# Acid-base titrations

Acid-Base Titrations February 21, 2013 Abstract: With the given volume measurement of KHP and NaOH, the students experimented to calculate the mass of unknown acid and its equivalence point using a method called titration and titration curve. The known volume of an acid solution would be titrated by slowly adding drops of solution into NaOH, and the volume of base needed to react with the acid is measured throughout. By using those data, the students are able to find what they were trying to calculate. On this lab, the students found the unknown acid, 4. 36 x 10-5, and the equivalence point of 8. 4, as well as the value of pH. Purpose: The purpose of the experiment was to utilize the techniques of titrations to measure the concentration of an acid or base in solution, to calculate molar mass of an unknown acid or base, and to determine the equilibrium constant of a weak acid or weak base. Materials: NaOH Unknown acid Distilled water Balance Erlenmeyer flask Funnel pH sensor or pH meter Ring stand and buret clamp 250-mL beaker 50-mL buret KHP Procedure: Part 1: Standardization of a Sodium Hydroxide Solution 1. Obtain a sample of pre-dried potassium hydrogen phthalate (KHP) 2. Weigh 0. 4-0. 6 grams of KHP, then record in Data Table 1. 3. Transfer the KHP into the Erlenmeyer flask using the funnel, then use the spray bottle to get any remaining solid into the flask. 4. Add 40 mL of water to the flask and whirl until thoroughly dissolved. 5. Obtain 75 mL of sodium hydroxide (NaOH) solution. 6. Clean a 50-mL buret, then rinse it with three small portions of NaOH. 7. Fill the buret to above the zero mark with the NaOH solution. 8. Open the buret stopcock to allow any air bubbles to escape from the tip. Close the stopcock when the liquid level is between 0 and 10- mL 9. Measure the precise volume of the solution in the buret and record this value in Data table 1 as the " initial" volume. 10. Position the buret over the Erlenmeyer flask so that the tip of the buret is within the flask but at least 2 cm above the liquid surface. 11. Add three drops of phenolphthalein solution to the KHP solution in the flask. 12. Begin the titration by adding 1. 0 mL of the NaOH to the Erlenmeyer flask. Close the buret stopcock and swirl the flask to mix the contents. 13. Repeat step 12 until 15 mL of the NaOH solution have been added to the flask. Be sure to continuously swirl the flask. 14. Reduce the incremental volumes of NaOH solution to . 05 mL until the pink color starts to persist for 15 seconds, while constantly swirling. 15. Measure the volume of NaOH solution remaining in the buret, estimating to the nearest . 01 mL and record this value as the " final volume" in Data table 1 16. Repeat the standardization titration two more times. Rinse the flask thoroughly between each trial with water. Part 2: Determination of the Equivalent Mass of an Unknown Acid 1. Weigh about 0. 3-0. 4 g of a sample of the unknown acid and record the precise mass in Data table 2. 2. Dissolve the unknown acid in 40 mL of water and titrate to the phenolphthalein endpoint as above in steps 5-16. 3. Record the initial and final volumes of NaOH in Data table 2. 4. Repeat once more, choose a mass for the second sample so that the volume of NaOH needed will be about 45 mL when using the 50 mL buret. Part 3: Determination of the pKa of the Unknown Acid 1. Set up a pH meter and electrode. Calibrate the pH meter isong a pH 7. 0 buffer. Rinse the electrode well with water. 2. Weigh a sample of unknown acid that will require approximately 20 mL of titrant. 3. Dissolve the acid in approximately 100 mL water in a 250-mL beaker. 4. Fill the buret with the standardized NaOH solution used in part 1. Record initial volume as the " initial buret reading" in Data table 3. 5. Stir the beaker containing the unknown acid solution. Submerge the pH electrode into the acid solution. 6. When the pH reading has stabilized, record the initial pH of the solution in Data table 3. 7. Add about 1 mL of sodium hydroxide solution. Record the exact buret reading and the pH in Data table 3. 8. Record the pH of the solution next to the buret reading in data table 3. 9. Add another 1 mL increment of sodium hydroxide solution. Record both the buret reading and the pH. 10. Continue adding sodium hydroxide in 1 mL protions, recording buret reading and pH. 11. When the pH begins to increase by more than 0. 3 pH units after 1 mL of NaOH, decrease the amount of sodium hydroxide added to about . 2-mL 12. Continue adding sodium hydroxide in about . 2 — mL increments, recording the buret reading and the pH 13. When the pH change is again about 0. 3 pH units, resume adding the sodium hydroxide in 1 mL increments, still recording buret reading and pH 14. Stop the titration when the pH of the solution is greater than 12. Record the final volume of the solution is the buret and the final pH. 15. Graph the data, with pH on the y-axis and the volume of NaOH on the x-axis. Data: Data Table 1 Trial 1 Trial 2 Trial 3 Mass KHP, g 0. 56 0. 54 0. 50 Final volume, mL 35. 7 35. 9 35. 4 Initial volume, mL 1. 20 1. 60 5. 70 Volume of NaOH added, mL 34. 5 34. 3 29. 7 Molarity NaOH (average) 0. 217 M Data Table 2 Trial 1 Trial 2 Mass of unknown acid, g 0. 35 0. 70 Final volume, mL 26. 5 45. 2 Initial volume, mL 5. 4 5. 6 Volume of NaOH added, mL 21. 1 39. 6 Equivalent Mass (average) 78. 96 g/mol Mass of unknown acid 0. 39 Standard NaOH concentration 0. 01 Initial buret reading 4. 3 Initial pH 3. 90 Buret reading (mL) pH Buret reading (Cont) pH 5. 3 3. 90 21. 5 4. 80 6. 3 3. 95 22. 5 4. 89 7. 3 4. 00 23. 6 5. 01 8. 2 4. 05 24. 6 5. 09 9. 5 4. 11 25. 6 5. 25 10. 4 4. 16 26. 5 5. 44 11. 5 4. 20 27. 6 5. 76 12. 5 4. 25 27. 7 5. 82 13. 4 4. 30 28. 0 5. 95 14. 5 4. 36 28. 3 6. 13 15. 5 4. 41 28. 5 6. 42 16. 4 4. 47 28. 7 7. 07 17. 4 4. 53 28. 9 7. 30 18. 7

