JOURNAL OF
EMERGING INVESTIGATORS

Article

Determining degree of dissociation through conductivity

Abigail H. Dunn[1], Heather W. Zimmer[2]

[1] Milton Academy, Milton, Massachusetts

[2] Department of Science, Milton Academy, Milton, Massachusetts

SUMMARY

When ionic compounds dissolve in a polar solvent, they dissociate into separate ions and spread uniformly throughout the solution. The greater the degree of dissociation of an ionic compound, the greater the conductivity of the solution will be. The goal of this experiment was to assess how the degree of dissociation of an ionic compound is impacted by the molarity of the compound in solution, the magnitude of the charge of the ion, and the number of ions it produces upon dissociation. Conductivity was measured as a quantifiable representation of the degree of dissociation. We hypothesized that as the concentration of ions in the solution decreased, the conductivity of the solution would also decrease due to fewer interactions between the ions, resulting in a weaker flow of electrical current. We tested our hypothesis by measuring the conductivity of 0.25 M, 0.20 M, 0.15 M, and 0.10 M solutions of aluminum chloride, potassium chloride, and calcium chloride. The data showed that as the molarities of the solutions decreased, the conductivities of the solutions decreased. We also found that the conductivity of each solution did not decrease by the theoretical factor predicted by the number of ions each compound produced upon full dissociation, leading us to conclude that the impact of the charge magnitude of the ions on the degree of dissociation of the ionic compounds must also be considered.

INTRODUCTION

Understanding the properties of ionic compounds that increase the degree of dissociation makes it possible to identify strong electrolytes and use them for many purposes. Strong electrolytes serve crucial aspects in biological needs. Ionic compounds that result in strong electrolytes, like calcium, potassium, and chloride, are an important part of a healthy diet. When consumed, these ionic compounds dissociate into positive and negative charges, preserving water balance, aiding muscles in contracting, and balancing voltage across membranes (1).

Ions are molecules or atoms that have gained or lost one or more electrons to become negatively or positively charged. Ionic compounds are ions held together by the electrostatic forces between positively charged ions, called cations, and negatively charged ions, called anions (2). Coulomb's Law (**Equation 1**) quantifies the electrostatic attraction between two charged particles:

$$F_e = k_e \frac{q_1 q_2}{r^2}$$ **[Equation 1]**

F_e (newtons) is the electrostatic force between the charged particles, k_e is Coulomb's constant (newtons * coulombs squared per meter squared), q_1 and q_2 (coulombs) are the charge magnitudes of the first and second particles, and r (meters) is the distance between the charged particles (3). Based on Coulomb's Law, a greater magnitude of ionic charges results in greater electrostatic forces between the ions. Ionic compounds in the aqueous state act as electrolytes, dissociating into hydrated ions (4). **Equation 2** shows how the number of ions a compound produces upon dissociation can be determined by the chemical formula of the ionic compound: electrolyte A_xB_y dissociates into 'x' number of A cations and 'y' number of B anions in solution (5).

$$A_xB_y(aq) \rightleftharpoons xA^+(aq) + yB^-(aq)$$ **[Equation 2]**

A water molecule is polar, with positively charged regions at the hydrogen atoms and a negatively charged region at the oxygen atom, so ion-dipole forces between the ions and water molecules allow ionic compounds to dissociate in water (6). Cations attract the negative dipoles of water molecules in the solvation process, while anions attract the positive dipoles of water molecules. The water molecules surround each ion in three dimensions, breaking its attraction to the ionic structure and allowing the ions to dissociate in solution (7). In the aqueous state, hydrated ions are not held in a fixed position but rather are free to move throughout the solvent. Therefore, a flow of electrical current is possible between the dissociated ions, making an aqueous ionic solution electrically conductive (8). The conductivity of an ionic solution is a measure of the number of dissociated ions present in solution per mole of solvent, which can be quantified using a ratio known as the Van't Hoff factor, defined in **Equation 3** (9):

