Name: ____________________________________ Date Due: ______________

Lab #27: Determination of Molarity and pH of a Sample of Norwegian Precipitation

Pre-Lab Activities

1. Read the **Introduction,** talking to the text and completing tasks along the way.
2. Re-read the entire **Introduction**. As you read, write down any remaining questions you have in the margins.
3. Pick 3 new terms in the **Introduction**, underline them, and define them in your own words in the margins. These terms do not need to be "science" words.
4. Read the **Materials and Methods,** talking to the text and completing tasks along the way.
5. Re-read the **Materials and Methods,** and think about the following question, "What is the relationship among *pH*, *concentration*, and *molarity*?" Write this relationship below:

6. Set up your lab notebook for Lab #27: Fill in the heading and write the purpose of Lab #27.
7. Familiarize yourself with the new lab techniques during your Lab Techniques Discussion. Record any notes to yourself.
8. As a lab team, practice your titration techniques. Record any pertinent notes on this handout.
9. Read relevant MSDS for this lab and record "Safety Precautions" in your lab notebook.
10. Reread the entire lab handout. Revisit any tasks in the first sections of the **Introduction** that you may not have completed, and complete tasks in the remaining sections of the **Introduction**.
11. As a group, decide which data you will be gathering during this lab, and construct appropriate data tables in your lab notebooks.

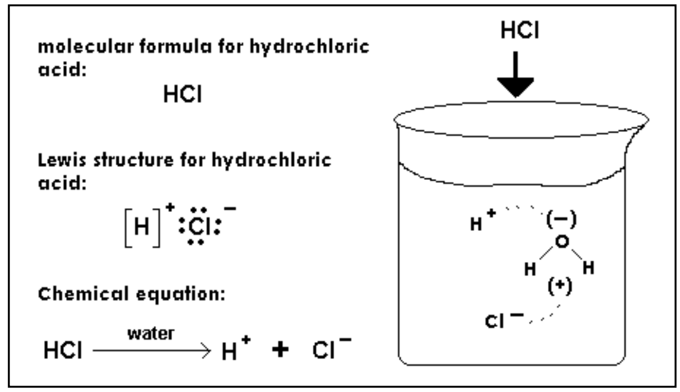
Introduction

The term *acid rain* appears in the media very often. For example, some countries such as Norway have claimed that other countries are sources of acid rain, and this acid rain travels through the atmosphere to harm their local environment. Is this possible? Lab #27 investigates the chemistry of acid rain in order to better understand how it impacts ecosystems. To do that, we need to revisit the concepts of acids and pH.

Acids

We defined *acids* in Lab #11: **An acid is a compound that will dissociate in solution to produce Hydrogen ions, H^+**. An example is hydrochloric acid, HCl. When HCl salts are placed in water, the hydrogen ions and chlorine ions *dissociate,* or separate from each other; each ion is attracted to oppositely charged ends of the polar water molecule (See **Figure 1**). This dissociation is the process of dissolving.

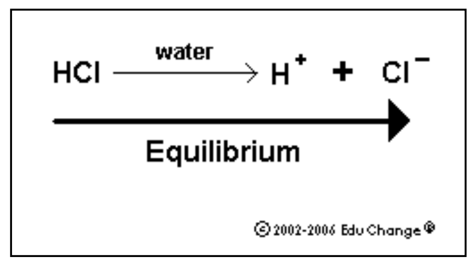
Figure 1: The Dissociation of Hydrochloric Acid



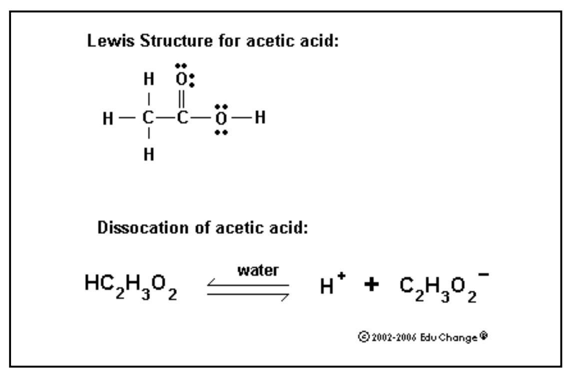
<u>Write your own definition of an acid below.</u>