4. 61 29. 2 8. 34 19. 6 4. 67 29. 5 9. 81 20. 6 4. 73 29. 8 10. 5 Analysis: From the standardization data in Part 1, calculate the molarity of the sodium hydroxide solution for each trial. Average the values and enter the average in Data Table 1. Molar mass of KHP= 39. 10 + 1. 01 + 30. 97 = 71. 08 g/mol (0. 56 g)/(71. 08 g/mol) = 0. 0079 mol (0. 0079 mol)/(0. 0345 L) = 0. 229 M (0. 54 g)/(71. 08 g/mol) = 0. 0076 mol (0. 0076 mol)/(0. 0345 L) = 0. 220 M (0. 50 g)/(71. 08 g/mol) = 0. 0070 mol (0. 0070 mol)/(0. 0345 L) = 0. 203 M From the equivalent mass of the unknown acid for each trial. Average the values and enter the average in Data Table 2. 0. 35/(21. 1/1000 x 0. 217) = 76. 45 g 0. 70/(39. 6/1000 x 0. 217) = 81. 46 g From the titration curve of pH versus volume of NaOH added in Part 3, determine the pKa of the unknown acid. Calculate the value of Ka for the unknown acid. Antilog 4. 36 = 4. 36 x 10-5 Discussion: In this laboratory, the students were given certain volumes of KHP, and NaOH by teacher, as well as unknown acid, and they were told to calculate molarity of the NaOH, equivalent mass of the unknown acid, and the pKa of the unknown acid. Using the method of titration, the students were able to calculate each one of the values that they were told to find. Those values were 0. 229 M, 81. 46 g, and 4. 36 x10-5. This tells the experimenters that the calculated concentration of a substance react with another substance until a change in color occurs. Once the reaction of definite and known proportion were calculated, the students used those measurements to calculate the unknown concentration of an acid. Acid-base titration theory was utilized in the lab. The titrant is added to the titer to neutralize the solutions, and then the chemical change will occur. The neutralization point occurs at a pH of 7. 0, and when that happens, the color for this lab should turn baby pink. If the color is too dark, then that means

the students have over-titrated the solution, which happens quiet often. The point at which added solution neutralize the unknown concentration of an acid is called the equivalence point, and its solution's pH would be 7. 0. Using the pH data, it can be graphed to make titration curve, producing a sloping curve that is particularly steep around the equivalence point. Since pH is measure of the H3O+ concentration, adding one or two drops to a neutral solution will greatly change its concentration while doubling the amount added does not. That is why the titration curve is so steep in one area, making the equivalence point to be identified easily by students. Some students get frustrated with acid-base titration because it is difficult to titration two solutions together and might wonder why they are doing this lab that doesn't do anything to the world. This is not true. The entire pharmacist utilizes this method; imagine the pharmacists who use titration every day to mix compound drugs successfully. In addition, the doctors apply titration to determine the correct proportion of different medicine, and it also is sued to monitor blood glucose levels, as well as in pregnancy test. Conclusion: The experiment was successful, because I was able to use the techniques of titrations to measure the concentration of an acid, and calculate the molar mass of an unknown acid, which turned out to be 4. 36 x 10-5, as well as the equilibrium constant, which is around 8. 4. Questions: Why is the answer obtained in Question #2 the equivalent mass of the acid rather than the molar mass? The answer obtained in Question #2 the equivalent mass of the acid rather than the molar mass, because it is the mass of the acid with one mole of hydrogen ions. Why must the KHP and the acid samples be dried? If they are not dried, would the calculated molarity of the NaOH be higher or lower than the actual value? Explain. If the KHP and acid samples are not dried, the calculated molarity of the NaOH be higher than the actual value, because we are going to calculate bigger moles than we should have got. Why must NaOH be standardized? Why can't an exact solution of NaOH be prepared? NaOH must be standardized, because when we weigh the solution, the weight will include not only the NaOH, but also an unknown amount of water that tends to absorb water from the air in the sample. Therefore, you have to standardize NaOH in order to find exact solution. Why is the equivalence point in the titration of the unknown acid with sodium hydroxide not at pH 7? The equivalence point in the titration of the unknown acid with sodium hydroxide is not a pH 7. When titrating NaOH against weak acid, at the equivalence point the weak acid has released its entire negative ion, and the negative ions from weak acids are that acid's strong conjugate base; therefore, the equivalence point is expected to be basic.