Aqueous Ammonia or Ammonium Hydroxide?
Identifying a Base as Strong or Weak

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Aqueous ammonia, which is traditionally used as a cleanser and grease remover, can be purchased at any grocery store. When the same solution is purchased from a chemical supply store, however, it is labeled as ammonium hydroxide (Figure 1). Ammonia is commonly recognized as a base, and the source of these two names can be seen from the base-ionization reaction of ammonia:

$$\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \leftrightarrow \text{NH}_4^+(\text{aq}) + \text{OH}^- (\text{aq})$$  (1)

The goal of this experiment is to help students determine which name is more accurate. To answer this question, students must determine which chemical species (neutral ammonia and water or ammonium ions and hydroxide ions) are present in greater quantities in these solutions.

Several articles in this Journal and other sources have tried to determine whether ammonia solutions contain mostly neutral ammonia molecules (alone or weakly associated with a neutral water molecule) or ammonium ions and hydroxide ions (1–7). The general consensus among chemists—derived from low-temperature X-ray powder diffraction and infrared spectral studies (1, 2), thermodynamic data (4), and Gaussian calculations (7)—is that these solutions contain very few ammonium and hydroxide ions compared to neutral ammonia molecules, and are best described as aqueous ammonia. Even with this general consensus about the chemical species present, there is no agreement among chemists or chemical supply companies on naming this system as aqueous ammonia (4, 6).

One major advantage of this experiment is that it asks students to make multiple comparisons of the chemical species present using different sets of data (including titration, pH, and electrical conductivity data). These methods include (i) comparing the pH of the solution to the pH expected for a strong base with the same concentration, (ii) calculating a value for the base-ionization constant, $K_b$, (iii) calculating the percent ionization using the pH and concentration, (iv) comparing the electrical conductivity data for the solution, sodium hydroxide, and methanol at the same concentration, and (v) calculating the percent ionization using the conductivity data. Students can use these different comparison techniques to corroborate each other, giving the students more confidence that their answer is correct. Instructors can choose to include some or all of these methods when they use this experiment. We recommend this experiment for students enrolled in a high school or college-level introductory chemistry course. After the authors tested this experiment to see if useful data could be collected, the first author field-tested this experiment in a 3-h laboratory for a second-semester introductory chemistry course at the college level.

Initial Instructor Testing of the Experiment

Data Collection

The first three methods require students to know the concentration of the NH$_3$/NH$_4$OH solution, and students can determine this value via titrations. Ammonia solutions are usually back-titrated (8–10), in which a stoichiometric excess of an acid such as hydrochloric acid is added to the ammonia solution, and the excess HCl is titrated with a strong base such as sodium hydroxide (11). The back-titration method is recommended for NH$_3$/NH$_4$OH solutions because the ammonia molecules are volatile and may escape during the titration if they are not converted to the nonvolatile ammonium ion. For this experiment, we purchased a bottle of 0.1 M ammonium hydroxide from Flinn Scientific, Inc. Using the back-titration method, each author titrated the NH$_3$/NH$_4$OH solution in triplicate, yielding a concentration of 0.1000 M for the NH$_3$/NH$_4$OH solution purchased from Flinn. Sample calculations using our data appear in the supporting information. We also titrated a sample of solid potassium hydrogen phthalate (KHP) dissolved in water with the NH$_3$/NH$_4$OH solution using a pH meter. The titration curve for this titration appears in Figure 2; because the pH at equivalence is close to 7.0, this titration can be performed without the pH meter using a chemical indicator whose color changes around a pH of 7.0 (e.g., bromothymol blue). Using the KHP pH titration data, we calculated a concentration of 0.0972 M (a 3% error) for the NH$_3$/NH$_4$OH solution. These calculations also appear in the supporting information. This lower value for the concentration could be due to the evaporation of some ammonia molecules during the titration. Still, this method gives a reasonable answer and could be used (adding minimal error) by students whose instructor does not want to introduce the concept of back-titration. It should be noted that students cannot use the pH at half-equivalence to determine the $K_a$ of KHP or the $K_b$ of NH$_3$ because this is the titration of a weak acid with a weak base. If instructors are interested in having their students interpret strong acid–weak base titration curves, the students could be asked to titrate a solution of HCl using the NH$_3$/NH$_4$OH solution (instead of titrating a solution of KHP with the NH$_3$/NH$_4$OH solution).

