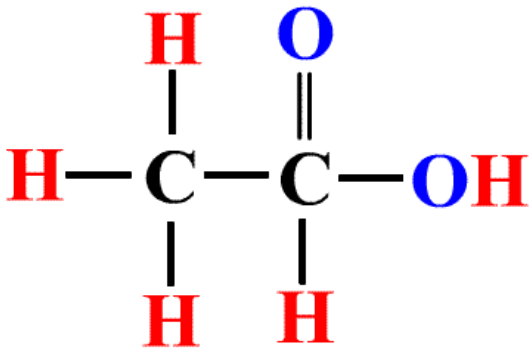
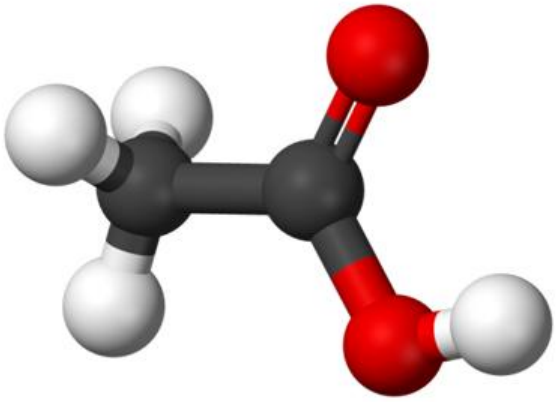
The Analysis of Vinegar

Vinegar is essentially a solution of acetic acid, CH_3CO_2H, in water. Acetic acid is an example of a carboxylic acid. Its structure is:





Acetic acid reacts with sodium hydroxide, a base, according to the reaction:

$$CH_3CO_2H + NaOH \rightarrow H_2O + Na\,CH_3CO_2 \quad \text{(Equation 1)}$$

acetic acid	sodium hydroxide	water	sodium acetate

This is an example of an acid-base neutralization reaction in which an acid and a base react to produce water plus a salt.

Purpose and Introduction to the Method

The purpose of this experiment is to determine the concentration of acetic acid in a sample of vinegar and to compare it with the federally required minimum concentration of 4 g of acetic acid per 100 mL of vinegar. The analytical method to be used makes use of the neutralization reaction between acetic acid and sodium hydroxide mentioned above.

The method to be used is the **titration method.** In this method, a sodium hydroxide solution of known molarity is contained in a **buret** and the acetic acid solution is contained in an **erlenmeyer flask**. See Figure 1 on the next page.

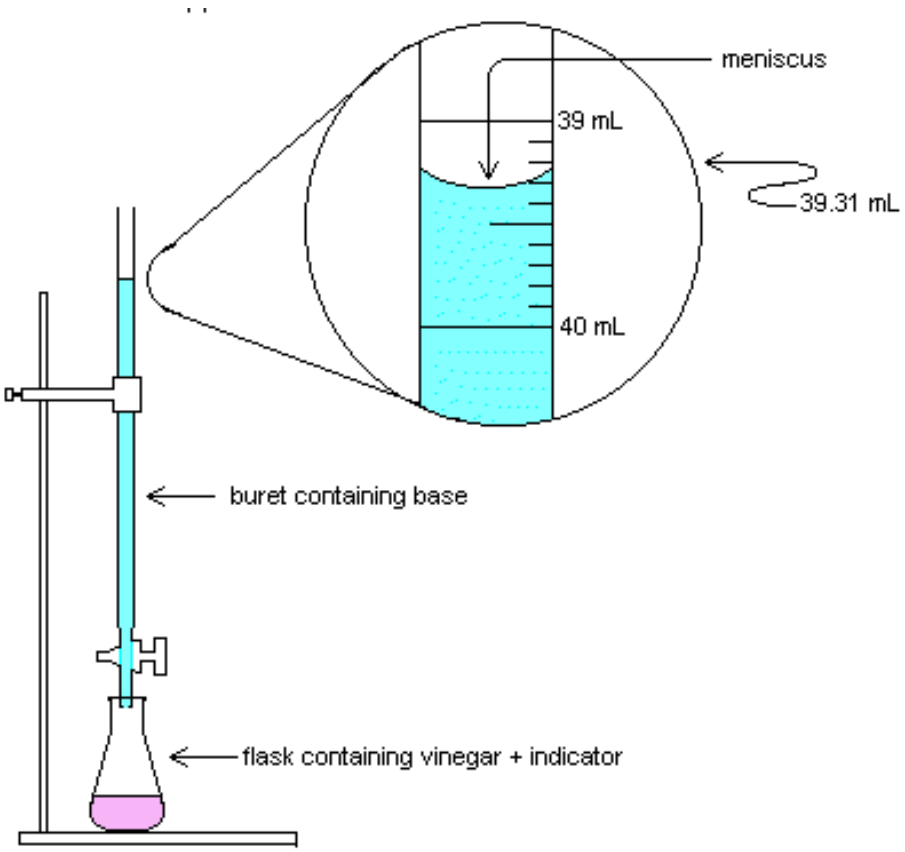


Figure 1.