<u>What Reference Table lists Common Acids?</u>

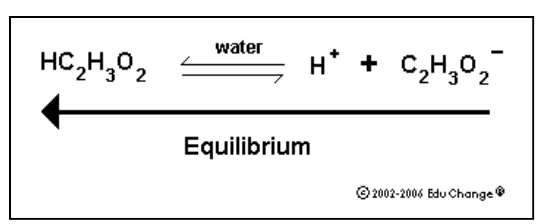
How much an acid dissociates in a particular solution is determined by the amount and type of solvent (such as water), and the amount and type of acid (such as HCl). Some acids tend to dissociate more readily in solution than others, and are classified as *strong acids*. The term strong refers <u>not</u> to the pH of the acid, since pH is entirely dependent upon the concentration of H^+, but instead to how completely an acid dissociates. Strong acids are defined as acids whose hydrogen ions (cations) completely dissociate from the anions and remain that way as long as they are in solution. This means that the reaction's equilibrium *lies towards the right*. In other words, the right side of the equation tends to describe the nature of the resulting solution. So once a strong acid dissociates, it stays that way.



Just as there are strong acids, there are *weak acids*. Again, this has nothing to do with the pH of the acid solution. Instead, weak acids refer to those acids that do not completely dissociate in solution. An example of this is acetic acid (CH_3COOH), otherwise known as vinegar. When acetic acid is placed in water, the first hydrogen atom will dissociate (See inset).



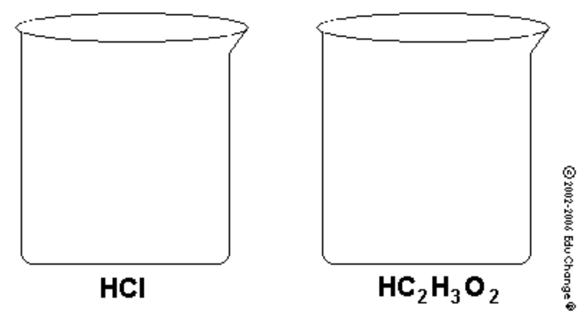
Unlike HCl, acetic acid has a tendency to return to its neutral, non-dissociated form rather than remain dissociated. Thus, the majority of the molecules will remain intact and only a few will dissociate. In this situation, we say that the reaction equilibrium lies to the left.



<u>Explain the difference between a strong acid and a weak acid in your own words in the space below.</u>

Before drawing the solutions in **Figure 2**, pretend that you are snapping a digital photo of each beaker to capture a single "moment in time." *In the beakers in Figure 2, sketch a solution of HCl and of $HC_2H_3O_2$, each containing 15 molecules.*

Figure 2: The Dissociation of Acidic Solutions



The "strength" of an acid is unrelated to the pH of that acid.

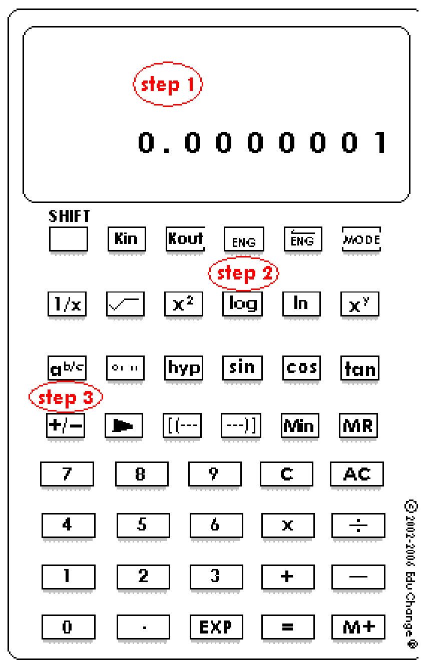
pH is a measure of the concentration of H^+ ions in a given solution. Values for ion concentrations are so low that they are expressed in scientific notation. A scientist named S.P.L Sorenson developed the concept of a pH value in 1909 to alleviate the tedious necessity of comparing exponential values of H^+ concentrations in different solutions. He chose to use the logarithm function in mathematics in order to express 10^{-7} as 7. The mathematical relationship is expressed as follows:

