Conductance of Strong and Weak Electrolyte Solutions

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Abstract

In this experiment, the differing behaviors of strong and weak electrolytes with regard to conductivity was studied. Hydrochloric acid (HCl) and acetic acid (HAc) were chosen for our strong and weak electrolytes respectively. Using these two solutions proved useful to show that strong electrolytes are stronger conductors than weak electrolytes. HCl was used to show that strong electrolytes follow the Onsager model when approaching infinite dilution due to full dissociation and HAc was used to show that the trend is drastically different for weak electrolytes due to only partially dissociating. While our experimental trend for HCl was opposite of expected, we have used theoretical data to help explain and gain insight on the correct behavior of strong electrolytes. We found that the equivalent conductance for HCl at infinite dilution is 391.93 cm$^2$ S M$^{-1}$ which falls within 8% of literature value. It was also found that HAc has an equilibrium constant of $1.55 \times 10^{-5}$ which is within 16% of the tabulated value from NIST.$^5$

Introduction

Solutions that contain free charges allow electrical currents to pass through them when electrically stimulated. Electrolytic solutions behave very similarly to an electrical circuit. When an electrolytic solution is supplied with a current, the conductive properties of the charged ions will carry the charge between terminals acting as a circuit following Ohm's Law$^1$:

$$V = I \times R$$

Where $V$ is the voltage, $I$ is the provided current, and $R$ is the resistance through the circuit.

The higher the concentration of electrolytes, the more charged ions, the easier it is for the current to pass through. Therefore pure solvents act as strong resistors and as concentration of electrolyte increases, resistance decreases. Since specific conductance is just the inverse of resistance, this points to specific conductance increasing as concentration increases. But specific conductance doesn't work as a legitimate comparison because it's not holding the amount of ions constant. We need to look at a more fundamental unit, equivalent conductance. Equivalent conductance can be thought of as the conductance contributed by one equivalent of charged ions contained in a liter of solution.$^1$

Much like acids, the strength of electrolytes is determined by the degree of dissociation from the neutral species.$^1$ In this experiment we study the difference between a strong electrolyte and a weak electrolyte pertaining to conductance. Our strong electrolyte, hydrochloric acid (HCl), will fully dissociate while the weak electrolyte, acetic acid (HAc), will only partially dissociate.

To properly explore the affects of concentration, the specific conductance was measured using a Vernier conductivity probe at various concentrations of the two compounds approaching infinite dilution. We approach infinite dilution so we can then gain insight on the effect of strong versus weak electrolyte solutions pertaining to conductance. The conductivity probe gave greater accuracy than a manual conductivity cell, such as a Wheatstone bridge, that takes precision in set up or data could severely be altered.$^2$
Since the Vernier probe measures specific conductance, we have to calculate the equivalent conductance using the following relationship found in Sime.

\[ \Lambda = \frac{1000 * \left( \frac{k}{R} \right)}{c} \]  

(2)

Where \( \Lambda \) is the equivalent conductance, \( k \) is the calculated cell constant, \( R \) is the resistance of the solution, and \( c \) is the concentration. Resistance of the solution is simply the inverse of our measured specific conductance.

The cell constant \( (k) \) is said to be 1, but was verified using a known concentration solution of potassium chloride (KCl) to check the measured conductance value to that tabulated in Sime.\(^1\)

HCl is a strong electrolyte and fully dissociates. This means it should theoretically follow the Onsager model when the equivalent conductance calculated from equation 1 is plotted against the square root of concentration while approaching infinite dilution. The Onsager model is an existing numerical fit following this equation.

\[ \Lambda = \Lambda_o - (60.32 + 0.2289\Lambda_o) * \sqrt{C} \]  

(3)

Where \( \Lambda \) is the equivalent conductance at a specific concentration, \( C \) is the concentration, and \( \Lambda_o \) is the equivalent conductance at infinite dilution. Notice that this equation fits that of a line which is what the Onsager model proposes strong electrolytes follow when approaching infinite dilution.

