An Investigation of Electrolyte Solutions Using a Simple Conductivity Apparatus

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Abstract: A student laboratory exercise in qualitative analysis has been developed to address student misconceptions associated with electrolyte and nonelectrolyte solutions. This exercise uses a previously reported, inexpensive, home-built conductivity meter to identify acid solutions. An additional short preparatory exercise provides students the opportunity to classify known solutions as strong, weak, or nonelectrolytes and gives students experience using a conductivity meter. These activities have been used successfully with both high school and introductory college chemistry students.

Introduction

Students usually somewhat grudgingly accept the concept that ionic compounds dissociate into ions in an aqueous solution. The grudging acceptance may be the result of presenting this topic somewhat closely following presentation of the concept that the formation of ionic substances is due to strong electrostatic attractions. Our experience has been that once the idea of ionic dissociation is introduced, students then often view all compounds, ionic or covalent, as potential sources of ions in solution. Student-generated chemical equations are often incorrect and seem to contain arbitrarily dissociated compounds. This problem is exacerbated by the paucity of common instructional experiments that demonstrate the differences between strong, weak, and nonelectrolytes. Experimental evidence for the ionization of some compounds and not others is usually limited to an instructor demonstration or video presentation of aqueous solutions that permit a light bulb to either light or not light.

Two safe and inexpensive conductivity measuring devices were introduced by Katz and Willis [1]. One of these indicates the relative conductivity of a solution using a simple circuit containing red and green LEDs, a 9-V battery, two resistors, and copper wires (as electrodes). A student using this device can distinguish five levels of conductivity from the intensities of the two LEDs. Katz and Willis also presented a number of clever experiments that use their home-built conductivity device.

In this report we extend the use of Katz and Willis's device and develop an activity that provides students the opportunity to investigate electrolyte and nonelectrolyte solutions within the context of a qualitative analysis experiment. The experiment described has been successfully used with both high school and introductory college chemistry students. Students need very little exposure to the theory of solution conductivity other than to the fact that a solution with more ionic species will generally conduct electricity better.

Methods

The conductivity apparatus used was developed by Katz and Willis [1]. In order to make data obtained from the device more reproducible, we have made some minor modifications. First, the copper wire electrodes have been inserted in predrilled holes in rubber stoppers in order to maintain a constant separation. This eliminates discrepancies in the observed conductivity due to possible variations in the distance between the electrodes. Second, the suggested 24-well plate [1] to hold the solutions has been replaced by a 96-well plate. One well on a 96-well plate can be filled completely with 6–7 drops (~0.3 mL) of the solution. This change assures both that a significant portion of the electrodes are immersed in the solution and that they are immersed to a consistent depth in each solution. Third, problems associated with differential ionic mobility can be circumvented by the appropriate choice of solutes (vide infra), and the measurement thus made more quantitative.

Experiment: Using Conductivity for Qualitative Analysis

This experiment provides an alternative to qualitative analysis laboratories in which the formation of insoluble substances is used as the determining factor in the identification of unknowns. Distinguishing all five levels of light intensity of the conductivity device's display is essential for successful qualitative identification of the solutions.

Students are given samples of five colorless liquids (pure isopropyl, pure water, and 0.010 M solutions of HCl, H_2SO_4, and CH_3CO_2H) and asked to identify the solutions based upon their conductivities. The unknown substances in this experiment include acids for a number of reasons: (1) the emphasis placed on strong and weak acids in introductory chemistry courses, (2) the scarcity of qualitative analysis experiments with acids, and (3) the relationship of the solution conductivity to the $[H^+]$. The third reason is particularly important; a similar experiment could not easily be used to identify ionic salt solutions because the conductivity of an individual solution is a function not only of the concentration and charge of the ions, but also of the ionic mobility. The mobility of H^+(aq) is much greater than that of most anions (except HO^-) because of the unique mechanism for hydrogen ion conductivity [2]. The conductivity of the acid solutions is therefore dominated by $[H^+]$ and can be directly related to this quantity. Acetic acid was chosen because students are usually aware from class discussions that this is a weak monoprotic acid.

