Conductometry is a measurement of electrolytic conductivity to monitor a progress of chemical reaction. Conductometry has notable application in analytical chemistry, where conductometric titration is a standard technique. Conductometric titration is a type of titration in which the electrolytic conductivity of the reaction mixture is continuously monitored as one reactant is added. The equivalence point is the point at which the conductivity undergoes a sudden change. Marked increases or decrease in conductance are associated with the changing concentrations of the two most highly conducting ions, namely $H^+$ and $OH^−$.

Acid-base titrations and redox titrations are often performed in which indicators are commonly used to signal the end point by changing color e.g., methyl orange or phenolphthalein for acid base titrations and starch solutions for iodometric type redox process. However, electrical conductance measurements can also be used as a tool to locate the end point, e.g., in a titration of a $H_2SO_4$ solution with the base $Ba(OH)_2$.

The conductivity of a solution is dependent on several factors, including the concentration of the solute, the degree of dissociation of the solute, the valence of the ion(s) present in the solution, the temperature, and the mobility of the ions in the solution.

As the titration progresses, the $H^+$ ions are neutralized to form water by the addition of $OH^−$ ions or vice versa. For each amount of $OH^−$ added equivalent amount of hydrogen ions is removed. Effectively, the faster moving $H^+$ cation is replaced by the slower moving $Ba^{2+}$ ion, and the conductivity of the titrated solution as well as the measured conductance of the solution fall. This continues until the equivalence point is reached, at which we have a solution of barium sulfate, $BaSO_4$. If more acid or base is added an increase in conductivity or conductance is observed, since more ions are being added and the neutralization reaction no longer removes an appreciable number any of them. Consequently, in the titration of a strong acid with a strong base, the conductance has a minimum at the equivalence point. This minimum can be used instead of an indicator dye to determine the endpoint of the titration.

In this experiment, you will monitor conductivity during the reaction between sulfuric acid, $H_2SO_4$, and barium hydroxide, $Ba(OH)_2$, in order to determine the equivalence point. From this information, you can find the concentration of the $Ba(OH)_2$ solution.

In this reaction, the total number of dissociated ions in solution is reduced dramatically during the reaction as a precipitate is formed. As 0.100 $M$ $H_2SO_4$ is slowly added to $Ba(OH)_2$ of unknown concentration, changes in the conductivity of the solution will be monitored quantitatively using a conductivity probe or qualitatively using a conductivity apparatus. When the probe is placed in a solution that contains ions, and thus has the ability to conduct electricity, an electrical circuit is completed across the electrodes that are located on either side of the hole near the bottom of the probe body. This results in a conductivity value that can be read by the interface. The unit of conductivity used in this experiment is the microsiemens per cm, or $µS/cm$.

**OBJECTIVES**

In this experiment, you will

- Measure the conductivity of the reaction between sulfuric acid and barium hydroxide.
- Use conductivity values as a means of determining the equivalence point of the reaction.
- Calculate the molar concentration of a barium hydroxide solution.
Conductometric Titration

MATERIALS

Data Collection Device 0.10 M sulfuric acid, H₂SO₄, solution
Conductivity Probe or Conductivity Apparatus Barium hydroxide, Ba(OH)₂, solution of unknown concentration
ring stand 50 mL buret
250 mL beaker utility or buret clamp
magnetic stirrer and stirring bar if available 50 mL graduated cylinder
balance, ±0.01 gram accuracy (or better) distilled water

PROCEDURE

1. Obtain and wear goggles.

2. Dispense approximately 10 mL of the Ba(OH)₂ solution of unknown concentration into a clean, dry 250 mL beaker. **Record the precise volume of the analyte. CAUTION: The barium hydroxide solution is caustic. Avoid spilling it on your skin or clothing.**

3. Place the beaker on a magnetic stirrer and add a stirring bar. If no magnetic stirrer is available, you will stir with a stirring rod during the titration.

4. Set up the data collection system for events with entry. If using a conductivity probe set the selector switch on the Conductivity Probe to the 0-2000 µS/cm range and set it up as shown in Figure 1. If using a conductivity apparatus, follow the directions provided by your teacher.