$$i = 1 + \alpha(n - 1)$$ **[Equation 3]**

The Van't Hoff factor (i) is determined by the values of two variables: the degree of dissociation of the compound, α, a ratio between moles of substance dissociated and moles of total substance, and n, the number of ions produced when that compound dissociates. The degree of dissociation depends on the nature of the solvent and the solubility of the solute at a given temperature (10). Ionic compounds with larger ionic charges are less soluble, or have lower degrees of dissociation, because the electrostatic forces between the ions are stronger, making it harder for the ion-dipole forces between the water molecules and the ions to overcome the electrostatic forces between the ions to dissociate them (11).

The larger the degree of dissociation of the solute and the greater the number of ions it produces upon dissociation, the larger the dissociation factor of the solute will be (9). All other factors being equal, solutes with greater dissociation factors have greater conductivities in aqueous solution.

While the number of ions an ionic compound produces upon complete dissociation is easily calculated, it is difficult to calculate its degree of dissociation. To better understand the effect of the degree of dissociation on conductivity, we compared the theoretical conductivities of aluminum chloride, potassium chloride, and calcium chloride if fully dissociated in water to the measured conductivity of each compound. This comparison revealed a higher degree of dissociation in compounds with lower charges.

RESULTS

Potassium chloride, calcium chloride, and aluminum chloride were each dissolved in water to create three 0.25 M ionic solutions. We measured the conductivity of each solution using a conductivity probe. Conductivities were measured in a lab with an average temperature of 20.78°C. Measurements show that the temperature of the lab varied by an average of ±0.5°C across testing. As the molarities in solution of aluminum chloride, calcium chloride, and potassium chloride increased from 0.10 M to 0.25 M, their conductivities increased as well ($p<0.0001$, ANOVA, **Figure 1).** Assuming full dissociation, aluminum chloride produces four ions upon dissociation, calcium chloride produces three, and potassium chloride produces two. At every molarity value, aluminum chloride, the compound producing the greatest number of ions upon dissociation, displayed the greatest conductivity values, followed by calcium chloride, then potassium chloride ($p<0.01$, ANOVA, **Figure 1).** A comparison of the conductivities of ionic solutions containing compounds with different numbers of ions per mole at a range of molarities revealed that solutions with greater numbers of ions per mole and greater molarities would display greater conductivities.

Comparing the theoretical conductivity values to the measured conductivity values revealed that the conductivity of the solutions did not increase with molarity by the factor expected had full dissociation taken place **(Figure 2).** Notably, the greatest disparities between measured and theoretical conductivity, assuming full dissociation, were observed in solutions containing ions with higher charges and a higher molarity. It is also interesting to note that while the theoretical conductivities show an increase in separation at higher molarities, the actual data showed a convergence of conductivity values as molarities increased.

DISCUSSION

In our experiment, we measured the conductivities of three different ionic solutions at various molarities to determine how the degree of dissociation of a compound is impacted by the molarity in solution, the number of ions produced upon dissociation, and the charge magnitude of the ions. We used conductivity as a quantifiable measure of the degree of dissociation.

We hypothesized that the conductivity of a compound should increase at a proportional rate to the molarity because as the concentration of the compound in solution increases, the number of dissociated ions, and thus the conductivity, should also increase. Our results indicate that conductivity increased with molarity but at a disproportionate rate. Instead of increasing linearly with molarity, conductivity values increased at a decreasing rate. This could be attributed to the fact that as the molarity of a solution increases and ion-ion interactions become more frequent than ion-solvent interactions, the degree of dissociation of the compound, and thus the conductivity of the compound, will increase at a decreasing rate (12). This would also explain why measured conductivity values converge as molarity increases, suggesting that their conductivities begin to plateau once solutions reach a certain molarity value.

