# EXPERIMENT 4

*Acid Strength*

## Introduction

Many common substances are either acids or bases. Some acids, like stomach acid are necessary for our health, while others, like sulfuric acid are dangerous and can cause burns and other injuries. Baking soda is a common, weak base used in our homes, while sodium hydroxide, a strong base, is hazardous to skin and eyes. Traditionally, solutions were labeled as being acidic or basic based on their taste and texture. Those that tasted sour were said to be acidic and solutions that tasted bitter and were slippery to the touch were said to be basic. Thus, substances such as lemon juice and vinegar were identified as acids, and solutions of lye and caustic soda as bases. However, in chemistry, we need to have a more definite concept of what makes an acid "acidic" and a base "basic". We also need a means of measuring the relative acidity of different materials. The acidity scale developed is measurement of acidity based on the logarithmic concentration of the hydronium ion, $H_3O^+$ and assigns a mathematical value to the degree of acidity of different compounds.

**Strong Acids and Bases**

In addition to the observed properties of tasting sour and producing a burning sensation when touched, acids also have the following chemical properties. First, they do not conduct electricity when pure, but are conductive when dissolved in water. Second, they react with metal and metallic compounds. Both of these observations are the result of an acid dissociating in an aqueous solution to produce hydrogen ions, $H^+$, or a hydrated hydrogen ion, $H_3O^+$. Any compound that donates a hydrogen ion is said to be an **acid**.

$$HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

Bases undergo a similar dissociation in water. When in an aqueous solution, these compounds dissociate to produce hydroxide ions, $OH^-$. Any compound that dissociates in water to produce hydroxide ions is said to be a **base**.

$$KOH\ (aq) \rightarrow K^+\ (aq) + OH^-\ (aq)$$

The degree of dissociation determines the strength of the acid or base. Acids or bases that dissolve completely in water are considered to be **strong**, while, those that do not completely dissociate from their anions are designated as **weak**. Strong acids produce more hydronium ions for a given concentration of acid than weak acids. **The higher the concentration of hydronium ions produced when adding an acid to water, the stronger the acid solution.**

Like acids, bases produce differing amounts of ions as a result of incomplete dissociation. Strong bases such as KOH or NaOH completely dissociate, producing large quantities of OH- ions while weaker bases do not. The reason hydroxide ion production is so important is that hydroxide ions react with hydronium ions to form water.

$$H_3O^+(aq) + OH^-\ (aq) \rightarrow 2H_2O\ (l)$$

Thus, for every hydroxide ion in solution, you will remove 1 hydronium ion. As hydronium ions are removed from the solution, the hydrogen ion concentration, [H+] decrease and your solution becomes less acidic.

When we are discussing relative acidity of different solutions, we are looking at the formation of hydrogen ion in aqueous solutions only. Water is considered to be **amphoteric**; it can act as either an acid or a base. To understand this, consider the following dissociation of water:

$$H_2O\ (l) + H_2O\ (l) \leftrightarrows H_3O^+(aq) + OH^-\ (aq)$$

When 2 water molecules collide, a hydronium ion, $H_3O^+$, forms when a proton, $H^+$, is transferred from one $H_2O$ molecule to the other. Thus while one water molecule functions as an acid by forming a hydronium ion, $H_3O^+$, the other water molecule acts as a base by forming a hydroxide ion, $OH^-$. This reaction results in equal numbers of $H_3O^+$ ions and $OH^-$ ions and hence the solution is neutral. In a sample of pure $H_2O$, the concentrations of $H_3O^+$ and $OH^-$ ions at 25 °C are $1.0 \times 10^{-7}$ M.

$$[H_3O^+] = [OH^-] = 1.0 \times 10^{-7} \text{ M}$$

As soon as an acid or a base is added to the water, the concentrations of $H_3O^+$ and $OH^-$ change. To determine the exact degree of change relative to neutral water, we use a pH scale.

