Background

Acids, Bases, and pH

Acids and bases are powerful chemicals that are often encountered in the laboratory. One common definition of acids and bases focuses on the ions that are produced when an acid or a base dissolves or dissociates in water. An acid is a substance that readily donates its hydrogen cation, H^+, or proton, when placed in water forming the polyatomic ion hydronium, H_3O^+. To recognize an acid, look for ionic compounds that have hydrogen as the cation. A base is the counterpart that accepts a proton from the acid. The most recognizable bases have hydroxide, OH^-, as the anion because the hydroxide anion can accept the proton to become water.

$$\text{Acid: } \underset{\text{H}^+\text{ donor acid}}{\mathrm{HCl}} + \underset{\text{H}^+\text{ acceptor base}}{\mathrm{H_2O}} \longrightarrow \underset{\text{Hydronium ion}}{\mathrm{H_3O^+}} + \underset{\text{Chloride ion}}{\mathrm{Cl^-}} \qquad \text{Base: } \underset{\text{Ionic compound}}{\mathrm{NaOH}} \xrightarrow{\text{water}} \underset{\text{Cation}}{\mathrm{Na^+}} + \underset{\text{Hydroxide ion}}{\mathrm{OH^-}}$$

Strong acids completely dissociate in water to their respective anions and the hydronium ion. For weak acids this dissociation is partial and reversible. The extent of dissociation is captured in the value of its acid ionization constant, K_a. Water is a substance that can act as an acid, donating protons, and as a base, accepting protons, and is said to be amphoteric. The autoionization constant for water is K_w, and the water dissociation equilibrium is:

$$2\mathrm{H_2O} \rightleftharpoons \mathrm{H_3O^+} + \mathrm{OH^-}$$

The value of K_w is 1.0×10^{-14} at 25°C. In pure water, the molar concentrations of hydronium and hydroxide ions are equal because water is a neutral solution. The calculation of K_w is shown in Equation 1, where the molar concentrations of H_3O^+ and OH^- are indicated by the square brackets.

$$K_w = [\mathrm{H_3O^+}][\mathrm{OH^-}] \qquad \text{Equation 1}$$

$$K_w = [1.0 \times 10^{7}\ \mathrm{M}][1.0 \times 10^{7}\ \mathrm{M}]$$

$$= 1.0 \times 10^{14}$$

In this experiment, you will determine how acidic or basic different solutions are by measuring the pH of each solution. The pH of a solution is the negative log of the hydronium ion concentration, in other words, the pH of a solution is a measure of its concentration of the hydronium ion.

$$\mathrm{pH} = -\log[\mathrm{H_3O^+}] \qquad \text{Equation 2}$$

The pH scale ranges from 0 to 14. A pH of 7 is neutral. Below pH 7, the solution is acidic with a larger hydronium ion concentration than a neutral solution, and a pH above 7 is a basic solution with a smaller hydronium ion concentration than a neutral solution.

$$\text{pH} < 7\text{: acidic} \quad \text{pH} = 7\text{: neutral.} \quad \text{pH} > 7\text{: basic}$$

If the concentration of either the hydronium ion or the hydroxide ion is known for a solution, Equation 1 can be used to calculate the unknown value and Equation 2 can be used to calculate the pH of the solution.

Example:

Suppose the concentration of hydronium ion in solution is 1.04×10^{-3}M, calculate the concentration of hydroxide and determine the pH of the solution. First calculate the concentration of hydroxide ion.

$$[\text{OH}^-] = \frac{1.0 \times 10^{-14}}{[\text{H}_3\text{O}^+]} = \frac{1 \times 10^{-14}}{1.04 \times 10^{-3}} = 9.62 \times 10^{-12}\text{M}$$

Now calculate the pH of the solution. The hydronium ion concentration is given in the question.

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log[1.04 \times 10^{-3}] = 2.98$$

Looking at the results, this solution is acidic because the pH is below 7 and the hydroxide ion concentration is much smaller than the hydronium ion concentration, as expected.

Buffers

Many everyday solutions need to have their pH controlled, usually near neutral; from contact solutions to your blood. It is important to understand how pH behaves for different types of acids and bases as well as how to maintain a constant pH. Buffers are solutions that are able to resist a change in pH over a particular range by reacting with and neutralizing small amounts of added acids or bases.

