
- 100 mL beakers, 3
- 125 mL Erlenmeyer flask
- double buret clamp
- NaOH solution of unknown molarity
- phenolphthalein indicator
- ring stand
- wash bottle filled with deionized water

**Always wear safety goggles, gloves, and a lab apron to protect your eyes and clothing.** If you get a chemical in your eyes, immediately flush the chemical out at the eyewash station while calling to your teacher. Know the locations of the emergency lab shower and eyewash station and the procedures for using them.

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**Call your teacher in the event of a spill.** Spills should be cleaned up promptly, according to your teacher's directions.

**Never put broken glass into a regular waste container.** Broken glass should be disposed of properly.

## Procedure

1. Set up the apparatus as shown in **Figure A.** Label the burets *NaOH* and *HCl.* Label two beakers *NaOH* and *HCl.* Place approximately 80 mL of the appropriate solution into each beaker.
2. Pour 5 mL of NaOH solution from the beaker into the NaOH buret. Rinse the walls of the buret thoroughly with this solution. Allow the solution to drain through the stopcock into another beaker and discard it. Rinse the buret two more times in this manner, using a new 5 mL portion of NaOH solution each time. Discard all rinse solutions.

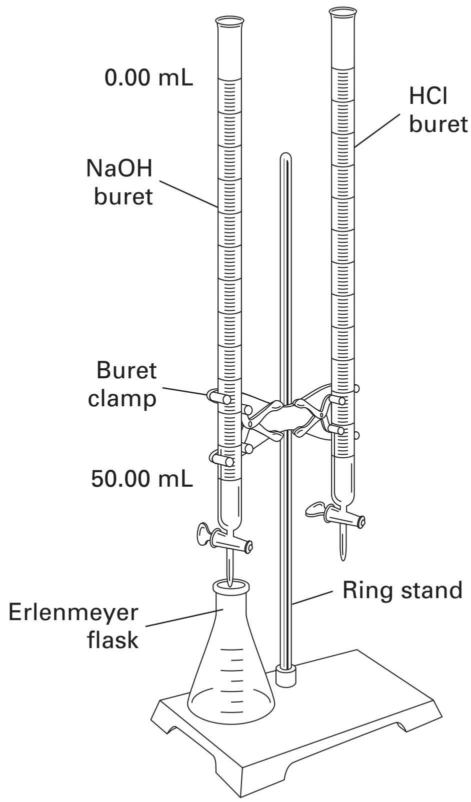


**Figure A**

Name ______________________ Class ______________ Date ______________



**3.** Fill the buret with NaOH solution to above the zero mark. Withdraw enough solution to remove any air from the buret tip, and bring the liquid level down within the graduated region of the buret.

**4.** Repeat steps **2** and **3** with the HCl buret, using HCl solution to rinse and fill it.

**5.** For trial 1, record the initial reading of each buret, estimating to the nearest 0.01 mL, in the **Data Table.** For consistent results, have your eyes level with the top of the liquid each time you read the buret. Always read the scale at the bottom of the meniscus.

**6.** Draw off about 10 mL of HCl solution into an Erlenmeyer flask. Add some deionized water to the flask to increase the volume. Add one or two drops of phenolphthalein solution as an indicator.

**7.** Begin the titration by slowly adding NaOH from the buret to the Erlenmeyer flask while mixing the solution by swirling it, as shown in **Figure B.** Stop frequently, and wash down the inside surface of the flask, using your wash bottle.

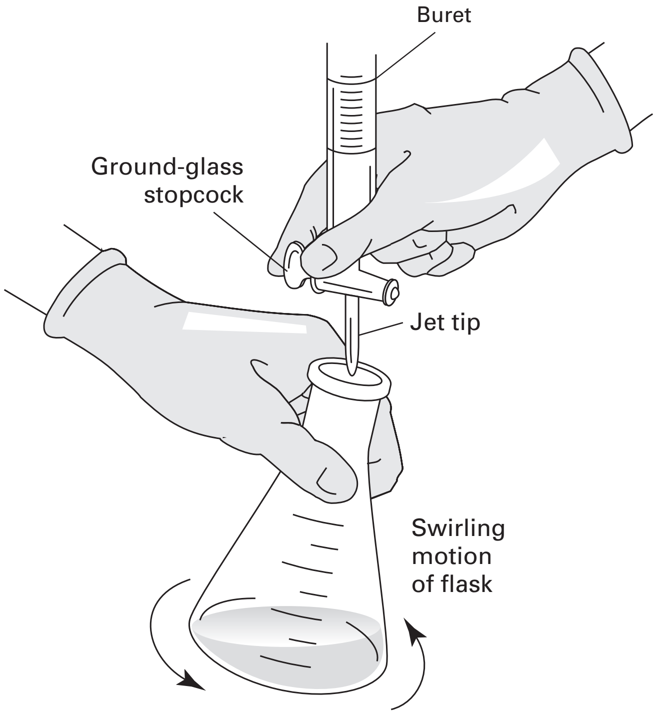


**Figure B**

8. When the pink color of the solution begins to appear and linger at the point of contact with the base, add the base drop by drop, swirling the flask gently after each addition. When the last drop added causes the pink color to remain throughout the whole solution and the color does not disappear, stop the titration. A white sheet of paper under the Erlenmeyer flask makes it easier to detect the color change.

