Name: ______________________________ Section: ____________________

Experiment 15 – 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 Brønsted-Lowry acid-base theory.

Brønsted-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:

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

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>

Your instructor will demonstrate the conductivity of various solutions and reactions. Be sure to follow along and record observations.

Name: ________________________________ Section: ____________________

Observations for Experiment 15

Place an "X" on the label that properly describes each compound below:

	Non-Electrolyte	Strong Electrolyte	Weak Electrolyte
1. Tap water			
2. Deionized 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. In 5a. to 5c., what happens to the glacial acetic acid as it is diluted? How does this explain the changes in light intensity?

2. In the final demonstrate, write a balanced chemical reaction (with states) for the mixing of aqueous solutions of barium hydroxide and sulfuric acid.

3. What type of reaction is this? Hint: More than 1 term may apply

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

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