**Acids, Bases, and Buffers**

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**CHEM 104L**

**03/24/2015**

**Abstract**

A pH meter and anion and cation tests were used to determine the identity of two unknown solutions, 14A and 14B. It was concluded that 14A was nitric acid () and 14B was ammonium hydroxide (). Through titration, the concentrations of both these solutions were discovered. The nitric acid had a concentration of 0.955 M and the ammonium hydroxide had a concentration of 0.813 M. The goal after these findings was to create a buffer using one of the solutions. To create the buffer, ammonium hydroxide was chosen because it was considered the weak specimen, and it was determined that 1.087 grams ammonium chloride () was the conjugate acid to be used. Titrating it with a strong acid and then a strong base tested the effectiveness of the buffer. Compared to the control titration done where water was titrated with a strong base and a strong acid, the buffer was considered effective because it was able to resist a pH change.

**Introduction**

The objective of this experiment was to determine the identity and concentrations of two unknown solutions using a pH meter, anion and cation tests, and titration. One of these solutions was then used to create a buffer, and titration was used once again to test the effectiveness of the buffer.

An acid is defined as a molecular compound that is able to donate an ion when dissolved in water, and therefore increases the concentration of , and a base is a molecular compound that either accepts ions or donates ions.1 On a pH scale, acidic solutions will have a pH of less than seven and will have a greater concentration of ions, while basic solutions will have a pH of greater than seven and will have a greater concentration of ions. A solution that is neutral will produce a pH equal to seven. pH is simply the measure of acidity, or mathematically, 1 If an acid and base react together, and the concentration of ions is equal to the concentration of ions, it can result in a neutralization reaction, where and the solution is therefore neutral. However, if the concentration of the ions is unknown in either the acid or base, acid-base titration can be used to determine it.

The titration works by adding an acid (or base) of known concentration to a base (or acid) of unknown concentration.2 The acid (or base) of known concentration is contained in the burette, and slowly added to the base (or acid) of unknown concentration. At one point, the concentration will switch from having an excess of ions to having an excess of ions (or vice versa). Right in the middle of this switch, is the equivalence point, where []=[], and it is at this point where the moles of the acid (or base) of known concentration is equal to the moles of the base (or acid) of unknown concentration.2 The unknown concentration can then be deducted through this information.

A titration curve can be created, and based on the way the curve looks, it can be determined if an acid (or base) is weak or strong. Weak acids (or bases) have a curve that declines fairly quickly at the beginning, then levels out, and then drops a large amount, and then levels out again. A strong acid (or base) has a curve that is consistent at the beginning, and then has a large jump, and then levels out. These are used to determine which solution to use for a buffer.

A buffer resists pH change by neutralizing added acid or added base.1 A buffer must contain significant amounts of a weak acid (or base) and its conjugate base (or acid).1 A buffer works because there is an acid and a base component to it already, so whether an acid or base is added to it, there is something already there to react with it, and therefore neutralize it.1

**Experimental**

Two unknown solutions were obtained, 14A and 14B, and a properly calibrated pH meter was used to determine the acidity of the solutions. These solutions were then run through multiple tests to determine the components that make up each solution.

To evaluate the cation component of the unknown substances a flame test was done. The flame test can only be used for metal cations. The wire was decontaminated with HCl and the HCl was burned off in the flame until no color remained and the flame was completely blue. The wire was then dipped in the standard solution. It was then heated in the flame and a color was produced. Unknown 14A and unknown 14B were then tested using the same process. The colors they produced were compared to the standards originally tested to determine what cation they contained. An ammonium test was also done to test for the chance the compounds may contain a nonmetal cation. 1mL of 6 M NaOH was added to a test tube that contained 1mL of the standard solution. Moist litmus paper was then placed over the top of the test tube. This procedure was repeated for unknown 14A and unknown 14B, but instead of the 1mL of the standard solution, there was 1mL of the unknown substances, in separate test tubes.

Several tests were completed to determine the anion portion of the unknown compounds. For the halogen test 1 mL of the standard solution was added to a test tube, followed by 1 mL of 6 M and 1 mL of 0.1 M . This method was repeated for both unknown substances, but 1 mL was added of the unknown substances (in different test tubes) instead of the 1 mL of the test solutions.

Next, the presence of sulfate was tested. 1 mL of the standard solution was added to a test tube along with 1 mL of 6 M HCl and 1 mL of 0.1 M . This technique was repeated for unknown 14A and unknown 14B in place of the standard solution.

