EXPERIMENT 11
An Introduction to Acids and Bases

Introduction

Two essential ions found in aqueous solution are H_3O^+ and OH^-. These ions are always present in water because of the autoionization reaction:

$$H_2O + H_2O \rightleftarrows H_3O^+ + OH^-$$

$$[H^+]\,[OH^-] = K_w = 1.00 \times 10^{-14}$$

Because so many chemical reactions are performed in aqueous solution, and both H_3O^+ (frequently abbreviated as H^+) and OH^- react readily with many dissolved species (including the biochemicals present in living things), understanding and quantifying their behavior is an essential part of chemistry. Fortunately, the ideas used to analyze general chemical equilibria can be applied to the study of acids and bases. The only major difference is the specialized nomenclature, such as "pH," used when working with acids and bases.

Discussion

This experiment introduces three methods for measuring the concentration of H^+, otherwise expressible as pH. These methods are indicators, pH paper, and pH meters. We will examine the chemical properties of the indicator bromocresol green, which can be used to measure pH. We will also measure the pH of eight solutions with pH paper, bromocresol green, and a pH meter, to understand the advantages and drawbacks of each method. Finally, we will see how compounds that contain no H^+ or OH^- ions can affect the pH of an aqueous solution.

Acids and Bases

Any aqueous solution can be classified as acidic, neutral, or basic based on the relative concentrations of H^+ and OH^- ions present:

$$[H^+] > [OH^-] \quad \Rightarrow \quad \text{acidic}$$
$$[H^+] = [OH^-] \quad \Rightarrow \quad \text{neutral}$$
$$[H^+] < [OH^-] \quad \Rightarrow \quad \text{basic}$$

Pure water is neutral. Some substances, when added to water, increase the concentration of H^+, making the solution acidic. We can classify these substances as follows:

Dissociating acid – These compounds dissociate to form H^+ (*aq*):

$$H_2SO_4 \rightleftarrows H^+ + HSO_4^-$$

Acidic cation – These compounds react with water to free H^+ in a reaction that is frequently referred to as "hydrolysis" of a metal cation:

$$Fe^{3+} + H_2O \rightleftarrows FeOH^{2+} + H^+$$

Other substances, when added to water, increase the concentration of OH^-, making the solution basic. We can classify these substances as follows:

Basic anion – The conjugate base B^- of a weak acid HB can remove H^+ from H_2O to form a basic solution:

$$F^- + H_2O \rightleftarrows HF + OH^-$$

Basic neutral – These uncharged compounds also react with water to form OH^-:

$$NH_3 + H_2O \rightleftarrows NH_4^+ + OH^-$$

Finally, some ions, such as Na^+ and NO_3^-, exhibit no acid or base tendencies, and thus do not affect the acidity of a solution to which they are added. In this experiment you will encounter examples of each of the above categories by studying eight solutions. The solutions will be 0.10 *M* in each of the following:

1. $CH_2ClCOOH$	3. $C_8H_9NH_2$	5. $Al(NO_3)_3$	7. $NaHSO_4$
2. CH_3COOH	4. $Zn(NO_3)_2$	6. Na_2CO_3	8. HCl

Introduction to pH

The pH of an aqueous solution is the negative base 10 logarithm of the H^+ ion concentration:

$$pH = -\log_{10}[H^+]$$

We already know that at 25°C, $[H^+]\,[OH^-] = 1.00 \times 10^{-14}$. If the solution is neutral, then $[H^+] = [OH^-] = 1.00 \times 10^{-7}$, and the pH = 7.0. Thus we commonly refer to 7.0 as a neutral pH. Acidic solutions, therefore, will have a pH less than 7.0, while basic solutions will have a pH greater than 7.0. Although $[H^+]$ can exceed 10^0 and can be less than 10^{-14}, it is not customary to use the pH scale to describe these very strongly acidic or basic conditions. In usual practice, the pH scale extends from 0 to 14.

Chemical Behavior of Indicators

Some chemicals have a pH-dependent color. These compounds, called *indicators*, can be used to measure the pH of a solution, within a specific range. Most indicators are colored weak acids that dissociate to form conjugate bases of a different color. For example, the yellow acidic form of bromocresol green, denoted by HIn, dissociates to the blue basic form In^-.

