AIM

The main objective of this experiment is to determine weak acids in an unknown solution by using titrimetry.

INTRODUCTION

Neutralization of hydronium or hydroxide ion to form water is widely used as the basis for volumetric determinations of acids, bases and salts of weak acids. The reaction is characterized by a rapid change in pH near the equivalence point, a change that is readily detected by an acid-base indicator or that can be followed electrically by use of a pH meter. Neutralization titrations are performed with standard solutions of strong acids or strong bases. A standard solution (or standard titrant) is a reagent of exactly known concentration.

Standard solutions play a central role in all volumetric methods of analysis. The ideal standard solution for a volumetric method should

- be sufficiently stable so that it will not be necessary to determine its concentration (to standardize it) so frequently.
- react rapidly with the analyte so that the time required between each addition of the titrant is minimized.
- react more or less completely with the analyte so that satisfactory end points are realized.
- undergo a selective reaction with the analyte that can be described by a simple balanced equation.

The accuracy of a volumetric method is closely related with the accuracy of the concentration of the standard solution used in the titration. Two basic methods are used to determine the concentration of standard solutions: In the direct method, a carefully weighed quantity of a **primary standard** is dissolved and diluted to an exactly known volume in a volumetric flask. In the second method, the solution is standardized by titrating:

- a weighed quantity of a primary standard
- a weighed quantity of a secondary standard or
- a measured volume of another standard solution.

As a result, **standardization** is a process of determining the concentration of a substance in solution by adding a standard reagent of known concentration to it in carefully measured amounts until a reaction of definite and known proportion is completed and then calculating the unknown concentration.

A **titrant** that is standardized against a primary standard or against a standard solution which is called as a secondary solution. A secondary standard is less desirable than a primary standard solution, because the concentration of the former is subject to greater uncertainty.

A **primary standard** is a highly purified compound that serves as a reference material in all titrimetric methods. Important requirements for a primary standard are:

- High purity
- Stability in air
- Absence of hydrate water so that the composition doesn't change with variations in relative humidity
- Reasonable solubility in the titration medium
- Reasonably large formula weight so that the relative error associated with weighing is minimized
- Readily available at a modest cost

**The Effect of Carbon Dioxide on Standard Base Solutions**

In the preparation of acid and base solutions, atmospheric CO₂ may cause some problems. Atmospheric CO₂ is in equilibrium with aqueous carbonic acid whose concentration is about $1.5 \times 10^{-5}$ M at ordinary temperature. So this small concentration of carbonic acid leads to no significant error. On the other hand, distilled water is sometimes supersaturated with the CO₂ gas and thus contains sufficient carbonic acid to cause detectable errors. In order to test the distilled water to be used in neutralization titrations, take about 500 mL from the source of production. Add 5 drops of phenolphthalein and titrate with ~0.10 M NaOH. Less than 0.2 to 0.3 mL of base should be used to see the first faint pink color. If a larger volume is needed, standard solutions should be prepared from water that has been boiled for 2 to 3 minutes to remove CO₂ and then cooled to room temperature.

CO₂ (from the atmosphere) reacts with the hydroxides of sodium, potassium, and barium (solution or solid), producing carbonates:

$$
\text{CO}_2(g) + 2\text{OH}^- \rightarrow \text{CO}_3^{2-} + \text{H}_2\text{O}
$$

Absorption of CO₂ by a standardized solution of sodium or potassium hydroxide leads to a negative systematic error (called **carbonate error**), in analyses when a basic range indicator is used. Here, each carbonate ion (produced from two hydroxide ions) reacts with only one hydronium ion when the indicator changes its color:

$$
\text{CO}_3^{2-} + \text{H}_3\text{O}^+ \rightarrow \text{HCO}_3^- + \text{H}_2\text{O}
$$

However, no systematic error occurs when an acidic range indicator is used. Each carbonate ion produced will react with two hydronium ions of the acid:

$$
\text{CO}_3^{2-} + 2\text{H}_3\text{O}^+ \rightarrow \text{H}_2\text{CO}_3^- + 2\text{H}_2\text{O}
$$

The amount of hydronium ion consumed by this reaction equals the amount of hydroxide lost during formation of the carbonate ion, thus, no error occurs.