**Acid Concentration and pH**

The acidity of a solution is based on its hydrogen ion concentration. Unfortunately, these concentrations are not convenient to work with, as they are in the range of 1 to $1 \times 10^{-14}$ M. A logarithmic scale, called a pH scale, was developed to more easily determine the relative acidity of solutions and make the numbers more manageable. **Thus, the pH of a solution is defined as the negative logarithm of $[H_3O^+]$ or negative logarithm of $[H^+]$ in the solution**;

$$\mathbf{pH = -\log [H_3O^+] \quad or \quad pH = -\log [H^+]}$$

Therefore, if you know the pH of a solution, you can calculate the hydrogen ion concentration as follows;

$$\mathbf{[H_3O^+] = 10^{-pH}}$$

Similarly, the concentration of hydroxide ions, $[OH^-]$, of a solution is commonly expressed in terms of the pOH of the solution, which is defined as the negative logarithm of $[OH^-]$

$$\mathbf{pOH = -\log [OH^-]}$$

The hydroxide ion concentration can be obtained from the pOH of the solution using the equation:

$$\mathbf{[OH^-] = 10^{-pOH}}$$

Additionally, the pH and the pOH of any aqueous solution are related, as are the hydrogen and the hydroxide ion concentrations. The equations for these relationships are:

$$\mathbf{[H^+] \times [OH^-] = 1.0 \times 10^{-14}}$$

$$\mathbf{pH + pOH = 14}$$

The following are 2 examples to show the relationships that exist between the pH, the pOH, the hydronium ion concentration and the hydroxide concentration.

***Example (1)***

What is the pH and pOH of a solution that contains $3.50 \times 10^{-5}$ M hydronium ions?

$$pH = -\log [H_3O^+] = -\log (3.50 \times 10^{-5}) = 4.46$$

$$pOH = 14 - pH = 14 - 4.46 = 9.54$$

***Example (2)***

Calculate the hydronium ion and hydroxide ion concentrations of a solution that has a pOH of 4.40.

$$pH = 14 - pOH = 14 - 4.40 = 9.60$$

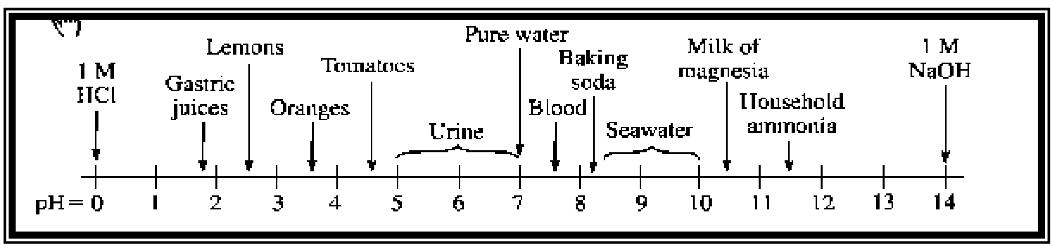
$$[H_3O^+] = 10^{-pH} = 10^{-9.60} = 2.51 \times 10^{-10}\ M$$

$$[OH^-] = 10^{-pOH} = 10^{-4.40} = 3.98 \times 10^{-5}\ M$$

In water, $[H_3O^+]$ is equal to $1.0 \times 10^{-7}$ M, so the pH is 7.0. Because $[H_3O^+] = [OH^-]$ in water, it is neither acidic nor basic.

**pH = 7.0 ( neutral)** **pH < 7.0 (acidic)** **pH > 7.0 (basic)**

The pH scale has a range of 0.0 (very strongly acidic) to 14.0 (very strongly basic).




***Weak Acids, Weak Bases and Buffers***

Many acids and bases do not completely dissociate into cations and anions when in a water solution. We call these compounds weak acids and bases.

Acetic Acid: $HC_2H_3O_2 + H_2O\ (l) \leftrightarrows H_3O^+(aq) + C_2H_3O_2^-(aq)$

Ammonia: $NH_3 + H_2O\ (l) \leftrightarrows NH_4^+(aq) + OH^-(aq)$

The double arrow means that not all of the acid or base forms products, some of it stays in the original form. This means that the base or acid concentration is less than you would expect from the original concentration. Thus, the pH solutions of weak acids or bases are found toward the middle of the pH scale, rather than on either end.

One of the reasons that weak acids and bases are very commonly used is that they have buffering ability. This means that if a small amount of a strong acid or base were added to a buffered solution, no real pH change is seen. A **buffer** is a solution that contains equal amounts of a weak acid and its conjugate base. The conjugate base is the anion that is produced when the weak acid dissociates.

