Goals

- Measure the pH of several solutions using a natural indicator, an indicator paper, and a pH meter.
- Calculate the pH of a solution from the $[H^+]$ or $[OH^-]$.
- Observe changes in the pH when acids or bases are added to a buffer solution.

Materials

- Red cabbage leaves
- Knife
- 400 mL beaker
- 10 mL graduated cylinder
- Hot plate
- Beaker clamp
- Wire gauze
- Stirring rod
- Test tubes
- Test tube rack
- Set of buffer solutions with pH range from 1 to 13
- pH meter with small beaker for rinsing electrode
- Calibration buffers for pH meter
- Wash bottle with deionized water
- Samples to test for pH: colorless or lightly colored juices and beverages, shampoo, conditioner, mouthwash, detergents, liquid soap, vinegar, household cleaners, etc.
- Small beakers or vials for test samples above
- Universal pH paper (e.g., pHydrion®) and color chart
- High pH buffer (9 – 11)
- Low pH buffer (3 – 4)
- 0.1 M NaCl with dropper
- 0.1 M HCl with dropper
- 0.1 M NaOH with dropper
- Optional: colored pencils

Discussion

Acids and bases are encountered frequently in everyday life. Foods, cleaning products, and biological fluids are examples of substances that may be categorized as acids or bases. Several different methods are available to determine whether a substance is acidic or basic. In this experiment, a natural indicator extracted from red cabbage, a commercially available universal indicator paper, and an electronic pH meter will be used to test for acids and bases.

There are several definitions of acids and bases, some are narrow while others are broader. One of the narrowest definitions was suggested by Svante Arrhenius. *Acids* dissociate when dissolved in water and produce hydrogen ions (H^+). *Bases* dissociate when dissolved in water and produce hydroxide ions (OH^-). For example, hydrogen chloride (HCl) is an Arrhenius acid, while sodium hydroxide (NaOH) is an Arrhenius base, as shown below:

$$HCl_{(g)} \xrightarrow{H_2O} H^+_{(aq)} + Cl^-_{(aq)}$$

Acid

$$NaOH_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + OH^-_{(aq)}$$

Base

The Brønsted-Lowry theory is a more general definition of acids and bases. *Acids* are hydrogen-containing compounds that can donate protons (H^+) to another substance. *Bases* are substances that accept a proton (H^+). In an *aqueous* (water) solution, nitric acid (HNO_3) donates a hydrogen ion that reacts with a water molecule to form a *hydronium ion* (H_3O^+). Ammonia (NH_3), a base, receives a hydrogen ion from a water molecule to form a hydroxide ion (OH^-). The balanced equations are:

$$HNO_3(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + NO_3^-(aq)$$

Acid

$$NH_3(aq) + H_2O(l) \longrightarrow NH_4^+(aq) + OH^-(aq)$$

Base

Acids may also be defined as substances that increase the hydronium ion (H_3O^+) concentration, while bases are substances that increase the hydroxide ion (OH^-) concentration. Water also produces hydronium ions and hydroxide ions in a process called *self-ionization*. The balanced equation showing the reaction of the very few water molecules undergoing this reaction is:

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

It is not surprising that water is *amphiprotic* (can act as an acid or a base). In the reaction with nitric acid above, water acts as the base and receives a proton from the acid. In the reaction with ammonia, water acts as the acid and donates a proton to the base.

In pure water, the concentration of the hydronium ions equals the concentration of the hydroxide ions. Their concentrations can be measured experimentally:

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

The product of these two concentrations is a constant and it is called the *ion product constant for water*. The equation for the ion product constant for water is:

$$K_w = [H_3O^+] \times [OH^-] = \left(1.0 \times 10^{-7}\right)\left(1.0 \times 10^{-7}\right) = 1.0 \times 10^{-14}$$

The term H_3O^+ is frequently abbreviated as H^+. Thus, the equation above also may be written as $K_w = [H^+] \times [OH^-]$.

Since the numerical value for K_w must always equal 1.0×10^{-14}, any addition of hydronium ions to the water (in the form of an acid) will increase the $[H_3O^+]$ and decrease the $[OH^-]$. The addition of a base will cause a decrease in the $[H_3O^+]$ and an increase in the $[OH^-]$. If the concentration of either the $[H_3O^+]$ or the $[OH^-]$ is known, the other concentration can be calculated by rearranging the K_w equation.

