Names ____

Exploring Ocean Acidification: Chemistry-Dependent pH Change

When acid is added to an aqueous solution, the pH decreases. However, the pH decreases to different extents in different types of solutions. In this activity, we will consider how adding acid to solutions with different characteristics changes the pH of those solutions and how the solutions differ from each other. In the Models explored below, acid is added to each system in the form of carbon dioxide gas (CO_2). This is what happens during ocean acidification and what is causing the pH of the oceans to decrease.

Model 1: Adding CO_2 to Pure Water – How Does The pH Change?

A sample of pure water is exposed to CO_2 in the air for the first time. Prior to this exposure, there was no CO_2 gas dissolved in the water. Figure 1 shows a beaker containing water, CO_2 gas in the air, and CO_2 gas dissolved in water.

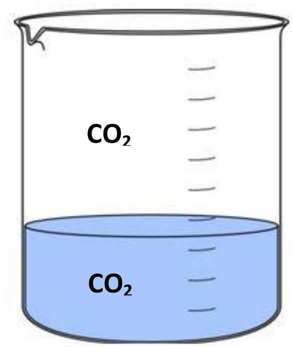

Figure 1

1. Draw an arrow in Figure 1 that indicates the net reaction direction of CO_2 movement when CO_2 gas is first introduced to the air space above the beaker.

2. Write the balanced equation that describes the equilibrium between CO_2 gas in the air and the CO_2 gas in water.

3. Use Le Châtelier's Principle to explain the net reaction direction for the reaction that you wrote for Q2 when CO_2 gas is first introduced to the air space above the beaker.

4. Once in the water, dissolved CO_2 gas reacts with water molecules to form carbonic acid (H_2CO_3). Write the balance equation for this reaction.

5. Once formed, carbonic acid is short-lived and quickly dissociates. Write the <u>two</u> balanced equations for the dissociation of this polyprotic acid in water. [HINT: One proton (H^+) is lost at a time.]

6. Use some or all of your answers from Q1 through Q5 to explain what happens to the pH of distilled water when it is first exposed to CO_2 gas. [HINT: No formal calculation is required for this explanation.]

Model 2: Adding CO_2 to Tap Water – How Does The pH Change?

Tap water is characterized as either hard or soft, depending on the ion concentrations. Soft water has lower ion concentrations than hard water. In contrast, almost all ions have been removed from distilled water. Ions are present in the tap water that we drink because it is derived from natural sources, such as rivers and groundwater, where water comes into contact with minerals such as $CaCO_3$ and $MgCO_3$.

How does the presence of ions affect pH changes in tap water as compared to distilled water? In this section, you will explore this question.

7. A sample of tap water is exposed to CO_2 in the air for the first time. Prior to this exposure, there was no CO_2 gas dissolved in the water. Some of the chemical equilibria involved are the same as those for distilled water. Re-write the balanced equations for those equilibria here (from Q2, Q4, and Q5):

(a) Dissolution of CO_2 gas in water:

(b) Reaction of dissolved CO_2 with water:

(c) Dissociation of carbonic acid (two equations):

8. In tap water, there are additional equilibria. Assume that the sample of tap water was exposed to only three types of minerals in the rock that it flowed through: $CaCO_3$ (calcium carbonate), $MgCO_3$ (magnesite), $MgCa(CO_3)_2$ (dolomite). Write the balanced equations for these solubility equilibria here:

(a) Limestone:

(b) Magnesite:

(c) Dolomite:

9. Use the balanced equations from Q7 and Q8, and your knowledge of Le Châtelier's Principle and acid-base chemistry, to determine whether the pH of tap water would change **more** or **less** than the pH of distilled water when first exposed to CO_2 gas? (Circle either **more** or **less**)

10. Explain your answer to Q9.

Model 3: Adding CO_2 to Ocean Water – How Does The pH Change?

Ocean water contains even more dissolved ions than tap water because all of the rivers of the world ultimately flow into the ocean carrying ions that have been weathered from the minerals found in all the rocks on all of the continents. When water evaporates from the ocean, these ions are left behind. In fact, the “saltiness” of ocean water is a result of the high ion concentrations.

11. Based on your observations from Model 1 and Model 2, would the pH of the ocean change **more** or **less** than the pH of tap water when first exposed to CO_2 gas? (Circle either **more** or **less**)

12. Explain your answer to Q11.

13. Humans have added extra CO_2 to the atmosphere through fossil fuel burning and deforestation activities. Just as in Models 1 and 2, this additional CO_2 has changed the pH of the ocean. Changes in ocean pH have resulted in impacts on marine life, particularly organisms that use the minerals calcium carbonate ($CaCO_3$), $MgCO_3$ (magnesite), and $MgCa(CO_3)_2$ (dolomite) to build their shells. Using the balanced equations from Q7 and Q8, explain how changes in ocean pH have affected these organisms.

14. If these shelled organisms were living in distilled water rather than salty ocean water, how would ocean acidification affect them differently?

Synthesis

15. It should be clear after exploring Models 1, 2, and 3 that ions dissolved in natural waters affect the extent of pH change when acids are added. In your own words, come up with a general theory that describes how dissolved ions regulate changes in solution pH.

Follow Up Question

16. The temperature of ocean water will increase as CO_2 accumulates in the atmosphere and warms the global climate. How will higher ocean water temperatures affect ocean pH changes?