Name: ______________________________ Section: ____________________

Experiment 7 – Ionization and the Nature of Acids, Bases, and Salts

Discussion

Compounds were defined by Sven Arrhenius to be acids if they release H^+ ions in solution when dissolved. This modern definition replaced older definitions based on taste (i.e acids tend to be sour tasting) or if they changed litmus paper's color. Bases (which tend to taste bitter) were defined as compounds that give up OH^- (hydroxide) ions in water. This definition was limited to compounds in water and gives way to Brϕnsted-Lowry acid-base theory.

Brϕnsted-Lowry acid-base theory keeps the definition of an acid as something that donates an H^+ ion and defines bases as anything that accepts the H^+ ion. Acids become proton donors; bases become proton acceptors. In any acid-base equation, there will be one acid and one base *on each side* of the equation. Which compound is an acid depends on whether that compound is donating or accepting a proton.

$$\underset{\text{Base}}{NH_3} + \underset{\text{Acid}}{H_2O} \rightarrow \underset{\text{Conj. Acid}}{NH_4^+} + \underset{\text{Conj. Base}}{OH^-}$$

$$\underset{\text{Acid}}{HCl} + \underset{\text{Base}}{H_2O} \rightarrow \underset{\text{Conj. Base}}{Cl^-} + \underset{\text{Conj. Acid}}{H_3O^+}$$

Water can function as both an acid and a base, depending on the other reagents!

HCl(aq)	Hydrochloric acid	H_2SO_4	Sulfuric acid
HBr(aq)	Hydrobromic acid	$HC_2H_3O_2$	Acetic acid
HI(aq)	Hydroiodic acid	H_2CO_3	Carbonic acid
H_3PO_4	Phosphoric acid	HNO_3	Nitric acid

Many common strong bases contain hydroxides (OH^-) and a metal.

NaOH	Sodium hydroxide
KOH	Potassium hydroxide
$Ca(OH)_2$	Calcium hydroxide
$Mg(OH)_2$	Magnesium hydroxide
NH_4OH	Ammonium hydroxide (best written as $NH_3 \cdot H_2O$)

Solutions that contain bases are called *alkali* or *alkaline*, from an Arabic word for "ashes". Campfire ashes ("bitter ashes") contain hydroxides and carbonates of potassium and sodium, which form basic or alkaline solutions. Compounds from plants that dissolve in water to form alkaline solutions are called *alkaloids*. A common example of a bitter-tasting alkaloid is caffeine.

Name: ________________________________ Section: ____________________

The term pH is used to measure the concentration of an acid in water. Thus, it is important to remember that one acid can produce a range of pH values, depending upon the amount of acid relative to the volume of solution. pH is defined by the equation pH = –log [H^+]. Therefore, a solution of 1.0 M HCl will produce 1.0 M H^+ ions, assuming the HCl breaks up entirely. Since log [1.0] = 0, the pH of this solution is 0. The pH of pure water will be 7.0, while the pH of a very basic solution can be above 14.

pH < 7 acidic solutions	pH = 7 neutral solution	pH > 7 basic solution

When acids react with bases, the H^+ from the acid and the OH^- from the bases "cancel" each other and form water molecules ("HOH"). The anions of the acid and the cations from the base combine to form ionic compounds or salts. For example, consider the reaction of sulfuric acid with sodium hydroxide:

$$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$

Reactions of Oxides with Water

The oxides of elements often react with water to form new compounds. Depending upon which family the element is in, the new compound may be acidic or basic. For example, sulfur can be oxidized to form sulfur trioxide, which reacts with water to make sulfuric acid. Consider the following balanced equations:

$$S + O_2 \rightarrow SO_2$$

$$2SO_2 + O_2 \rightarrow 2SO_3$$

$$SO_3 + H_2O \rightarrow H_2SO_4$$

Carbon dioxide reacts with water to form carbonic acid as follows:

$$CO_2 + H_2O \rightarrow H_2CO_3$$

The metal oxides react with water to form basic compounds. Calcium oxide reacts with water to form calcium hydroxide, while magnesium oxide reacts with water to form magnesium hydroxide:

$$CaO + H_2O \rightarrow Ca(OH)_2$$

$$MgO + H_2O \rightarrow Mg(OH)_2$$

Name: ______________________________ Section: ____________________

Electrolytes

Surprisingly, pure distilled water does not conduct electricity. In order for a charge to pass through water, it needs to be carried by positive and negative charges. The more charges, the more current can pass. If the charges cannot move, as in solid salts with no water present, then electricity cannot be conducted.

