Experiment #10. Hydrolysis and Buffers

Introduction

Hydrolysis is a chemical reaction of a compound with water. It is different from simple dissolution. When an ionic compound (salt) is dissolved in water, the resulting solution can be either acidic, basic or neutral depending on the ions contained in the salt. The ions in the solution undergo hydrolysis and determine the pH of the solution. Recall that Group IA and IIA cations (which are the conjugate acids of strong bases such a NaOH and KOH) are inert as are the conjugate bases of the strong monoprotic acids (i.e., anions such as Cl^-, I^-, NO_3^-, etc.). These ions do not undergo hydrolysis in water and a solution containing these ions (such as a solution of NaCl) has a neutral pH.

If the anion is the conjugate base of a weak acid, however, it will react with the water to form a basic solution. This is because conjugate bases of weak acids are strong enough to remove the protons from water leaving excess OH^- in the solution, which makes the solution basic.

For example: When sodium benzoate, $NaC_6H_5COO(s)$, is added to water, the salt dissociates into Na^+ and $C_6H_5COO^-$. Na^+ is inert, but the benzoate ion reacts with water:

$$\mathbf{C_6H_5COO^-(aq) + H_2O(\mathit{l}) \rightarrow C_6H_5COOH(aq) + OH^-(aq)}$$

A basic solution with a pH > 7 will result.

If the cation is the conjugate acid of a weak base, then it will react with the water to form an acidic solution. For example, when the salt of methylamine, $CH_3NH_3Cl(s)$, is added to water, the salt dissociates into $CH_3NH_3^+$ and Cl^-. Cl^- is inert, but $CH_3NH_3^+$ reacts with water to make an acidic solution with pH < 7.

$$\mathbf{CH_3NH_3^+(aq) + H_2O(\mathit{l}) \rightarrow CH_3NH_2(aq) + H_3O^+(aq)}$$

If the salt contains both the cation of a weak base and anion of a weak acid then the pH of the resulting solution will depend on the relative strengths of the related acid and base.

Buffers

A buffer solution is a solution that resists a change in its pH upon the addition of small quantities of either a strong acid or a strong base. Buffers are usually made by mixing a weak acid and its conjugate base, or a weak base and its conjugate acid. For example, a solution containing NH_4Cl with NH_3 will be a buffer solution. In this lab we will make a buffer using $CH_3COOH(aq)$ and $NaCH_3COO(aq)$.

Buffers work by neutralizing the added strong acid or base. An added strong acid reacts with the base component in the buffer solution:

$$\mathbf{CH_3COO^-(aq) + H_3O^+(aq) \rightarrow CH_3COOH(aq) + H_2O}$$

This removes the hydronium ions created by the addition of the acid and changes the equilibrium amounts of the weak conjugate base (here, acetate) and its corresponding acid (here, acetic acid). This will hold the pH relatively stable.

When a strong base is added to a buffer it will react with the acid component in the buffer solution:

$$\mathbf{CH_3COOH(aq) + OH^-(aq) \rightarrow CH_3COO^-(aq) + H_2O}$$

This removes the added hydroxide ions and changes the equilibrium amounts of weak acid and its conjugate base. Again, the pH of the solution is relatively stable.

In this way the increased amounts of hydronium or hydroxide ion are removed from solution and replaced by the weaker acid molecule or weaker base ion. In today's experiment, you will be measuring the pH of a few salt solutions and study the buffering action of the acetic acid/sodium acetate buffer.

Equipment

Eight 50 or 100 ml beakers   25 ml and 50 ml graduated cylinders

Chemicals

$[Al(H_2O)_6]Cl_3$,	NH_4Cl,	Na_2CO_3	$NaHCO_3$	$CH_3COONa{\cdot}3H_2O$
0.10 M CH_3COOH (aq) Acetic Acid	0.10 M NaOH(aq)	0.10 M HCl (aq)	Deionized water	

Spill/Disposal: All solutions from this experiment may be disposed of into the sink.

Procedure

Part I. Hydrolysis – Salts of Weak Acids or Weak Bases

1. Before coming to lab, Calculate the mass of each of the solids needed to prepare 25.0 mL of 0.10 M solution. Enter this mass on your data sheet and check with your instructor to make sure that you have the right mass.

2. Clean, dry and label four 100 mL beakers. Make 25.0 mL solutions of 0.10 M $NaHCO_3$, Na_2CO_3, $[Al(H_2O)_6]Cl_3$, and NH_4Cl by dissolving the correct mass of each in 25.0 mL of deionized water. Use graduated cylinders to measure the volume of water needed.