**Acid-Base Indicators**

In all neutralization titrations, the accuracy of the results are depend upon the selection of the proper indicator. The indicators most commonly used in neutralization titrations are highly colored organic compounds which have the property of changing color when the hydrogen ion concentration of the solution is changed over a certain range. The most widely used are phenolphthalein, methyl orange, methyl red and bromocresol green.

The hydronium ion concentrations, $\text{H}_3\text{O}^+$, at the point of color change (end point) are widely different for various indicators. Phenolphthalein is colorless in solutions in which [$\text{H}_3\text{O}^+$]$\geq 10^{-8}$ M and pink in solutions in which [$\text{H}_3\text{O}^+$]$\leq 10^{-10}$ M. If the hydronium ion concentration changes from $10^{-10}$ to $10^{-8}$ M, the color of the solution will change from pink to colorless. If the solution has a hydronium ion concentration within these two values, the color is an also intermediate between
deep pink and colorless. In other words, phenolphthalein shows a color change in the basic side of neutrality. Methyl orange, methyl red and bromocresol green give color changes on the acid side of neutrality, in the region of hydronium ion concentration $10^{-3}$ to $10^{-6}$ M.

A list of common indicators with pH range is given in the following table.

**Table 1. Some important acid-base indicators**

<table>
<thead>
<tr>
<th>Common Name</th>
<th>Transition Range, pH</th>
<th>Color Change*</th>
<th>Type**</th>
</tr>
</thead>
<tbody>
<tr>
<td>Thymol blue</td>
<td>1.2-2.8</td>
<td>R-Y</td>
<td>1</td>
</tr>
<tr>
<td>Methyl orange</td>
<td>3.1-4.4</td>
<td>R-O</td>
<td>2</td>
</tr>
<tr>
<td>Bromocresol green</td>
<td>3.8-5.4</td>
<td>Y-B</td>
<td>2</td>
</tr>
<tr>
<td>Methyl red</td>
<td>4.2-6.3</td>
<td>R-Y</td>
<td>2</td>
</tr>
<tr>
<td>Bromothymol blue</td>
<td>6.0-7.6</td>
<td>Y-B</td>
<td>2</td>
</tr>
<tr>
<td>Phenolphthalein</td>
<td>8.3-10.0</td>
<td>C-P</td>
<td>1</td>
</tr>
<tr>
<td>Tymolphthalein</td>
<td>9.3-10.5</td>
<td>C-B</td>
<td>1</td>
</tr>
<tr>
<td>Alizarin yellow GG</td>
<td>10-12</td>
<td>C-Y</td>
<td>2</td>
</tr>
</tbody>
</table>

*B=blue, C=colorless, O=orange, P=purple, R=red, Y= yellow

** (1) acid type:  \[ \text{HIn} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{In}^- \]

(2) base type: \[ \text{In}^+ + \text{H}_2\text{O} \rightleftharpoons \text{InH}^+ + \text{OH}^- \]

The proper indicator for a titration is the one that will exhibit a color change in the hydronium ion concentration range overlapping with at the equivalence point of the titration.

**Standard Solutions for Neutralizations Titrations**

Standard solutions used for neutralization titrations are usually solutions of strong acids or bases. The most frequently used acids are HCl and H$_2$SO$_4$, either of which is suitable for use as a permanent reference standard.

Solutions of sodium, potassium and barium hydroxides are the most frequently used as basic standard solutions. Bases, however, are not as good as acids as permanent standards, because they absorb carbon dioxide whenever they come in contact with air. Solutions of bases also react with glass during long storage period. Base solutions can, however, be stored in pyrex or plastic bottles (such as polyethylene) when the bottle is equipped with a siphon and the inlet tube is equipped with a tube of soda lime (mixture of Ca(OH)$_2$ and Na$_2$CO$_3$) to remove CO$_2$ from the entering air.