Compounds can be divided into strong electrolytes, weak electrolytes, and non-electrolytes depending upon how well they conduct electricity when dissolved in solution. Remember that compounds that don't dissolve in the solvent shouldn't be called electrolytes at all. For example, iron bars, wood, or plastics are not electrolytes regardless of whether they conduct electricity or not.

In a strong electrolyte, the compound breaks up into cations or anions in a process called "dissociation". In a weak electrolyte, some of the compound dissociates into ions, even though the entire compound dissolves. In non-electrolytes, the compound dissolves but does not break up at all.

<u>Procedure</u>

A. Electrolytes

   In this part of the experiment, your instructor will demonstrate the conductivity of various solutions and reactions.

B. Investigating Acids

   1. Reactions of Acids with Metals

      a. Take four separate test tubes and place 5 mL of 6 M HCl in tube #1, 3 M H_2SO_4 in tube #2, 6 M HNO_3 in tube #3, and 6 M acetic acid in tube #4.

      b. Put roughly a 2 cm strip of magnesium metal into each tube. Record the results.

      c. As the metal is still bubbling, place a glowing piece of wood (splint) into the test tube.

   2. Measurement of pH and Acidity

      a. Place 5 mL of water in a test tube and add 2 drops of a phenolphthalein indicator solution in it. Add a few drops of dilute hydrochloric acid and record what happens.

      b. There are three solutions of HCl prepared in front of the classroom. The most concentrated, 0.1 M HCl, is one hundred times more concentrated than the weakest solution, the 0.001 M HCl. Use the pH meter to record the pH's of the three solutions.

3. Reactions of Acids with Carbonates and Bicarbonates

   a. Take a 100 mL beaker and just cover the bottom with a thin layer of sodium bicarbonate (baking soda). Add about 4 to 5 mL of diluted (6 M) HCl to the beaker. Record the results. Lower a lit match into the beaker and record what happens.

   b. Try the above reaction again with a chip of calcium carbonate (limestone, marble). Let the reaction go for about 2 minutes before lowering a lit match into it. When completed, throw them in the labeled water container; DO NOT CLOG THE DRAIN!

4. Neutralizing Acids with Base: Using Indicators

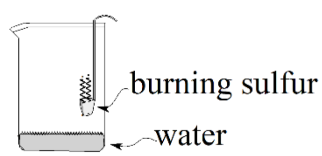
   In this experiment, you will make water acidic and then basic to see how the pH affects a common indicator solution.

   Add 25 mL of water and 3 drops of a phenolphthalein solution to a 100 mL beaker, and then add 5 drops of 6 M hydrochloric acid. To this solution, add 10 percent sodium hydroxide solution drop by drop until the indicator changes color. Once you've gotten this color change, reverse it by adding more dilute acid dropwise.

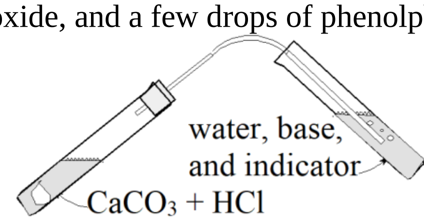
5. Reaction of a Non-Metal Oxide and Water

   In this section, you'll investigate what happens when an oxide of a non-metal, sulfur, reacts with water.

a. This part of the experiment must be done in the fume hood! Place a small lump of sulfur in a deflagrating spoon (which looks like a ladle with a long handle) and set it on fire with a Bunsen burner. Once the sulfur is burning, lower the spoon into a bottle containing 15 mL of water; this will allow the fumes of combustion to fill the air space of the bottle. After 2 minutes, remove the sulfur and cover the bottle with a glass plate. Shake the bottle to mix the gas and water. Is the water acidic or basic?



b. In a test tube, generate carbon dioxide gas by treating marble chips with hydrochloric acid (see section 3b). Bubble the gas into another beaker containing 10 mL of water, 2 drops of 10% sodium hydroxide, and a few drops of phenolphthalein indicator.



Name: ______________________________ Section: ____________________

C. Properties of Bases and Basic Solutions

1. Properties of ammonium and sodium hydroxides

   a. Place three drops of concentrated ammonium hydroxide (used in “Windex” cleaners) in 10 mL of water in a test tube. In another test tube, place three drops of concentrated sodium hydroxide (used in “Drano” pipe cleaners) in 10 mL of water. Rub a few drops of the diluted solution from each test tube onto your fingers. What is the difference in feeling between the two solutions? Wash your hands with water afterwards until your skin feels normal.

   b. Test the two solutions with red and blue litmus papers and record the changes you see.

   c. Add two drops of phenolphthalein indicator to each test tube and record the changes you see.

