

## Experiment 8

# Acidity and Alkalinity of Drinking Water

(Standard Methods for the Examination of Water and Wastewater,  
APHA, AWWA, WPCF, 16th edition, p265)

### A. Acidity

#### General Discussion

While pH is a measure of a water sample's deviation from a neutral value of 7.00, it provides little information about a water sample's ability to neutralize acids or bases. The standard procedures for "acidity" and "alkalinity", which have been modified here to fit the time and equipment requirements of a student laboratory, quantitatively provide that information.

The acidity of a water sample is its capacity to neutralize hydroxide ions. Acidity may be caused by mineral acids such as sulfuric acid or hydrochloric acid or by dissolved carbon dioxide. Most commonly in drinking water, carbon dioxide is the principal cause of acidity. Acidity increases the corrosive behavior of water. Drinking water with a high acidity is likely to be corrosive to copper water pipes and to the solder which joins those pipes. High levels of copper and lead in drinking water often occur when acidic water stands in pipes for extended periods of time (such as over night). In addition to creating a possible health hazard due to dissolved metal ions, acidity in water can cause copper plumbing to develop pin hole leaks after a few years.

Acidity is generally measured by titration with sodium hydroxide to an accepted pH value. Phenolphthalein is an acid-base indicator which changes from colorless to a pink (magenta) at a pH of about 8.3. Generally, acidity is measured by titration of a water sample to pH 8.3 with NaOH titrant. Metacresol purple also changes color at pH 8.3, but gives a sharper color change than phenolphthalein. If available, its use is recommended over phenolphthalein. If a water sample is at the alkaline color of the indicator before any titrant is added, then the acidity is zero and the alkalinity of the water should be tested.

Because CO<sub>2</sub> is the most likely cause of acidity in water, the water sample should be collected within a few hours of the time of analysis. The container used to collect the water should be filled completely and closed with an air-tight seal. A clean plastic soft drink bottle with screw cap is suitable for water samples tested in this procedure.

#### Apparatus

- a. **pH meter and calibration standards.**
- b. **Magnetic stirrer and stirring bar.**
- c. **Volumetric flask:** 1000-mL and 100-mL.
- d. **Buret:** 50-mL, glass.
- e. **Polyethylene bottle:** 1-L.

### Reagents

- a. **Carbon dioxide-free water:** For all stock and standard solutions and for dilutions and titrations, water which is free of CO<sub>2</sub> should be used. Boil purified or deionized water for 15 minutes. Allow to cool to room temperature while allowing nitrogen (if available) to slowly bubble through the water as it cools. The stream of nitrogen serves to displace air from the container and prevent carbon dioxide from re-entering the water.
- b. **Potassium hydrogen phthalate (KHP), 0.05 N:** Primary standard grade KHP should be dried for 2 hours at 110 - 120 °C and allowed to cool in a desiccator. Weigh, in a small beaker, about 1.0 g to the nearest 0.1 milligram and transfer quantitatively to a 100-mL volumetric flask by rinsing with CO<sub>2</sub>-free water through a funnel into the flask. Dilute to the mark with CO<sub>2</sub>-free water. The equivalent weight of KHP is 204.23. Calculate the exact normality of the KHP solution based on the weight diluted in the flask.
- c. **Standard sodium hydroxide titrant, 0.025 N:** For most drinking water samples, relatively dilute NaOH is required. Prepare the solution by weighing 1.0 g of solid NaOH and diluting to 1 liter with CO<sub>2</sub>-free water. Standardize against the KHP solution by pipetting 20.00 mL of KHP into an erlenmeyer flask, adding 20 mL CO<sub>2</sub>-free water, and titrating with NaOH to a phenolphthalein (colorless to pink) or metacresol purple endpoint. Determine the normality of the NaOH using the equation:

$$(\text{mL NaOH}) \times (\text{N NaOH}) = (\text{mL KHP}) \times (\text{N KHP})$$

- d. **Phenolphthalein indicator solution, 0.5%:** Dissolve 0.5 g of phenolphthalein in 50 mL ethanol and add 50 mL purified or deionized water.
- e. **Metacresol purple indicator solution, 0.1%:** Dissolve 0.100 g of metacresol purple in 100 mL water.

**Procedure**

Begin by obtaining one or more water samples as suggested or provided by the instructor. Measure and record the pH of the water with a calibrated pH meter.

Following cleaning of a 50-mL buret, rinse it with purified water followed by several rinses with 0.025 N NaOH. Fill the buret with the NaOH solution, make sure there are no air bubbles in the tip, and make sure the meniscus is readable at close to 0.00 mL on the buret scale. Measure 100.0 mL of the water sample to be analyzed into a 250-mL erlenmeyer flask with as little splashing or turbulence as possible. A pipet is preferred but a graduated cylinder may be used if no large-volume pipet is available.

Titrate to a phenolphthalein or metacresol-purple endpoint. If the water is highly acidic, smaller volumes of the sample may be titrated as seems appropriate. Do at least duplicate (preferably triplicate) titrations on each sample being investigated. If the alkaline color of the indicator is observed before any titrant is added, report zero acidity and go on to measurement of alkalinity.

**Calculation**

Acidity is expressed in terms of milligrams of calcium carbonate per liter. These are the same units that are used to express alkalinity and water hardness. For the procedure described here, acidity would be reported as: "The acidity to pH 8.3 =          mg CaCO<sub>3</sub>/L."

**Rinse burets with water before storing them.**

$$\text{Acidity} = \frac{(\text{mL NaOH titrant}) \times (\text{normality NaOH}) \times (50,000)}{(\text{mL water sample})}$$

Calculate the mean value for each sample investigated.

