The alkalinity of water is a measure of how much acid it can neutralize. If any changes are made to the water that could raise or lower the pH value, alkalinity acts as a buffer, protecting the water and its life forms from sudden shifts in pH. This ability to neutralize acid, or H⁺ ions, is particularly important in regions affected by acid rain.

In the diagram below, for example, the lake on the right has low alkalinity. When acid rain falls, it is not neutralized, so the pH of the water decreases. This drop in the pH level can harm or even kill some of the aquatic organisms in the lake. The lake on the left, however, has high alkalinity. When acid rain falls in this lake, the acid is partially neutralized and the pH of the water remains fairly constant. In this way, a high alkalinity level helps maintain the health of the water and the organisms that live there.

Alkalinity should not be confused with pH. The pH of a solution is a measure of the concentration of acid, or H⁺ ions, in the water. Alkalinity is a measure of the water's capacity to neutralize an acid, or H⁺ ions, thereby keeping the pH at a fairly constant level.
The alkalinity of surface water is primarily due to the presence of hydroxide, OH⁻, carbonate, CO₃²⁻, and bicarbonate, HCO₃⁻, ions. These ions react with H⁺ ions by means of the following chemical reactions:

\[
\begin{align*}
\text{OH}^- + \text{H}^+ & \rightarrow \text{H}_2\text{O} \\
\text{CO}_3^{2-} + \text{H}^+ & \rightarrow \text{HCO}_3^- \\
\text{HCO}_3^- + \text{H}^+ & \rightarrow \text{CO}_2 + \text{H}_2\text{O}
\end{align*}
\]

Most alkalinity in surface water comes from calcium carbonate, CaCO₃, being leached from rocks and soil. This process is enhanced if the rocks and soil have been broken up for any reason, such as mining or urban development. Limestone contains especially high levels of calcium carbonate.

Alkalinity is significant in the treatment of wastewater and drinking water, because it will influence treatment processes such as anaerobic digestion. Water may also be unsuitable for use in irrigation if the alkalinity level in the water is higher than the natural level of alkalinity in the soil.

### Expected Levels

Alkalinity is reported in units of mg/L CaCO₃, because the carbonate ion, CO₃²⁻, is its primary constituent. Alkalinity levels will vary across the country. Some sample data are shown in Table 1. In general, water in the eastern half of the United States will have a higher alkalinity than water in the west because of a higher occurrence of limestone. Areas in the extreme northeast that have had the limestone scoured away by glacial action will often have a lower alkalinity.

### Summary of Method

Alkalinity is measured by titrating a water sample with sulfuric acid. The Vernier pH Sensor is used to monitor pH during the titration. The equivalence point will be at a pH of approximately 4.5, but will vary slightly, depending on the chemical composition of the water. The volume of sulfuric acid added at the equivalence point of the titration is then used to calculate the alkalinity of the water.
**Materials Checklist**

- **computer**
- **Vernier computer interface**
- **Logger Pro**
- **Vernier pH Sensor**
- **sampling bottle**
- **25 or 50 mL buret**
- **ring stand**
- **two 250 mL beakers**
- **wash bottle with distilled water**
- **100 mL graduated cylinder**
- **2 utility clamps**
- **0.0100 M H₂SO₄ solution**
- **magnetic stirrer and bar (if available)**

**Collection and Storage of Samples**

1. This test should be conducted in the lab. Collect at least 300 mL of sample water so that several 100 mL trials could be run, if needed.

2. It is important to obtain the water sample from below the surface of the water and as far away from the shore as is safe. If suitable areas of the stream appear to be unreachable, samplers consisting of a rod and container can be constructed for collection. Refer to page Intro-4 of the Introduction of this book for more details.

3. If the testing cannot be conducted within a few hours, place the samples in an ice chest or a refrigerator.

**Testing Procedure**

1. Obtain and wear goggles.

2. Using a 100 mL graduated cylinder, carefully add 100.0 mL of sample water to a clean 250 mL beaker. **Note:** The sample water should be near room temperature when this test is conducted.

3. Place the beaker on a magnetic stirrer and add a magnetic stirring bar. Set the stirrer to a speed that mixes the sample well, but does not splash. If no magnetic stirrer is available, you will need to stir the solution with a stirring rod during the titration.

4. Position the computer safely away from the water. Keep water away from the computer at all times.

5. Plug the pH Sensor into Channel 1 of the Vernier interface.

6. Prepare the computer for data collection by opening the file “11 Alkalinity” from the Water Quality with Vernier folder of Logger Pro.

7. Obtain a clean buret and rinse it with a few mL of the 0.0100 M H₂SO₄ solution. **CAUTION:** Sulfuric acid, H₂SO₄, is corrosive. Avoid spilling it on your skin or clothing. Dispose of the rinse solution as directed by your teacher. Use a utility clamp to attach the buret to the ring stand, as shown. Fill the buret a little above the 0 mL level with the H₂SO₄ solution. Drain a
small amount of the solution so it fills the buret tip and leaves the $\text{H}_2\text{SO}_4$ solution at the 0 mL level.