For the nitrate test, 1 mL of the standard solution was placed in a test tube, followed by 3 mL of concentrated and was then mixed together. After the test tube had cooled to room temperature, it was inclined at 45 degrees on a three-fingered clamp and 2 mL of was slowly added down the side of the test tube. Once again, this test was repeated separately for unknown 14A and unknown 14B substituted for the standard test solution.

The last test completed tested for the presence of acetate. 2 mL of standard solution was placed in a test tube, and one drop of concentrated sulfuric acid was added, followed by 1 mL of ethanol. The solution was then heated in a water bath for several minutes. This procedure was repeated for unknown 14A and 14B.

Serial dilutions of 10-, 100-, 1,000-, and 10,000-fold were then done of both 14A and 14B. From the stock solution, 10 mL was taken and it was then diluted with 90 mL of water to make the 10-fold, and the pH was taken of this solution. 10 mL of this new solution was then taken and diluted with 90 mL of water to make the 100-fold solution, and the pH was taken once again. This was repeated for the 1,000- and 10,000-fold solutions, each time using the diluted solution that was just created in the step before.

The concentrations of the unknown solutions were then determined. 14B was titrated by 1.01 M HCl. HCl was added in 0.5 mL increments until it reached 8 mL, and the pH was taken after every 0.5 mL added. The increments were then decreased to 0.1 mL until the pH of the solution leveled off. The process was repeated for 14A, but 14A was titrated by 0.974 M NaOH instead.

Using 14B, a buffer was then prepared by adding . Based on the concentration of 14B, it was determined that 1.087 g of should be added to create the buffer. The buffer was then tested, using the same process as listed above, by titrating it with a strong acid (1.01 M HCl) and then a strong base (0.974 M NaOH). A control titration was then done using the same process, but water was used in place of the buffer.

**Results and Discussion**

A pH test and various cation and anion tests were used to determine the identities of 14A and 14B. Using a calibrated pH meter, the pH of 14A was determined to be acidic with a pH reading of 0.5, and the pH of 14B was determined to be basic at 11.6. Acidic pH readings mean that there are more ions than ions, and the opposite is true for basic pH readings. The anion tests were then conducted. In the halogen test, the standard solution resulted in white precipitate being formed. When 14A and 14B were tested, there was no reaction for 14A, and while 14B had an exothermic reaction, there was no white precipitate formed so it was concluded that the unknowns did not contain any halogens. A sulfate test was then conducted, and the standard once again resulted in a white precipitate being formed. 14A had no reaction and 14B had an exothermic reaction, but once again no white precipitate was formed so it was determined there was no sulfate component in these solutions. The standard for the acetate test resulted in a fruity smell, but when 14A and 14B were tested, there was no fruity smell. Once again it was determined that this anion is not present in either unknown solution. The nitrate test standard resulted in brown ring being formed in between the solutions. When 14B was tested, there was no reaction, but when 14A was tested, a brown ring did form so it was concluded that 14A contained nitrate.

The cation component was then tested. The flame test was used to test for potassium and sodium. The standard for potassium turns purple when it burns off in the flame, and the standard for sodium is orange. 14A and 14B both had negative results for both standards. The ammonium test was then run. The presence of ammonium is indicated by a moist piece of litmus paper turning blue when placed over the top of the test tube containing the solution. 14A had no reaction when this test was conducted with it, but 14B resulted in the litmus paper turning blue.

The anion component of 14A and the cation component of 14B were identified, but the other portion of the compound still had to be determined. Because 14A was acidic, it was concluded that to account for this, there was to be an abundance of ions in its chemical formula. This information was used in combination with the anion component already known to determine that the identity of 14A was , or nitric acid. The same idea was used to identify 14B; because it was basic it was determined that there had to be an abundance of ions. With this information, and with the cation component already identified, it was concluded that the identity of 14B was ammonium hydroxide ().

A 10-, 100-, 1,000-, and 10,000-fold dilution were done with these solutions, and the pH was recorded. The stock solution of 14A, the nitric acid, had a pH of 0.24, for the 10-fold dilution the pH was 0.70, the 100-fold was 1.52, the 1,000-fold was 2.37, and the 10,000-fold was 3.15. While the pH slowly approached to a neutral reading, there were still much greater amounts of water compared to the acid but the pH of the solution did not go up very much, and still stayed acidic. For 14B, the ammonium hydroxide, the pH of the stock solution was 11.64, for the 10-fold it was 10.98, the 100-fold was 10.40, the 1,000-fold was 9.84, and the 10.000-fold was 9.35. The same was true for this solution, even with large amounts of water, the pH did not go down an extreme amount and the solution was still basic.