   d. Determine the pH of each solution using a pH meter. *Wash the electrode with dilute acetic acid and then distilled water to clean it between every reading and after you’re done.*

2. The Reaction of Metal Oxides and Water

   a. In three test tubes, place 10 mL of water, 2 drops of phenolphthalein, and a pinch of calcium hydroxide, magnesium hydroxide, or calcium oxide. Record the color changes.

   b. In this last section, you will explore the reaction that occurs when you heat limestone (“slaking lime”) to make a compound known as “quicklime”, which is used in the manufacture of concrete:

      Take a small piece of iron wire and wrap it around a small chip of calcium carbonate (marble chip). Heat the chip until it is white hot with a Bunsen burner, for about 2 minutes. Let the chip cool and drop it into a beaker with 15 mL of water and a few drops of phenolphthalein. Compare this result to an unheated chip.

Name: ______________________________ Section: ____________________

Data and Calculations for Experiment 7

A. Electrolytes and Instructor Demo
Place an "X" on the label that properly describes each compound below:

	Non-Electrolyte	Strong Electrolyte	Weak Electrolyte
1. Tap water			
2. Distilled water			
3. Sugar solution			
4. NaCl solution			
5a. Pure (glacial) acetic acid			
5b. Diluted acetic acid			
5c. Twice diluted acetic acid			
6a. 1 M acetic acid			
6b. 1 M HCl			
6c. 1 M NH_4OH			
6d. 1 M NaOH			
7a. $NaNO_3$			
7b. NaBr			
7c. $Ni(NO_3)_2$			
7d. $CuSO_4$			
7e. NH_4Cl			

1. What reaction occurs when barium sulfate and sulfuric acid are mixed?

2. Explain why the light becomes dimmer as two strong electrolytes are mixed with each other.

3. Why does the light come back on after more of the electrolyte is added?

4. What happens to the glacial acetic acid as it is diluted? How does this explain the changes in light intensity?

Name: ______________________________ Section: ____________________

B. Properties of Acids

1. Reactions of Acids with Metals

   a) Which acids reacted with the magnesium?

   b) Represent the reaction between the metal and ONE acid that occurred with an equation.

2. Measurement of pH and Acidity

   a) Acids turned the red litmus paper ____________ .

   b) Acids turned the blue litmus paper ____________ .

   c) What is the color of phenolphthalein in acidic solution? ____________

   d) What is the pH of the 0.1 M solution? ____________

      What is the pH of the 0.01 M solution? ____________

      What is the pH of the 0.001 M solution? ____________

   e) Which solution has the greatest concentration of H^+?

   f) Calculate the H^+ concentration of a pH = 4.6 solution. Write the answer in scientific notation.

3. Reactions of Acids with Carbonates and Bicarbonates

   a) What is the name and formula of the gas formed in this reaction?

   b) What happened to the burning stick when it was placed in the beaker?

c) Write out the products of the reactions in a balanced equation:

$NaHCO_3$ + HCl $\rightarrow$

$CaCO_3$ + HCl $\rightarrow$

4. Neutralizing Acids with Base: Using Indicators

a) Write a balanced equation for the reaction of HCl and NaOH.

b) What happened when the acid was all neutralized?

5. Reaction of a Non-Metal Oxide and Water

a) Write a balanced equation for the reaction of sulfur and oxygen.

b) What happens when the product of the above reaction reacts with water? Write a balanced equation that represents this reaction.

c) Write a balanced equation for the reaction of carbon dioxide and water.

d) How do you know that the product in the reaction above is acidic?

Name: ______________________________ Section: ____________________

C. Properties of Bases

1. Properties of ammonium and sodium hydroxides

   a) What did the sodium hydroxide feel like?

   b) What did the ammonium hydroxide feel like?

   c) Bases turned the red litmus paper ____________ .

   d) Bases turned the blue litmus paper ____________ .

   e) What is the pH of the ammonium hydroxide solution? ________

   f) What is the pH of the sodium hydroxide solution? ________

   g) What is the concentration of H^+ in the more basic solution? ________

2. The Reaction of Metal Oxides and Water

   a) What is the color of phenolphthalein with CaO? ______________

      What is the color of phenolphthalein with MgO? ______________

      What is the color of phenolphthalein with $Ca(OH)_2$? ______________

   b) Write the balanced equations for the following reactions:

      $CaO \quad + \quad H_2O \quad \rightarrow$

      $MgO \quad + \quad H_2O \quad \rightarrow$

   c) Marble is calcium carbonate ($CaCO_3$). Write a balanced equation for the reaction that occurs when you heat the marble chip.

   d) Write a balanced equation for the reaction that occurs when you put the heated marble chip in water.