8. Remove the pH Sensor from the storage bottle, and rinse the tip with distilled water from the wash bottle. Use the second beaker to catch the rinse water. Clamp the sensor in place and position it in the sample water so that it is not struck by the stirring bar.

9. You are now ready to perform the titration. This process goes faster if one person manipulates and reads the buret while another person operates the computer and enters volumes.
   a. Click $\text{Collect}$ to start data collection.
   b. Monitor the pH value on the computer screen. Once it has stabilized, click $\text{Keep}$.
   c. Type 0 (the buret volume in mL) in the edit box, then press ENTER.
   d. Add a small quantity of $\text{H}_2\text{SO}_4$ titrant (enough to lower the pH about 0.2 pH units). When the pH stabilizes, click $\text{Keep}$.
   e. Type the current buret reading (to the nearest 0.01 mL) in the edit box, then press ENTER.
   f. Continue adding $\text{H}_2\text{SO}_4$ solution in increments that lower the pH by about 0.2 pH units and enter the buret reading after each increment. When the graph shows the pH value beginning to drop more quickly (at approximately pH 5.5), change to one-drop increments. Enter a new buret reading after each addition. **Note:** It is important that all additions of acid in this part of the titration be exactly one drop in size.
   g. When the pH values start to flatten out (approximately pH 4), again add larger increments that lower the pH by about 0.2 pH units, and enter the buret level after each increment.
   h. Continue for two or three more additions, or until the graph clearly shows that the pH has leveled off again.
   i. Click $\text{Stop}$ when you have finished.

10. Dispose of the beaker contents as directed by your teacher. Rinse the pH Sensor with distilled water from the wash bottle. Use the second beaker to catch the rinse water. Return the sensor to the storage solution bottle and tighten the cap.

**Calculations**

1. Determine the volume of $\text{H}_2\text{SO}_4$ added at the *equivalence point* of the titration. The equivalence point is the point where the titration curve makes the steepest drop in pH.
   a. Examine the data points along the displayed graph of pH *vs.* $\text{H}_2\text{SO}_4$ volume. As you move the cursor right or left, the volume (X) and pH (Y) are displayed below the graph. To determine the equivalence point, go to the region of the graph with the steepest drop in pH. In the screen shot at the right, the cursor is shown at the equivalence point.
   b. Find the $\text{H}_2\text{SO}_4$ volume just *before* this jump.
   c. Find the $\text{H}_2\text{SO}_4$ volume just *after* this jump.
d. Calculate the average of these points by adding them together and dividing by two. Record this number, which represents the exact volume of H₂SO₄ added at the equivalence point, on the Data & Calculations sheet (round to the nearest 0.1 mL).

2. Calculate the alkalinity of the sample by multiplying the volume of H₂SO₄ added at the equivalence point by a conversion factor of 10.0. Record this value on the Data & Calculations sheet.

**Optional Calculations**

3. Calculate the moles of H₂SO₄ used to reach the equivalence point.

4. The reaction occurring in this titration is

   \[ \text{H}_2\text{SO}_4 + \text{CaCO}_3 \rightarrow \text{H}_2\text{O} + \text{CO}_2 + \text{CaSO}_4 \]

   Based on the mole ratio of H₂SO₄ to CaCO₃, calculate the moles of CaCO₃ reacted at the equivalence point.

5. Calculate the mass in grams of CaCO₃ in the sample. Convert to milligrams.

6. Calculate alkalinity in mg/L CaCO₃. Compare this value to your answer in Problem 2.
DATA & CALCULATIONS

Alkalinity

Stream or lake: _____________________________ Time of day: _____________________________
Site name: _________________________________ Student name: __________________________
Site number: _______________________________ Student name: __________________________
Date: _____________________________________ Student name: __________________________

<table>
<thead>
<tr>
<th>Column</th>
<th>A</th>
<th>B</th>
</tr>
</thead>
<tbody>
<tr>
<td>Reading</td>
<td>Volume of H₂SO₄ at the equivalence point (mL)</td>
<td>Alkalinity (mg/L CaCO₃)</td>
</tr>
<tr>
<td>1</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Average Alkalinity (mg/L CaCO₃)

Column Procedure:

A. Record the volume of H₂SO₄ at the equivalence point.
B. Multiply Column A by 10.0 as described in the procedure to obtain alkalinity (B = A × 10.0).

Field Observations (e.g., weather, geography, vegetation along stream) _____________________________
____________________________________________________________________________________
____________________________________________________________________________________
____________________________________________________________________________________

Test Completed: ____________ Date: ______

Water Quality with Vernier
Vernier Lab Safety Instructions Disclaimer

THIS IS AN EVALUATION COPY OF THE VERNIER STUDENT LAB.

This copy does not include:

- Safety information
- Essential instructor background information
- Directions for preparing solutions
- Important tips for successfully doing these labs