We also hypothesized that the conductivities of aluminum chloride, calcium chloride, and potassium chloride at a

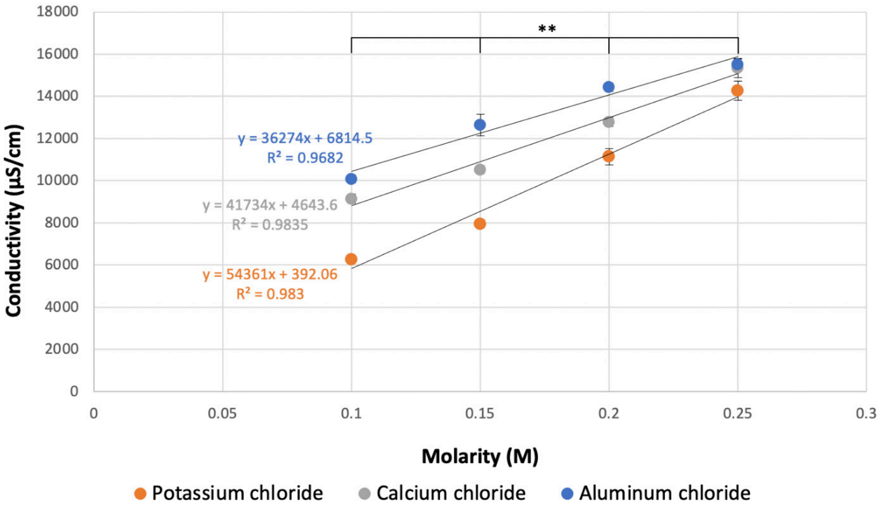

Figure 1: Conductivity of aqueous aluminum chloride (blue), calcium chloride (gray), and potassium chloride (orange) solutions of different molarities. Conductivity was measured using a Vernier conductivity meter placed in each ionic solution. Error bars represent standard deviation. Statistical analysis by multiple-comparison Tukey HSD Test comparing conductivity values across molarities for each chemical, **$p<0.01$. N = 5 for potassium chloride and calcium chloride. N = 6 for aluminum chloride.

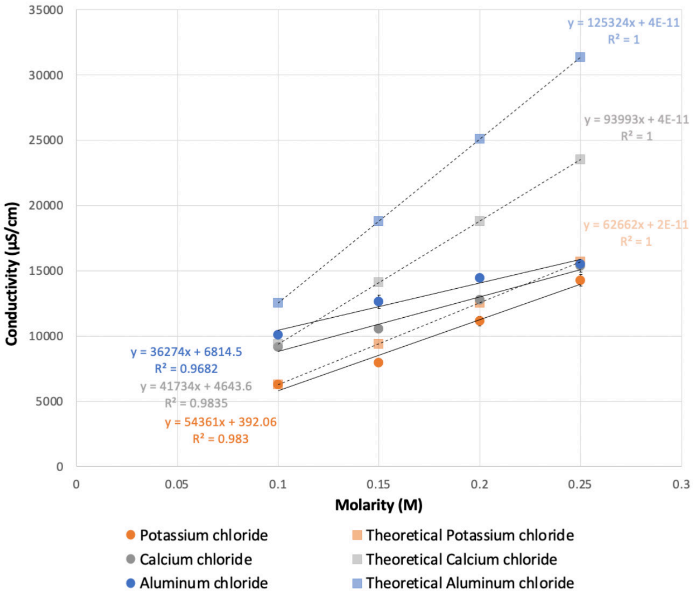

Figure 2: Theoretical and measured conductivity (µS/cm) of KCl (orange), CaCl2 (gray), and AlCl3 (blue) solutions of different molarities. Theoretical conductivity is represented by translucent squares, while experimentally measured conductivity is represented by opaque circles. Assuming the conductivity of a compound depends solely on the number of ions in solution and the molarity of the solution, theoretical conductivity was based on the experimental conductivity of 0.10 M potassium chloride.

constant molarity should be proportional to the number of ions each compound produces upon dissociation. However, our results revealed that aluminum chloride, calcium chloride, and potassium chloride displayed lower conductivity values than predicted at each molarity. Aluminum chloride, the compound with the greatest electrostatic forces between its ions, displayed the greatest disparities between its measured and theoretical conductivity values. We observed similar but not as extensive disparities for calcium chloride and minimal disparities for potassium chloride in the measured conductivity values. We hypothesize that the greater the electrostatic forces between ions, the fewer the number of ions that can fully dissociate into solution, and thus the lower the degree of dissociation of the compound.

