8.6: Acid-Base Indicators

pH indicators are weak acids that exist as natural dyes and indicate the concentration of H^+ (H_3O^+) ions in a solution via color change. A pH value is determined from the negative logarithm of this concentration and is used to indicate the acidic, basic, or neutral character of the substance you are testing.

Introduction

pH indicators exist as liquid dyes and dye-infused paper strips. They are added to various solutions to determine the pH values of those solutions. Whereas the liquid form of pH indicators is usually added directly to solutions, the paper form is dipped into solutions and then removed for comparison against a color/pH key.

pH	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**
Color								

Very Acidic Acidic Neutral Basic Very Basic

See Figure 1 and 2 to see a color range (1) of a universal indicator (2).

The Implications of the Indicated pH via the Equation

Recall that the value of pH is related to the concentration of H^+ (H_3O^+) of a substance. pH itself is approximated as the cologarithm or negative logarithm of the H^+ ion concentration (Figure 3).

$$pH \approx -log[H_3O^+] \tag{3}$$

A pH of 7 indicates a neutral solution like water. A pH less than 7 indicates an acidic solution and a pH greater than 7 indicates a basic solution. Ultimately, the pH value indicates how much H^+ has dissociated from molecules within a solution. The lower the pH value, the higher concentration of H^+ ions in the solution and the stronger the acid. Likewise, the higher the pH value, the lower the concentration of H^+ ions in the solution and the weaker the acid.

How the Color Change of the Indicator Happens

The color change of a pH indicator is caused by the dissociation of the H^+ ion from the indicator itself. Recall that pH indicators are not only natural dyes but also weak acids. The dissociation of the weak acid indicator causes the solution to change color. The equation for the dissociation of the H^+ ion of the pH indicator is show below (Figure 4).

$$HIn + H_2O \rightleftharpoons H_3O^+ + In^- \tag{4}$$

with

- HIn is the acidic pH indicator and
- In^- is the conjugate base of the pH indicator

It is important here to note that the equation expressed in figure 4 is in equilibrium, meaning Le Chatelier's principle applies to it. Thus, as the concentration of H_3O^+ (H^+) increases or decreases, the equilibrium shifts to the left or right accordingly. An **increase** in the HIn acid concentration causes the equilibrium to shift to the **right** (towards products), whereas an **increase** of the In^- base concentration causes the equilibrium to shift to the **left** (towards reactants).

pH Ranges of pH Indicators

pH indicators are specific to the range of pH values one wishes to observe. For example, common indicators such as phenolphthalein, methyl red, and bromothymol blue are used to indicate pH ranges of about 8 to 10, 4.5 to 6, and 6 to 7.5 accordingly. On these ranges, phenolphthalein goes from colorless to pink, methyl red goes from red to yellow, and bromothymol blue goes from yellow to blue. For universal indicators, however, the pH range is much broader and the number of color changes is much greater. See figures 1 and 2 in the introduction for visual representations. Usually, universal pH indicators are in the paper strip form.

Graphing pH vs. the H^+ (H_3O^+) Concentration

It is important to note that the pH scale is a **logarithmic scale**: hence an increase of 1 pH unit corresponds to a ten times increase of H_3O^+. For example, a solution with a pH of 3 will have an H^+ (H_3O^+) concentration ten times greater than that of a solution with a pH of 4. As pH is the negative logarithm of the H^+ (H_3O^+) concentration of a foreign substance, the lower the pH value, the higher the concentration of H^+ (H_3O^+) ions and the stronger the acid. Additionally, the higher the pH value, the lower the H^+ (H_3O^+) concentration and the stronger the base.

Indicators in Nature

pH indicators can be used in a variety of ways, including measuring the pH of farm soil, shampoos, fruit juices, and bodies of water. Additionally, pH indicators can be found in nature, so therefore their presence in plants and flowers can indicate the pH of the soil from which they grow.

Hydrangeas

Nature contains several natural pH indicators as well: for example, some flower petals (especially Roses and Hydrangeas), certain fruits (cherries, strawberries) and leaves can change color if the pH of the soil changes. See figure 7.

