THE BUFFER ZONE
ACID-BASE CHEMISTRY IN THE WORLD'S OCEANS

Step 1

Consider the following information in preparation for the rest of this lesson:

What is pH? What do alkalinity and buffering mean? How do these concepts relate to each other and to global climate change? Why do we care?

In this lesson you will learn about the concepts of pH, acid, base, alkalinity, and buffering by completing a simple experiment and reading an article about scientists' discoveries related to global climate change and our oceans.

You will also discuss with your classmates the effects that changes in the ocean may have on humans.

Acids, Bases, and pH

You may already be familiar with the concepts of acids, bases, and pH. There are many examples of acids and bases in our daily lives. For example, vinegar, an acid, is composed of about 5% acetic acid in about 95% water. Baking soda, a base, is a white crystalline solid composed of sodium bicarbonate.

So what makes vinegar an acid and baking soda a base? The difference is how they behave in water once they dissolve. Vinegar and other acids release positive hydrogen ions (H^+) in water; the plus sign refers to a net positive charge. Baking soda and other bases release negative hydroxide ions (OH^-) in water.

Just as distance can be measured in meters and mass in grams, pH is a measure of whether a chemical behaves like an acid (H^+) or a base (OH^-), or whether it is neutral ($H^+ = OH^-$). Instruments such as meters and special chemical test strips measure pH.

pH is reported as a range from 0 to 14. One way to think about the pH scale is to start in the middle, at 7. If you place a pH instrument in a solution (a mixture of water and other chemicals) and obtain a value of 7, that means the ratio of H^+ to OH^- is equal, and the pH is neutral. If the pH is less than 7, there is more H^+ than OH^- in the water, and the pH is acidic (acid = pH < 7). Conversely, when there is more OH^- in the water, the pH is basic, or alkaline (basic = pH > 7).

Salts

When learning about acids and bases, it is important to know about salts. You may be familiar with table salt— sodium chloride (NaCl)—but there are many different forms of salt, such as calcium chloride (CaCl), magnesium sulfate ($MgSO_4$, commonly called Epsom salts), and sodium bicarbonate ($NaHCO_3$).

Salts are formed when an acid and a base are combined. For example, when a strong acid, such as hydrochloric acid (HCl), is mixed with a strong base, such as sodium hydroxide (NaOH), salt and water are produced, resulting in a net neutral charge. This chemical reaction can be shown as

$$HCl + NaOH \rightarrow NaCl + H_2O.$$

The $\rightarrow$ indicates a 1-way reaction, meaning once the reaction occurs, it is done; new products are formed, and the reaction is irreversible.

In this example, when you mix the same ratio of a strong acid and a strong base, a net neutral charge occurs ($H^+ = OH^-$, so the pH = 7). However, not all acids and bases are strong; some are weak. "Strong" and "weak" refer to how completely the H^+ or OH^- disassociates from its parent chemical and doesn't necessarily reflect where the acid or base falls on the pH scale. Strong acids and bases have complete disassociation of H^+ or OH^-, whereas weak acids and bases have varying degrees of disassociation. Mixing strong acids and weak bases (or strong bases and weak acids) still produces salt, but those salts are not neutral; they behave like a weak acid or a weak base when mixed with water. Sodium bicarbonate is an example of a salt that behaves like a weak base in water.

The pH of Life

So why are acids, bases, and salts important? The bottom line: Life depends on them!

Most living things can live only in or near a neutral pH of 7. For example, the human body is constantly regulating its pH to make sure blood stays at a pH of around 7.4. If the pH of our blood changes to below 6.8 or above 7.8, we could die. When we exercise, our body produces lactic acid, which is what makes our muscles sore. Our body uses weak basic salts such as sodium bicarbonate to help neutralize (or "buffer") the acid. Keeping proper salt levels in our body during heavy exercise (when acids can build up quickly) is one reason why various salts are added to sports drinks.

The acid–base buffering is a constant back-and-forth process called conjugation. Here, the acid becomes its conjugate base when the acid loses its H^+ ion (it then has basic properties), and the base becomes its conjugate acid when the base accepts the H^+ ion. If there are equal amounts of acid and base, that equilibrium results in a neutral pH. If a weak acid is added to the system, a buffer can absorb a certain amount of the acid again, bringing the pH closer to neutral again. But if too much acid is added, the buffer is "used up," and the pH starts to drop.

An important buffering system that occurs both in our blood and in the ocean is between bicarbonates (HCO_3^-) and carbonic acid (H_2CO_3). Bicarbonates and carbonic acid come from minerals and carbon dioxide (CO_2), respectively. The bicarbonate–carbonic acid buffering equation in water is

$$H_3O^+ + HCO_3^- \leftrightarrow H_2CO_3 + H_2O.$$

The $\leftrightarrow$ indicates a back-and-forth, reversible, 2-way reaction.

Now that you have a general conceptual understanding of acids and bases and how they support life, you are going to conduct a simple acid–base experiment. What you learn from the experiment will later be applied to the real-world situation of global warming from excess CO_2 in the environment and its impact on the oceans. You will then write a hypothesis about what you think could happen to the pH of the oceans if CO_2 continues to be released into the atmosphere at current levels.