Acetic Acid: $HC_2H_3O_2 + H_2O\ (l) \leftrightarrows H_3O^+(aq) + C_2H_3O_2^-(aq)$ (conjugate base)

Sodium Acetate: $NaC_2H_3O_2 + H_2O\ (l) \rightarrow Na^+(aq) + C_2H_3O_2^-(aq)$ (Sodium salt of conjugate base)

If an strong acid is added to a buffer, the conjugate base neutralizes it, so very little $H_3O^+$ is produced.

Add Acid: $H_3O^+(aq) + C_2H_3O_2^-(aq) \rightarrow HC_2H_3O_2 + H_2O$

If a strong base added to a buffered solution, it is neutralized by the weak acid.

Add Base: $HC_2H_3O_2 + OH^-(aq) \rightarrow H_2O(aq) + C_2H_3O_2^-(aq)$

Either way, there is little change on the pH since there is little change in the hydrogen ion concentration. Buffers are widely used in pharmaceutical products to stabilize the active ingredients; they are also critical to living systems, resulting in precise pH control within the body. In your lab, you will create a buffer and add small amounts of strong acid and strong base to both a buffer and pure water to observe the buffering effects.

**Indicators**

While we can often determine a rough idea if something may be acidic or basic from its properties, a more precise means of measuring pH is often needed. The easiest way to determine pH is with an indicator. Compounds that undergo color changes when the pH of a solution containing the compound changes are called indicators (Table 1). Indicators are used to provide information about the degree of acidity of a solution being tested. The easiest way to determine if a substance is acidic or a basic is to use an indicator. Indicators are molecules that change color in an acid or a base. The most common acid/base indicator paper is called litmus paper, so a litmus test is the first test used to determine acidic or basic properties. If the litmus paper is red the solution is acidic, if it is blue it is basic.

The strength of an acid or base is measured in pH, which is the concentration of the hydrogen ion ($H^+$). A high pH indicates a strong base, while a low pH indicates a strong acid. A pH of ~7 indicates a neutral substance (like water). pH paper consists of strips of filter paper which have been soaked in an indicator. A drop of an unknown solution can be placed on the pH paper, and the resulting color compared to a chart. By matching the color of the paper to a color on the chart, the pH of the solution can be determined. In addition, an indicator may be in liquid form that is added to a solution to monitor pH changes during chemical reactions. In your lab, you will be investigating the pH of common household substances and monitoring buffer effects using a variety of indicators.

## Chemical Hazards

***<u>0.1M Hydrochloric Acid Solution</u>***

NFPA RATING: HEALTH: 0 FLAMMABILITY: 0 REACTIVITY: 0

**DERMAL EXPOSURE:**
Wash off with soap and plenty of water.
**EYE EXPOSURE:**
Flush with copious amounts of water for at least 15 minutes. Contact your TA immediately.

***<u>0.10M Sodium Hydroxide Solution</u>***

NFPA RATING: HEALTH: 0 FLAMMABILITY: 0 REACTIVITY: 1

**ORAL EXPOSURE**
If swallowed, wash out mouth with water provided person is conscious. Do not induce vomiting. Contact your TA immediately.
**DERMAL EXPOSURE**
In case of extensive skin contact, flush with copious amounts of water for at least 15 minutes. Remove contaminated clothing and shoes. Contact your TA immediately.
**EYE EXPOSURE**
In case of contact with eyes, flush with copious amounts of water for at least 15 minutes. Assure adequate flushing by separating the eyelids with fingers. Contact your TA immediately.

***<u>1.0M $HC_2H_3O_2$ Solution</u>***

NFPA RATING: HEALTH: 1 FLAMMABILITY: 0 REACTIVITY: 0

**INHALATION:**
Move person into fresh air immediately. Contact TA immediately.
**DERMAL EXPOSURE:**
Wash off with soap and plenty of water.
**EYE EXPOSURE:**
Flush with copious amounts of water for at least 15 minutes. Assure adequate flushing by separating the eyelids with fingers. Contact your TA immediately.

***<u>1.0M $NaC_2H_3O_2$ Solution</u>***

NFPA RATING: HEALTH: 1 FLAMMABILITY: 0 REACTIVITY: 1

**INHALATION:**
Move person into fresh air immediately. Contact TA immediately.
**DERMAL EXPOSURE:**
Wash off with soap and plenty of water.
**EYE EXPOSURE:**
Flush with copious amounts of water for at least 15 minutes. Assure adequate flushing by separating the eyelids with fingers. Contact your TA immediately.

