

Acid Deposition

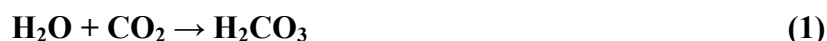
Objectives

1. To calibrate a pH meter and measure the pH of rain water.
2. To observe the effects of CO₂ in water and acid on limestone.

Introduction

Acid rain is a serious environmental problem. It has been implicated in the decline of soil fertility and crop yields, the disappearance of fish from lakes, the destruction of forests, and the disintegration of limestone statues.

Natural rainfall is slightly acidic and has a pH of 5.6. Carbon dioxide in the atmosphere dissolves in water vapor in clouds to produce the weak acid carbonic acid (H₂CO₃) (Equation 1).



The small fraction of H₂CO₃ that dissociates to give hydronium ions is sufficient to reduce the pH of rain to 5.6 (Equation 2).

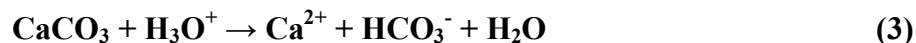


Rainwater with a pH less than 5.6 is described as acid rain. At the present time, the average pH of rainfall in many parts of northeastern North America is below 4.6, and rain with a pH as low as 2.9 has been recorded.

The primary anthropogenic causes of acid rain are emission of sulfur dioxide (SO₂) and oxides of nitrogen (NO_x) from coal- and oil-burning power plants, and emissions of NO_x from automobiles. In the atmosphere, these emissions are converted into sulfuric acid (H₂SO₄) and nitric acid (HNO₃), which accumulate in cloud droplets and fall to earth in rain and snow.

Neutralization

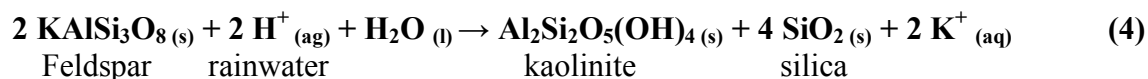
When acid rain falls in regions where the underlying rock is limestone, the acid is neutralized. Limestone, which consists primarily of the base calcium carbonate (CaCO₃), dissolves in the acid, and the hydronium ions are neutralized (Equation 3).



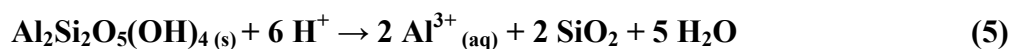
Impacts

In areas where the underlying rock is mainly granite or basalt, which are composed of silicate minerals such as feldspar, the acid is not neutralized. In the normal process of weathering, feldspar, is partially dissolved by naturally acidic rain, and converted to the clay kaolinite and silica. Potassium ions go into solution (Equation 4).

Acid Deposition



Other silicate minerals weather in the same general way to produce clays and metal ions. In the presences of an increased concentration of acid, the clay partially dissolves, and aluminum ions, which are toxic to plants and wildlife, are released into the water (Equation 5).



Lakes in many parts of the world are now abnormally acidic, and their fish populations have declined. The effects on aquatic life are caused both by increasing acidity and by the presence of toxic metal ions, particularly Al^{3+} .

Since 1980, trees in many forests in the eastern United States and parts of Europe have suffered severe damage from acid deposition. Needed nutrients, such as K^+ , Ca^{2+} , and Mg^{2+} ions are leached from the soil by acid, often being replaced by toxic Al^{3+} ions. Unusually acidic soil damages fine root hairs and destroys beneficial microorganisms. Acid deposition can also damage leaves and pine needles.

Building materials, particularly limestone and marble, which are composed of CaCO_3 , are readily eroded by acid, and many ancient statues in Europe have been severely eroded over the last 50 years.

In this experiment you will dissolve carbon dioxide in pure water and observe the effects of the carbon dioxide on the pH of the water. You will observe the effect of acidic solutions on limestone. You will also determine the pH of rain samples you have collected.