Example 1:

If a solution has $[OH^-] = 1.0 \times 10^{-5}$ M, what is the $[H_3O^+]$?

$$[H_3O^+][OH^-] = 1.0 \times 10^{-14}$$ Rearrange this equation…

$$[H_3O^+] = \frac{1.0 \times 10^{-14}}{[OH^-]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-5}} = 1.0 \times 10^{-9}\ \text{M}$$

Example 2:

Sufficient acidic solute is added to a quantity of water to produce a solution with $[H_3O^+] = 4.0 \times 10^{-8}$ M. What is the $[OH^-]$ in this solution?

$$[H_3O^+][OH^-] = 1.0 \times 10^{-14}$$ Rearrange this equation…

$$[OH^-] = \frac{1.0 \times 10^{-14}}{[H_3O^+]} = \frac{1.0 \times 10^{-14}}{4.0 \times 10^{-8}} = 2.5 \times 10^{-7}\ \text{M}$$

<u>pH of solutions</u>

The pH of a solution is a logarithmic measure of its hydronium ion concentration (see Figure 1). The equation to calculate pH is:

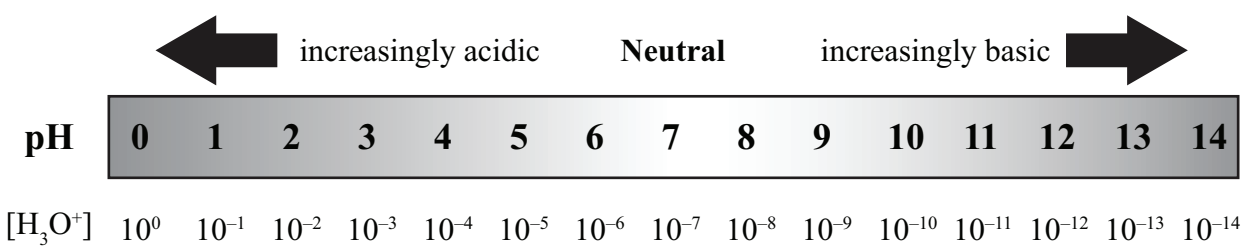
$$pH = -\log[H_3O^+] \quad \text{and may also be written as} \quad pH = -\log[H^+]$$

The rule for the number of significant figures in a logarithm is: the number of digits to the right of the decimal point in a logarithm is equal to the number of significant figures in the original number.

$$[H_3O^+] = \underbrace{4.5}_{\text{Two significant figures}} \times 10^{-3}\ \text{M}$$

$$pH = \underbrace{2.35}_{\text{Two digits}}$$

Figure 1 ***Graphical representation of the pH scale***



It is important to realize that a difference of 1 pH unit represents a ten-fold difference in the hydronium ion concentration.

Measuring pH

For accurate numerical values of pH, an electronic pH meter should be used. A pH meter should also be used to test colored solutions, since it may be difficult to see the color change of an indicator compound. A pH meter is an electronic device with a special electrode that generates a small voltage proportional to the $[H_3O^+]$ of the solution. The electrical signal is converted to pH and displayed on a meter or a digital readout. Be sure to use the correct number of significant figures when reading a pH meter. The electrode is very fragile and expensive. It must be cleaned after each use. When not in use, it is submerged in a buffer solution to keep the tip moist.

Another common way to measure pH is to use an *indicator*. An indicator is a complex organic compound that changes color at a certain pH. Many natural substances contain compounds that act as indicators. For example, the hydrangea bush flower is blue in acidic soil and pink/red in basic soils. Red cabbage changes colors at several different pH's and may be used as a wide range indicator.

pH indicator papers such as litmus paper or pHydrion® paper have been impregnated with indicators. These papers are easy to store and simple to use. They change color and the color shown by the paper is then compared to a color chart provided by the manufacturer. Dipping the indicator paper into a solution will contaminate the solution. It is preferable to use a clean, dry stirring rod to transfer a drop of the solution to the paper strip.

Buffers

There are a number of different systems in the body and they all have specific pH ranges. For example, the stomach is acidic, due to the production of hydrochloric acid. The small intestine is usually slightly basic. Urine is slightly acidic in the morning (pH = 6.5–7.0), generally becoming more alkaline (basic) by the evening (pH = 7.5–8.0).

The bloodstream is the most sensitive to pH changes and it must be kept in a narrow range from 7.35 to 7.45. Small departures from this range can cause serious illness, and death can result from pH changes of only a few tenths of a pH unit. To protect against pH changes, the bloodstream contains three different buffer systems to keep the pH of blood in the proper range.

