11

Many flowers, fruits, and vegetables contain organic compounds that change color with pH.

USING VEGETABLE INDICATORS to DETERMINE pH

Hydrangeas will be blue or pink, depending on the acidity of the soil.

Acids and bases were originally classified by their physical properties such as taste: acids tasting sour and bases tasting bitter. It was also noticed that many natural substances contained pigments that changed color when exposed to acids and bases. One of the earliest methods for determining the pH of a solution was to use chemical compounds that are derived from plants and change color with the pH of the solution. These substances are called **indicators.** One such indicator, litmus, will turn blue in base and red in acid.

In the 1920s a more systematic definition of acids and bases was proposed by Brønsted and Lowry. They defined an acid as a substance that is capable of donating a proton; conversely, a base is a substance that accepts protons. An example of this is shown below:

$$HCl + H_2O \rightarrow H_3O^+ + Cl^-$$

In this example, hydrochloric acid, HCl, is the acid and donates a proton (H^+) to the base, H_2O, water. Hydrochloric acid is defined as a *strong* acid; in other words, it is an excellent proton donor. When hydrochloric acid is mixed with water, the above reaction essentially goes to completion, with all the HCl being converted to H_3O^+, hydronium ion, and Cl^-, chloride ion.

Acetic acid (which is found in vinegar) is said to be a *weak* acid:

$$CH_3COOH + H_2O \leftrightarrows CH3COO^- + H_3O^+$$

pH = $-\log[H_3O^+]$—the brackets indicate that the amount of hydronium ion is stated in molarity (*M*).

A weak acid is not a good proton donor, so when acetic acid is mixed with water, few hydronium ions are formed. If equal molar amounts of acetic acid and hydrochloric acid are put into separate containers of water, the solution of acetic acid has a low concentration of hydronium ions, whereas the solution of hydrochloric acid has a high concentration of hydronium ions.

To measure the strength of an acidic solution, we measure the concentration of H_3O^+. The scale used to denote the acidity of a solution is **pH.** The pH of a solution is equal to $-\log[H_3O^+]$. Because this log is negative, a high concentration of H_3O^+ corresponds to a low pH. Thus, a 0.1 molar (0.1 *M*) solution of HCl, a strong acid, has a pH of 1 ($-\log 1 \times 10^{-1}$). A 0.1 *M* solution of acetic acid, a weak acid, has a pH of 3, ($-\log 1 \times 10^{-3}$). This pH is higher because, of the 0.1 moles of acetic acid present in each liter of solution, only 0.001 moles of the total amount has reacted to form hydronium ion. This means in there is a significant quation of unreacted (undissociated) acetic acid in the solution.

In pure water, there is a slight dissociation of the water molecules into hydronium ions and hydroxide ions:

$$2\,H_2O \leftrightarrows H_3O^+ + OH^-$$

In pure water, the concentration of hydronium ions is 1×10^{-7} *M*, and because every time a H_3O^+ is formed an OH^- is formed, the concentration of OH^- is also 1×10^{-7} *M*. Pure water has a pH of 7 ($-\log 1 \times 10^{-7}$). Because the concentration of $OH^- = H_3O^+$, pure water is neither acidic nor basic but is considered neutral.

Most familiar bases are ionic compounds that contain a hydroxide ion, such as sodium hydroxide (NaOH) or potassium hydroxide (KOH). These bases accept protons from acids to form water:

$$H_3O^+ + OH^- \leftrightarrows 2\,H_2O$$

Again, a strong base such as NaOH will completely dissociate in water to give a high concentration of hydroxide ions. Some bases do not contain hydroxide ions themselves but form hydroxide ions when in aqueous solution through a reaction with water. Ammonia (NH_3) is an example of such a base:

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

The pH system is also used to measure the strength of a basic solution. How can this be done when pH is determined by the concentration of hydronium ions but a base usually involves hydroxide ions? Remember that the concentrations of hydronium ion and hydroxide ion in pure water are both equal to 1×10^{-7} *M*. The product of the concentration of hydronium ion and hydroxide ion is called the ion product of water, K_w.

$$K_w = [H_3O^+][OH^-] = [1 \times 10^{-7}][1 \times 10^{-7}] = 1 \times 10^{-14}$$

The ion product of water is a constant (at constant temperature) and is the same for all aqueous solutions. Because this parameter is constant, when the concentration of hydronium ion increases, the concentration of hydroxide ion must decrease and *vice versa*. When a base is present, the concentration of hydroxide ion increases and the concentration of hydronium ion decreases. Because the pH is $-\log[H_3O^+]$, when the concentration of hydronium ion decreases, the pH increases.