$$\mathbf{pH = -log[H^+]}$$

Let's try this on the calculator. Calculate the pH for all of the concentration values in the table.

Steps for Calculating pH on a Calculator

1. Press the "+/-" key. (change of sign)
2. Press the "log" key.
3. Enter the [H+]. (Molarity)



Concentration of Hydrogen Ions $[H^+]$	pH = -log$[H^+]$
1.0×10^{-7} M	
1.0×10^{-6} M	
1.0×10^{-5} M	
1.0×10^{-4} M	
1.0×10^{-3} M	
1.0×10^{-2} M	
1.0×10^{-1} M	
1 M	0

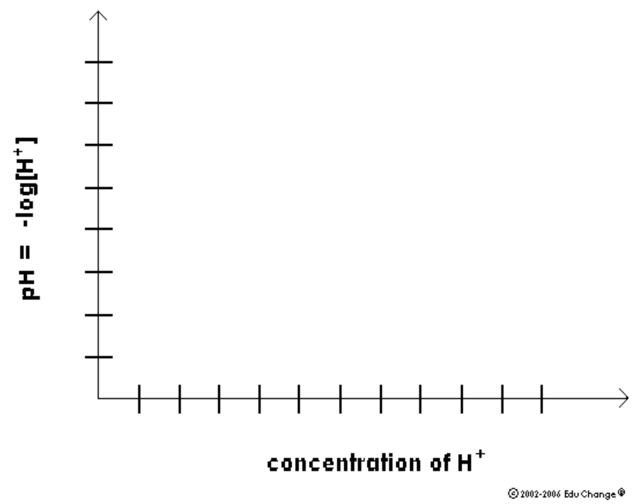
According to the table above, a solution that has a greater concentration of H^+ ions has a lower pH, or more acidic, solution. <u>In the space below, state the term you would use to describe the type of relationship between $[H^+]$ and pH?</u>

The pH scale is logarithmic, meaning that because our concentration units are expressed in scientific notation, each number on the pH scale actually relates to a base factor of 10, or ten times the concentration of H+ as you go down the pH scale.

Since molarity is the measure of the amount of a solute in a solvent, 1 M HCl has 10 times as many H^+ as 0.1 M HCl in the same volume of solution. *Use a calculator to determine the pH values of the solutions below. Then use the values to plot a graph of pH vs. Hydrogen Ion Concentration in* ***Figure 3****.*

Concentration of Solution	pH
1.0 M solution of HCl	
1.0×10^{-1}M solution of HCl	
1.0×10^{-2}M solution of HCl	
1.0×10^{-3} M solution of HCl	
1.0×10^{-4} M solution of HCl	
1.0×10^{-5} M solution of HCl	
1.0×10^{-6} M solution of HCl	
1.0×10^{-7} M solution of HCl	

Figure 3: pH Verses Concentration of H^+



What do you notice about the slope and shape of the graph plotted in Figure 3? Why does the graph look like this?

*** **To review:**

- The pH of a solution can be calculated by knowing the molarity of that solution.
- The value of pH increases or decreases by a factor of 10.
- The value of pH is indirectly proportional to the concentration of hydrogen ions in the solution: the greater the concentration of hydrogen ions, the lower the pH value.

The pH of Acid Rain

A common misconception about acid rain is that is has a very low pH. In fact, the pH of acid rain is only slightly more acidic (lower pH) than your average rain sample. Rain has an average pH of 5.5. Rain water is often a solution of H_2O and dissolved $CO_{2(g)}$. When carbon dioxide dissolves in water, it creates a small amount of carbonic acid that slightly lowers the pH. Thus compared to pure H_20, rainwater does not have a neutral pH of 7.