Buffers usually consist of a weak acid and its salt. This pair of compounds differs only by a proton. In solution, the weak acid reacts with excess base, and the anion of the salt picks up the excess proton. One example of a buffering system in the human body is the blood. Blood maintains a pH between 7.35 and 7.45. The weak acid in blood is carbonic acid, H_2CO_3, and its salt is the hydrogen carbonate ion, HCO_3^-. When base enters the blood, the hydroxide ions react with the carbonic acid to produce hydrogen carbonate and water.

$$\text{H}_2\text{CO}_3{}^-(\text{aq}) + \text{OH}^-(\text{aq}) \longrightarrow \text{HCO}_3{}^-(\text{aq}) + \text{H}_2\text{O}(\text{l})$$

If an acid enters the blood, the hydronium ion, or proton, reacts with the carbonate ion and reforms carbonic acid.

$$\text{H}_2\text{CO}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \longleftarrow \text{H}_3\text{O}^+(\text{aq}) + \text{HCO}_3{}^-(\text{aq})$$

In this part of the experiment, you will study how buffered solutions and a unbuffered solutions are affected by the addition or acid or base.

Acid-Base Titration

Neutralization Reactions

When acids and bases are mixed the hydrogen ions from the acid combine with the hydroxide ions from the base to form water and a neutral salt. This is called a neutralization reaction. One example is the reaction between hydrochloric acid and sodium hydroxide.

$$\underset{\text{acid}}{\mathrm{HCl(aq)}} + \underset{\text{base}}{\mathrm{NaOH(aq)}} \longrightarrow \underset{\text{salt}}{\mathrm{NaCl(aq)}} + \underset{\text{water}}{\mathrm{H_2O(l)}}$$

In this reaction, the cation from the base combines with the anion from the acid to form the salt, NaCl, and the proton from the acid and the hydroxide ion form water. This reaction can also be written to show each ion in solution.

$$\mathrm{H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \longrightarrow Na^+(aq) + Cl^-(aq) + H_2O(l)}$$

Ions that appear on both sides of the equation are spectator ions and can be removed. The net ionic equation remains.

$$\mathrm{H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)}$$

In a complete neutralization reaction, the amounts of H^+ and OH^- are equal.

Titration

A titration is a method to determine the concentration of an unknown solution (analyte) using a solution of known concentration (titrant). In this part of the experiment, you will perform a titration to determine the concentration, molarity, of acetic acid in vinegar. Acetic acid, CH_3COOH, is a weak acid that will be neutralized by the strong base sodium hydroxide, NaOH, to form a salt and water.

$$\mathrm{CH_3COOH(aq) + NaOH(aq) \longrightarrow NaCH_3COO^-(aq) + H_2O(l)}$$

The reaction will proceed until all of the acid in the vinegar is neutralized. This is the endpoint of the reaction. In order to see when this occurs, the indicator phenolphthalein will be used. In acidic solutions phenolphthalein is colorless and in basic solutions it is pink. Once the endpoint is reached, no more base will be added to the acid and the volume of base needed to neutralize the acid is determined.

Molarity of Acetic Acid

The molarity of acetic acid can be calculated from the volume of base needed to neutralize the acid. Molarity is moles of solute per liter of solution. Equation 3 shows how to calculate the moles of NaOH used from the volume of NaOH used and the molarity of the solution.

$$\text{moles}_{\text{NaOH}}(\text{mol}) = V_{\text{NaOH}}(\text{L}) \times \frac{\text{moles NaOH (mol)}}{\text{1 L NaOH}} \qquad \text{Equation 3}$$

Use the molar ratio between sodium hydroxide and acetic acid to determine the moles of acetic acid present. In this reaction, the mole ratio is 1:1, so at the endpoint of the reaction the moles of NaOH used equal the moles of CH_3COOH in the vinegar.

$$\text{moles}_{\text{acetic acid}}(\text{mol}) = \text{moles}_{\text{NaOH}}\ (\text{mol}) \times \frac{1\ \text{mol}\ CH_3COOH}{1\ \text{mol NaOH}}$$

Now the molarity of acetic acid can be calculated using the volume of acetic acid used and the moles of CH_3COOH calculated.

$$M_{\text{acetic acid}}\left(\frac{\text{mol}}{\text{L}}\right) = \frac{\text{moles}_{\text{acetic acid}}(\text{mol})}{V_{\text{acetic acid}}(\text{L})} \qquad \text{Equation 4}$$