Just as there are strong and weak acids, there are also strong and weak bases. In a 0.1 *M* solution of NaOH, the *strong* base is completely dissociated; therefore, the $[OH^-] = 0.1 = 10^{-1}$. Since the product of $[OH^-]$ and $[H_3O^+]$ must be 10^{-14} for all aqueous solutions, $[H_3O^+]$ for this basic solution must equal $10^{-14}/10^{-1}$, or 10^{-13}. The pH of a 0.1 *M* solution of NaOH is thus $-\log[10^{-13}] = 13$. In a 0.1 *M* solution of ammonia (NH_3), a weak base, only about 1/100 of the ammonium hydroxide is dissociated at any point in time so $[OH^-] = 0.1/100 = 10^{-3}$. The pH of a 0.1 *M* ammonia solution is $-\log[10^{-14}/10^{-3}] = -\log[10^{-11}] = 11$.

How does one determine the pH of a solution? One can use plant indicators, like those previous mentioned. Some indicators have a wide range of color changes. In this experiment, you will extract colored substances from red cabbage and use them as an indicator. First, you will determine the color the indicator will be at a specific pH, using buffer solutions of known pH provided for you in the lab. Then, you will take various household substances and common laboratory solutions and determine the pH of their aqueous solutions.

MATERIALS

Equipment

100-mL graduated cylinder
500-mL beaker
test tubes and rack
eyedropper
grease pencil
hotplate
mortar and pestle

Chemicals

red cabbage
buffer solutions
various household chemicals
0.1 *M* HCl
0.1 *M* acetic acid
0.1 *M* ammonium hydroxide

CAUTION:
In case of spills, wash your skin thoroughly with water and clean up the lab bench.
Use caution with the *household chemicals*. Familiarity tends to breed carelessness.

PROCEDURE

1. Place several purple cabbage leaves in a 500-mL beaker and cover the leaves with water. Boil the cabbage leaves to remove the pigment.
2. While the indicator cools, set up an array of buffer solutions of different pHs. Label ten small test tubes to correspond with the pHs of the buffer solutions (pHs 2 to 11). Fill each tube approximately one-half full (~5 mL) with the appropriate buffer from the stock solution bottle.
3. Arrange the tubes in order of increasing pH values in your test-tube rack. Add several drops of the cabbage indicator to each tube with an eyedropper. You should get a nice array of colors from the vegetable dye. Record the color of each pH. It's fine to add more of the cabbage indicator to intensify the color but be sure to add the same amount to each test tube.
4. Test the pH of the household products available in the lab by placing a 5-mL sample in a test tube and adding cabbage indicator. Solid samples such as vitamin C or antacid tablets should be prepared by crushing about one-quarter of the tablet and dissolving it in ~5 mL of deionized water. Compare the resulting colors with your buffer array to determine an approximate pH.
5. With the cabbage indicator, test the pH of several laboratory reagents.

Name: ______________________________

DATA AND RESULTS

1. In the table below describe the indicator color of the buffers tested.

Buffer pH	Color
2	
3	
4	
5	
6	
7	
8	
9	
10	
11	

2. List the household products and chemicals tested and the pH of the solution.

Solutions tested **pH**

QUESTIONS

1. List the household products in order of increasing pH. Does this order indicate increasing acidity or increasing basicity?

2. Calculate the concentration of H_3O^+ in each of the household product solutions.

3. What effect would the initial color of a tested solution have on your determination of the pH? For example, what effect would the purple color of grape juice have on your determination?