**The commonly used solutions and solids used in Part 1 are considered to be nonhazardous.**

## Experimental Procedures

Your TA must sign your data page(s) before leaving lab or you will not receive credit for attending the lab.

**Part 1: Determining the pH of common substances**

In the hood, you will find many different common household substances. Test the pH of each of these substances in the following manner:

1. Add the substance to the well plate that corresponds to the label on each well.
   a. For liquids, place 5 drops of the solution to be tested in the well
   b. If the substance is a solid, crush a tiny amount first and add the powder to the well. Then add 5 drops of water and stir with your spatula.

2. Test each substance to see if the solution is acidic or basic by using litmus paper. Do the procedure for blue litmus paper first and then use the red litmus paper. You can use 1 piece of the paper for several spots. You are only transferring a tiny drop of liquid to the paper. Do not dip the paper into the well.
   a. Touch the tip of your spatula to the liquid in a well and then touch it to the indicating paper.
   b. Record the color and approximate pH in your chart. (Blue is basic, Red is acidic)
   c. Wipe off your spatula with a paper towel before going to the next well.
   d. Repeat for all substances.

3. Test each substance for its approximate pH range by using the Alkacid paper. Follow the procedure you used above with the litmus paper. Look on the vial of Alkacid paper for the pH range that best fits the color you see on your paper.

4. Determine an actual pH value for the substance by using pH Hydrion paper. Use this paper sparingly as it is very expensive. Follow the previous procedures to record a value for the pH from the pH Hydrion vial.

5. Dispose of solutions as directed by your TA.

**Part 2: Determining the pH of Strong and Weak Acids**

1. Use your 10 mL graduated cylinder to measure out 2.0 mLs of each of the following solutions to the corresponding wells on your well plate.
   a. A1: 0.1M HCl
   b. A2: 0.1M NaOH
   c. A3: 1.0M $HC_2H_3O_2$
   d. A4 1.0M $NaC_2H_3O_2$
   e. B2: distilled water
   f. B4: distilled water

2. Add exactly 1.0 mL of each of the following solutions to the corresponding wells on your well plate.
   a. B1: 1.0M $HC_2H_3O_2$ and 1.0M $NaC_2H_3O_2$ (1 mL each)
   b. B3: 1.0M $HC_2H_3O_2$ and 1.0M $NaC_2H_3O_2$ (1 mL each)

3. Add 1 drop of bromthymol blue indicator to each of the 8 wells.
4. Record the color of the solution for each well.
5. Use the chart posted on the wall to record the approximate pH of all solutions.

## Part 3: Buffer Effects

The liquid in wells B1 and B3 is a buffer solution consisting of equal amounts of acetic acid, $HC_2H_3O_2$, and its conjugate base, $NaC_2H_3O_2$. The water in wells B2 and B4 has no buffering ability. You will study the buffering effects of this solution by monitoring the color change as small amounts of strong acid and base are added.

### *Addition of Strong Acid*

1. Record the color and pH of the solution in B1.
2. Add 1 drop of the 0.10M HCl in well A1 to well B1.
3. Record the color of the solution and the approximate pH after adding the HCl.
4. Record the color and pH of the solution in B2.
5. Then add 1 drop of the 0.10MHCl in well A1 to well B2.
6. Record the color of the solution and the approximate pH.

### *Addition of Strong Base*

1. Record the color and pH of the solution in B3.
2. Add 1 drop of the 0.10M NaOH in well A2 to well B3.
3. Record the color of the solution and the approximate pH.
4. Record the color and pH of the solution in B4.
5. Add 1 drop of the 0.10M NaOH in well A2 to well B4.
6. Record the color of the solution and the approximate pH.

## Cleanup

1. Use a pipet or an eyedropper to transfer the solutions to a waste beaker.
2. Pour waste solution into waste container in hood.
3. Rinse all equipment, dry and return to lab drawer.

## Laboratory Data

Create the following data tables in your lab notebook before coming to class.
Record the data shown in your laboratory notebook during lab in pen.
Include the **signed** copy of this data when you turn in your lab report.
Include the correct number of significant figures for each measurement.