The useful pH range of an indicator depends on its equilibrium constant:

$$K_i = \frac{[In^-][H^+]}{[HIn]}$$

For bromocresol green the equilibrium constant $K_i = 1.3 \times 10^{-5}$. Because indicators are strongly colored, a very small concentration of indicator still gives a clearly visible color. Since only a small amount of indicator is used, the indicator itself does not affect the H^+ concentration. Instead, the $[H^+]$ of the solution determines the color of the indicator through the ratio:

$$\frac{[H^+]}{K_i} = \frac{[HIn]}{[In^-]}$$

For cases where the $[H^+]$ is much greater than K_i, the indicator exists almost entirely in its acid form ($[HIn]/[In^-]$ is much greater than 1). This means that at a high enough $[H^+]$, the indicator's color becomes unresponsive to substantial changes in $[H^+]$. Similarly, at the other extreme, when $[H^+]$ is much less than K_i the indicator exists entirely in its basic form, and significant pH changes fail to produce any observable color change. Therefore, an indicator is most sensitive for $[H^+]$ near K_i where the species HIn and In^- are present in similar amounts, and the color is somewhere between the colors of the pure acid and base forms.

Conductivity Measurements

You may recall from Experiment 3 last semester that the conductivity of a solution is indicative of the number of dissolved ions present in solution. Since H^+ and OH^- ions also conduct electricity, there may be some relationship between the conductivity of a solution and its acidic or basic nature. We will explore these relationships with the eight solutions.

pH Titrations

For part 2 of your experiment you will perform a pH titration to determine the concentration of a solution of sulfuric acid (H_2SO_4). In this titration, you will slowly add a solution of 0.1 *M* NaOH to the acid and measure the pH with a pH meter. You will also follow the titration with the bromocresol green indicator. You should review the section of your textbook that deals with acid/base titrations.

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TF: ________________ Name: ________________________

Experiment 11: Procedure, Lab Report and Prelab

Before You Come to Lab:

- Read the entire lab report, including the previous introduction and discussion, and the entire procedure.
- Complete the Prelab, which is the last page of the lab report, and turn the prelab in to your TF as you enter the lab.

Safety in the Laboratory

- Safety glasses or safety goggles and lab coats must be worn at all times while in the laboratory. We will be using some strong acids and strong bases in this lab, so use extra caution when handling any chemicals.
- Nitrile gloves must be worn at all times while performing experiments or working with chemicals. In this experiment we will be working with acids and bases so extra care must be taken.
- Use caution when working with glass. Any broken glassware must be disposed of in the glassware container.

Waste Disposal and Cleanup

- Empty all chemicals into the "Used Chemicals" beaker at your lab bench.
- When you are done with the lab, empty the "Used Chemicals" beaker into the waste collection bucket in the back of the lab.
- Leave everything else at your lab bench.
- When you leave the lab, your lab bench should look exactly as it did when you arrived.

Before You Leave the Lab

- Have your TF check your lab bench for cleanup.
- Submit your data and lab report to your TF. This page and all subsequent pages must be stapled and turned in.
- Wash your hands before leaving the lab.

Grading:

Prelab:	_____ / 10
Lab Report:	_____ / 20
Safety:	_____ / 3
Cleanup:	_____ / 2
Total:	**_____ / 35**

TF: __________________ Name: __________________________

Procedure

Waste Disposal

As you are working, collect all chemicals into the "Used Chemicals" beaker at your lab bench. When you are done with the lab, empty the "Used Chemicals" beaker into the waste collection container in the back of the lab.

Use of the pH Meter

We will use an instrument called a pH meter to obtain precise pH measurements of these eight solutions, as well as for part 2 of this experiment. A pH meter measures the pH of a solution with an electrochemical cell that produces a voltage dependent on the outside hydrogen ion concentration. That voltage is amplified and translated into a pH measurement.

The pH meters we will use are connected to the computers and the Logger Pro software. Your pH meter should be in a small container of buffer solution. Log in to the computer (User: Physical Sciences; Password: ps1) and start Logger Pro. You should be able to read the pH in the bottom left corner of the screen. For this lab, we will just be reading the pH off the computer screen and recording the readings by hand into the lab report.

pH meters are notoriously delicate, so note these precautions before using one:

- The glass probe (which contains the electrochemical cell) is very fragile. This is especially true of the tip of the probe. Do not drop or strike the probe, and never support it by resting the end on the bottom of a beaker.

- Always rinse the probe with distilled water from a wash bottle or a small beaker before transferring the probe to a new solution. When you rinse the probe, be sure to do so over your "Used Chemicals" beaker. Never rinse the probes over the sink.

- When you have finished with the meter, rinse the probe with distilled water and **leave it in the small container of buffer solution**. Make sure the tip of the probe is submerged in the solution. Never allow the probe to dry out; it can become irreparably damaged.

To measure pH, rinse the probe with distilled water from a wash bottle over the "Used Chemicals" beaker, and then carefully lower the probe into the desired solution. Read the pH from the computer screen. It might take a little time for the pH meter to give a stable reading (stirring the solution sometimes helps). Rinse the probe with distilled water again after using it. If you are unable to obtain a relatively stable reading, or if you think there is a problem with your pH readings, ask your TF to show you how to recalibrate the probe in the buffer solution.