Titration Apparatus

In the titration method, base is added to the acetic acid solution until just enough base has been added to completely react with all of the acid. The point where just enough base has been added to neutralize the acid is called the **equivalence point.** According to reaction (1), one mole of base reacts with one mole of acid. Therefore, at the equivalence point we have the relation

moles of base added = moles of acid initially present

moles of base added = molarity of base x volume of base added

and, therefore:

moles of acid initially present = molarity of base x volume of base added (2)

Example calculation:

A student titrates a **25.00 mL sample of vinegar** with 1.000 molar NaOH. The volume of base needed to reach the equivalence point is17.00 mL. What is the concentration of acetic acid in the vinegar in units of grams per 100 mL?

Solution:

From equation (2):

$$\text{moles of acid initially present} = 1.000\ \text{mol/L} \times 17.00\ \text{mL} \times 1\text{L}/(1000\text{mL})$$

$$= 0.0170\ \text{mol acetic acid}$$

This is the number of moles of acetic acid in 25.00 mL of vinegar.

The molecular formula of acetic acid is CH_3CO_2H. The molar mass is given by

$$\text{molar mass} = 2\text{C} \times 12.0\ \text{g/mol} + 4\text{H} \times 1.01\ \text{g/mol} + 2\text{O} \times 16.0\ \text{g/mol} = 60.0\ \text{g/mol}$$

The grams of acetic acid in 25.00 mL of vinegar is:

$$60.0\ \text{g/mol} \times 0.0170\ \text{mol} = 1.02\ \text{g acetic acid}$$

1 mL of vinegar contains 1.02g /25 grams per mL

The grams of acetic acid in 100 mL of vinegar is:

$$(1.02\text{g} / 25\text{mL})\ \text{g/mL} \times 100\ \text{mL} = 4.08\ \text{g}$$

therefore, this sample of vinegar meets the federal requirement of a minimum of 4 g of acetic acid per 100 mL of vinegar.

How do we know when the equivalence point has been reached? That is, how do we know when to stop adding base? The answer is that we add an **indicator**. An indicator is a substance whose color depends on the acidity of a solution. In this experiment we will be using the indicator **phenolphthalein**. Phenolphthalein is colorless in acidic solutions and pink in colorless solutions. The phenolphthalein is added to the vinegar solution before starting the titration. Near the equivalence point the indicator changes from colorless to pink.

Laboratory Procedure

Make sure that your buret is clean. Rinse it with approximately 10 - 15 mL of the 1.0 M NaOH solution to be used in the titration. Drain the solution through the buret tip. Fill the buret with the 1.0 M NaOH solution; make sure there are no air bubbles in the tip of the buret or just above the stopcock. Run the base out of the buret until the level is at 0.00 or below. Record the the level of the base, estimating the reading to two decimal places. Putting a white piece of paper with a thick black line behind the buret will help you see the **meniscus,** the curved surface of the liquid in the buret. See Figure 1 for a sample buret reading.

Pour about 30 mL of vinegar into a clean, dry beaker. Use a pipet to transfer 10 mL of vinegar to a clean 250 ml erlenmeyer flask. Add approximately 25 ml of distilled water to the vinegar sample and then add 3 drops of phenolphthalein to the vinegar sample.

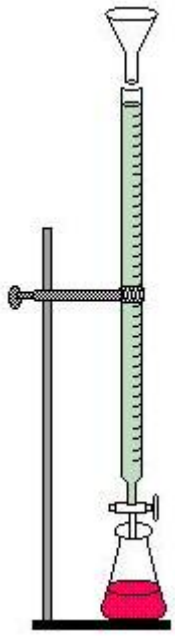
Slowly run the base out of the buret into the vinegar solution, swirling the flask and contents. As you approach the equivalence point, the area in the vinegar where the drop of base falls will turn pink; then the pink color will disappear as the solution becomes mixed. From this point on, add the base dropwise with constant swirling. Occasionally wash down the sides of the flask with water from your wash bottle. The equivalence point is where 1 drop (or less) of base causes the solution to become *very pale* pink throughout. essential ingredients of a successful titration include care and patience, so don't try to hurry. Record the final buret reading, estimating it to the nearest 0.01 mL.

Repeat the titration two more times - thoroughly rinse the flask between each trial. After the first titration, the others should go more quickly since you now have some idea of how much base is required per 10 mL sample of vinegar. The base may be added quickly until you are within 2 or 3 mL of the equivalence point; then change to dropwise addition.

For each titration, calculate the grams of acetic acid per 100 mL of vinegar. Calculate the average value. Show all calculations.

Vinegar Analysis

REPORT SHEET



NAME_____**DATE**_____

1. Create a data table for your data.
2. Show your calculation of the average grams of acetic acid per 100 mL of vinegar for each trial conducted.
3. Does the vinegar meet the federal requirement of 4 g of acetic acid per 100 mL of vinegar?