Because we want students to identify the substances based on solution conductivity and not on other properties, such as odor, we spike all the aqueous samples with 3% isopropanol without significantly affecting the solutions' conductivities. This gives all samples a similar odor so that the distinct odors of acetic acid and isopropanol do not lead to their identification.

By carefully observing the intensities of the LEDs and relating them to the overall conductivity of the solution, students should be able to distinguish the solutions of isopropyl alcohol, pure water, and acetic acid, from hydrochloric acid and sulfuric acid. Although these latter two strong acids do give slightly different conductivity results (as evidenced by careful observation of the relative light intensities), they are quite difficult to positively identify.

Although students are generally taught that both sulfuric acid and hydrochloric acid are strong acids, students are not normally introduced to the fact that there are differences in their dissociation constants (K_a). Although not known accurately, the K_a value for hydrochloric acid is estimated to be 10^3–10^5 times greater than that of sulfuric acid [3, 4]. In addition, although sulfuric acid is diprotic, the K_a for the second dissociation is only 1.2×10^{-2} [5]. Hence, while the diprotic nature of the sulfuric acid leads many students to predict an increased [H^+], and therefore increased conductivity over HCl, this is not necessarily observed because of differences in K_a and other more subtle effects such as ionic strength and ionic mobility.

We have found that students use one of the following two contradictory rationalizations, which show that the students are thinking of the problem as chemists might.

> The solution with the higher conductivity should contain sulfuric acid. Because both acids are strong acids, the first proton is completely dissociated from both, giving a 0.010 M contribution to the [H^+]. Since the sulfuric acid is diprotic, the small amount of H^+ from the second dissociation should give [H^+] = 0.010 M + x M, while for hydrochloric acid, [H^+] = 0.010 M.

> Consulting a table in a textbook [6] or other reference material [7], a 0.10 M solution of HCl has a slightly lower pH than that of 0.10 M sulfuric acid. If this still holds at 0.010 M, the hydrochloric acid solution should have greater conductivity, due to its greater [H^+].

This experiment requires less than one hour to perform and is therefore quite useful for short laboratory periods. It has been used in both secondary school and introductory college level chemistry laboratories with success.

Preparatory Activity: Which Substances Dissociate in Water?

Because many students have difficulty writing equations for ion formation in aqueous solutions, we have used this short exercise to address this problem as well as to prepare students for the qualitative analysis experiment. The use of this exercise represents a minor variation from the experiment proposed by Katz and Willis [1].

Students are given labeled 0.010 M solutions of a variety of strong, weak, and nonelectrolytes, as well as deionized water. The students' task is to predict whether the solutions will conduct electricity (i.e., whether the LEDs will light). Once this is done the students should discuss their predictions with other students. Only after this should they measure the relative conductivities of the solutions. At this point they should reconcile their predictions with their results, and write chemical equations that account for what they observed.

Although simple, this exercise gives students first-hand experience measuring the relative conductivities of solutions, provides an activity with visible and immediate feedback (the LED intensities), and gives each a personal database upon which to draw when performing the qualitative analysis experiment. This exercise also helps the development of careful experimental technique; one common student problem is cross-contamination of solutions. The two-step process of prediction and rationalization in chemical terms provides the students (with some facilitation by the instructor) the opportunity to recognize their misconceptions about electrolyte solutions. Depending on the amount of interstudent discussion and the abilities of the students to interpret results and write chemical equations, one to two hours are needed to complete this activity.

Discussion

We have found these experiments to be extremely valuable for helping students learn first-hand about ionic solutions and construct their own knowledge. They provide students with experience designing experiments, making careful observations, developing general laboratory skills, and making deductions based on experimental data. The amount of intra- and intergroup discussion that we have experienced during the laboratory period is quite satisfying, as students grapple to make connections between the experimental data and what they think they know about electrolytes. The extra emphasis placed on understanding the dissociation of strong and weak electrolytes and the use of the constructivist approach should help the long-term retention of these concepts.

The Katz and Willis conductivity device is inexpensive (about $3, without battery), so many can be made without a significant budgetary impact. The amount of each solution used is minimized through the use of the 96-well plates. The low costs of reagent purchase and disposal provide an added advantage to these experiments.

References

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