Materials Needed

Chemicals	Equipment
<ul style="list-style-type: none"> • Minimum of 10-mL of rain water • 25-mL of distilled water • Crushed dry ice ($\text{CO}_2(\text{s})$) • Limestone (CaCO_3); a piece the size of a pea • 3-mL of xM Nitric Acid (HNO_3) • Standard pH solutions, pH 4, 7 for each lab group 	<ul style="list-style-type: none"> • 2 liter plastic soda bottle • Small plastic sample bottle • Wash bottle • pH probe & screwdriver • Kimwipes • 10 mL, 50 mL Graduated cylinders • 150-mL beaker • Tweezers • Hot plate • Tongs • Matches • Test tube

Acid Deposition

Procedure

Collection of Rain Sample

1. Obtain a clean, dry 2-liter plastic soda bottle and a small plastic sample bottle from your instructor. Cut the soda bottle in half with a kitchen knife. Throw away the top half and use the bottom half as a wide-mouth collection bottle for rain.
2. To collect rain, place the 2-liter bottle in an unobstructed location. The bottle should be raised off the ground to minimize contamination with dirt.
3. You will need a minimum of 10-mL of rain for your experiment. After you have collected the sample, transfer it to the small plastic sample bottle.
4. Label the sample bottle with your name, and the date and location of collection.
5. Store the sample in the refrigerator.

Calibration of the pH probe

1. Rinse the electrode with pure water and immerse in a standard buffer solution of pH 7.0. Stir for 30 seconds, let stand undisturbed for 30 seconds, and then use a small screwdriver to turn the pH 7 Calibration Trimmer until the display shows "7.0". Turn the pH probe off before removing the probe from the solution.
2. Rinse the electrodes thoroughly with distilled water (a plastic wash bottle is convenient for this). Then immerse the electrodes in a second standard buffer of pH 4.0, stir for 30 seconds, let stand 30 seconds, and adjust the pH4/pH10 Calibration Trimmer until the display shows "4.0". Turn the pH probe off before removing the probe from the solution.
3. Finally, rinse the electrode thoroughly and repeatedly with pure water to remove all traces of the buffer solutions. The pH probe is now calibrated and ready to use.

pH of Rain Sample

Measure the pH of the rainwater using a calibrated pH probe. Record the pH. Rinse the electrode with distilled water, blot it dry with a kimwipe, and return to the buffer storage solution.

The Effect of Dissolving CO₂ in Water

1. Using a graduated cylinder, measure 50-mL of distilled water, and pour it into a 150-mL beaker.
2. Measure the pH of the distilled water, and record the value on your report sheet.
3. Using tweezers, add a small piece (the size of a lima bean) of dry ice (solid CO₂) to the beaker. When the CO₂ has finished bubbling, measure the pH of the solution, and record the value on your report sheet. *Caution: Dry ice can cause burns. Do not touch it with your hands.
4. Place the beaker on a hot plate and boil the contents gently for 3 to 4 minutes. Remove the beaker from the hot plate with tongs.
5. When the beaker has cooled to room temperature, measure the pH of the solution, and record the value.

Acid Deposition

The Effect of Acid on Limestone

1. Obtain a sample of limestone, CaCO_3 . Place a piece about the size of a pea in a medium size test tube.
2. Add 2-mL of water, measured with a graduated cylinder. Measure the pH with a calibrated pH probe. Record your value.
3. Measure the pH of your nitric acid (HNO_3) solution. Record your value.
4. Add 3-mL of x M HNO_3 to the limestone sample. *Caution: HNO_3 is a Strong Acid, which can cause chemical burns and destroy clothing. If you spill acid on yourself, wash the contaminated area thoroughly with tap water, and report the incident to your instructor.
5. After the reaction is complete, measure the pH with a calibrated pH probe. Record your value.
6. Add your pH values to the class data on the board.
7. Light a match, and place it over the top of your test tube. Record your observations.

Waste Disposal

Discard leftover HNO_3 in a designated container provided by your instructor.

Discard other wastes in the sink and flush down the drain with tap water.

Adapted from *Chemistry Fundamentals: An Environmental Perspective*, James E. Girard and Phyllis E. Buell

Acid Deposition

Name: _____

Date: _____

Sample Site Description:

Day and time of collection:

pH

Rain water _____

Distilled water _____

After Dry Ice addition _____

After Boiling _____

CaCO₃ + water _____HNO₃ solution _____After HNO₃ addition _____Observations after addition of HNO₃ to limestone:

Observations after lighting the match:

Questions

Acid Deposition

Why did the pH change when the CO_2 was added?

Why did the pH change again when the water was boiled?

Why is the pH of the distilled water below 7?

How does your rain pH compare with locally reported values?

What happened when you added HNO_3 to limestone? How did your values compare with your classmates? Explain any trends.

What happened to the match? What gas was given off?

Write the equation for the reaction of limestone with nitric acid.

Acid Deposition

Laboratory Preparation:

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