Acid-base titration was then used to determine the concentrations of the two solutions. Ammonium hydroxide, because it was a base and contained more ions, was titrated with 1.01 M HCl, which is a strong acid. When a solution is neutral it will have a pH of 7, and the amount of and ions will be equal, but as soon as there’s more of one than the other, it will shift to either being basic (more ions) or being acidic (more ions). Because the concentration of the HCl was known, it was possible to determine the concentration of the ammonium hydroxide. The ammonium hydroxide was placed in the flask and the HCl was placed in the burette, and added in small increments, while the pH was measured after every addition. When the pH went from being basic to being acidic, it was determined that the equivalence point was in that range, and at that point the moles of HCl equaled the moles of . On the titration curve, which can be seen in Figure One: Titration of Ammonium Hydroxide, the equivalence point is the midway point of the Sigmoidal curve, which is where the major jump of pH value takes place.

It took 8.05 mL of 1.01 M HCl to reach this point. With this information, it was determined that the concentration of the ammonium hydroxide was 0.813 M. Sample calculation 1 below shows how this concentration was determined. It was also possible to determine that ammonium hydroxide was a weak base, based on the fact the curve drops a little a the beginning, then stays fairly consistent, and then drops even more again before leveling off. The same process was repeated for the nitric acid, but the nitric acid was placed in the flask and it was titrated with NaOH, which is a strong base. The results for this titration can be seen in Figure Two: Titration of Nitric Acid. Based on this figure, it can be determined that nitric acid is a strong acid based on the fact that the curve is consistent at the beginning and then has one large jump in pH value before leveling off.

For this titration, it took 9.80 mL of 0.974 M NaOH to reach the point where the moles were equal of the solutions. Based on this, it was determined that the nitric acid had a concentration of 0.955 M. Once again, sample calculation 1, which can be seen below shows the process that the concentration was determined by.

**Sample Calculation 1**

Because ammonium hydroxide is the weak specimen, it was chosen to create the buffer, and ammonium chloride (was determined to be the conjugate acid for it because it differed by only one proton. The buffer also needed to have an equal amount of and ions and based on the number of moles in 25 mL of (amount of solution in the flask), it was determined that 1.087 grams of ammonium chloride was needed in order to achieve this. Sample calculation 2 demonstrates how this mass was concluded.

**Sample Calculation 2**

Titration was then used to determine the effectiveness of this buffer. The buffer was first titrated with HCl, and then titrated with NaOH. A control titration was also run, in which water was titrated with HCl and NaOH. When the water was titrated with HCl, its starting pH was 8.97, and within 0.5 mL of HCl being added, the pH dropped to 2.16. When it was titrated with NaOH, its starting pH was 3.65, and within 0.5 mL of the NaOH being added, the pH increased to 11.64. The results from these titrations were used in reference for the effectiveness of the buffer created. When the buffer was titrated with HCl, the pH slowly dropped after each addition of HCl to it. Between the 7.50 and 8 mL addition, the pH went from 6.49 to 2.05. It was at this point that the acidity of the solution became too much for the buffer, and the acid was able to overcome it. The buffer was then titrated with NaOH. Once again there was no immediate major change in the pH of the solution. And while there was no major jump in pH like there was with the HCl, the greatest difference of pH occurred between 7.5 mL being added and 8.0 mL being added. The pH value went from 11.26 to 12.01. The effect of the titrations on the pH value can be seen in Figure Three: Titrations of the Buffer.

Compared to the control titration of water, the buffer proved to be very effective in resisting pH change. However, it should be noted that the in order for the buffer to be effective, there needs to be about 10 times as much of the buffer than of the acid or base that is to be added.2 The buffer capacity is broken once there is no longer sufficient amounts of the buffer in comparison to the base or acid added.

**Conclusion**

` Unknown solution 14A was identified as nitric acid and 14B was identified as ammonium hydroxide through the use of a pH meter and various anion and cation tests. The nitric acid had a concentration of 0.955 M and the ammonium hydroxide 0.813 M, which was discovered through acid-base titration. The ammonium hydroxide was determined to be a weak base, and was used with 1.087 grams of ammonium chloride, which was determined to be the conjugate acid, to create a buffer. The buffer was titrated with a strong acid and a strong base and proved to be more effective at resisting a pH change than compared to the control experiment, which used water in place of the buffer.

**References**

1. Meyers, J. Prelab lecture. March 2015.
2. Meyers, J. Lecture notes. March 2015.