Name\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Chapter 17 – Practice Problems with Buffers

|  | Composition of Solution |
| --- | --- |
| Solution #1 | 0.025 *M* HOCl(*aq*) |
| Solution #2 | 0.025 *M* HOCl(*aq*) and 0.015 *M* NaOCl(*aq*) |

1. Solution #1 is a 0.025 *M* solution of hypochlorous acid, HOCl (*Ka* = 3.0 × 10–8). Calculate each of the

following quantities for Solution #1.

 (a) [H+] in Solution #1

 (b) The pH of Solution #1 =

 (c) The % ionization of HOCl in Solution #1 =

2. Consider the following equilibrium as you make predictions about Solution #2.

HOCl(*aq*) ⇄ H+(*aq*) + OCl–(*aq*)

The major difference between Solution #2 and Solution #1 is the presence of additional hypochlorite

(OCl–) ions in Solution #2. Due to the common ion effect, the presence of additional ClO– ions has an

effect on the equilibrium shown above.

 (a) The presence of additional OCl– ions in the solution causes the equilibrium position to shift

 toward the (reactants ; products).

 (b) The value of [H+] in Solution #2 should be (less than ; greater than) the value of [H+] in

 Solution #1.

 (c) The pH of Solution #2 should be (less than ; greater than) the pH of Solution #1.

 (d) The % ionization of HOCl in Solution #2 should be (less than ; greater than) the % ionization

 of HOCl in Solution #1.

| Composition of Solution #2 |
| --- |
| 0.025 *M* HOCl(*aq*) and 0.015 *M* NaOCl(*aq*) |

3. In Question #2, you made several predictions about Solution #2. Now it is time to do the calculations to verify that your predictions were correct. Calculate each of the following quantities for Solution #2.

 (a) [H+] in Solution #2

 (b) The pH of Solution #2 =

 (c) The % ionization of HOCl in Solution #2 =



4. Which of the solutions shown above would behave as a better buffer solution? Justify your answer.

5. The value of *Ka* for HNO2 is equal to 4.5 × 10–4.

 (a) The value of p*Ka* for HNO2 is equal to \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ .

 (b) Make predictions about the pH of each of the following solutions.

| Solution | Predict the pH of each solution bychoosing one of the following options.pH < p*Ka* pH = p*Ka* pH > p*Ka* |
| --- | --- |
| a mixture of 1.0 *M* HNO2 and 1.0 *M* NaNO2 |  |
| a mixture of 0.75 *M* HNO2 and 0.55 *M* NaNO2 |  |
| a mixture of 0.83 *M* HNO2 and 1.1 *M* NaNO2 |  |

6. In Question #5, you made predictions about the pH values for three different solutions. Now it is time to do the calculations to verify that your predictions were correct. Calculate the pH of each solution, and

show the set-up for each calculation.

 (a) a mixture of 1.0 *M* HNO2(*aq*) and 1.0 *M* NaNO2(*aq*)

 (b) a mixture of 0.75 *M* HNO2(*aq*) and 0.55 *M* NaNO2(*aq*)

 (c) a mixture of 0.83 *M* HNO2(*aq*) and 1.1 *M* NaNO2(*aq*)

7. The pH range of a buffer is the pH range over which the buffer acts effectively. A buffer is usually

chosen in which the weak acid has a p*Ka* close to the desired pH of the buffer. Buffers that contain equimolar quantities of weak acid and conjugate base have a usable pH range within ±1 pH unit of the p*Ka*.

The pH range of an equimolar mixture of HNO2(*aq*) and NaNO2(*aq*) is between pH \_\_\_\_ and pH \_\_\_\_\_.

8. A student prepares a buffer solution by combining 100 mL of 1.0 *M* HNO2(*aq*) and 100 mL of 1.0 *M*

KNO2(*aq*). Write the net-ionic equation that represents the reaction that best helps to explain why adding

a few drops of 1.0 *M* HCl(*aq*) to the buffer does not significantly change the pH of this solution.



9. A student prepares a buffer solution by combining 100 mL of 1.0 *M* HNO2(*aq*) and 100 mL of 1.0 *M*

KNO2(*aq*). Write the net-ionic equation that represents the reaction that best helps to explain why adding a few drops of 1.0 *M* NaOH(*aq*) to the buffer does not significantly change the pH of the solution.





10. Which of the buffer solutions shown above will be more resistant to changes in pH when either a strong acid or a strong base is added to the buffer? Justify your answer.

11. Determine the volume, in mL, of 1.0 *M* NaOH(*aq*) that should be added to 100. mL of 1.0 *M* HNO2(*aq*)

in order to create a buffer solution that has a pH of 3.35. Justify your answer.

| Substance | Formula | *Ka* |
| --- | --- | --- |
| butanoic acid | HC4H7O2 | 1.5 × 10–5 |
| dihydrogen phosphate | H2PO4– | 6.2 × 10–8 |
| ammonium | NH4+ | 5.6 × 10–10 |

12. Use the information in the table above to calculate the pH for each of the following solutions.

 (a) mixture of 1.0 *M* HC4H7O2(*aq*) and 1.0 *M* NaC4H7O2(*aq*)

 (b) mixture of 1.0 *M* NaH2PO4(*aq*) and 1.0 *M* Na2HPO4(*aq*)

 (c) mixture of 1.0 *M* NH3(*aq*) and 1.0 *M* NH4Cl(*aq*)

13. Hydrofluoric acid, HF, is a weak acid with a Ka of 6.8 × 10–4 and a pKa of 3.17. Three different buffer solutions are prepared in the laboratory that contain mixtures of HF(aq) and NaF(aq). A particulate representation of a small representative portion of each buffer solution is shown in the table below. Cations and water molecules are not shown. Make predictions about the pH of each of these buffer solutions.



| BufferSolution | Particulate Diagramof the Buffer Solution | Predict the pH of each buffer solution bychoosing one of the following options.pH < 3.17 pH = 3.17pH > 3.17 |
| --- | --- | --- |
| #1 |  |  |
| #2 |  |  |
| #3 |  |  |
| #4 |  |  |