# THE THEORY OF pH

The pH scale was suggested by Sorensen almost a century ago. We can think of it as a sort of shorthand for specifying the concentration of protons (H+ ions) in (usually aqueous) solution. Sorensen defined pH = -log[H+]. (Remember that [ ] means concentration or molarity.) In water, pH values from 0 (very acidic) to 14 (very basic) are possible. Thus pH 7, being right in the middle of the range, is neutral, that is, neither acidic nor basic. There are two tricky things about this scale: it's nonlinear and it seems "to go the wrong way" (high H+ corresponds to low pH). Both of these odd features stem from the fact that pH involves logarithms. First of all, a negative logarithm indicates an inversion or reciprocal of a number; that is why the scale seems "to go the wrong way." Also because of logarithms, the numbers seem much smaller: $\log(10^7) = 7$, $\log(10^{-2}) = -2$, log(4.87x10) = 3.68. So a change the pH from 7 to 5, increases the proton concentration by 100 times!

Why does pH warrant its own rather weird scale? Mostly its due to the central role that water plays in chemistry, biochemistry, and biology. In as much as water is the solvent in which most chemical reactions occur, it's very important to consider ionization of water. In addition to un-ionized $H^20$, $OH^-$ and H+ can be present. The concentration of $OH^-$ ions can be expressed by pOH which is defined as pOH = -log $[OH^-]$. There is a simple relationship between these ions: pH + pOH =14. The pH specifies the ionic state of the most important solvent. Another reason for the widespread use of the pH scale is the ease of constructing electrochemical devices ("pH meters") to measure it.

The pH is a scale of numbers that simplifies manipulation, calculation, and discussions involving the wide range of hydrogen ions in a solution. Determination of pH is c a determination of emf (voltage) difference in a pH cell. pH is dependent upon temperature; a temperature compensator varies the instrument definition of a pH unit from 54.20 mV to 66.10 mV at 60°C.

A pH meter consists of an electrode to sense pH, a reference electrode to complete the circuit while not affecting the pH measurement, an electronics module to link the electrode to the readout, and a display unit which could be a meter or a computer screen. Some pH meters also have a third electrode to sense the temperature in order to compensate for thermal effects. Often the electrodes are packaged into a "combination electrode" which is more convenient to use. The most frequent cause of pH meter malfunction is poorly-maintained electrodes.

In modern chemistry the "p" ("negative log of') notation has been extended to other species. Thus a chemist may use $pS^-$, $pN02^-$, pNa+, and even p(urea). So-called ion-selective electrodes have been developed to measure all of those and many more.

DRAIN CLEANERS

# ANALYSIS OF DRAIN CLEANERS

## Introduction:

Most common household cleaners contain acids or bases. Acidic cleaners, such as toilet bowl cleaners, often contain hydrochloric acid or sodium bisulfate to remove alkaline deposits or stains. Basic cleaners are designed to dissolve grease, hair, and food. In this experiment, you will use a pH meter to monitor the titration of a drain cleaner. This kind of reaction is referred to as a potentiometric titration and does not require the use of an acid-base indicator.

When the data has been plotted, two endpoints will be determined. Since sodium hydroxide is a stronger base, it will expend the acid first.  Therefore, the first endpoint will allow the determination of the amount of sodium hydroxide in the sample. The second endpoint indicates the amount of sodium hypochlorite in the sample. When

## Purpose:

The purpose of this experiment is to determine the percentages of sodium hydroxide and sodium hypochlorite in a commercial drain cleaner.

## Equipment / Materials:

drain cleaner
droppers
distilled water
pH meter and
electrode top loader
balance
pH buffer solutions (pH 7,

stirring bar
1 mL pipet or automatic
pipet stirring rod
50 mL buret and
clamp 0.10 M HCl
250 ml,

Procedure:

1. Plug in the pH meter and allow it to warm up for about 10 munites.

2. The temperature knob should be set between 20 and 25°C.

3. Remove the cap from the electrode, and rinse the electrode with DI water. Blot the end with a Kimwipe.

4. Place the electrode in the pH 7 buffer, turn the knob to pH, and adjust the pH to 7.00 with the standardization knob.

5. Place the instrument on standby. Rinse and blot the electrode.

6. Place the electrode in pH 10.00 buffer, turn the knob to pH, and adjust the pH to 10.00 with the slope knob.

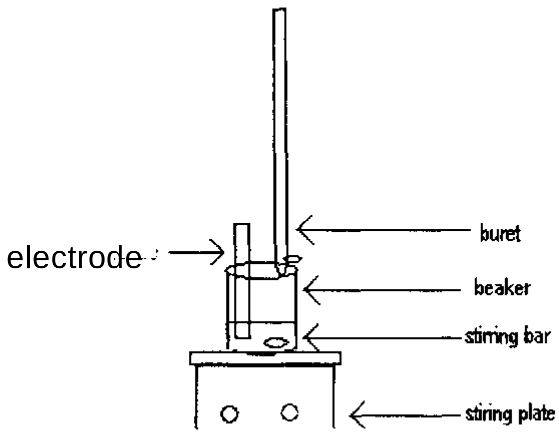
7. Place the instrument on standby. Rinse and blot the electrode.