### Table 1: pH of Common Substances

| | Cell Code | Color Blue Litmus | Color Red Litmus | Acidic/Basic | Color Alkacid | Relative Acidity | Color pH Hydrion | pH |
|---|---|---|---|---|---|---|---|---|
| Pepsi | A1 | | | | | | | |
| Vinegar | A2 | | | | | | | |
| Glass cleaner | A3 | | | | | | | |
| Distilled water | A4 | | | | | | | |
| dish soap | A5 | | | | | | | |
| Household Ammonia | A6 | | | | | | | |
| Instant coffee | B1 | | | | | | | |
| Baking soda | B2 | | | | | | | |
| Antacid | B3 | | | | | | | |
| Aspirin | B4 | | | | | | | |
| Rocksalt | B5 | | | | | | | |
| Acetaminophen | B6 | | | | | | | |

### Table 2: pH of Strong and Weak Acids and Bases

| Substance | Cell Code | Color | pH |
|---|---|---|---|
| 0.1M HCl | A1 | | |
| 0.1M NaOH | A2 | | |
| 1M $HC_2H_3O_2$ | A3 | | |
| 1m $NaC_2H_3O_2$ | A4 | | |
| 1M $HC_2H_3O_2$ + 1M $NaC_2H_3O_2$ | B1 | | |
| Water | B2 | | |
| 1M $HC_2H_3O_2$ + 1M $NaC_2H_3O_2$ | B3 | | |
| Water | B4 | | |

### Table 3: Buffer Effects

| Substance | Cell Code | Color before addition of acid or base | PH before addition of acid or base | Color after addition of acid or base | pH after addition of acid or base |
|---|---|---|---|---|---|
| 1M $HC_2H_3O_2$+ 1M $NaC_2H_3O_2$ + HCl | B1 | | | | |
| Water + HCl | B2 | | | | |
| 1M $HC_2H_3O_2$+ 1M $NaC_2H_3O_2$ + NaOH | B3 | | | | |
| Water + NaOH | B4 | | | | |

## Calculations

Remember to show one example of each calculation either handwritten or typed on a separate sheet.
Include correct significant figures and units
Units can be included in column or row title a demonstrated in tables below
Brackets, [xx], can be used to designate concentration in molarity.

**Part 1: pH of Common Substances**

***Hydronium Ion Concentration from the pH of the solution***

Plug in the pH value that you recorded from your pH hydrion for the value of the pH shown below to find the hydronium ion concentration of the solution. Repeat for all of your solutions from part 1.

$$[H_3O^+] = 10^{-pH}$$

**Part 2: pH of Strong and Weak Acids and Bases**

***pH of the HCl Solution***

HCl is a strong acid and will fully dissociate in water, the concentration of the acid, 0.1M is the concentration of the hydronium ion. Use this value for the concentration of the hydronium ion in the equation below to fine the pH of the solution.

$$pH = -\log [H_3O^+]$$

***pOH of the the NaOH Solution***

NaOH is a strong base and will fully dissociate in water, the concentration of the base, 1.0M is the concentration of the hydroxide ion. Use this value for the concentration of the hydroxide ion in the equation below to fine the pOH of the solution.

$$pOH = -\log [OH^-]$$

***pH of the NaOH Solution***

Since pOH is not commonly used for expressing acidity, convert the pOH of the NaOH to pH using the following equation.

$$pOH + pH = 14$$

**Part 3: No calculations required**

## Laboratory Results

Create a table with all the results from parts 1 and 2 of the experiment. Be sure to have correct units and significant figures included

## Discussion Questions

**You must use your data to answer these questions.  If the answer you give does not reflect YOUR data and results, you will receive no credit for the question.**

1. Which has a higher concentration of $H^+$, an acid or a base? If you have a solution with a pH of 4.5, calculate the $H^+$ concentration of this solution.  Show your work.

2. Identify 2 of the common household materials that would be considered acids and give their pH values.  Identify 2 substances that were considered bases.  List the 4 items in the order of increasing acidity going from top to bottom. (basic to acidic)

3. How does a 10-fold increase in the concentration of $H^+$ affect the number on the pH scale? Briefly explain your answer.

4. Why would it be important for humans to have a balanced pH in blood. Briefly explain how a human being maintains a balanced pH in their blood?

5. Why was the pH of the 0.1M HCl lower than the more concentrated acetic acid solution?  Briefly explain.