The main source of uncertainty in our experiment is the temperature of the solution. The temperature in the classroom varied by an average of ±0.5°C across testing. To eliminate this uncertainty, solutions could have been prepared from the same initial container of water allowed to reach room temperature. Alternatively, all testing could have been completed in one testing period to minimize the time available for the temperature of the solvent to change.

While the three ionic compounds used in the experiment produced different numbers of ions upon dissociation and had different charge magnitudes, the solvent remained constant. While varying solute characteristics provided an understanding of the properties of a solute that result in a higher or lower degree of dissociation, using a uniform solvent prevented us from understanding the characteristics of a solvent that result in a higher or lower degree of dissociation. Our next potential step includes investigating how the nature of a solvent impacts the degree of dissociation of the solute. To do so, the same solute could be dissolved in different solvents or different temperature solvents to understand how the chemical structure and temperature of the solvent impact the degree of dissociation and conductivity.

METHODS & MATERIALS

Procedure

Laboratory-grade aluminum chloride (Flinn Scientific Inc, A0179), anhydrous calcium chloride (Fisher Science Education, S25223), and reagent-grade potassium chloride (Fisher Science Education, S25484A) were dissolved in distilled water to create three 0.25 M solutions. A Vernier conductivity meter was used to measure the conductivity of the solutions. Each solution was then diluted to 0.20 M, 0.15 M, and 0.10 M, and the conductivities were recorded. Five replicate tests were conducted for potassium chloride and calcium chloride, while six replicate tests were conducted for aluminum chloride to improve line fit quality.

Theoretical Calculations

After experimentation, we used the measured potassium chloride value at 0.10 M as a base to calculate the theoretical potassium chloride values, assuming the conductivity of the compound depended solely on the number of potassium chloride ions in the solution and the molarity of the solution. For example, the theoretical conductivity at 0.15 M should be

Molarity (M)	Potassium chloride (µS/cm): 2 ions	Calcium chloride (µS/cm): 3 ions	Aluminum chloride (µS/cm): 4 ions
0.10 M	6266.2 = x	9399.3 = 1.5x	12532.4 = 2x
0.15 M	9399.3 = 1.5x	14098.9 = 1.5(1.5x)	18798.6 = 2(1.5x)
0.20 M	12532.4 = 2x	18798.6 = 1.5(2x)	25064.8 = 2(2x)
0.25 M	15665.5 = 2.5x	23498.3 = 1.5(2.5x)	31331.0 = 2(2.5x)

Table 1: Calculations of theoretical conductivity values (µS/cm). All values are based on the experimental conductivity of 0.10 M potassium chloride.

1.5-fold the conductivity at 0.10 M **(Table 1)**. We predicted the theoretical calcium chloride and aluminum chloride values based on the theoreticized potassium chloride values, again assuming conductivity depends solely on the number of ions the compound produces upon dissociation and the molarity of the solution **(Table 1).**

Statistical Analysis

To determine that a statistically significant difference existed between the average conductivity values across the range of molarities for each compound, we performed separate multiple-comparison Tukey HSD Tests comparing conductivity values across molarities for each chemical, **$p<0.01$.

ACKNOWLEDGEMENTS

We would like to thank Omar Hamoda and Imran Serifovic at Milton Academy for contributing to data collection.