What is the molecular formula of carbonic acid? *(Hint: Use the* ***NYS Chemistry Reference Table K****)*

While average rainwater has a pH of about 5.5, acid rain has a pH in the range of 4.3-5.0. This may not seem like a big difference, but remember that these values represent factors of 10. So, while the pH may not appear significantly different between normal rain and acid rain, acid rain has many more hydrogen ions. These H^+ ions have the ability to damage the environment when they are in high concentrations.

Pollutant Acids

Acid rain results from pollutants dissolving into the rain as it condenses to form precipitation. When carbon dioxide dissolves in rainwater it creates a slightly acidic solution. Other dissolved pollutants, such as sulfur dioxide (SO_2) and nitrogen oxides (NO_x), cause acid rain. Acid rain forms when these gases react in the atmosphere with water, oxygen, and other chemicals.

*Find the formulas for sulfuric and nitric acid in your Reference Tables. Record 1) the table where you found their formulas and 2) the molecular formula in **Figure 4**.*

Figure 4: Sulfuric Acid and Nitric Acid

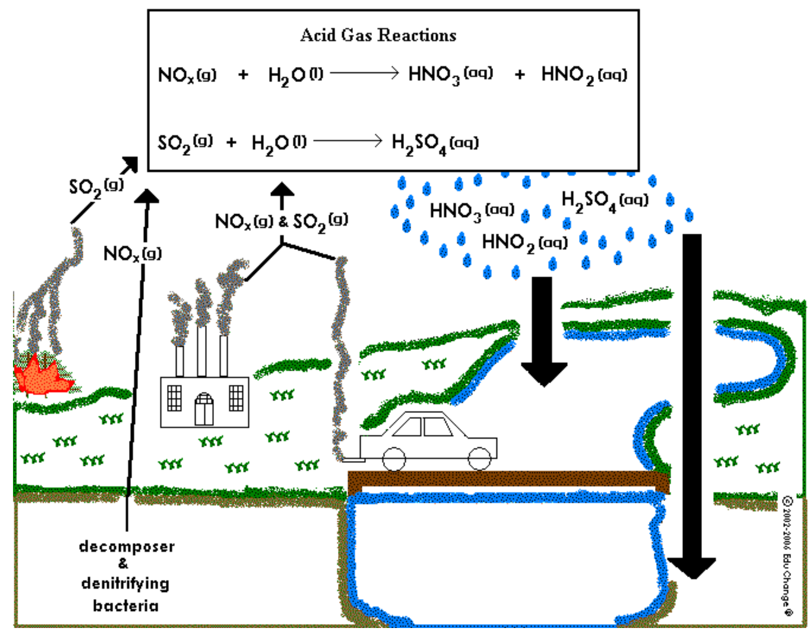
sulfuric acid: ____________

Reference Table with formulas:

nitric acid: ____________

Remember from Lab #18 that fossil fuels, such as coal and oil, have sulfur and nitrogen stored in them because they are formed from the decomposition of organic matter that initially contained sulfur and nitrogen. Factories and vehicles emit SO_2 and NO_x into the atmosphere, where they are converted into acid forms that later precipitate back to the ground or into waterways (See **Figure 5**).