TF: __________________ Name: ___________________________

Comparison of pH Determination Methods

Your TF will assign you two of the eight 0.1 *M* solutions listed in the Discussion. Collect about 20 mL of each solution in two 50-mL beakers. Label the beakers, and record the identity of the two solutions. For each of your two solutions, perform the following pH measurements, and measure the conductivity as follows:

pH Meter: Measure the pH of the solution with a pH meter. Follow the instructions given on the previous page for proper use of the pH meter.

Indicator: Add a few drops of bromocresol green indicator into a test tube. Next, transfer about 5 mL of the solution into a test tube. Note the color. You will find at the center bench a number of test tubes showing the color of bromocresol green at different pH values. Can you make an estimate of the pH using this indicator alone?

pH Paper: Take a strip of pH paper about 1 inch (2 cm) long. Moisten the paper by dipping it into the remaining solution. Compare the color with the color on the dispenser, and estimate the pH using the guide.

Measuring Conductivity

The conductivity measurements must be made on more dilute solutions, and it is important that they not be contaminated. Take your 250-mL beaker and rinse it with distilled water. Add 100 mL of distilled water. Check that the conductivity of the water is not greater than 75 μS/cm. Then add 1.0 mL of your solution (you should get this fresh from the stock bottles on the center bench). Measure the conductivity of this dilute solution, stirring with the meter.

Data, Observations, and Notes

Solutions assigned:

Solution #	Identity

Using a pH meter:

Solution Identity	pH Measurement

Using bromocresol green indicator:

Solution Identity	Color	Est. pH (range)

Using pH paper:

Solution Identity	Color of pH paper	Est. pH

Conductivity measurements:

Solution Identity	Conductivity (μS/cm)

When you are through with all of these measurements, write your results in the appropriate columns on the blackboard for your section.

TF: __________________ Name: __________________________

Performing a pH Titration

Obtain a clean 250-mL beaker. Add about 75 mL of distilled water. Go to the center bench and dispense 10.0 mL of the H_2SO_4 solution into this beaker. Add a few drops of bromocresol green, and measure the initial pH of this solution; you may need to mix the solution with the pH probe.

The buret at your lab bench is already filled with 0.100 *M* NaOH. (There should be at least 20 mL of NaOH remaining in the buret. If your buret needs to be refilled, let your TF know.) Run a small amount of NaOH from the buret into your "Used Chemicals" beaker to make sure the tip of the buret is filled with NaOH and does not contain any air.

Titrate the NaOH solution into the sulfuric acid in 1.0-mL increments, recording the pH at each interval. Note that it may take some time for the meter to stabilize after each addition; be patient! You can leave the pH probe in the solution throughout this titration, and you can use the pH probe to stir the solution. Be careful not to dispense base onto the probe—make sure the base goes directly into the solution. In addition, be sure to hold the beaker at all times. Do not leave the plastic beaker standing with the probe resting in it because the beaker may tip over. Record the point at which the indicator seems to change color.

Stop the titration once the pH rises above 11.0 (highly basic solutions can damage the pH probe). Note the final solution color.

Data, Observations, and Note

Initial color of acidic solution with indicator:

Titration data:

mL NaOH added	pH
0.0 (initial reading)	
1.0	
2.0	
3.0	
4.0	
5.0	
6.0	
7.0	
8.0	
9.0	
10.0	
11.0	
12.0	
13.0	
14.0	
15.0	
16.0	

Volume of NaOH required to reach bromocresol endpoint:

Final color of basic solution:

TF: ________________ Name: ________________________

Lab Report

1. Fill in the following chart using the collected data from your section. In the column "category," place the species into one of the four categories: dissociating acid, acidic cation, basic anion, or basic neutral.

Substance	Predicted pH (from prelab)	Average pH (from meter)	Average Conductivity	Category
1. $CH_2ClCOOH$				
2. CH_3COOH				
3. $C_8H_9NH_2$				
4. $Zn(NO_3)_2$				
5. $Al(NO_3)_3$				
6. Na_2CO_3				
7. $NaHSO_4$				
8. HCl				

2. Do you notice any relationship between the measured pH, the conductivity, and the category of each substance? Explain.

3. Compare the three pH determination methods in terms of speed, precision, and accuracy. (When considering accuracy, look at the pH predicted in your prelab.)