Step 2

Now you will conduct the buffering experiment. You will need the following materials for each group of students:

- pH test (4 colorimetric test strips or pH meter)
- alkalinity test (colorimetric test liquid or strips)
- 20 g sea salt
- 600 mL carbonated water
- 600 mL distilled water
- 4 500-mL containers
- 3 500-mL beakers
- scale
- labels
- stirring rod or spoon

a. Obtain the materials for your group.

b. Label your four containers and (if using them) your pH test strips:

- #1 distilled water
- #2 distilled water with sea salt
- #3 carbonated water
- #4 carbonated water with sea salt

c. Measure two quantities of 300 mL of distilled water and add to container #1 and to container #2.

d. Measure two quantities of 300 mL of carbonated water and add to container #3 and to container #4.

e. Measure 10 g of sea salt and add to container #2 ("distilled water with sea salt"). Stir until dissolved. (NOTE: This ratio of sea salt to water is similar to that in the ocean.)

f. Measure 10 g of sea salt and add to container #4 ("carbonated water with sea salt"). Stir until dissolved.

g. Measure the pH of the solutions in the four containers by following the instructions on the pH test kit or meter. Record the pH of each in the table below.

h. Measure the alkalinity of the solutions in the four containers by following the instructions on the alkalinity test kit.

Record the alkalinity or alkalinity range of each in the table below.

	1 Distilled Water	2 Distilled Water with Sea Salt	3 Carbonated Water	4 Carbonated Water with Sea Salt
pH				
Alkalinity				

NOTE: You may be using colorimetric tests to test for alkalinity and pH. Colorimetric tests use chemical reactions that result in color changes to loosely measure the amount of the substance you are testing. The term "loosely" is used for two reasons. F irst, these tests are not precise in their measurements; instead, you must use your judgment to match the color against a scale (so two people looking at the same pH test strip may interpret different shades of color). Second, some tests, such as the alkalinity test, provide a range rather than a single quantity. For example, the comparison chart for your alkalinity test may show a gradation of color from blue to yellow. If your sample turns blue, the alkalinity of the water you are testing is in the "high" range of 2.9–3.6 milliequivalents/liter (meq/L). If your sample is bluish-yellow, the alkalinity is somewhere in the range of 1.7–2.8 meq/L (which is considered normal for a saltwater aquarium environment). Although these colorimetric tests are not precise, they still provide useful information.

Metered instruments can provide precise quantitative measurements rather than a range. In the case of a pH meter, the tip of the meter's probe (what you stick in the water) is actually measuring the activity of the H^+ ions (the meter measures the change in free energy as ions diffuse across a membrane).

Step 3

Answer the following questions:

a. Refer to the pH data you recorded above. Write whether each experimental parameter below is acidic, basic, or neutral.

Distilled water:

Distilled water with sea salt:

Carbonated water:

Carbonated water with sea salt:

b. What gas is used to make carbonated water? (Read the bottle label or search online for information.)

c. CO_2 is the gas that is released during combustion, or burning, of carbon-based chemicals such as wood, gasoline, coal, and propane (the gas used in some barbecues). As more humans use more fossil fuel–based energy, more CO_2 is released into the atmosphere. The oceans naturally absorb CO_2.

Based on what you have learned about CO_2, acids, bases, salts, and buffers, generate a hypothesis about what you think could happen to the pH of the oceans if CO_2 continues to be released into the atmosphere at current levels. Be sure to incorporate the following concepts into your hypothesis: pH, acid, basic, salt, buffer, and CO_2. Write your hypothesis below.

Step 4

Read **In Hot Water: Global Warming Takes a Toll on Coral Reefs** and answer the following questions:

a. Describe the symbiotic relationship between coral and zooxanthellae, including what happens to coral if zooxanthellae die.

b. How do coral reefs benefit humans?

c. List two ways excessive CO_2 can contribute to an ocean's decline.

d. How does the experiment you conducted simulate the relationship between CO_2 and the oceans?

e. Based on the information you read in the article, is your hypothesis supported or not supported? Explain.

Step 5

Answer the following questions, then discuss your answers as a class or within your group.

a. How did the scientific process help scientists untangle the variables of temperature, coral bleaching, disease, and coral death?

b. The scientific process involves many people testing different hypotheses in an attempt to explain a single event or process. What are some hypotheses explaining the relationship between rising ocean temperature and coral bleaching?

What are some of the variables mentioned in the article that may affect coral health?

c. Observing and describing the natural environment can be particularly challenging because there are many unidentifiable and uncontrollable variables. The article states that "links between warming and disease have only recently gained traction among scientists, although that picture too is complicated; diseases don't always follow bleaching, and sometimes they occur in the absence of warmer water."

The environment is very complex, and similar types of habitats may vary widely depending on their geographic location. For example, a coral reef near Florida has variables different from those of a coral reef near Australia. The Florida reef may receive more wastewater runoff, whereas the Australian reef may be more protected from runoff. Or the Florida reef has different types or amounts of fish compared with the reef in Australia. However, large-scale events such as rising ocean temperatures may produce similar overall effects. This is called variability.

d. What are some ways to reduce coral bleaching and death?

RESOURCES:

Environmental Health Perspectives, News by Topic page, **http://ehp03.niehs.nih.gov/article/browsenews.action**
Choose Climate Change/Global Warming, Marine Science

Casiday R, Frey R. Blood, sweat, and buffers: pH regulation during exercise. Acid–base equilibria experiment. St. Louis, MO: Department of Chemistry, Washington University. **http://www.chemistry.wustl.edu/~edudev/LabTutorials/Buffer/Buffer.html**

Holmes-Farley R. 2002. Calcium and alkalinity. Reef Keeping 1(3). **http://reefkeeping.com/issues/2002-04/rhf/feature/index.php**

National Resources Defense Council. Ocean Acidification: The Other CO_2 Problem. **http://www.nrdc.org/oceans/acidification/default.asp**

State of the Science FACT SHEET: Ocean acidification, **nrc.noaa.gov/plans_docs/2008/Ocean_AcidificationFINAL.pdf**