**REAGENTS AND APPARATUS**

- Unknown HCl and CH$_3$COOH solutions (2 replicates for each student)
- 0.5 L of 0.10 M NaOH for 2 students
- 1.0 L of 0.10 M HCl for 2 students*
- Dried potassium hydrogen phthalate, KHC$_8$H$_4$O$_4$ (KHP) as primary standard
- Dried sodium carbonate, Na$_2$CO$_3$ as primary standard
- Bromocresol green, phenolphthalein, methyl orange indicators, bromothymol blue indicator in droppers (ready)
- 1 bottle vinegar for 4 students (students` responsibility is to bring a bottle of vinegar to the lab.)
- Buret
• 2 Erlenmeyer flasks of 250 mL
• 1.0 L brown bottle for 2 students
• 0.25, 0.5 L and 1.0 L volumetric flasks for 2 students
• Wash bottle

* Do not pour standardized HCl at the end of this experiment. It will be later used in the analysis of carbonate mixture. Store HCl in brown colored 1 L bottle. **Do not forget to note the concentration of HCl** which will be used in the analysis of carbonate mixtures experiment as a standard solution.

**PROCEDURE**

**A. Preparation and Standardization of 0.10 M HCl**

1) Prepare 1.0 L of 0.10 M HCl solution from the concentrated HCl solution (for **two students**).

2) Transfer 0.20 to 0.25 g (**±0.1 mg**) of **dried** primary standard Na$_2$CO$_3$ to a 250 mL Erlenmeyer flask and add 50.0 mL of distilled water and 5 to 6 drops of bromocresol green indicator. **Make sure that Na$_2$CO$_3$ dissolves in water.**

3) Titrate the solution with 0.10 M HCl until the solution just changes its color from **blue** to **green** ($V_1$). The net reaction during titration can be written as follow:

$$\text{CO}_3^{2-} + 2\text{H}^+ \rightarrow \text{CO}_2 + \text{H}_2\text{O} \quad (\text{H}_2\text{CO}_3)$$

4) Place the flask on a wire gauze and heat just to gentle boiling for two or three minutes to remove CO$_2$ produced. Cool to the room temperature and complete the titration.

- If there is CO$_2$ in the solution, the indicator should change from green to blue as CO$_2$ is removed during heating. When you see blue color, titrate this solution with 0.10 M HCl till green color is observed and this additional amount of HCl has to be added to $V_1$ for finding the final HCl volume which is used to determine end point.

- As shown below, two end points are observed in the titration of sodium carbonate. The first at about pH 8.3, corresponding to the conversion of carbonate to hydrogen carbonate, the second one involving the formation of carbon dioxide, is observed at about pH 3.8.

The second end point is always used for standardization as the change in pH is larger than that of the first. An even sharper end point can be achieved by boiling the solution to eliminate carbonic acid and carbon dioxide. The sample is titrated to the first appearance of the acid color of the indicator (such as bromocresol green or methyl orange). At this point, it contains a large amount of dissolved carbon dioxide and small amounts of carbonic acid together with unreacted hydrogen carbonate. Boiling destroys this buffer produced by eliminating the carbonic acid:

$$\text{H}_2\text{CO}_3(\text{g}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$$

Solution then becomes alkaline again due to the residual hydrogen carbonate ion.

5) **Repeat the titration with another replicate.**

6) Your partner will also perform the standardization experiment twice. Record the results of all titrations, 4 titrations in total and calculate the individual molarity of the HCl solution for four replicates using the following equation.
\[ V_{\text{HCl}} \times M_{\text{HCl}} = 2 \times \frac{m_{\text{Na}_2\text{CO}_3}}{f_{\text{wNa}_2\text{CO}_3}} \]

where:
- \( V_{\text{HCl}} \) is the volume of HCl solution (in L).
- \( M_{\text{HCl}} \) is the molarity of HCl solution to be found.
- \( m_{\text{Na}_2\text{CO}_3} \) is the mass of \( \text{Na}_2\text{CO}_3 \) taken (in g).
- \( f_{\text{wNa}_2\text{CO}_3} \) is the molar mass of \( \text{Na}_2\text{CO}_3 \), 105.99 g/mol.

and then calculate the **average molarity of HCl solution**.

**B. Preparation and Standardization of 0.10 M NaOH**

1) Boil ca. 500 mL of distilled water and then cool it. Use this distilled water to prepare 500 mL of 0.10 M NaOH solution **(for 2 students)**.

