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## Ph measurement and buffer capacity lab report

How well can a buffer resist pH change? In diluted aqueous solutions, weak acids are slightly dissociated. They produce a small concentration of hydronium ( $\text{H}_3\text{O}^+$ ) and an equal concentration of the conjugate base of the acid. Such dissociation reactions are equilibrium, and equilibrium mathematics can be used to calculate concentrations of the species contained in the solution. Consider formic acid ( $\text{CH}_2\text{O}_2$ ): Dissociation constant ( $K_a$ ) is by the way, formic acid is what red ants inject when they bite. The concentration of  $\text{H}_3\text{O}^+$  present in a 0.010-M solution of formic acid can be calculated based on the equilibrium expression and a reaction table. ( 1 )  $\text{HCOOH}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{HCOO}^-(\text{aq})$  initial 0.010 0  $\Delta -x +x$  equilibrium 0.010  $-x +x$  ( 2 )  $K_a = 1.7 \times 10^{-4} = \frac{[\text{H}_3\text{O}^+][\text{HCOO}^-]}{[\text{HCOOH}]}$ . For this calculation, the square formula was used. If you make the simplified assumption that  $x$  is small compared to  $[\text{HCOOH}]$ , the calculated value is 0.0013 M.† Expressing  $[\text{H}_3\text{O}^+]$  as pH, ( 3 )  $\text{pH} = \log \frac{1}{[\text{H}_3\text{O}^+]} = -\log(0.0012) = 2.92$ . Although a weak acid dissociates only slightly in water, the soluble salt of a weak acid (such as sodium formate) is a strong electrolyte and completely dissociates. ( 4 )  $\text{HCOONa}(\text{s}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{Na}^+(\text{aq}) + \text{HCOO}^-(\text{aq})$  The salt, if added to the weak acid solution, produces a large amount of formation compared to what is produced by the acid dissociation. Adding formation to the equilibrium of the equation 1  $\text{HCOOH}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{HCOO}^-(\text{aq})$  emphasizes the system by adding a product. According to Le Châtelier's principle, the equilibrium will shift to the left (against reactants) and the concentration of  $\text{H}_3\text{O}^+$  will decrease (and the pH will increase). The suppressed dissociation caused by adding an ion already present in the solution is called the usual ion effect. One can make a solution with the help of a weak acid and conjugate base (common ion). The resulting solution will withstand major changes in pH when an acid or base is added to the mixture. Such solutions are called buffer solutions. Consider what would happen in a solution that contains both formic acid and sodium formate when adding acid or base. Adding acid (a source of  $\text{H}_3\text{O}^+$ ) underscores the system by adding a product. The equilibrium of the equation 1  $\text{HCOOH}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{HCOO}^-(\text{aq})$ , shifts towards the reactants, consumer formation and some of the additional  $\text{H}_3\text{O}^+$ . The result: a slight reduction in pH. Adding a base causes hydronium ion to neutralize the base. This emphasizes the system by removing a product. Some formic acid dissociates to replace  $\text{H}_3\text{O}^+$  and equilibrium of the equation 1  $\text{HCOOH}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{HCOO}^-(\text{aq})$  shifts towards the products. The result: a slight increase in pH. The concentrations of the acid and conjugate base in a buffer will determine how much additional acid or base can be added to the solution before the buffering capacity is exhausted. This is called the buffer capacity. The higher the concentrations of acid and conjugate base, the greater the buffer capacity. The previous discussion will also apply if a buffer is prepared using a weak base and conjugate acid. However, buffers cannot be made with strong acids or strong bases and their conjugates. No buffer capacity exists in such solutions because there is no equilibrium; everything has completely dissociated in ions. PH of a buffer solution can be calculated with the Henderson-Hasselbalch equation: ( 5 )  $\text{pH} = \text{p}K_a + \log \frac{[\text{A}^-]}{[\text{HA}]}$  or  $\text{pH} = \text{p}K_a + \log \frac{[\text{base}]}{[\text{acid}]}$ . The diversion of this equation follows from the general dissociation expression of a weak acid, and includes the assumption that  $[\text{H}_3\text{O}^+]$  is small compared to  $[\text{HA}]$ . ( 6 )  $\text{HA} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{A}^-$  ( 7 ) Solution for  $[\text{H}_3\text{O}^+]$  provides: ( 8 ) Taking the negative logarithm on both sides sets the equation in the form of pH: ( 9 )  $\text{pH} = -\log [\text{H}_3\text{O}^+] = \log K_a - \log \left( \frac{[\text{HA}]}{[\text{A}^-]} \right)$ . By definition,  $\text{p}K_a = -\log K_a$  and  $-\log \left( \frac{[\text{HA}]}{[\text{A}^-]} \right) = \log \left( \frac{[\text{A}^-]}{[\text{HA}]} \right)$ . Replace these terms in equation 9  $\text{pH} = -\log K_a - \log \left( \frac{[\text{HA}]}{[\text{A}^-]} \right) = -\log K_a + \log \left( \frac{[\text{A}^-]}{[\text{HA}]} \right)$ . : ( 10 )  $\text{pH} = \text{p}K_a + \log \left( \frac{[\text{A}^-]}{[\text{HA}]} \right)$ . Equation 5  $\text{pH} = \text{p}K_a + \log \frac{[\text{base}]}{[\text{acid}]}$  shows that pH can be found using either concentrations of acid and base, or the number of moles of each. This follows from the fact that the volume amount is the same for the acid and conjugate base, and interrupts in the calculation. In the pre-lab exercise, you will prepare various acetate buffer solutions by direct method and measure the pH of each solution. In the direct method, the conjugate acid and the base are added together in the solution to get the desired base to acid conditions. For example, acetic acid and sodium acetate will be combined in solution. In the buffer survey you will try to prepare an acetate buffer of pH 5.00 and determine the buffer capacity with respect to strong acid and strong base. Measure When you complete this survey, you will: 1 Prepare a buffer with a pH of  $\sim 5.00$ . 2 Check the buffer pH. 3 Determine the buffer capacity with respect to strong acid and strong base. Materials available for use Volumetric bottles Graduated cylinders Pipettor and tips Beaker erlenmeyer flask cage Funnel Analytical balances Vernier LabPro system pH probe Probe calibration solutions 6.0 and 1.0 M acetic acid ( $\text{HC}_2\text{H}_3\text{O}_2$ ) 0.50 and 0.050 M NaOH(aq) 0.50 and 0.050 M HCl(aq) Caution: Acetic acid, HCl, and NaOH is corrosive. They can attack the skin and cause permanent damage to the eyes. If one of these solutions splashes into your eyes, use the eye wash station immediately. Keep your eyes open and rinse with water. If contact with skin or clothing occurs, rinse the affected area with water. Have your lab partner notify your teacher of the spill. All solutions can be rinsed down the sink with water. 1 What reagents do you want to use to achieve the desired pH? 2 What amounts of reagents do you want to use to produce the buffer? 3 titration techniques in preparation for determining the buffer capacity. 1 You need to prepare 100 ml of the buffer solution using the correct volume of 1.0 M of acetic acid and the right mass of sodium acetate. You will want the concentrations relatively diluted, so make the acid concentration 0.10 M or 0.010 M. 2 Check the pH of your buffer. It should be within  $\pm 0.2$  units. 3 Buffer capacity is defined as the number of moles with strong base or strong acid required to cause a device change in the pH of a buffer solution. a Determine the buffer capacity for strong acid by titrating 25.0 ml of the buffer with 0.50 M HCl for the 0.1M buffer or 0.050 M HCl for the 0.01 M buffer until pH is reduced by 1 unit. Report the capacity as mole acid per L buffer. b Determine the buffer capacity for strong base by titrating 25.0 ml of the buffer with 0.50 M NaOH for the 0.1M buffer or 0.050 M NaOH for the 0.01 M buffer until the pH increases by 1 unit. Report the capacity as moles base per L buffer. After your group completes your work, you can prepare a whiteboard that you can use to share and justify your ideas. See the handout provided for details about this process. After you complete the survey, you must prepare a survey report consisting of three sections. This report may require more than 2 pages of data tables. This report must be entered, and any charts, shapes, or tables should be embedded in the document. Part 1: What concept did you examine, and how does it relate to the guiding question? How well can a buffer resist pH change? Define and explain relevant terms, buffer, pH, and buffer capacity. See the questions that get started, as well as the introduction. Part 2: Outline the calculations used to determine the amounts of materials you used in the buffer solution. Describe the steps you used to prepare the buffer and measure buffer capacity. Section 3: What's your argument? Present data for pH and buffer capacity. View a buffer capacity prediction calculation. How was the experimental pH and buffer capacity of the prepared buffer compared to the calculated values? What would cause variations? What range of pH can you buffer with the proposed system? How did the system compare with other groups, especially those working on another buffer system? This third part is where you not only present your data, but use the values you get as proof in your reasoning. Statements that see data table for values are not acceptable! Remember: An argument is not just an answer to the question. It is a claim or conclusion supported by evidence with a justification for why the evidence supports the claim or conclusion. Conclusion.