**B. ALKALINITY**

**General Discussion**

Alkalinity is the measure of a water sample's ability to neutralize hydrogen ions (its acid-neutralizing ability). Alkalinity may be caused by dissolved strong bases such as sodium hydroxide or potassium hydroxide (and other hydroxide-containing compounds), and it may also be caused by dissolved carbonates, bicarbonates, borates, and phosphates. The measured alkalinity is the total of all of these species found in a water sample. For the sake of simplicity, it is expressed in terms of mg  $\text{CaCO}_3/\text{L}$  although many species other than dissolved calcium carbonate may actually contribute to the alkalinity.

One important environmental consequence of alkalinity is the ability of a body of water to withstand acidification due to acidic precipitation or atmospheric deposition. A body of water may have a fairly neutral pH, but if its alkalinity is low, it will be readily acidified. A body of water with the same pH but with higher alkalinity will have a greater buffer capacity and, consequently, a greater resistance to acidification.

### **Apparatus**

- a. **Magnetic stirrer and stirring bar.**
- b. **Volumetric flask:** 1000-mL and 100-mL.
- c. **Buret:** 50-mL, glass.

### **Reagents**

- a. **Sodium carbonate solution ( $\text{Na}_2\text{CO}_3$ ), 0.05 N:** Primary standard grade  $\text{Na}_2\text{CO}_3$  should be dried for 4 hours at  $250^\circ\text{C}$  and allowed to cool in a desiccator. Weigh, in a small beaker, about 0.25g to the nearest 0.1 milligram and transfer quantitatively to a 100-mL volumetric flask by rinsing with purified or deionized water through a funnel into the flask. Dilute to the mark with purified or deionized water. The equivalent weight of  $\text{Na}_2\text{CO}_3$  is 53.00. Calculate the exact normality of the  $\text{Na}_2\text{CO}_3$  solution based on the weight diluted in the flask.

- b. Standard hydrochloric acid titrant, 0.02 N:** Transfer 8.3 mL of concentrated (12 N) reagent-grade HCl to a 1000-mL volumetric flask and dilute to the mark with purified or deionized water. This solution has a concentration of about 0.1 N. Dilute 200 mL of 0.1 N HCl to 1000 mL to prepare a solution which is about 0.02 N.

Standardize by pipetting 10.00 mL of 0.05 N Na<sub>2</sub>CO<sub>3</sub> solution into a 250-mL erlenmeyer flask and diluting to about 40 mL with purified or deionized water. Bromcresol green changes in color from blue to yellow as it is acidified. The endpoint is intermediate between blue and yellow and appears as a distinct green color. Titrate using bromcresol green as an indicator until a color change from blue to green appears. Determine the normality of the HCl using the equation: (mL HCl) x (N HCl) = (mL Na<sub>2</sub>CO<sub>3</sub>) x (N Na<sub>2</sub>CO<sub>3</sub>)

- c. Bromcresol green indicator solution, 0.1%:** Dissolve 0.100 g of the sodium salt of bromcresol green in 100 mL water.

### Procedure

Rinse the 50-mL buret with several rinses with 0.02 N HCl. Fill the buret with the HCl solution, make sure there are no air bubbles in the tip, and make sure the meniscus is readable at close to 0.00 mL on the buret scale. Measure 100.0 mL of the water sample to be analyzed into a 250-mL erlenmeyer flask.

Titrate to a bromcresol green (pH = 4.5) endpoint. If the water is high in alkalinity, smaller volumes of the sample may be titrated as seems appropriate. Do at least duplicate (preferably triplicate) titrations on each sample being investigated. **Rinse burets with water before storing them.**

### Calculation

Alkalinity is expressed in terms of milligrams of calcium carbonate per liter.

$$\text{Alkalinity} = \frac{(\text{mL HCl titrant}) \times (\text{normality HCl}) \times (50,000)}{(\text{mL water})}$$

Calculate the mean value for each sample investigated.

### Waste Disposal

All solutions may be rinsed down the drain with water.

### A. Report for Acidity of WATER

If the normality of the NaOH was student-determined, record all data and calculations.

Record water sample sources and pH of each sample.

**Water Acidity** (Be sure to show calculations in notebook)

1. Volume of NaOH \_\_\_\_\_ mL  
Acidity: \_\_\_\_\_ mg CaCO<sub>3</sub>/L
2. Volume of NaOH \_\_\_\_\_ mL  
Acidity: \_\_\_\_\_ mg CaCO<sub>3</sub>/L
3. Volume of NaOH \_\_\_\_\_ mL  
Acidity: \_\_\_\_\_ mg CaCO<sub>3</sub>/L

### B. Report for Alkalinity of WATER

If the normality of the HCl was student-determined, record all data and calculations.

Record Water Sample Sources and pH of each sample

**Water Alkalinity** (Show calculations)

1. Volume of HCl \_\_\_\_\_ mL  
Alkalinity: \_\_\_\_\_ mg CaCO<sub>3</sub>/L
2. Volume of HCl \_\_\_\_\_ mL  
Alkalinity: \_\_\_\_\_ mg CaCO<sub>3</sub>/L
3. Volume of HCl \_\_\_\_\_ mL  
Alkalinity: \_\_\_\_\_ mg CaCO<sub>3</sub>/L

Does the pH of the water samples tested allow a prediction of the acidity or alkalinity of the water? Why or why not? What do the results of this experiment allow you to conclude about the water samples tested (compare them)?