2) Transfer 0.100 to 0.200 g (±0.0001 g) of potassium hydrogen phthalate (\( \text{KHC}_8\text{H}_4\text{O}_4 \)/KHP) into a 250 mL Erlenmeyer flask. Dissolve the salt in 50 mL of distilled water and then add 5 to 6 drops of phenolphthalein indicator. **Make sure that KHP dissolves in water before starting the titration.**

3) Titrate the solution with 0.10 M NaOH prepared until a faint pink color appears and persists at least 30 seconds. The net reaction during titration can be written as follow:

\[ \text{HP}^- + \text{OH}^- \rightarrow \text{H}_2\text{O} + \text{P}^{2-} \]

Where, \( \text{HP}^- \) is hydrogen phthalate ion coming from KHP, \( \text{OH}^- \) is hydroxide ion coming from \( \text{NaOH} \) and \( \text{P}^{2-} \) is phthalate ion.

4) **Repeat the titration with another replicate.**

5) Your partner will also perform the standardization experiment **twice**. Record the results of all titrations, 4 titrations in total and calculate the individual molarity of the NaOH solutions for four replicates using the following equation:

\[ V_{\text{NaOH}} \times M_{\text{NaOH}} = \frac{m_{\text{KHP}}}{f_{\text{wKHP}}} \]

where:
- \( V_{\text{NaOH}} \) is the volume of NaOH (in L).
- \( M_{\text{NaOH}} \) is the molarity of NaOH solution to be found.
- \( m_{\text{KHP}} \) is the mass of KHP taken (in g).
- \( f_{\text{wKHP}} \) is the molar mass of KHP, 204.24 g/mol.

and then calculate the **average molarity of NaOH solution**.

**C. Selection of Proper Indicator(s)**

In the titration of strong acids and bases the pH at the stoichiometric point is **always** 7.00. Changing in pH around equivalence point becomes very large in case of high concentration of reactants as shown in the following figure. For this type of titration (**Curve A**) either phenolphthalein (base range indicator) or bromocresol green (acid range indicator) can be used in addition to neutral range indicator, like bromotymol blue. However, as the concentrations of
the analyte and titrant decrease (Curve B) the magnitude of the pH break decreases. In this case, transition ranges of phenolphthalein and bromocresol green lie outside the equivalence region and only bromothymol blue turns to be a suitable indicator.

![Figure 1](image)

**Figure 1.** Curve for the titration of HCl with standard NaOH. **Curve A**: 50.00 mL of 0.0500 M HCl with 0.1000 M NaOH. **Curve B**: 50.00 mL of 0.000500 M HCl with 0.00100 M NaOH.

The following experiment is designed to demonstrate the effects of titrant/analyte concentrations on the magnitude of pH change around the equivalence point.

1) Prepare 250 mL of 0.0010 M HCl and 0.0010 M NaOH solutions from 0.10 M HCl and 0.10 M NaOH, respectively for two students. Do the following three titrations (i-iii).

   **i-**
   1. Take 30.0 mL of 0.0010 M NaOH into a 250 mL conical flask.
   2. Add 5 drops of phenolphthalein indicator (pink).
   3. Titrate with 0.0010 M HCl solution until the color disappears.
   4. Record the volume of HCl solution required ($V_{phth}$).

   **ii-**
   1. Take 30.0 mL of 0.0010 M NaOH solution into a 250 mL conical flask.
   2. Add 5 drops of bromocresol green indicator (blue).
   3. Titrate with 0.0010 M HCl solution until green color is observed.
   4. Record the volume of HCl solution required ($V_{bcg}$).

   **iii-**
   1. Take 30.0 mL of 0.0010 M NaOH solution to the conical flask.
   2. Add 10 drops of bromothymol blue indicator (blue-green).
   3. Titrate with 0.0010 M HCl until green color is observed.
   4. Record the volume of HCl solution required ($V_{btb}$).

2) Calculate the theoretical volume of 0.0010 M HCl ($V_{HCl}$) required to neutralize 30.0 mL of 0.0010 M NaOH solution (equivalence point) using the following equation:

$$V_{HCl} \times M_{HCl} = V_{NaOH} \times M_{NaOH}$$

3) Compare $V_{HCl}$(theoretical) and experimental $V_{phth}$, $V_{bcg}$ and $V_{btb}$ values and select the most proper indicator(s). Discuss your reason(s) of choosing that indicator.