(7)

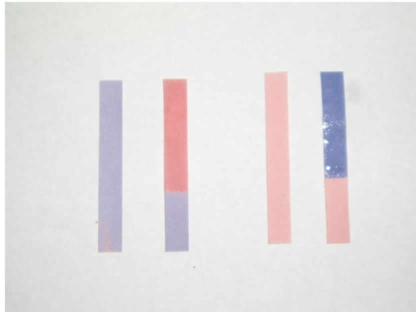
Lemon juice

In the lemon juice experiment, the pH paper turns from blue to vivid red, indicating the presence of H_3O^+ ions: lemon juice is acidic. Refer to the table of Universal Indicator Color change (figure 1 in the introduction) for clarification.

Cleaning Detergent

The household detergent contained a concentrated solution of sodium bicarbonate, commonly known as baking soda. As shown, the pH paper turns a dark blue: baking soda (in solution) is basic.Refer to the table of Universal Indicator Color change (figure 1 in the introduction) for clarification.

Here is a closer look of the pH papers before and after dipping them in the lemon juice and cleaning detergent (Figure 10):



neutral acidic neutral basic

Figure 10:

Cabbage Juices

Here is a simple demonstration that you could try in the lab or at home to get a better sense of how indicator paper works. Make sure to always wear safety glasses and gloves when performing an experiment!

Materials

- 1 cabbage
- cooking pot
- white paper coffee filters
- strainer
- water
- a bowl

Procedure

1. Peal the cabbage leaves and place them into the pot.
2. Add water into the pot, making sure the water covers the cabbage entirely.
3. Place the pot on the stove and allow to cook at medium heat for about 30 to 35 minutes.
4. Allow it to cool, then pour contents into the bowl using the strainer.
5. Soak your coffee filters in the cabbage juice for about 25 to 30 minutes.
6. Allow the filters to fully dry, then cut them into strips.
7. Now start your pH testing (starts out blue, changes to green [basic], and **red** [acidic]).

Practice Problems

1. A hair stylist walks into a store and wants to buy a shampoo with slightly acidic/neutral pH for her hair. She finds 5 brands that she really likes, but since she never took any introductory chemistry classes, she is unsure about which one to purchase. The first has a pH of 3.6, the second of 13, the third of 8.2, the fourth of 6.8 and the fifth of 9.7. Which one should she buy?

Answer: The brand that has a pH of 6.8 since it's under 7 (neutral) but very close to it, making it slightly acidic.

2. You decide to test the pH of your brand new swimming pool on your own. The instruction manual advises to keep it between 7.2-7.6. Shockingly, you realize it's set at 8.3! Horrified, you panic and are unsure whether you should add some basic or acidic chemicals in your pool (being mindful of the dose, of course. Those specific chemicals are included in the set, so no need to worry about which one you have to use and (eek!) if they are legal for public use). Which one should you add?

Answer: Since the goal is to lower the pH to its ideal value, we must add acidic solution to the pool.

3. Let's say the concentration of Hydronium ions in an aqueous solution is 0.033 mol/L. What is the corresponding pH of this solution, and based on your answer identify whether the solution is acidic, basic or neutral.

Answer: Using the formula $pH \approx -log[H_3O+]$

pH= -log[0.033]= **1.48** : The solution is highly acidic!

4. Now let's do the inverse: Say you have a solution with a pH of 9.4. What is the H_30^+ ions concentration?

Answer: $[H_3O^+]$= $10^{-9.4}$= **3.98E-10 mol/L**. Seem too low to be true? Think again, if the pH is >7, the solution will be basic, hence the hydronium ions will be low compared to the hydroxide (OH^- ions).

5. A more trickier one: 0.00026 moles of acetic acid are added to 2.5 L of water. What is the pH of the solution?

Answer: M=n/L : $M_{acetic\ acid}$= 0.00026/2.5 =**1.04E-4 mol/L**

pH= -log[1.04E-4]= **3.98**

Outside Links

- external link:http://www.epa.gov/acidrain/measure/ph.html
- external link:www.krampf.com/experiments/Science_Experiment16.html
- There are many common household products and garden plants that can be used or made into pH indicators. For more information on these common house hold indicators visit http://chemistry.about.com/cs/acidsandbases/a/aa060703a.htm
- commons.wikimedia.org/wiki/Ca...ndicator_paper

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