3. Carefully stir each salt until it has completely dissolved in the water. Measure the pH of each solution. Your instructor will have standardized the pH meter. Do not take a reading until the electrode is immersed completely in the salt solution

4. When you have finished all of your pH measurements, discard all solutions down the sink.

Part II. Buffering Action of Acetic Acid-Sodium Acetate

1. Before coming to lab, calculate the weight of $NaCH_3COO \cdot 3\,H_2O$ (s) that you will need to make 50.0 mL of a 0.10M solution. The molar mass of the sodium acetate must include the waters of hydration. Check your value with that of the instructor.

Adding strong acid and base to water

2. Clean two 100 mL beakers and place 25.0 mL of deionized water in each. Use a graduated cylinder to measure the required volume of water.

3. It is difficult to determine the pH of deionized water, so assume that the pH = 7.

4. Add 2.5 mL of 0.10 M HCl to one beaker. Stir and measure the pH.

5. Add 2.5 mL of 0.10 M NaOH to the other beaker. Stir and measure the pH.

Adding strong acid and base to a buffer

6. Measure 50.0 mL of 0.10 M CH_3COOH (aq) (acetic acid, a weak acid) into a clean dry 100 mL beaker using a graduated cylinder.

7. Weigh out the calculated amount of $NaCH_3COO \bullet 3\ H_2O$ (s) and add it to the 50.0 mL CH_3COOH (aq) (acetic acid) in the beaker. Carefully stir the solution until all of the salt has dissolved. You have just prepared a buffer where the concentration of the weak acid (acetic acid) and the conjugate base (Sodium acetate) are each 0.10 M.

8. Divide the solution into two equal 25.0 mL portions and measure the pH of each portion.

9. Add 2.5 mL of 0.10 M HCl (aq) to one of the 25.0 mL portions of the buffer. Stir and measure the pH.

10. Add 2.5 mL of 0.10 M NaOH (aq) to the other 25.0 mL portion of the buffer. Stir and measure the pH.

11. When finished, discard all solutions down the sink.

Name___________________________

CHM112 Lab – Hydrolysis and Buffers – Grading Rubric

Criteria	Points possible	Points earned
Lab Performance		
Lab work performed correctly. Proper safety procedures followed and waste disposed of correctly. Work space and glassware cleaned up. Participated actively in performing the experiment.	3	
Lab Report		
Part 1. Solutions correctly made, pH correctly measured, and hydrolysis reactions correctly written.	6	
Part 2. Buffer correctly made and pH accurately measured for all solutions.	4	
Question 1 (work shown in detail with units)	2	
Question 2 (work shown in detail with units)	2	
Question 3 (work shown in detail with units)	3	
Total	20	

Subject to additional penalties at the discretion of the instructor.

Part I: Hydrolysis of salts.

Salt solution	Mass of solid weighed to make 25 ml of 0.10 M (show calculations clearly)	pH
$NaHCO_3$		
Na_2CO_3		
NH_4Cl		
$[Al(H_2O)_6]Cl_3$		

Hydrolysis

	write ions formed, circle ion that is acidic or basic	write hydrolysis reaction for circled ion + H_2O. Neutral ions are not included. pH in table above should align with your reaction.
$NaHCO_3$		
Na_2CO_3		
NH_4Cl		
$[Al(H_2O)_6]Cl_3$	$\mathbf{[Al(H_2O)_6]^{3+}(aq)}$ + $3Cl^-(aq)$	$[Al(H_2O)_6]^{3+}(aq) + H_2O(l) \rightleftharpoons [Al(OH)(H_2O)_5]^{2+}(aq) + H_3O^+(aq)$

Part II: Buffers

Mass of the solid $NaCH_3COO \cdot 3\ H_2O$ needed to make 50. ml of 0.10 M buffer = ______________________

Solution	**pH**
Deionized water	**7**
Deionized water + 2.5 mL 0.10 M HCl	
Deionized water + 2.5 mL 0.10 M NaOH	
Buffer (0.10 M CH_3COOH w/ 0.10 M CH_3COO^-)	
25.0 ml Buffer + 2.5 mL 0.10 M HCl	
25.0 ml Buffer + 2.5 mL 0.10 M NaOH	

1. Calculate the pH of 0.10 M $NaHCO_3$. For H_2CO_3 **($Ka_1 = 4.30 \times 10^{-7}$ and the $Ka_2 = 5.62 \times 10^{-11}$.)**

(Make sure you know if $NaHCO_3$ is acting as an acid or a base!)

Then compare your calculated pH to your measured pH and calculate the percent error.

2. Calculate the pH of the acetic acid and sodium acetate buffer solution (look up K_a for acetic acid). Compare this pH with the measured value.

3. Calculate the pH of the buffer after NaOH (aq) has been added. (Don't forget the new volume of the solution!) How does this calculated pH compare to the measured value? Calculate the percent error.