Buffers minimize changes to the pH of a solution when small amounts of excess acid or base are added to the solution. This does not mean that the pH is neutral. Rather, the buffer stabilizes the solution at a <u>certain</u> pH. That pH depends upon the buffer system chosen. A buffer may consist of a weak acid and its salt such as acetic acid and an acetate salt. A buffer may also consist of a weak base and its salt such as ammonium hydroxide and an ammonium salt.

The principal buffer in the bloodstream is carbonic acid/bicarbonate ion (H_2CO_3/HCO_3^-). Carbon dioxide (CO_2) is produced in the body as a byproduct of metabolic reactions. The CO_2 dissolves in water to form carbonic acid (H_2CO_3), which is a weak acid. The carbonic acid, in turn, dissociates to produce the bicarbonate ion (HCO_3^-). The two equilibria are shown below:

$$CO_{2(g)} + H_2O_{(l)} \rightleftharpoons H_2CO_{3(aq)} \underset{}{\overset{H_2O_{(l)}}{\rightleftharpoons}} HCO_3^-{}_{(aq)} + H_3O^+{}_{(aq)}$$

carbonic acid — bicarbonate ion

principal buffer in blood

When small amounts of acids or bases enter the bloodstream, the buffer reacts with them to counter their impact on blood pH. The following equations show the reaction when excess acid or base is added to the buffer solution:

Excess base: $H_2CO_3(aq) + OH^-(aq) \rightleftharpoons HCO_3^-(aq) + H_2O(l)$

Excess acid: $HCO_3^-(aq) + H_3O^+(aq) \rightleftharpoons H_2CO_3(aq) + H_2O(l)$

Excess base reacts with the carbonic acid component of the buffer and excess acid reacts with the bicarbonate ion component of the buffer. Notice that no new reactants or products are formed. All of the substances, H_2CO_3, HCO_3^-, and H_2O, were present in the original buffer solution. But the excess acid or base has been "removed" from the solution, keeping the pH relatively unchanged.

Buffer solutions do not have the ability to neutralize unlimited quantities of additional acid or base. The amount of $[H_3O^+]$ or $[OH^-]$ that can be neutralized is called the *buffering capacity*.

If the pH of the bloodstream is too low (more acidic), the condition is called *acidosis*. If the pH of the bloodstream is too high (more basic), the condition is called *alkalosis*. There are two general types of acidosis and alkalosis, one type resulting from changes in metabolic processes and the other type resulting from changes in respiratory processes. Both types can be understood in terms of Le Châtelier's principle and equilibrium shifts.

Experimental Procedures

Eye protection and appropriate clothing must be worn at all times.

Many of these chemicals are corrosive to the skin and toxic if ingested. Wash any contacted skin area immediately with running water and alert the instructor.

Discard all wastes properly as directed by the instructor. Dispose of indicator papers and cabbage in the trash containers, NOT in the sinks.

A. Red cabbage indicator

Optional: ***The pH reference sets may be prepared by the instructor before the lab meeting or students may prepare them.***

If prepared by the instructor, do not do Procedure A. Instead, continue to Procedure B. Two or more "pH reference sets" should be available for the class.

If prepared by the students, the instructor will determine how many students should work together to prepare their "pH reference set."

1. Chop several leaves of red cabbage and place in a 400 mL beaker. Fill the beaker half full with deionized water to cover the leaves. Heat gently on a hot plate until the solution turns dark purple. Turn off the burner. Use a beaker clamp to set the beaker on a wire gauze to cool.
2. Place 13 test tubes in a test tube rack. Label the test tubes #1 through #13.
3. Pour 3–4 mL of the buffer solution labeled "pH = 1" into the first test tube. Repeat this step with the remaining buffer solutions and their corresponding numbered test tubes.
   Add 2–3 mL of the *cooled* red cabbage indicator to each test tube. Stir each solution with a clean, dry stirring rod.

B. Colors of the pH reference set

1. Record the colors of the pH reference set (use words and color the wedges). It is helpful to hold a piece of white paper behind the test tubes to observe the colors.
2. Do not dispose of the reference set until the entire lab is completed.