9. Add HCl solution dropwise just until the pink color disappears. Add NaOH again, dropwise, until the pink color remains. Go back and forth over the end point several times until one drop of the basic solution just brings out a faint pink color. Wash down the inside surface of the flask, and make dropwise addition of NaOH, if necessary, to reestablish the faint pink color. Read the burets to the nearest 0.01 mL, and record these final readings in the **Data Table.**

10. Discard the liquid in the flask, rinse the flask thoroughly with deionized water, and run a second and third trial.

11. Record the known concentration of the standard HCl solution in the **Data Table.**

### DISPOSAL

12. Clean all apparatus and your lab station. Return equipment to its proper place. Dispose of chemicals and solutions in the containers designated by your teacher. Do not pour any chemicals down the drain or in the trash unless your teacher directs you to do so. Wash your hands thoroughly before you leave the lab and after all work is finished.

**Data Table**

| Buret readings (ml) | | | | |
|---|---|---|---|---|
| | HCl | | NaOH | |
| **Trial** | **Initial** | **Final** | **Initial** | **Final** |
| 1 | **0.70** | **10.90** | **5.08** | **13.58** |
| 2 | **10.90** | **20.80** | **13.58** | **21.80** |
| 3 | **20.80** | **28.81** | **21.80** | **28.33** |

**Molarity of HCl** **0.500 M**

## Analysis

1. **Organizing Data** Calculate the volumes of acid used in the three trials. Show your calculations and record your results below.

Trial 1: Volume of HCl = **10.90 mL − 0.70 mL = 10.20 mL**

Trial 2: Volume of HCl = **20.80 mL − 10.90 mL = 9.90 mL**

Trial 3: Volume of HCl = **28.81 mL − 20.80 mL = 8.01 mL**

Name ______________________ Class ______________ Date ____________

## Titration with an Acid and a Base *continued*

**2. Organizing Data** Calculate the volumes of base used in the three trials. Show your calculations and record your results below.

Trial 1: Volume of NaOH = **13.58 mL − 5.08 mL = 8.50 mL**

Trial 2: Volume of NaOH = **21.80 mL − 13.58 mL = 8.22 mL**

Trial 3: Volume of NaOH = **28.33 mL − 21.80 mL = 6.53 mL**

**3. Organizing Data** Use equation **3** in the introduction to determine the number of moles of acid used in each of the three trials. Show your calculations and record your results below.

Trial 1: Moles of acid = **0.500 M × 0.01020 L = 0.00510 mol**

Trial 2: Moles of acid = **0.500 M × 0.00990 L = 0.00495 mol**

Trial 3: Moles of acid = **0.500 M × 0.00801 L = 0.00401 mol**

**4. Relating Ideas** Write the balanced equation for the reaction between HCl and NaOH.

$\mathbf{HCl + NaOH \rightarrow NaCl + H_2O}$

**5. Organizing Ideas** Use the mole ratio in the balanced equation and the moles of acid from **Analysis** item **3** to determine the number of moles of base neutralized in each trial. Show your calculations and record your results below.

Trial 1: Moles of acid = **0.00510 mol = *moles of base***

Trial 2: Moles of acid = **0.00495 mol = *moles of base***

Trial 3: Moles of acid = **0.00401 mol = *moles of base***

**6. Organizing Data** Use equation **2** in the introduction and the results of **Analysis** item **2** and **5** to calculate the molarity of the base for each trial. Show your calculations and record your results below.

Trial 1: Molarity of NaOH $\dfrac{\textit{moles}\ \text{NaOH}}{\textit{volume}\ \text{NaOH}} = \dfrac{0.00510\ \text{mol}}{0.00850\ \text{L}} = 0.600\ \text{M}$

Trial 2: Molarity of NaOH $\dfrac{0.00495\ \text{mol}}{0.00850\ \text{L}} = 0.602\ \text{M}$

Trial 3: Molarity of NaOH $\dfrac{0.00401\ \text{mol}}{0.00653\ \text{L}} = 0.614\ \text{M}$

**7. Organizing Conclusions** Calculate the average molarity of the base. Show your calculations and record your result below.

$$\textbf{Average molarity of NaOH} = \frac{0.600\ \text{M} + 0.602\ \text{M} + 0.614\ \text{M}}{3} = 0.605\ \text{M}$$

Name ______________ Class ______________ Date ______________



## Conclusions

1. **Analyzing Methods** In step **6,** you added deionized water to the HCl solution in the Erlenmeyer flask before titrating. Why did the addition of the water not affect the results?

   **Diluting with water did not change the number of moles of HCl in the flask. NaOH was added until the number of moles of NaOH equaled the number of moles of HCl.**

2. **Analyzing Methods** What characteristic of phenolphthalein made it appropriate to use in this titration? Could you have done the experiment without it? How does phenolphthalein's end point relate to the equivalence point of the reaction?

   **Phenolphthalein changes from colorless in acid solution to pink in basic solution. Without an indicator such as phenolphthalein, there would have been no visual way to determine when the equivalence point had been reached. Phenolphthalein's end point is close to the equivalence point (neutral pH) of this titration.**