Calculations Involving Buffers

You have seen how buffered solutions can be made by adding a conjugate salt to a weak acid or base, and how the pH of the buffered solution can be found using the ICE method or the Henderson-Hasselbalch equation.

Practice:

Calculate the pH of a buffered solution containing .700 M methylamine and .500 M methylaminium chloride.

Two important characteristics of a buffer include:

1. pH Range
   - The pH range is the range of pH values over which a buffer system works effectively.
   - The pH range of an acid or base can be seen as the region of little pH change vs. volume of base/acid added on a titration curve.
   - It is best to choose an acid with a $pK_a$ close to the desired pH.

2. Buffering Capacity
   - The buffering capacity of a buffered solution represents the amount of acid or base that can be absorbed by the solution without a significant change in pH.
   - A buffer with a large capacity contains large concentrations of buffering components.

Therefore the capacity of a buffered solution is determined by the magnitudes of $[HA]$ and $[A^-]$.

Using the Henderson-Hasselbalch equation, along with these relationships, we can create a buffer system at almost any pH.

\[
pH = pK_a + \log \frac{[Base]}{[Acid]}, \quad pOH = pK_b + \log \frac{[Acid]}{[Base]}
\]

1. Calculate the mass of sodium benzoate that must be added to 1 L of 0.40 $M$ benzoic acid ($K_a = 6.4 \times 10^{-5}$) solution to buffer at a pH of 4.5. (Assume no change in total volume)
When Strong Acids or Bases Are Added to a Buffer...

...it is safe to assume that all of the strong acid or base is consumed in the reaction.

In the addition of a strong acid or base to a buffered solution, there are two considerations:

1. How does the added acid-base react with the acid-base present in the solution (neutralization)?

2. How does the neutralization of hydronium or hydroxide affect the equilibrium of the weak acid-base present in the solution?

Addition of Strong Acid or Base to a Buffer

1. Determine how the neutralization reaction affects the amounts of the weak acid and its conjugate base in solution.

2. Use the Henderson–Hasselbalch equation to determine the new pH of the solution.

Calculating pH Changes in Buffers

Example Problem:

- A buffer is made by adding 0.300 mol HC$_2$H$_3$O$_2$ and 0.300 mol NaC$_2$H$_3$O$_2$ to enough water to make 1.00 L of solution. The pH of the buffer is 4.74. Calculate the pH of this solution after 0.020 mol of NaOH is added.

Calculating pH Changes in Buffers

Before the reaction, since

$$\text{mol} \ HC_2H_3O_2 = \text{mol} \ C_2H_3O_2^-$$

$$\text{pH} = pK_a = -\log (1.8 \times 10^{-5}) = 4.74$$
Calculating pH Changes in Buffers

The 0.020 mol NaOH will react with 0.020 mol of the acetic acid:

\[
\text{HC}_2\text{H}_3\text{O}_2(aq) + \text{OH}^-(aq) \rightarrow \text{C}_2\text{H}_3\text{O}_2^- (aq) + \text{H}_2\text{O}(l)
\]

<table>
<thead>
<tr>
<th></th>
<th>HC_2H_3O_2</th>
<th>C_2H_3O_2^-</th>
<th>OH^-</th>
</tr>
</thead>
<tbody>
<tr>
<td>Before reaction</td>
<td>0.300 mol</td>
<td>0.300 mol</td>
<td>0.020 mol</td>
</tr>
<tr>
<td>After reaction</td>
<td>0.280 mol</td>
<td>0.320 mol</td>
<td>0.000 mol</td>
</tr>
</tbody>
</table>

Now use the Henderson–Hasselbalch equation to calculate the new pH:

\[
\text{pH} = 4.74 + \log \left( \frac{0.320}{0.200} \right)
\]

\[
\text{pH} = 4.74 + 0.06
\]

\[
\text{pH} = 4.80
\]

2. Calculate the change in pH that occurs when 10.00 mL of 0.500 M perchloric acid is added to the buffer from the example, after the base was added.

3. Calculate the change in pH that occurs when 0.10 mol of HCl gas is added to 1 L of a buffered solution containing 0.25 M ammonia and 0.40 M ammonium bromide.