Received: December 11, 2022
Accepted: August 26, 2023
Published: January 26, 2024

REFERENCES

1. «Electrolytes and Electrolyte Balance.» *World of Anatomy and Physiology*, Gale, 2007. Gale in Context: Science. link.gale.com/apps/doc/CV2430500123/SCIC?u=mlin_s_miltona&sid=bookmark-SCIC&xid=5c41b19f
2. Flowers, Paul, et al. "Ionic and Molecular Compounds." *Chemistry 2e*, e-book ed., OpenStax, 2019. https://openstax.org/books/chemistry-2e/pages/2-6-ionic-and-molecular-compounds
3. «2.1: Coulomb›s Law & The Electrostatic Potential.» *Chemistry LibreTexts*, Oregon Institute of Technology, 5 June 2019. chem.libretexts.org/Courses/Oregon_Institute_of_Technology/OIT%3A_CHE_202_-_General_Chemistry_II/Unit_2%3A_Electrons_in_Atoms/2.1%3A_Coulomb›s_Law_and_the_Electrostatic_Potential
4. «18.3 Electrolytes, Ionisation and Conductivity.» *Physical Sciences Grade 10*, e-book ed., Siyavula Education, 2012. siyavula.com/read/za/physical-sciences/grade-10/reactions-in-aqueous-solution/18-reactions-in-aqueous-solution-03
5. "15.8 Dissociation." *Chemistry LibreTexts*, Oregon Institute of Technology, 5 June 2019. chem.libretexts.org/Bookshelves/Introductory_Chemistry/Introductory_Chemistry_(CK-12)/15%3A_Water/15.08%3A_Dissociation#:~:text=Dissociation%20is%20the%20separation%20of,chemical%20formulas%20of%20ionic%20 compounds
6. «Salt, Chemical Compound.» *The Columbia Electronic Encyclopedia™*, Columbia UP, 2023. Gale in Context: Middle School. link.gale.com/apps/doc/A69231240/MSIC?u=mlin_s_miltona&sid=bookmark-MSIC&xid=bc912356
7. Flowers, Paul, et al. «Ionic Electrolytes.» *Chemistry 2e*, e-book ed., OpenStax, 2019. openstax.org/books/chemistry-2e/pages/11-2-electrolytes#term-00006
8. «Indicators: Conductivity.» *National Aquatic Resource Surveys, United States Environmental Protection Agency.* 11 July 2022. epa.gov/national-aquatic-resource-surveys/indicators-conductivity
9. «11.7: Colligative Properties of Ionic Solutes.» *Chemistry LibreTexts*, 5 June 2019. chem.libretexts.org/Bookshelves/Introductory_Chemistry/Beginning_Chemistry_(Ball)/11%3A_Solutions/11.07%3A_Colligative_Properties_of_Ionic_Solutes
10. «New Chemical and Engineering Data Findings from Tianjin University Described (Correlation and Thermodynamic Analysis of Solubility of Mesotrione in Pure Solvents).» *News of Science*, 19 Apr. 2020, p. 419. Gale in Context: Middle School. link.gale.com/apps/doc/A620259960/MSIC?u=mlin_s_miltona&sid=bookmark-MSIC&xid=cb9789aa
11. Greenough, John D., et al. «Trace Element Geochemistry of Nova Scotia (Canada) Maple Syrup.» *Canadian Journal of Earth Sciences*, vol. 47, no. 8, Aug. 2010, p. 1093+. Gale in Context: Science. https://doi.org/ 2048/10.1139/E10-055.
12. "8.10.9C: Weak and Strong Electrolytes." *Chemistry LibreTexts*, Oregon Institute of Technology, 5 June 2019. https://chem.libretexts.org/Bookshelves/General_Chemistry/Chem1_(Lower)/08%3A_Solutions/8.10%3A_Ions_and_Electrolytes/8.10.9C%3A_8.10.9C%3A__Weak_and_Strong_Electrolytes