C. Measuring pH

Universal pH paper (pHydrion®)

1. Pour 3–4 mL of each substance to be tested into clean, dry test tubes. List these substances in **alphabetical order** in Table B of the report sheet.
2. Obtain a pH paper color chart and approximately 15 cm of the universal pH paper. Place this strip of paper on a clean dry paper towel.
3. Dip a clean, dry stirring rod into the first test tube and place a drop of the substance close to one end of the strip of pH paper. Compare the resulting color with the color on the color chart. Record the pH. Save the test tube and solution for the next section of the experiment.
4. Continue to use the same strip of pH paper. Repeat the procedure for each of the household substances provided, spacing each drop approximately 1 cm apart. Save all of the test tubes and solutions for the next section of the experiment.

Natural indicator

5. To each of the test tubes prepared in the universal pH paper section, add 2–3 mL of the red cabbage solution.
6. Mix each of the solutions with a clean dry, stirring rod. Compare the resulting color of each of the solutions with the reference set prepared in Procedure A. Record the color and pH for each solution.

pH meter

7. The lab instructor will demonstrate how to calibrate and use the pH meter.
8. Select four substances tested in the previous sections. Choose the substances with the highest pH, the lowest pH, the pH closest to neutral, and one other substance of your choice.
9. Obtain new samples of each substance **without** the cabbage juice, approximately 10 mL of each one. Test the pH of each substance by pouring it into a small beaker or vial that will accommodate the electrode of the pH meter.
10. Measure the pH of the four substances using the pH meter. Use a squeeze bottle and a waste beaker to rinse the electrode with deionized water between each measurement.

D. Buffer solutions

Addition of an acid to a buffer solution

1. Obtain four test tubes and label them #1– #4.
2. Using a 10 mL graduated cylinder, measure approximately 10 mL of each of the following solutions into a different test tube. Clean and dry the graduated cylinder between solutions.

   #1 deionized water
   #2 0.1 M NaCl
   #3 a high pH buffer solution
   #4 a low pH buffer solution

3. Add 2–3 mL of cabbage indicator to **each** test tube. Mix well.
4. For each test tube, determine the pH of the solution by comparing the color with the reference set prepared in Procedure A. Record the color and pH for each test tube.
5. Add 5 drops of 0.1 M HCl to **each** test tube. Stir and determine the pH. Record.
6. Add another 5 drops of 0.1 M HCl to **each** test tube. Stir and determine the pH. Record.
7. Determine the pH change for each test tube (pH after 10 drops minus original pH). Record.
8. Indicate whether the substance in each test tube is a buffer.

<u>Addition of a base to a buffer solution</u>

9. Prepare four new test tubes and follow the steps in Procedure C, using 0.1 M NaOH instead of HCl in steps 5 and 6.

Name _____ Date _____

Partner(s) _____ Section _____

_____ Instructor _____

1. Describe the correct procedure to test solutions with indicator paper.

2. What is the purpose of a buffer?

3. What should you do if you accidentally spill an acid or a base on your skin?

4. Complete the following table:

$[H_3O^+]$	$[OH^-]$	pH	Acidic, basic, or neutral?
		8.00	
2.5×10^{-2} M			
			neutral
	1.0×10^{-9} M		