Mass/Volume Percent %(m/v) of Acetic Acid

The mass/volume percent is another way to express concentration and is defined as the mass of solute in grams per milliliters of solution. To calculate this convert the moles of acetic acid to grams, using the molar mass.

$$\text{mass}_{\text{acetic acid}}(\text{g}) = \text{moles}_{\text{acetic acid}}(\text{mol}) \times \frac{\text{60.1 g acetic acid}}{\text{1 mol acetic acid}} \qquad \text{Equation 5}$$

Then, calculate the mass/volume percent.

$$\text{mass/volume percent}\ \%(\text{m/v}) = \frac{\text{mass}_{\text{acetic acid}}(\text{g})}{V_{\text{acetic acid}}(\text{mL})} \times 100 \qquad \text{Equation 6}$$

Objectives

- Measure the pH of solutions with a pH meter
- Observe effects of added acid or base to buffered and unbuffered solutions
- Calculate pH from [H_3O^+] and [OH^-] of solutions
- Perform a titration
- Calculate molarity and mass/volume percent of acetic acid in vinegar

Materials

- Test tubes and rack
- Stirring rod
- Solutions for testing pH
- Vinegar
- Calibration buffers
- NaCl, HCl, NaOH solutions of varying concentrations
- pH meter
- Wash bottle
- Graduated cylinder
- 150- and 150-mL beakers
- 250-mL Erlenmeyer flask
- 50-mL or 25-mL buret
- Ring stand, buret clamp
- Funnel

Safety goggles must be worn at all times! HCl and NaOH are corrosive and can cause severe skin burns and irritation. Handle these substances with care. Dispose of all solutions as instructed.

Check with TA or Instructor for any other supplies you may need.

Procedure

Measuring pH of Solutions

Determining pH with pH Paper

1. Obtain sample solutions, pH paper, test tubes, stirring rod, and a test tube rack.
2. Add ~2 mL of each liquid sample to separate test tubes. For viscous/solid samples, add ~2 mL of water and mix.
3. Take a stirring rod and dip it into your sample solutions. Add a drop of solution to the pH paper.
4. For each sample, record the color observed and compare the color to the scale on the container and record the pH.

Determining pH with a pH Meter

5. Prepare new samples of solutions as described in step 1.
6. Obtain a pH meter and rinse the electrode with DI water.

7. Place the electrode in the sample and record the pH.
8. Repeat for all samples, rinsing the electrode with DI water between each reading.
9. Determine and record each sample as acidic, basic, or neutral.

Buffers and pH

Effect of Adding Acid to a Buffer Solution

1. In four new test tubes, add 10 mL of the following solutions into separate test tubes: H_2O, 0.1 M NaCl, high pH buffer, low pH buffer.
2. Determine the pH of each solution using pH paper or a pH meter and record. (If you use a pH meter, be sure to rinse the electrode with DI water between each reading!)
3. Add ~5 drops of 0.1 M HCl solution to each test tube and mix by gently swirling.
4. Determine the pH of each solution using pH paper or a pH meter and record.
5. Repeat steps 4 and 5.
6. Determine the change in pH for each solution.
7. Identify the buffer solutions.

Effect of Adding Base to a Buffer Solution

8. In four new test tubes, add 10 mL of the following solutions into separate test tubes: H_2O, 0.1 M NaCl, high pH buffer, low pH buffer.
9. Determine the pH of each solution using pH paper or a pH meter and record. (If you use a pH meter, be sure to rinse the electrode with DI water between each reading!)
10. Add ~5 drops of 0.1M NaOH solution to each test tube and mix by swirling.
11. Determine the pH of each solution using reference solutions or a pH meter and record.
12. Repeat steps 9 and 10.
13. Determine the pH change of each solution.
14. Identify the buffer solutions.

Acid-Base Titration

Determining the Concentration of Acetic Acid in Vinegar

1. Obtain a 150-mL beaker and add ~20 mL of vinegar.
2. Record the brand of vinegar and the percent acetic acid listed on the label.
3. Add 5.0 mL of vinegar to an Erlenmeyer flask using a 10-mL graduated cylinder or a 5.0-mL pipet.
4. Add ~25 mL of DI water to the Erlenmeyer flask.
5. Add 2-3 drops of phenolphthalein indicator to the Erlenmeyer flask.