For HAc, the equivalent conductance was once again calculated using equation 1, but the trend does not follow the Onsager model. This is because HAc is a weak electrolyte and does not fully dissociate in solution. It follows a logarithmic function and thus we cannot experimentally determine acetic acids equivalent conductance at infinite dilution. To find \( \Lambda_o \) for acetic acid, we can follow the procedure in Sime as to use the sum of the individual ions that make up HAc.\(^1\)

\[ \Lambda_o(HAc) = \Lambda_o(HCl) + \Lambda_o(NaAc) - \Lambda_o(NaCl) \]  

(4)

HAc doesn’t fully dissociate and is stuck in dissociation equilibrium between its aqueous charged ions and the neutral molecule. Weak electrolytes are complex as they are always dissociating into charged ions and pairing up to form the neutral species but as concentration decreases, they approach full dissociation. This means we can look at the degree of dissociation of HAc and this will give us further insight into the equilibrium constant. The degree of dissociation can be calculated at a specific concentration using its equivalent conductance.

\[ \alpha = \frac{\Lambda}{\Lambda_o} \]  

(5)

Once the degree of dissociation is obtained, the \( K_c \) values could then be graphically fit to find the equilibrium constant \( K \) for HAc by studying the infinite dilution point. This is the intercept of the following equation established by Debye and Huckel.\(^1\)

\[ \log K_c = \log K + 1.018\sqrt{ca} \]  

(6)

Finding the equilibrium constant of HAc can provide us with insight into the amount of ions are going to be present at infinite dilution. This will then show us how HAc will act in terms of conductance at infinite dilution, giving us a way to compare it to that of a strong electrolyte such as HCl.
Experimental Procedures

First we needed starting solutions before we could start measuring specific conductance. But since we were going to be diluting the solutions to approach infinite dilution we needed accurate concentrations. This means solutions need to be standardized with simple acid-base titrations. The most simple way is to choose one to one reactions and use NaOH for the standardization of our acid electrolytes we used. NaOH was first made and titrated with potassium hydrogen phthalate (KHP) to find its exact molarity so we can then use it to titrate our acids to obtain a very accurate starting concentration.

Once starting solutions with known concentrations of around 0.1 M were obtained, they were diluted in a fashion to approach infinite dilution. Our strong electrolyte, HCl, was diluted 5 times following N/10, N/20, N/40, N/80, N/160, N/320, and N/640. Dilutions were performed using volumetric glassware.

The Vernier probe first had to be calibrated by calculating its cell constant (k) as previously mentioned. This was done by testing the specific conductance using a known KCl solution at 0.2 M. This was done in triplicate to obtain a cell constant of 1.032. While this was close to their claimed value of 1, the small 3% error would have affected our data significantly through calculations. Once k was determined, the solutions could then be tested using the Vernier conductivity probe in conjunction with the Vernier LabQuest. Each solution was then measured with the probe by swirling it in the solutions until the reading stabilized. This was done in triplicates for each solution so an average and standard deviation could be obtained. The Vernier probe was cleaned thoroughly with milliQ water between each reading.

Results

Table 1 shows the starting concentration of our strong and weak electrolyte and also the NaOH solution we used to standardize them. Each solution was standardized using the acid-base titration method mentioned in the experimental. We needed accurate standardizations because equivalent conductance is affected very heavily by concentration.

<table>
<thead>
<tr>
<th>Solution</th>
<th>Concentration (M)</th>
<th>St. Dev</th>
</tr>
</thead>
<tbody>
<tr>
<td>NaOH</td>
<td>0.0919</td>
<td>± 0.0016</td>
</tr>
<tr>
<td>HCl</td>
<td>0.0986</td>
<td>± 0.0011</td>
</tr>
<tr>
<td>HAc</td>
<td>0.0987</td>
<td>± 0.0016</td>
</tr>
</tbody>
</table>

Tables 2 and 3 are the measured values of specific conductance we found by using the Vernier probe. Each solution was measured three times to ensure repeatable data for correct calculations. The data was very repeatable with a small percent uncertainty and both followed the trend of increasing specific conductance with increasing concentration.
Table 2. HCl Measured Conductances