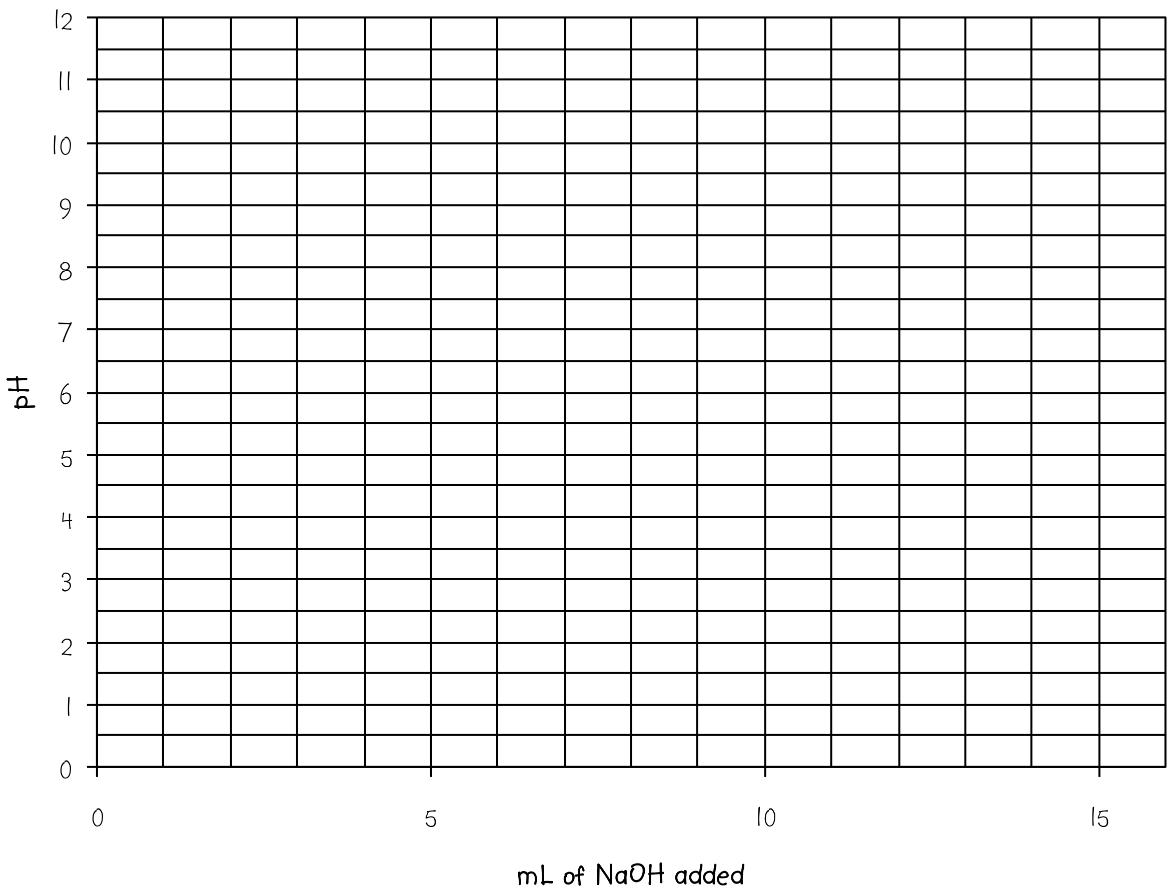
Indicator:

pH Paper:

pH Meter:

TF: __________________ Name: __________________________

4. Plot your titration data on the graph below. Indicate the point at which the indicator changed color.



5. What is the volume of NaOH at the endpoint of this titration?

Volume NaOH at endpoint:

6. Calculate the concentration of H_2SO_4 in the unknown H_2SO_4 solution you obtained from the center bench.

$[H_2SO_4]$ =

TF: ___________________ Name: ___________________________

Prelab

Here are the acid or base dissociation constants (whichever is applicable) for each of the eight species:

Substance	K_a	K_b
1. $CH_2ClCOOH$	1.4×10^{-3}	--------
2. CH_3COOH	1.8×10^{-5}	--------
3. $C_8H_9NH_2$	--------	5.6×10^{-4}
4. $Zn(NO_3)_2$	1.0×10^{-6}	--------
5. $Al(NO_3)_3$	1.0×10^{-5}	--------
6. Na_2CO_3	--------	1.8×10^{-4}
7. $NaHSO_4$	1.2×10^{-2}	--------
8. HCl	strong acid	--------

Calculate the pH of a 0.1 *M* solution of each of these substances. Simply write the *results* of your calculations in the table below; but attach all calculations when you turn in this prelab. Also transfer these results to the table in the lab report on page 9. In addition, predict using a calculation whether the bromocresol green indicator will appear yellow, green, or blue when added to the solution. (Recall that the indicator has $K_i = 1.3 \times 10^{-5}$.)

Substance	Predicted pH of 0.1 *M* solution	Predicted color of indicator
1. $CH_2ClCOOH$		
2. CH_3COOH		
3. $C_8H_9NH_2$		
4. $Zn(NO_3)_2$		
5. $Al(NO_3)_3$		
6. Na_2CO_3		
7. $NaHSO_4$		
8. HCl		

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