**Chemistry II—Lab #2: Determination of the Amount of Acetic Acid in Vinegar by Titration**

**Purpose:** The purpose of this lab is to ensure that the student understands the basic principles of titration, as well as how to properly use burettes and pipettes in the lab.

**Introduction:** A way of measuring the amount of a solute dissolved in aqueous solution is by **concentration**. There are several ways to indicate concentration. One method is to indicate the percent mass to volume ratio. This tells how many GRAMS of a substance are dissolved in **100 ml** of water. The **standard solution** in a titration is the solution of KNOWN concentration. You can determine the concentration of an unknown solute concentration by CAREFULLY controlling the addition of a measured amount of a standard solution that will completely react with the unknown solution. This is called **titration**.

Acid base titrations use an **indicator**. This is a weak acid or weak base that changes color depending on its state.

**H*In* ↔ H+ + *In-***

HIn and In- are different colors. Depending on the pH of the resulting solution, the proper indicator will tell you when you have reached the **equivalence point** in the reaction. In a neutralization reaction, this is where you have EQUAL amounts of **H3O+ and OH-** ions in solution. The **end point** is the point in the titration where the indicator changes color. If you are titrating an acid with a base, you KNOW the amount of **OH-** that was added, so therefore you also know the concentration of the ACID (unknown) at the EQUIVALENCE POINT! Phenolphthalein is an indicator that changes color at a pH GREATER than 7. This makes it a good indicator to use when titrating a weak acid with a strong base, since the salt solution will have a pH greater than 7. Phenolphthalein is colorless in acid, pink during the transition (equivalence point) and red in basic solutions.

In this experiment you will determine the concentration of acetic acid in vinegar by titrating the solution with a known amount of 1% NaOH (1 gram NaOH per 100 ml of water).

**Procedure:**

1. Obtain a 100ml beaker and fill with distilled water. Have EACH lab partner practice pipetting water into the sink using a 10 ml pipette. Having each lab partner practice this SEVERAL times will make sure that the pipettes are clean for the experiment. REMEMBER do NOT shake the pipette. Just touch the tip ONCE to the side of the container to get the last drop out. This will be exactly 10 ml of solution.

2. Rinse the pipette one last time with a little of the vinegar solution that you will use for the experiment.

3. Obtain a 50 ml burette, burette clamp and ring stand and set up apparatus as instructed by your teacher.

4. Be sure that the stopcock is in the CLOSED position (perpendicular to the burette) and rinse the burette several times with distilled water into the sink. Do NOT transfer any liquid into the burette when it is clamped in the apparatus. Hold the burette BELOW EYE LEVEL in the sink and fill it that way!

5. Rinse the burette ONCE with a little of the 1% NaOH solution that you will use in the experiment. Discard the rinse down the sink drain.

6. Fill the burette (IN THE SINK) to the zero mark with the 1% NaOH solution. Make your volume readings on the burette by reading the BOTTOM of the meniscus. You should estimate your readings to 1/100th of a ml. (Example volume = 5.73 ml)

7. Open stopcock and let the NaOH flow through until all the air bubbles are removed. Record the starting volume in your data table (this should NOT be zero!!!)

8. Using the 10 ml pipette, transfer 10 ml of vinegar into a CLEAN beaker. Add about 50 ml of distilled water (using a graduated cylinder) to the beaker.

9. Add 2 – 3 drops of phenolphthalein indicator to the beaker and stir. Place beaker on a white piece of paper on the ring stand under the burette.

10. SLOWLY add the NaOH from the burette while continuously stirring the beaker solution. When you begin to notice a color change, add the NaOH DROP BY DROP!

11. Continue adding the NaOH until the first tint of red appears and REMAINS after stirring. This is the **end point** of the reaction!

12. Read the burette volume again and record in your data table.

13. Repeat the titration (Steps 8 – 12) 2 more times so you have 3 volumes to average. Be sure that the beaker is CLEAN each time.

**DATA TABLE:**

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
|  | Trial 1 | Trial 2 | Trial 3 | Average |
| **1) Initial volume of NaOH in burette** |  |  |  | X |
| **2) Final volume of NaOH in burette** |  |  |  | X |
| **3) Volume of NaOH needed to neutralize 10ml of acetic acid**  **#2 - #1** |  |  |  |  |
| **4) Mass of NaOH needed to neutralize 10ml of acetic acid**  **#3 x 0.01** |  |  |  |  |
| **5) Mass of acetic acid neutralized by mass of NaOH**  **(#4 x 60) / 40** |  |  |  |  |
| **6) Mass of acetic acid in *100ml* of vinegar**  **#5 x 10** |  |  |  |  |
| **7) Percentage of acetic acid in the vinegar**  **(#6 / 100) x 100%** |  |  |  |  |

**Calculations and Questions:**

1. For each trial and then average, calculate the volume of 1% NaOH

needed to neutralize the acid in 10 ml of the vinegar.

Vol of NaOH needed = Final volume of NaOH – Initial volume of NaOH

2. For each trial and then average, calculate the mass of NaOH needed to

neutralize the acid in 10 ml of vinegar.

Since the NaOH is a 1% solution:

Mass of NaOH needed = (0.01)(Vol of NaOH needed)

3. Use a proportion to determine for each trial and then average the mass of

acetic acid (CH3COOH) that was neutralized by the mass of NaOH.

Mass of NaOH Mass of Vinegar

\_\_\_\_\_\_\_\_\_\_\_\_\_ = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

40 g NaOH/mol 60 g CH3COOH/mol

4. Determine for each trial and then average, the mass of acetic acid in

100ml of vinegar.

Since we know the mass now in 10 ml:

Mass of acetic acid in 100 ml = (Mass in 10 ml)(10)

5. Determine for each trial and then average, the percentage of acetic acid in

100 ml of vinegar.

Mass in 100 ml

% acetic acid = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ x 100%

100 ml

6. What was your final determination of the % of acetic acid in vinegar?

7. Look up the accepted value of acetic acid in vinegar and determine your

percent error.

8. What possible sources of error did you encounter in this experiment?

How can you remedy these errors the next time you do this experiment?

9. What have you learned from this procedure?