8. Use- 5 mL of the 0.10 M HCl solution to rinse the buret. Fill the buret with the acid solution past the 0.00 line and drain. If the 0.00 mL line is passed use a dropper to get the reading to 0.00 mL. Make sure the buret tip is filled with acid.

9. Mass a clean, dry 250-mL beaker and record the value. Using the pipet, add 3.0 mL of drain cleaner to the beaker. Remass the beaker and record the value.

10. Add - 75 mL of distilled water to the beaker.

11. Place the magnetic stirrer in the beaker and turn on the stirrer slowly. Immerse the electrode in solution, keeping the electrode avoce and to the side of the stirring bar. See diagram




12. Begin titrating, stopping to record the volume of acid and the pH on the data sheet at intervals of about 0.2 to 0.3 pH units or 1 mL, whichever comes first. Continue to add acid until a pH of less than 3.0 is obtained.

13. Plot the pH versus milliliters of acid. There will probably be two inflection points on the graph. The midpoints of these regions represent two different endpoints, the first for NaOH, and the second for NaOCl.

Name:______________________________

Name:______________________________

Period:_________

Date:__________

# ANALYSIS OF DRAIN CLEANERS

Data Table:

mass of dry beaker:____________

mass of beaker and cleaner:______________

name of product:_________________________

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Calculations:

Using the volume of acid to reach each endpoint, the concentration of the hydrochloric acid, and the formula weights of the components (NaOH and sodium NaOCl) calculate the percentage of each by weight.

***% NaOH***_____________ *% NaOCl*______________

*Questions:*

a. Why is water added to the beaker containing the drain cleaner?

b. Why does the volume of the water added not have to be exact?

c. What volume of acid is needed to neutralize the drain cleaner solution?

# ANALYSIS OF DRAIN CLEANERS
## TEACHER NOTES

Lab Time: about 45 minutes

Preparations:

Time: 20 minutes

Prepare one liter of 0.10 M HCl by diluting 8.3 ml, of concentrated HCl to one liter with distilled water.

Turn on the pH meters about 15 minutes before class to warm up.

Answers to Questions:

a. Why is water added to the beaker containing the drain cleaner?
*Water is added to increase the volume of the solution.*

b. Why does the volume of the water added not have to be exact?
*The water added does not affect the number of hydroxide ions in solution.*

What volume of acid is needed to neutralize the drain cleaner solution?

*Answers will vary depending upon data collected. The volume of acid used at a pH of 7.*

Considerations:

This lab allows the students to analyze a consumer product, not just an "unknown concentration of base". It also shows the students the effects of titrating two different bases in solution at the same time. Different products should be assigned to the lab groups for a variety of analyses and results.

Sample Calculations:

Endpoints: #1 -- 21 nil, #2 -- 45 mL

$$\% \text{ NaOH} = \frac{21\text{ mL HCl} \times \frac{1\text{ liter HCl}}{1000\text{ mL HCl}} \times \frac{0.1\text{ mol HCl}}{1\text{ liter HCl}} \times \frac{1\text{ mol HCl}}{1\text{ mol NaOH}} \times \frac{40.0\text{ g NaOH}}{\text{mol NaOH}} \times 100\%}{3.14\text{ g sample}} = 2.55\% \text{ NaOH}$$

$$\% \text{ NaOCl} = \frac{24\text{ mL NaOH} \times \frac{1\text{ liter HCl}}{1000\text{ mL HCl}} \times \frac{0.1\text{ mol HCl}}{1\text{ liter HCl}} \times \frac{1\text{ mol NaOCl}}{1\text{ mol HCl}} \times \frac{74.5\text{ g NaOCl}}{\text{mol NaOCl}} \times 100\%}{3.14\text{ g sample}} = 5.7\% \text{ NaOCl}$$

The values on the bottle were 2.4% NaOH and 6.0% NaOCI.

Reference: Journal of Chemical Education, February 1988.| *Potentiometric Titration of Acidic and Basic Compounds in Household Cleaners, p.* 184-85.