Figure 5: Acid Rain Production and Deposition

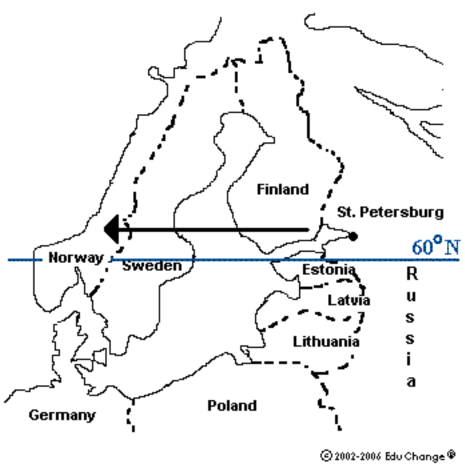


Global Reaches of Acid Rain

What is very interesting about acid rain is that often the geographic region sending the acid into the atmosphere is not the region suffering the impacts. For example, in the United States, weather tends to move from west to east and south to north. Like a well-adapted pollination strategy, a factory to the west can send its acid emissions into the air, only to touch ground miles and miles east of their starting point. Acid rain production verses deposition—and who is to blame—is an issue all over the world.

<u>Look at a world map</u>. If you travel west from Kodiak Island you reach areas with very few industrialized areas and they are a distance away. Kodiak Island has been spared massive acid rain deposition. However, if we travel across Asia to St. Petersburg and look at the distance separating it from Norway we will see a different situation. St. Petersburg has been industrialized for the last 300 years. In addition, weather patterns in the Baltic Sea often carry matter from Russia to the Scandinavian countries during the summer months. The Norwegians have been noting an increase in the acidity of their lakes and streams over the last 100 years. They believe that the lack of industrial pollution controls in St. Petersburg, as well as other parts of Eastern Europe, is part of their problem (See **Figure 6** on the next page).

Figure 6: Map of Eastern Europe



For a recent study on air pollution in St. Petersburg visit:
http://aix.meng.auth.gr/saturn/annualrep99/Genikhovich.pdf

In the space below, summarize why Norway has a greater acid rain problem than Kodiak Island.

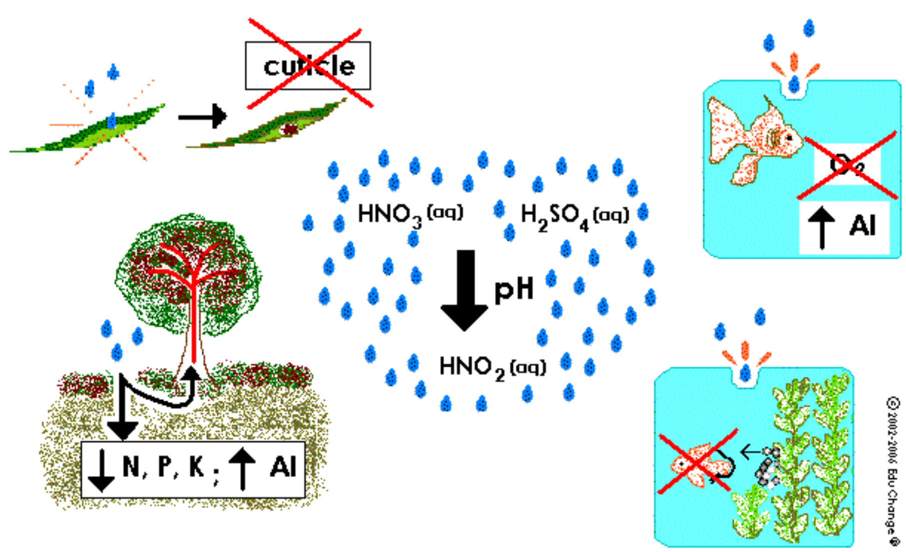
Acid Rain's Impact on Biotic Factors

As we've just studied in **Lab #26**, different particle composition affects a soil's capacity to absorb water. When acid is present, the soil may not have the *buffering capacity* to neutralize it. Buffering capacity is one of the ways that living systems maintain homeostasis. Soil can be seen as one main indicator of an ecosystem's homeostatic balance. Without a buffering capacity, soil pH would lower rapidly.

Low pH levels in the soil accelerate soil weathering and nutrient removal. When acid rain falls, the acidic rainwater dissolves the nutrients and helpful minerals in the soil. These minerals are then washed away before trees and other plants can use them to grow. Low soil pH creates conditions for higher metal solubility. This makes impact of the mineral runoff even more devastating.