5. Use a buret clamp to connect a 50 mL buret to the ring stand (see Figure 1). Rinse and fill the buret with 0.100 M H₂SO₄ solution. **CAUTION: H₂SO₄ is a strong acid, and should be handled with care.**

6. Conduct the titration carefully. A good way to proceed is to add 1.0 mL of 0.100 M H₂SO₄ at a time until the conductivity has dropped below 100 µS/cm, then add the acid drop-by-drop to best identify the equivalence point.

7. Examine the data on the displayed graph to determine the equivalence point; that is, the volume when the conductivity value reaches a minimum. Determine the H₂SO₄ volume of the point with the minimum conductivity value and record it in the data table.

8. When you have completed the titration, filter the reaction mixture as directed by your teacher.

9. Rinse the Conductivity Probe with distilled water in preparation for the second titration.

10. Save or print a copy of the original titration graph.

11. Repeat the necessary steps to conduct a second titration. Conduct a third trial, if needed. Record the results in the data table.
SAMPLE DATA

<table>
<thead>
<tr>
<th>Volume of Ba(OH)₂ sample</th>
<th>Trial 1</th>
<th>Trial 2</th>
<th>Trial 3</th>
</tr>
</thead>
<tbody>
<tr>
<td>Equivalence point (mL)</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

PRE-LAB QUESTIONS

1. Write the net ionic equation for the reaction taking place in this laboratory exercise.
2. What is the name and chemical formula of the analyte?
3. What is the name and chemical formula of the titrant?
4. What is the reading on the buret pictured right? Explain your reasoning.
5. Why is it important to rinse the buret before filling it with titrant?
6. List the four factors upon which the conductivity of a solution depends.
7. A student conducts a titration using a conductivity probe and collects the following data.

**Conductivity Titration Data for the Reaction of Ba(OH)₂ and H₂SO₄**

<table>
<thead>
<tr>
<th>H₂SO₄ added (mL)</th>
<th>Conductivity (µS/cm)</th>
<th>H₂SO₄ added (mL)</th>
<th>Conductivity (µS/cm)</th>
<th>H₂SO₄ added (mL)</th>
<th>Conductivity (µS/cm)</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.0</td>
<td>969</td>
<td>12.0</td>
<td>117.9</td>
<td>14.5</td>
<td>38.3</td>
</tr>
<tr>
<td>1.0</td>
<td>908</td>
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<td>55.8</td>
<td>15.0</td>
<td>64.1</td>
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<td>2.0</td>
<td>834.6</td>
<td>13.5</td>
<td>29</td>
<td>16.0</td>
<td>159.3</td>
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<tr>
<td>3.0</td>
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<td>13.6</td>
<td>20.7</td>
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<tr>
<td>4.0</td>
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<td>13.7</td>
<td>16.5</td>
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<td>13.4</td>
<td>19.0</td>
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<td>6.0</td>
<td>532.6</td>
<td>13.9</td>
<td>12.4</td>
<td>20.0</td>
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<tr>
<td>7.0</td>
<td>462.3</td>
<td>14.0</td>
<td>13.4</td>
<td>21.0</td>
<td>625.7</td>
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<tr>
<td>8.0</td>
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<td>14.5</td>
<td>22.0</td>
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<tr>
<td>11.0</td>
<td>184.1</td>
<td>14.4</td>
<td>32.1</td>
<td>25.0</td>
<td>968</td>
</tr>
</tbody>
</table>

(a) What is the lowest conductivity recorded?

(b) What volume of 0.100 M H₂SO₄ corresponds to the lowest conductivity recorded?

(c) Use your titration results to calculate the molar concentration (molarity) of the Ba(OH)₂ solution using the molar amount of H₂SO₄ used in this trial.
POST-LAB QUESTIONS AND DATA ANALYSIS

1. What are the two types of chemical reactions occurring simultaneously in this experiment?

2. Write the balanced net ionic equation for each of the types of reactions you identified in Question 1.

3. Use the space provided below to draw a rough sketch of the graph of your data for your best trial:

4. Use the titration results to calculate the moles of H₂SO₄ that were used to reach the equivalence point in each trial.

5. Use your titration results to calculate the molar concentration (molarity) of the Ba(OH)₂ solution using the molar amount of H₂SO₄ used in each trial.

6. A student fails to rinse the buret with titrant prior to filling it. What effect does this have on the calculated value of the Ba(OH)₂ solution?