**D. Determination of Weak Acids**

**Treat each unknown solution separately.**

1) Take two replicate acetic acid solution (CH$_3$COOH) with unknown concentration from your assistant.
2) Add 50.0 mL of distilled water and 5 to 6 drops of phenolphthalein and then titrate using standardized NaOH solution to the first permanent (~30 s) pink color. The net reaction during the titration is:

\[ \text{CH}_3\text{COOH} + \text{OH}^- \rightarrow \text{CH}_3\text{COO}^- + \text{H}_2\text{O} \]

3) Calculate the total acidity as mg CH$_3$COOH using the following equation for each unknown sample separately and then calculate the average of the results.

\[ V_{\text{NaOH}} \times M_{\text{NaOH}} = \frac{m_{\text{CH}_3\text{COOH}}}{f_{\text{mCH}_3\text{COOH}}} \]

where:

- $V_{\text{NaOH}}$ is the volume of NaOH (in L).
- $M_{\text{NaOH}}$ is the molarity of NaOH solution found in part B.
- $m_{\text{CH}_3\text{COOH}}$ is the mass of CH$_3$COOH in g to be found.
- $f_{\text{mCH}_3\text{COOH}}$ is the molar mass of CH$_3$COOH, 60.05 g/mol.

**E. Analysis of a HCl and CH$_3$COOH Mixture**

**Treat each unknown solution separately.**

1) Obtain two replicate unknown solution from your assistant.
2) Add 50.0 mL of distilled water and 3 to 5 drops of methyl orange indicator and then titrate with standard NaOH solution until color changes ($V_{\text{Mo}}$).
3) Now add 5 to 6 drops of phenolphthalein indicator to the same solution and then titrate until the appearance of pink color ($V_{\text{phth}}$).
4) Determine the amount of HCl in mg from the volume and molarity of NaOH used to obtain first end point (methyl orange end point, $V_{\text{Mo}}$).
5) Determine the amount of CH$_3$COOH in mg from the volume and molarity of NaOH.

**F. Determination of Acetic Acid in Commercial Vinegar**

1) Pipette exactly 2.0 mL of the commercial vinegar sample into a 250 mL Erlenmeyer flask and add 50.0 mL of distilled water.
2) Add 5 to 6 drops of phenolphthalein indicator and then titrate the acetic acid solution with the standard NaOH solution to a pale pink equivalence point. Record the burette readings.
3) Repeat the titration twice more using a fresh aliquot of vinegar. Results should agree within ±0.2 mL or additional titrations are required.
4) Calculate % (w/v) of acetic acid present in the vinegar.

**PRE-LAB STUDIES**

*Read Titrations in Analytical Chemistry and Principles of Neutralization Titrations from the textbook.*

1) What are the requirements for a primary standard?
2) What is the aim of standardization?
3) What is the difference between primary standard and standard solution?
4) Define and compare end-point and equivalence point.
5) Why weak acids cannot be used as a titrant in the titrations?
6) Why do the indicators change their colors at different pH's. (Hint: Consider the chemical behavior of the indicators.)
7) What is the difference between the shapes of titration curves for dilute and concentrated solutions for the same analyte and titrant? For which of them the indicator selection is more important? Why?
8) Describe the preparation of 1.0 L of 0.10 M HCl from concentrated HCl solution.*
9) Describe the preparation of 1.0 L of 0.10 M NaOH from solid NaOH.
   * Write the density and mass percent values from the concentrated HCl solution bottle in the lab. These values are also listed in the web page.

POST-LAB STUDIES

1) Explain the effect, if any, of each of the following sources of error on the molarity of the base as determined in the experiment; i.e., would the experimental value for molarity be too high or too low, and why?
   a) If Erlenmeyer flask in which the titration was performed contained several mL of distilled water from the rinsing at the time the KHP was weighed into it.
   b) If the tip of the buret was not filled with solution before the initial reading was taken.
   c) If a bubble appeared in the tip during the titration.
   d) If liquid splashed from the titration beaker before the end point had been reached.
   e) If the buret was not rinsed with the NaOH solution following the rinsing with distilled water.
2) Suppose that, instead of using NaOH, a base such as Ba(OH)₂ had been used. What changes in the calculations would then have to be done to determine the molar concentrations of the base? Answer this question in words, and illustrate your answer by calculating molarity from the following data:

   mol KHP used 0.040 mol
   Ba(OH)₂ Initial buret level = 0.020 mL
   Final buret level = 36.70 mL