Name _____ Date _____

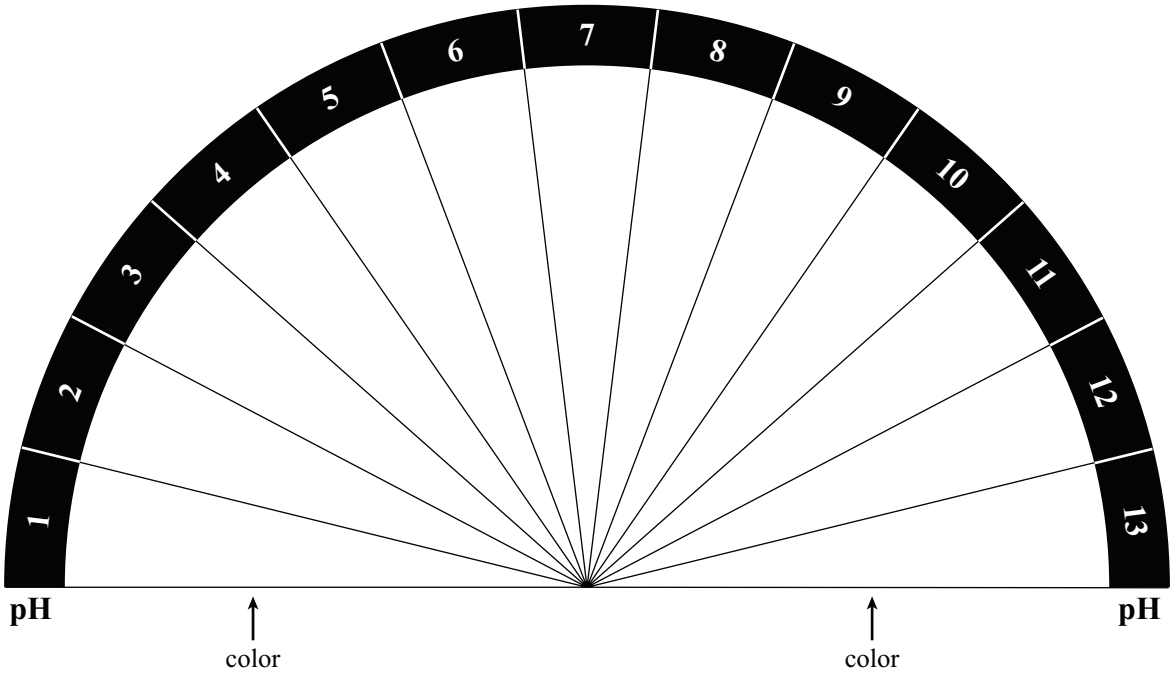
Partner(s) _____ Section _____

_____ Instructor _____

A. Red cabbage indicator

There is no data to record for this part.

B. Colors of the Cabbage pH reference set



C. Measuring pH

Substance	pH paper color	pH paper pH	cabbage color	cabbage pH	pH meter pH	Acidic, basic, or neutral?
tap water						
deionized water						

D. Buffer solutions

Addition of an acid to a buffer solution

Substance	pH of original solution	pH after 5 drops of HCl	pH after 10 drops of HCl	pH change	Buffer solution?
deionized water					
0.1 M NaCl					
high pH buffer					
low pH buffer					

Addition of a base to a buffer solution

Substance	pH of original solution	pH after 5 drops of NaOH	pH after 10 drops of NaOH	pH change	Buffer solution?
deionized water					
0.1 M NaCl					
high pH buffer					
low pH buffer					

Each table above used the same four solutions.

Of the four solutions, which solutions (plural) had the greatest change in pH? Why?

Of the four solutions, which solutions (plural) had the smallest change in pH? Why?

Problems and questions

1. Compare and contrast the three methods used to test pH (red cabbage indicator, pH paper, and a pH meter). Did they give similar results? Which is the most accurate? Explain.

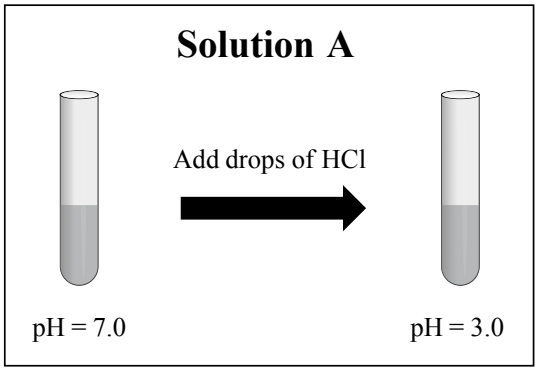
2. The hydroxide ion concentration of a solution is 2.5×10^{-8} M.

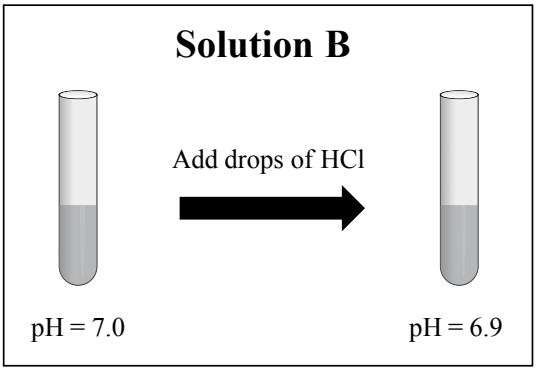
   a) What is the hydronium ion concentration?

   b) What is the pH?

   c) Is the solution acidic, basic, or neutral?

3. A few drops of HCl are added to two solutions, A and B. The pH of each solution is shown before and after the addition of HCl. Which solution is a buffer? Explain.





4. How many times more acidic is a solution with pH 2 relative to pH 4?

5. Circle the two substances that, when dissolved in water and mixed together, would make a buffer pair:

   HF K_2CO_3 HCl Na_2CO_3 NaF NaCl