For the first three comparison methods, students also need to measure the initial pH of the NH$_3$/NH$_4$OH solution. For the 0.1000 M NH$_3$/NH$_4$OH solution from Flinn, we measured a pH of 11.25 using a glass electrode pH probe that was calibrated using pH 7 and pH 10 standardized buffer solutions.

We used the conductivity probes by Vernier Software & Technology attached to a graphing calculator to measure the electrical conductivity values in microsiemens per centimeter ($\mu$S cm$^{-1}$). These probes were set at the highest range (0–20,000 $\mu$S cm$^{-1}$).
Comparing the Solution pH to the pH of a Strong Base

Once the students have calculated the concentration of the NH₃/NH₄OH solution, they are asked to calculate the pH expected for a strong base of this concentration. For the 0.1000 M NH₃/NH₄OH solution from Flinn, the pH of a strong base would be 13.00. Because the measured pH was less (11.25), the NH₃/NH₄OH solution is not a strong base. This tells us that the solution is not 100% ionized into ammonium and hydroxide ions. However, it does not tell us how much ammonia and how much ammonium hydroxide exists; as such, this is a more qualitative comparison. Students could argue that the solution is “about 2 pH units” less basic than a strong base, which would suggest that there are very few (ammonium and) hydroxide ions present compared to the neutral ammonia molecules, implying that this solution should be called aqueous ammonia.

Calculating a $K_b$ Value

If the students know the concentration of the NH₃/NH₄OH solution and the initial pH of the NH₃/NH₄OH solution, they can calculate a value for the base-ionization constant, $K_b$. For the 0.1000 M NH₃/NH₄OH solution with an initial pH of 11.25, the $K_b$ value is $3.2 \times 10^{-5}$. Sample calculations using our data appear in the supporting information. This is close to the accepted value ($1.8 \times 10^{-5}$) appearing in general chemistry textbooks (13). Students should be told that a calculated equilibrium constant within 1 order of magnitude of the accepted answer is usually considered to be good. If the instructor is interested in having students calculate the percent error for their answer, students should be asked to calculate the percent error for the $pK_b$ value and not the $K_b$. For our answer above, the percent error for $K_b$ is 77%; the percent error for the $pK_b$ value is 5%.

An equilibrium constant with a negative exponent tells students that the reaction is reactant favored and should have more reactant species (ammonia and water molecules) than product species (ammonium and hydroxide ions). Therefore, this solution should be called aqueous ammonia.

Calculating the Percent Ionization Using pH Data

At this point, the students already have the information needed to calculate the percent ionization of the NH₃/NH₄OH solution—the total NH₃/NH₄OH concentration and the concentration of the hydroxide ion in the NH₃/NH₄OH solution. The percent ionization is calculated as the concentration of NH₃/NH₄OH that has ionized (equal to the concentration of $OH^-$ ions calculated from the pH of the NH₃/NH₄OH solution) divided by the total concentration of the NH₃/NH₄OH solution, multiplied by 100%. For the 0.1000 M NH₃/NH₄OH solution with a pH of 11.25, the percent ionization of the NH₃/NH₄OH solution is 1.8% (98.2% un-ionized). Sample calculations using our data appear in the supporting information. This result implies that this solution should be called aqueous ammonia and not ammonium hydroxide, which is consistent with the other two pH methods.

Comparing Electrical Conductivity Values

The electrical conductivity values for the air ($38 \mu S cm^{-1}$) and the pure water ($34 \mu S cm^{-1}$) represent the value expected for a solution containing no ions. The electrical conductivity of methanol ($34 \mu S cm^{-1}$), a neutral molecule not expected to produce ions, matches these values. The electrical conductivity of sodium hydroxide ($21,557 \mu S cm^{-1}$), on the other hand, represents the electrical conductivity of a 1:1 ionic compound at 100% ionization. The electrical conductivity for the NH₃/NH₄OH solution is only $496 \mu S cm^{-1}$, bigger than the value of $34 \mu S cm^{-1}$ for the nonconducting solutions, but much smaller than the $21,557 \mu S cm^{-1}$ value for

Figure 1. Bottles of aqueous ammonia (left) purchased from a grocery store and ammonium hydroxide (right) purchased from a chemical supply store.

Figure 2. Titration curve of NH₃/NH₄OH solution added to a solution of potassium hydrogen phthalate.
the completely conducting solution. These qualitative results suggest that the NH₃/NH₄OH solution appears much closer to the nonconducting solutions than the completely conducting solutions. Therefore, one could infer that the NH₃/NH₄OH solution consists of mostly neutral molecules (ammonia and water) and very few ions (ammonium and hydroxide ions). So, this compound should be named aqueous ammonia and not ammonium hydroxide.