<table>
<thead>
<tr>
<th>HCl Concentration (M)</th>
<th>Specific Conductance (μS/cm)</th>
<th>Uncertainty (%)</th>
</tr>
</thead>
<tbody>
<tr>
<td>N/10</td>
<td>0.00986</td>
<td>4365.3</td>
</tr>
<tr>
<td>N/20</td>
<td>0.00493</td>
<td>2169.3</td>
</tr>
<tr>
<td>N/40</td>
<td>0.00246</td>
<td>987</td>
</tr>
<tr>
<td>N/80</td>
<td>0.00123</td>
<td>490.3</td>
</tr>
<tr>
<td>N/160</td>
<td>0.00062</td>
<td>240</td>
</tr>
</tbody>
</table>

The following figures are used to gain insight on our strong and weak electrolytes when approaching infinite dilution. With these trends we can study the effect of each chemical pertaining to conductance of solutions.

Figure 1. Plot of equivalent conductance against the square root of concentration for HCl. This trend is opposite of what is expected due to errors discussed below. The y-intercept is the equivalent conductance of HCl at infinite dilution.
Figure 2. Plot to show experimentally found trend versus the expected trend of the Onsager model. This further shows that our data is flawed.

Figure 3. Plot of expected trend shown against the theoretical Onsager model. This is used below to explain the chemistry behind maximum equivalent conductance at the infinite dilution point for our electrolytes. The y-intercept is said maximum equivalent conductance.
Table 3. HAc Measured Conductances

<table>
<thead>
<tr>
<th>HAc</th>
<th>Concentration (M)</th>
<th>Specific Conductance (μS/cm)</th>
<th>Uncertainty (%)</th>
</tr>
</thead>
<tbody>
<tr>
<td>N/10</td>
<td>0.00987</td>
<td>148.8</td>
<td>0.1</td>
</tr>
<tr>
<td>N/20</td>
<td>0.00493</td>
<td>104.2</td>
<td>0.1</td>
</tr>
<tr>
<td>N/40</td>
<td>0.00247</td>
<td>73</td>
<td>0.1</td>
</tr>
<tr>
<td>N/80</td>
<td>0.00123</td>
<td>50.5</td>
<td>0.1</td>
</tr>
<tr>
<td>N/160</td>
<td>0.00062</td>
<td>34.5</td>
<td>0.1</td>
</tr>
<tr>
<td>N/320</td>
<td>0.00031</td>
<td>23.2</td>
<td>0.1</td>
</tr>
<tr>
<td>N/640</td>
<td>0.00015</td>
<td>15.2</td>
<td>0.1</td>
</tr>
</tbody>
</table>

Figure 4. Plot of equivalent conductance versus the square root of concentration for our weak electrolyte HAc. This shows that weak electrolytes behave differently than strong electrolytes and do not follow the Onsager model. Since there is no y-intercept, equivalent conductance at infinite dilution for a weak electrolyte cannot be experimentally determined.

\[ y = -42.05 \ln(x) - 89.869 \]

\[ R^2 = 0.9535 \]
Figure 5. Plot of the log(Kc) versus the square root of the concentration multiplied by the degree of dissociation of HAc. This plot is used to determine the equilibrium constant of HAc.

Discussion

The data and trends found in this lab provide insight on how conductance varies between strong and weak electrolytes and differing concentrations among them. It shows that strong electrolytes have a higher conductance than weak electrolytes because of how they dissociate into solution.

As expected because of dissociation the specific conductances in tables 2 and 3 increase with concentration of the electrolyte. This is because as concentration increases the amount of charged ions increases in solution. But as said previously specific conductance doesn't work well as a comparison and we need to look at equivalent conductance.