*In **Figure 7**, circle the metal that is increasing in abundance and put a square around the other nutrients that are decreasing in abundance. Write the element names for each symbol inside **Figure 7**.*

Figure 7: Impacts of Acid Rain on Biotic Factors



It is interesting to compare the effects of acid rain on fish breathing with the effects of acids on human smokers. The pH level of fresh water and healthy human blood both falls between 6.0 and 8.0. Like acid rain, cigarettes decrease the pH of blood. For fish, when the surrounding environment of the gills/lungs becomes too acidic, mucus build-up prevents the assimilation of oxygen. How is this similar to what happens in the lungs of smokers?

Acid rain not only damages natural ecosystems, but also manmade materials and structures. Think back to the acid tests on stones like marble and sandstone that you conducted in **Lab #4**. Refer to the MSDS and think about the safety symbols and precautions you will use with corrosive chemicals. What is corrosive in the laboratory also eats away at buildings, monuments and statues. Ultimately, acid rain can cause structures to become unsafe or unusable.

In sum, acid rain impacts almost every substance that it comes into contact with. In recent years there have been many initiatives to reduce gas emissions, but it will take some time for the environment to respond to these improvements. *For more information, see the U.S. EPA Web site for the Acid Rain Program:* http://www.epa.gov/airmarkets/arp/overview.html.

In the space below, summarize some of the harmful effects of acid rain.

Materials and Methods

Titration: Purposes

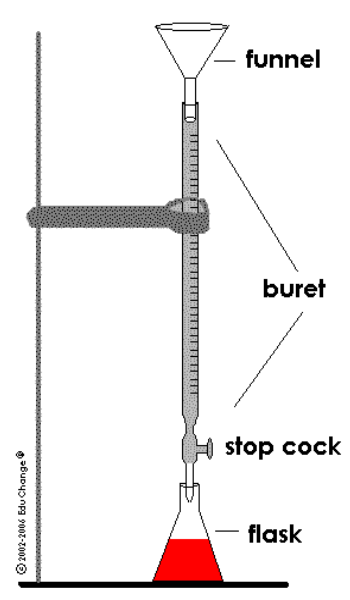
In Lab #27 we will determine whether or not the Norwegian Government has a reason to worry about their rain—or a case to make against the Russians. We will experimentally determine the pH of a Norwegian rain sample and decide whether or not it fits into the acid rain category.

We will be using a technique known as *titration*. Titration is a laboratory technique where we titrate, or add, a solution with known concentration into a solution of unknown concentration in an effort to determine the unknown concentration. The key to a titration is that the two solutions must chemically react with one another so that we can use balanced equations to "solve" the problem. Further, we use an indicator so that when the reaction occurs we can see it.

Titration: Methods

The titration technique relies upon known volumes and concentrations of *titrant,* the solution of known molarity, reacting completely with a sample of known volume but unknown concentration.

Figure 8: The Buret



The first technique is a buret titration (See **Figure 8**). We use a very precise piece of glassware known as a *buret* to deliver the titrant, in this case NaOH, to the unknown sample. A buret is graduated (has lines to indicate quantity) like a graduated cylinder so that it can be used to deliver exact volumes of a solution. At the bottom of the buret is a *stopcock* that allows you to control how much fluid from the buret is released; it functions like a valve. Just as with a graduated cylinder or a volumetric flask, you read a buret from the bottom of the meniscus (See **Figure 9**). We will be adding a base (NaOH) drop-wise from the buret into a flask containing the acid (rain samples). The chemical reaction between a base and an acid is known as a neutralization reaction.

Why do you think the reaction between an acid and a base is called a neutralization reaction?

By taking the volume of base (NaOH) added to our samples we will be able to calculate the molarity of the rain sample. We will also be able to determine the rain sample's pH. A simulation of a buret titration may be found at: http://oldmanhonda.com/Chemistry/WebLabs/Titration/Titration.html

When performing a titration you add drops until the *endpoint* of the chemical reaction is reached. This endpoint is the point at which all of the reactants have completely reacted. In the case of our neutralization reaction, it would be reached when all of the acid has reacted and the solution has reached a pH of 7 (neutral).