Calculating the Percent Ionization Using Electrical Conductivity Data

The percent ionization of the NH₃/NH₄OH solution is calculated in a manner similar to the method using the pH data, except that it uses the electrical conductivity data instead of the calculated concentrations. The percent ionization in this method is calculated as the electrical conductivity of NH₃/NH₄OH solution (NH₃/NH₄OH value minus the zero point, which equals 496 − 34 = 462) divided by the electrical conductivity of the NH₃/NH₄OH solution assuming it conducted completely (which is assumed to equal the sodium hydroxide value minus the zero point, which equals 22,557 − 34 = 22,523), multiplied by 100%. For the 0.1000 M NH₃/NH₄OH solution, its percent ionization is 2.1% (97.9% un-ionized). Although there are more assumptions associated with this calculation (e.g., that the conductivity of the NH₄⁺ ion would be the same as the conductivity of the Na⁺ or the nonexistent CH₃⁺ ion, that the reading for the sodium hydroxide solution that is past the suggested range of the probe is still accurate, etc.) than for the percent ionization calculated from the pH data, these two values are very similar. These results imply that this solution should be called aqueous ammonia and not ammonium hydroxide, which is consistent with every other method reported in this article.

Results of the Field Test Involving Students

One author (M.J.S.) field-tested this experiment using 37 students (working in groups of 2–3) enrolled in two sections of second-term general chemistry. These students determined the NH₃/NH₄OH solution concentration via titration using solid KHP and bromothymol blue indicator. In the interest of time, these students only performed the first three parts of this lab (comparing the NH₃/NH₄OH solution pH to the predicted value, calculating a KӨ value for the NH₃/NH₄OH solution, and calculating a percent ionization value for the NH₃/NH₄OH solution). The pH meter used in this laboratory was calibrated using pH 7 and pH 10 standardized buffer solutions.

In general, their data for this experiment were reasonable. Their calculated concentrations for the 0.100 M solution ranged from 0.092 to 0.105 M; their pH values for the solution ranged from 10.5 to 11.2; their calculated KӨ values were within 1 order of magnitude of the accepted values (most were 10⁻³ or 10⁻⁴); and their percent ionization values ranged from 0.3% to 4%.

The interpretation of the data provided more difficulty for the students, especially the comparison of the actual pH to the predicted pH for the strong base. Many recognized this lower pH value implied that the base was weak, but had difficulty converting that answer into a relative quantity of the chemical species present. The students had less difficulty with the interpretation of the KӨ and the percent ionization or percent un-ionized data. All of the students came to the conclusion that the base solution contained mostly NH₃ molecules and very few NH₄⁺ and OH⁻ ions, and all agreed that the base solution should be named “ammonia” or “aqueous ammonia”.

We have included some student comments from their lab reports below. Although this field-test did not specifically address the concept or measurement of electrical conductivity, a few students did mention it in their lab reports.

- These three methods complement each other and all reinforce the fact that the solution should be called ammonia. The initial method showed us that we were dealing with a weak base. The KӨ value told us that the solution was mostly ammonia and water.
- In this experiment I was able to practice my titration skills and learn a few things about how pH and equilibrium constants relate to a practical question such as what to name the base solution we used in today’s lab. The questions really made me think about how these different methods help us see that the base solution was mostly ammonia and water.
- Finally, because the base did not greatly dissociate and because we found that the base solution was weak, I would not expect the base solution to conduct electricity very well.

Hazards

Ammonia, hydrochloric acid, and sodium hydroxide are corrosive, cause skin burns, and are harmful if swallowed or inhaled. Methanol is flammable and is harmful if swallowed or absorbed through the skin. Potassium hydrogen phthalate can cause skin and eye irritation.

Summary

The five comparison or calculation methods used in this laboratory experiment consistently point to the fact that aqueous solutions of NH₃/NH₄OH contain neutral ammonia (and water) molecules at a much higher concentration compared to ammonium and hydroxide ions. As a result, students should conclude that these samples are more accurately called aqueous ammonia rather than ammonium hydroxide. A major advantage of this lab is that it allows students to make multiple comparisons to answer the same question, and instructors may choose to use some or all of these methods in their laboratory experiments, as time allows.

Literature Cited


**Supporting Information Available**

A sheet of instructions for students, an answer key to the student questions (including sample calculations), and notes to the instructor This material is available via the Internet at http://pubs.acs.org.