HCl is a strong electrolyte and should follow the Onsager model when approaching infinite dilution. Figure 1 shows that our data proves that it does indeed follow a linear trend, but it has an opposite slope than predicted by Onsager. This shows that our data is flawed and something must have gone wrong during the experiment due to human error. We think this has to do with our concentrations being misread and the dilutions being incorrectly done due to a misuse of volumetric glassware or other overlooked details. Even a slight change in the concentration that goes unaccounted for could cause a big error in the equivalent conductance value due to the mathematical relationship shown in equation 2. This would drastically affect the trend leading to something similar to what is shown in figure 1 and 2. Even though our data was flawed and our trend was opposite, figure 1 shows an intercept value of 391.93 cm² S M⁻¹. This value for the infinite dilution equivalent conductance for HCl falls within 8% error of the tabulated literature value of 426.16 in Sime.¹ Due to the major discrepancy in trend, we think this is just happenstance.
Figure 3 shows a theoretical trend that is similar to what we should have found within our experiment. This data was mocked up to help show and explain the concept of equivalent conductance and to show how it should actually behave while approaching infinite dilution. As shown, the slope should have been negative and the intercept closer to the tabulated value for infinite dilution equivalent conductance for HCl. The correct trend here shows that as concentration decreases, equivalent concentration increases, which is opposite of what was found for specific conductance. This means that the maximum equivalent conductance of a strong electrolyte is at the infinite dilution point. The phenomenon of ion shielding explains this. While volume is being held constant in equivalent conductance the amount of ions decreases as concentration decreases. This makes it harder for the ions to find each other and combine to make the neutral species and allows them to be more free to carry charge.

HAc on the other hand is a weak electrolyte and as previously stated only partially dissociates. This makes it behave much differently than our strong electrolyte HCl. While the trend of increasing equivalent conductance while concentration decreases still exists, it is in a different fashion. The trend of HAc follows a logarithmic function as shown in figure 4 and never intercepts with the y-axis. This is why we can't experimentally determine the infinite dilution equivalent conductance of HAc and must follow the procedure shown in equation 4. It also shows that the values were much lower than that of HCl, proving that weak electrolytes are weaker conductors than strong electrolytes.

Because HAc is in dissociation equilibrium, it never fully dissociates. HAc approaches full dissociation as it approaches infinite dilution, giving the trend shown in figure 4. This is why we must take another approach to study weak electrolytes by looking at their equilibrium constant. Using the data in table 3 and the equations 5 and 6, figure 5 was created to study the trend of equilibrium constants compared to concentrations of HAc. We want to study our electrolytes at infinite dilution for comparison, so we look at the y-intercept. While our linear trend fits poorly, it gives us an equilibrium constant of $1.55 \times 10^{-5}$ which falls within a 16% error of that tabulated by NIST of $1.76 \times 10^{-5}$.

**Conclusions**

This experiment portrays and correctly supports the idea that strong electrolytes are stronger conductors than weak electrolytes due to full dissociation.

It shows that strong electrolytes should follow the Onsager model and linearly increase in equivalent conductance when approaching infinite dilution. If extrapolated to the y-intercept you can find the exact value of the equivalent conductance at infinite dilution for a strong electrolyte.

Following the theory given previously, this lab provided insight on the dissociation equilibrium in weak electrolytes. It was shown that as weak electrolytes approach infinite dilution, they approach full dissociation and consequently their equivalent conductance at the infinite dilute point could not be experimentally determined. While looking at the degree of dissociation of weak electrolytes it is possible to determine their equilibrium constant at infinite dilution which will provide insight on their ability to conduct due to concentration of charged ions.

We take our found values for infinite dilution behaviors lightly because of the mishaps in our data. Most likely the problem lies within our concentrations. This is probably from our dilutions but it could
have stemmed from our initial standardization of solutions as well. Another possibility, although unlikely, is that our Vernier probe was having technical issues. Given more time, the lab should be started over and completed with better precision with regard to concentration. While doing so, more concentrations should be studied to fill in the trends and approach even closer to infinite dilution.
Reference

(1) Sime, R. J.; *Physical Chemistry: Methods, Techniques, and Experiments*; Saunders College Publishing, 1990