Endpoints can be indicated by any one of the basic signs of a chemical reaction. Since we cannot see a change of pH with the naked eye, we can use pH indicators to help us. We have used pH indicators to determine the relative pH of a solution (Lab #11, Lab #17). In a titration we can also use indicators to determine the endpoint of a reaction. In a neutralization titration, we can select an indicator that changes color at about pH 7. This will allow us to "see" when we have neutralized all of the given solution and therefore reached the endpoint of the reaction.

*Look at the **NYS Chemistry Reference Table M**. Select an indicator that will allow us to see the endpoint of a reaction between NaOH and HCl. Explain your selection.*

After completing the titrations you and your lab group will study the data, calculate the molarity and pH of the rain sample and then determine if the Norwegian Government has cause to be worried about the rain falling in their country.

Materials per Team

1, 50mL Burets
1 support stands
1 buret clamps
1 squirt bottle of distilled water
250mL Erlenmeyer flask
1 buret funnels
50mL of 1.0×10^{-3} M NaOH
25mL of Norwegian Rain
1 dropper bottle, phenolphthalein
Calculator

Procedure

PART A: Buret Titration

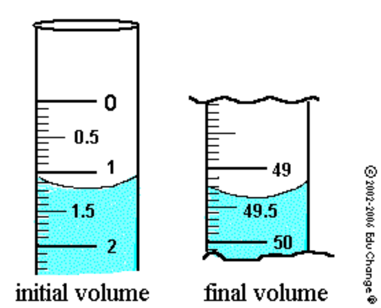
1. Obtain materials and set-up buret.
2. Make sure stop cock is clo<u>sed</u> and add about 10 mL of 1.0×10^{-3} M NaOH to base buret using funnel.
3. Carefully remove buret from the stand and carefully swirl the solution to rinse.
4. Drain rinse solutions and discard in WASTE beaker.
5. Carefully reclamp buret to the ring stand and check that the stopcock is <u>closed</u>.
6. Fill buret with 1.0×10^{-3} M NaOH. Use a funnel to prevent spillage.
7. Record starting volume of the NaOH.
8. Transfer 25 mL of rain sample to the 250 mL Erlenmeyer Flask.
9. Add 4-5 drops of phenolphthalein to the flask and swirl.
10. Place rain sample flask under the 1.0×10^{-3} M NaOH buret.
11. Add NaOH drop wise, swirling after each drop until the endpoint is reached.
12. Record the final volume of NaOH used.
13. Calculate the total volume of NaOH used for the titration. See **Figure 10**.
14. Empty contents of burets into WASTE beaker.
15. Return materials to prep area.
16. Wash hands and follow the directions in ***PART B*** to determine the Molarity and pH of the rain sample.

Figure 10: Determining the Volume of Titrant

final volume reading: 49.37mL

initial volume reading: 1.18mL

volume titrant: 48.19mL



PART B: Calculating the Molarity and pH of Rain Sample

17. Use the titration equation (see NYS Chemistry Reference Table T) to calculate the molarity of the rain sample. Record.

$$M_AV_A = M_BV_B$$

M_A = Molarity of Acid Solution
V_A = Volume of Acid Solution
M_B = Molarity of Base Solution
V_B = Volume of Base Solution

18. Then use pH = -log [H+] to calculate the pH of the rain sample. Remember: The $[H^+]$ is the molarity of the acid solution. Record.
19. Record molarity and pH for rain samples on class data table.
20. Copy the class data table into your notebook.
21. Calculate the average molarity and average pH of the class data.

Post-Lab Activities

1. Based on your data, should the Norwegian government be concerned about their rain? Why or why not?
2. What is the advantage of doing multiple trials in a titration? (Our lack of class time prohibited multiple trials with the buret titration that would normally occur.)