6. Obtain a buret, ring stand, buret clamp, and 100 mL of NaOH solution.
7. Record the molarity of the NaOH solution.
8. Rinse the buret with two 5-mL washes of the NaOH solution and discard.
9. Attach the clamp on the ring stand and clamp the buret. Place a beaker under the buret.
10. Make sure the stopcock is in the closed (horizontal) position. Place a small funnel at the top of the buret and carefully pour the NaOH solution into the buret.
11. Open the stopcock (vertical position) and slowly drain NaOH solution until the meniscus is at or below the 0.00 mL marking. Close the stopcock. Make sure the buret tip is full of solution and free of bubbles.
12. Record the volume on the buret. This is the initial volume. Notice that the numbers on the buret are lower at the top and higher at the bottom. Always read the volume such that the meniscus is at eye level as shown in Figure 1.
13. Remove the beaker with the waste solution and place the Erlenmeyer flask with the vinegar and indicator under the buret.

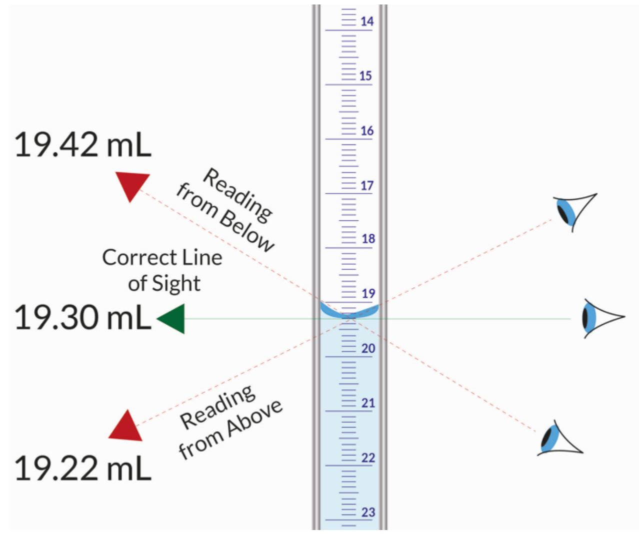

Figure 1 Reading a Buret

14. Slowly open and close the stopcock to add NaOH to the flask. Swirl the flask by holding it at the center joint and making a pendulum motion, or lasso shape.

15. Continue to add NaOH to the flask until the pink color becomes more persistent. When the pink color disappears more slowly, add the NaOH to the flask drop-by-drop. Once a faint pink color persists, the end point has been reached. Stop adding NaOH to the flask. To visualize this better, a white piece of paper can be placed under the flask.