3) Write the name of primary standards used for the standardization of HCl and NaOH? Could you use them interchangeably? Why should we pay attention to measure the mass of the primary standard precisely for standardization of NaOH and HCl?
4) Why do we boil and then cool distilled water before the preparation of NaOH solution? Why this process is not necessary for the preparation of HCl?
5) In the titration of a weak and strong acid mixture with a strong base, which acid is titrated first? Write the related titration reactions by drawing the corresponding titration curve.
6) For the titration of a strong acid (HCl), a neutral range indicator is used and for the titration of weak acids (CH₃COOH) a basic range indicator is used. However, for the titration of weak and strong acid mixture you used an indicator of acidic range for the first end-point. Why?
7) In section C, Selection of Proper Indicator(s):
   a) What is the reason of getting different end points (Vₚₜₜ ≠ V₂ₜₜ ≠ Vₜₜₜₜ)?
   b) Discuss which indicator you choose and why?
   c) After adding bromocresol green to 0.0010 M NaOH, what color did you observe? Explain the color change by relating the pH of the solution.
**REPORT SHEET**  
**NEUTRALIZATION TITRATIONS**

**A. Standardization of HCl using Na$_2$CO$_3$:**

<table>
<thead>
<tr>
<th>Replicate Number</th>
<th>Na$_2$CO$_3$ (g)</th>
<th>Volume of 0.10 M HCl (mL)</th>
<th>Concentration of HCl, M</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>2</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>3</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>4</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Average ($\bar{x} \pm s$), M

**B. Standardization of NaOH using KHP:**

<table>
<thead>
<tr>
<th>Replicate Number</th>
<th>KHP (g)</th>
<th>Volume of 0.10 M NaOH (mL)</th>
<th>Concentration of NaOH, M</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>2</td>
<td></td>
<td></td>
<td></td>
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<td>3</td>
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<tr>
<td>4</td>
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</tr>
</tbody>
</table>

Average ($\bar{x} \pm s$), M

**C. Selection of Proper Indicator(s)**

Volume of 0.0010 M HCl to reach end point using

i- phenolphthalein indicator, $V_{phth} =$…………………..mL

ii- bromocresol green indicator, $V_{bcg} =$…………………..mL

iii- bromothymol blue indicator, $V_{btb} =$…………………..mL
D. Analysis of Weak Acids

<table>
<thead>
<tr>
<th>Replicate Number</th>
<th>Volume of 0.10 M NaOH (mL) phenolphthalein end point</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td></td>
</tr>
<tr>
<td>2</td>
<td></td>
</tr>
</tbody>
</table>

Average mass of CH₃COOH in mg (\bar{x}±s) and RSD%=

The following information (true values) will be sent to your e-mail address:
Concentration of CH₃COOH in the unknown sample=
Volume of CH₃COOH unknown, mL=

% relative error=

E. Analysis of a HCl and CH₃COOH Mixture

<table>
<thead>
<tr>
<th>Replicate Number</th>
<th>Volume of 0.10 M NaOH (mL) methyl orange end point, V_MO</th>
<th>Volume of 0.10 M NaOH (mL) phenolphthalein end point, V_phth</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Average mass of HCl in mg (\bar{x}±s) and RSD%=
Average mass of CH₃COOH in mg (\bar{x}±s) and RSD%=

The following information (true values) will be sent to your e-mail address:
Concentration of HCl in the unknown sample=
Volume of HCl unknown, mL=
Concentration of CH₃COOH in the unknown sample=
Volume of CH₃COOH unknown, mL=

% relative error (HCl)=
% relative error (CH₃COOH)=

F. Determination of Acetic Acid in Commercial Vinegar

<table>
<thead>
<tr>
<th>Replicate</th>
<th>Volume of 0.1 M NaOH (mL) phenolphthalein end point</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td></td>
</tr>
<tr>
<td>2</td>
<td></td>
</tr>
<tr>
<td>3</td>
<td></td>
</tr>
</tbody>
</table>

% (w/v) of acetic acid in vinegar, (\bar{x}±s) and RSD%=

TA’s Name and Signature: