EXPERIMENT 11 – Acids, Bases, and pH

INTRODUCTION

The concept of **acidity** and **alkalinity** dates from ancient times. The word *acid* is derived from the Latin word *acidus*, meaning "sour." A common acid, acetic acid, is found in household vinegar (Latin *acetum* for "vinegar"). The word *alkaline* comes from the Arabic *Al-Qaly*, meaning "the calcined ashes." An example of calcination, the process of heating an ore or other solid that brings about its decomposition, is the calcination of calcium carbonate, forming calcium oxide and carbon dioxide gas. Calcium oxide and related oxides are well-known to react with water to form *hydroxides* such as calcium hydroxide, $Ca(OH)_2$. These solutions have a bitter taste and slippery feel. Thus for many years, substances have been classified as **acidic** or **alkaline** based simply on their taste and feel. Those that tasted sour were called acids and solutions that tasted bitter and were slippery to the touch were called alkalis. Lemon juice and vinegar were identified as acids, and solutions such as lye or "caustic soda" (NaOH) were identified as alkalis. Toward the end of the 19th century, the term "**base**" began to be used to label alkaline substances that are, in a chemical sense, the opposite of acids. A reaction between an acid and a base is known as a "**neutralization**" reaction.

A number of formal definitions of acids and bases have been proposed over the years. Depending upon the application, one or more of these definitions are applicable and useful. The three most common acid-base definitions are the **Arrhenius** definition (1887), the **Bronsted-Lowry** definition (1923), and the **Lewis** definition (1923).

In 1887, Svante August Arrhenius, a Swedish scientist, proved that dissolved salts and acids are *ionized*, existing as positive and negative ions in solution. In the case of acids, this discovery introduced the concept of the **hydrogen ion**, $\mathbf{H^+}$, as the positive ion produced by acids. **Arrhenius acids** are solutes that ionize in aqueous solution to produce hydrogen ions, H^+. Hydrogen ions can bond to water molecules in solution to form "**hydronium**" **ions**, $\mathbf{H_3O^+}$. Examples include hydrochloric acid and sulfuric acid.

Hydrochloric acid ionizes in aqueous solution in one step as follows:

$$HCl\,(aq) \longrightarrow H^+(aq) + Cl^-(aq) \quad \text{or} \quad HCl\,(aq) + H_2O\,(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

Sulfuric acid, a *diprotic* acid, ionizes in *two* steps;

1) $H_2SO_4\,(aq) \longrightarrow H^+(aq) + HSO_4^-(aq)$ (100% dissociation)

2) $HSO_4^-(aq) \rightleftharpoons H^+(aq) + SO_4^{2-}(aq)$ (partial dissociation at equilibrium)

Those solutes in aqueous solution that ionize or dissociate to produce hydroxide ions, OH^-, are called **Arrhenius bases**.

$$KOH\,(aq) \longrightarrow K^+(aq) + OH^-(aq)$$

Johannes Brønsted and **Thomas Lowry** expanded the definition of an *acid* to include all compounds that can be **proton (H^+ ion) donors** and the definition of a *base* to include all compounds that are **proton acceptors**. This definition is not restricted only to solutes based on their behavior in aqueous solution and

hence is applicable to a wider class of compounds than the Arrhenius definitions. In the Brønsted-Lowry theory, the ionization of an acid in water is viewed as a **proton transfer reaction** where the proton acceptor is a water molecule and hence is a Brønsted-Lowry base. The products that result from an acid-base reaction in this theory are called the "**conjugate acid**" of the base and the "**conjugate base**" of the acid. For example, consider again the ionization of hydrochloric acid in aqueous solution:

$$\underset{\text{acid}}{HCl\,(aq)} + \underset{\text{base}}{H_2O\,(l)} \longrightarrow \underset{\text{conjugate acid of } H_2O}{H_3O^+\,(aq)} + \underset{\text{conjugate base of HCl}}{Cl^-\,(aq)}$$

The conjugate acid is the species formed when a Brønsted-Lowry base accepts a proton from a Bronsted-Lowry acid. The conjugate base is the species that remains after the Brønsted-Lowry acid has lost a proton.

The **Lewis** definitions of acids and bases were proposed by Gilbert N. Lewis also in 1923 and are the broadest of the three theories. The Lewis theory defines an *acid* as an **electron-pair acceptor** and a *base* as an **electron-pair donor**. This definition includes reactions that contain neither hydrogen nor hydroxide ions.

$$\underset{\text{acid (e-pair acceptor)}}{BF_3} + \underset{\text{base (e-pair donor)}}{:NH_3} \longrightarrow F_3B{-}NH_3$$

In the context of the Arrhenius theory, water is termed as being neutral. The Bronsted-Lowry definition classifies water as an **amphoteric** species, one capable of behaving both as an acid and as a base. To understand this, consider the following dissociation of water, called the "*autoinization*" of water;

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

In this case, the hydronium ion, H_3O^+, forms when a proton, H^+, is transferred from one H_2O molecule to another. The other species that result from this process is the hydroxide ion, OH^-. Thus while one water molecule, the proton acceptor, functions as the base, the other plays the role of the acid. The ability of one water molecule to accept a proton from another water molecule or any acid is due to the two lone pairs of electrons on the oxygen atom of water molecule. This reaction results in equal molar amounts 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 determined from experiment to be 1.0×10^{-7} M each.

Neutral Solution: $[H^+]$ or $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$ M at 25°C

In terms of H^+ and OH^- concentrations, the solution is acidic if the hydrogen ion (or hydronium ion) concentration is greater than the hydroxide concentration, and the solution is basic if the hydrogen ion concentration (or hydronium ion) is less than the hydroxide ion concentration:

Acidic Solution: $[H^+] > [OH^-]$

Basic Solution: $[H^+] < [OH^-]$

Note that the brackets "[ ]" are shorthand for "concentration of" in molarity units (moles per liter).

THE pH SCALE OF ACIDITY

The concentration of hydrogen ion or hydronium ion in a solution is commonly expressed in terms of the **pH** of the solution, which is defined as the **negative base 10 logarithm of the H^+ concentration** or the **negative base 10 logarithm of the H_3O^+ concentration** in the solution.

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

Thus, the hydrogen ion concentration can be obtained from the pH of the solution as follows:

$$\mathbf{[H^+]} \quad \text{or} \quad \mathbf{[H_3O^+] = 10^{-pH}}$$

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

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

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

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

The pH and the pOH of any aqueous solution are related, as are the hydrogen and the hydroxide ion concentrations. The relevant equations are,

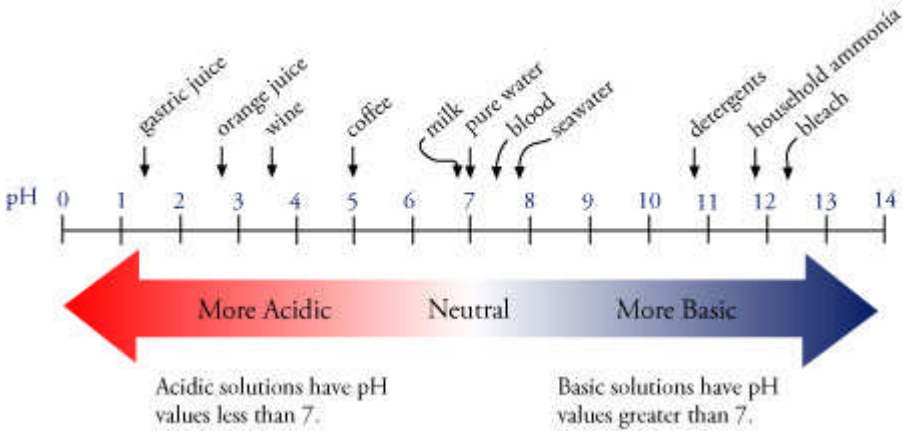
$$\mathbf{[H^+][OH^-] = 1.0 \times 10^{-14}} \quad \text{and} \quad \mathbf{pH + pOH = 14} \quad (\text{at } 25°C)$$

The pH of pure water at 25° is 7. This value is obtained by taking the negative logarithm of the concentration of H^+ ion, 1.0×10^{-7} M:

$$\mathbf{pH = -\log(1.0 \times 10^{-7}\ M) = -(-7.00) = 7.00}$$

A solution with a pH of **7** (at about 25°C, room temperature) is **neutral**. This provides a convenient "midpoint" pH that separates acidic solutions from basic solutions.

The pH values of some common substances are shown in the following diagram:



EXAMPLE 1. What is a) the pH and b) pOH of a solution with a hydrogen ion concentration of 3.50×10^{-5} M?

a) $pH = -\log [H^+] = -\log (3.50 \times 10^{-5}\ M) = \mathbf{4.46}$

b) $pOH = 14 - pH = 14 - 4.46 = \mathbf{9.54}$

EXAMPLE 2. Calculate the a) hydrogen ion and b) hydroxide ion concentrations of a solution that has a pH of 8.60.

a) $[H^+] = 10^{-pH} = 10^{-8.60} = \mathbf{2.51 \times 10^{-9}\ M}$

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

$$[OH^-] = \frac{1.0 \times 10^{-14}}{[H^+]} = \frac{1.0 \times 10^{-14}}{2.51 \times 10^{-9}} = \mathbf{3.98 \times 10^{-6}\ M}$$

USE OF INDICATORS TO MEASURE pH

Indicators, in chemistry, are natural or synthetic substances that change color in response to the nature of chemical environment. Litmus, for example, is a natural dye that turns red in acidic solutions and blue in basic solutions.

The pH of a solution can be estimated from the color that results when a few drops of different indicators mixed with the solution. For example, referring to Table 1 below, if a solution has a blue color with bromothymol blue, we know that the pH is 7.6 or higher. If the solution is colorless in the presence of phenolphthalein, we know that the pH is less than 8.2. Therefore, the pH of the solution must be between 7.8 and 8.2, or about 8.0 if we take the average value.

Special pH paper strips contain a mixture of different indicators whose colors combine to give a resultant color that allows one to estimate the pH of a solution without having to test the solution with the indicators separately.



TABLE 1. Indicators and their pH Ranges

Indicators	Acid Color	Base Color	pH Range of Color Change
Methyl Orange	Red	Yellow	3.1 – 4.4
Methyl Red	Red	Yellow	4.8 – 6.0
Bromothymol Blue	Yellow	Blue	6.2 – 7.6
Phenolphthalein	Colorless	Pink	8.2 – 10.0
Alizarin Yellow	Yellow	Red	10.1 – 12.0

In this experiment, you will observe the properties of acids and bases with suitable indicators. By the end of this experiment, you will be able to determine the pH of various solutions by observing the color of several indicators in these solutions and by using pH paper strips and a pH meter.

EXPERIMENTAL PROCEDURE

Part I. Determination of pH Using pH Paper

Place 5-10 drops of the solution to be tested in a spot plate. Wet a glass stirring rod with the solution, and touch the stirring rod on the pH paper. Compare the color to the color guide and record the approximate pH on the report form.

Part II. Determination of pH Using Indicators

Place 5-10 drops of each solution in the wells of a spot plate and add 1-2 drops of the first indicator to each. Record the colors on the report form and then repeat the process using the next indicator. After completing the table in the report form, refer to Table 1 to estimate the pH of each solution. Note: As you do this, it is convenient to "cheat" a little by referring to the pH values previously determined using the pH paper!

Part III. Determination of pH Using a pH Meter

Calibrate the pH meter using standard pH buffer solutions such as pH = 4.0, 7.0, and 10.0. Place about 10 mL of solution in a small beaker and determine the pH by immersing the pH probe into the solution. Record the pH reading on the report form, rinse the pH electrode using a wash bottle, and measure the remaining solutions in the same manner. Based on your pH readings, identify the substances as a **strong acid**, **weak acid**, **strong base**, **weak base**, or **neutral**. **Note:** Be careful not to bump the glass electrode on the rim of the beaker. The electrode is made of very thin glass at the bottom and is very fragile.

When finished, rinse the pH electrode well using the wash bottle and cap the electrode with a small amount of water in the cap to keep it wet during storage.



Desktop pH Meter



Pocket pH Meter

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REPORT FORM

Name _____

Instructor _____

Date _____

Part I. Determination of pH Using pH Paper

Type of pH paper used: _____

Solution	**Estimated pH**
1. Tap water	
2. 0.1 M HCl	
3. 0.1 M $HC_2H_3O_2$	
4. 0.1 M NaOH	
5. 0.1 M NH_3	
6. 0.1 M NaCl	
7. 0.1 M NH_4Cl	
8. 0.1 M $NaC_2H_3O_2$	

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Part II. Determination of pH Using Indicators

Solution	Methyl Orange	Methyl Red	Bromothymol Blue	Phenolph-thalein	Alizarin Yellow	Estimated pH
1. Tap water						
2. 0.1 M HCl						
3. 0.1 M $HC_2H_3O_2$						
4. 0.1 M NaOH						
5. 0.1 M NH_3						
6. 0.1 M NaCl						
7. 0.1 M NH_4Cl						
8. 0.1 M $NaC_2H_3O_2$						

Part III. Determination of pH Using a pH Meter

Solution	pH Reading	Strong / Weak Acid or Base or Neutral?
1. Tap water		
2. 0.1 M HCl		
3. 0.1 M $HC_2H_3O_2$		
4. 0.1 M NaOH		
5. 0.1 M NH_3		
6. 0.1 M NaCl		
7. 0.1 M NH_4Cl		
8. 0.1 M $NaC_2H_3O_2$		

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Pre-Laboratory Questions and Exercises

Due before lab begins. Answer in the space provided.

1. Provide definitions for the following terms:

a) Arrhenius acid _____

b) Arrhenius base _____

c) Brønsted-Lowry acid _____

d) Brønsted-Lowry base _____

2. A rainwater sample has a pH of 4.25.

a) What is the hydrogen ion concentration? _____

b) What is the hydroxide ion concentration? _____

c) What is the pOH of the rainwater? _____

3. A solution gives the following indicator colors. What is the approximate pH of the solution?

Methyl Orange: Yellow

Methyl Red: Yellow

Bromothymol Blue: Yellow

Phenolphthalein: Colorless

Alizarin Yellow: Yellow

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Post-Laboratory Questions and Exercises

Due after completing the lab. Answer in the space provided.

1. Which method that you used for determining pH was the least accurate? _____

Why was this so?

2. The hydrogen ion concentration in a a solution is measured to be 7.84×10^{-5} M. Calculate the following:

a) pH _____

b) pOH _____

c) $[OH^-]$, M _____

3. Fill in the blanks in the following table:

pH	**$[H^+]$, M**	**$[OH^-]$, M**	**pOH**
4.70	_____	_____	_____
_____	7.70×10^{-3} M	_____	_____
_____	_____	2.30×10^{-8} M	_____
_____	_____	_____	3.60