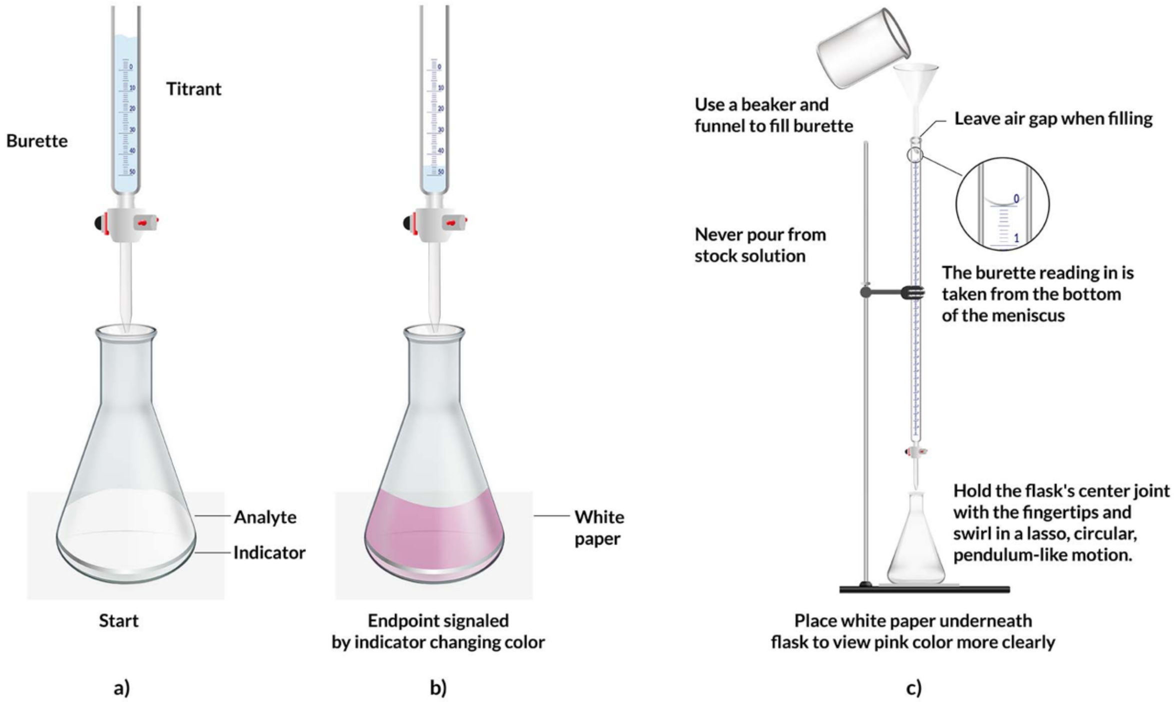

Figure 2 Titration

16. Record the volume in the buret. This is the final volume.
17. Repeat this titration two more times using new samples of vinegar each time and rinsing the buret with DI water before rinsing with NaOH solution.

Data Analysis

18. Calculate the volume of NaOH used to neutralize the acid for each trial. This is the final volume minus the initial volume.

$$V_{\mathrm{NaOH}}(\mathrm{mL}) = V_{\mathrm{final}}(\mathrm{mL}) - V_{\mathrm{initial}}(\mathrm{mL}) \qquad \text{Equation 7}$$

19. Calculate the average volume of NaOH needed to neutralize the acid.

$$V_{\mathrm{average}}(\mathrm{mL}) = \frac{V_{\mathrm{trial\ 1}} + V_{\mathrm{trial\ 2}} + V_{\mathrm{trial\ 3}}(\mathrm{mL})}{3} \qquad \text{Equation 8}$$

20. Convert the average volume of NaOH solution in milliliters to liters.

$$V_{\text{average}}(\text{L}) = V_{\text{average}}(\text{mL}) \times \frac{1\ \text{L}}{1000\ \text{mL}} \qquad \text{Equation 9}$$

21. Calculate the moles of NaOH needed to neutralize the acid in the vinegar for each trial using Equation 3. The number of moles of acetic acid is equal to the moles of NaOH.
22. Convert the volume of vinegar in milliliters to liters.
23. Calculate the molarity of acetic acid from the calculated moles and the volume of vinegar used in liters using Equation 4.
24. Calculate the mass, in grams, of acetic acid using Equation 5.
25. Calculate the percent (mass/volume) of acetic acid in vinegar from the calculated grams acetic acid and initial volume of vinegar using Equation 6.

Acids, Bases, Buffers, Titration Report Sheet

Team ________________________

Determining pH

Sample	Color with pH paper	pH Using pH paper	pH Using pH Meter	Acidic, Basic, Neutral
Vinegar				
Ammonia				
Lemon juice				
Shampoo				
Detergent				
Hair conditioner				
Mouthwash				
Antacid				
Aspirin				

Buffer Solutions

Effect of Adding Acid

Solution	Initial pH	pH after 5 drops HCl	pH after 10 drops HCl	pH change	Buffer (Y or N)
H_2O					
0.1 NaCl					
High pH buffer					
Low pH buffer					

Effect of Adding Base

Solution	Initial pH	pH after 5 drops NaOH	pH after 10 drops NaOH	pH change	Buffer (Y or N)
H_2O					
0.1 NaCl					
High pH buffer					
Low pH buffer					

Acid-Base Titration

Concentration of Acetic Acid in Vinegar

Brand of Vinegar ____________ Volume 5.0 mL % from label _____________

Molarity of NaOH ______________ M

	Trial 1	**Trial 2**	**Trial 3**
Initial Volume NaOH (ml)			
Final Volume NaOH (mL)			
Volume NaOH Used (mL)			
Average Volume (mL)			
Average Volume (L)			
Moles NaOH used (from average volume) *show calculations*			
Moles CH_3COOH (mol)			
Molarity (mol/L) CH_3COOH *show calculations*			
Mass CH_3COOH (g) *show calculations*			
%(m/v) CH_3COOH in vinegar *show calculations*			