# Buffers

A buffer is a solution that resist changes to it's pH. To do that it must be able to react with both added hydroxide and hydrogen ions. Buffers are formed by the mixing of a weak acid or base and it's conjugate. By having both the weak acid and it's conjugate the solution can react with both and added acid or base.

There are two types of problems that we see with buffers.

Pure buffered solutions are just equilibrium problems with weak acids or bases. This is the nothing new.

When a strong acid or base is added to a solution, we need to do this type of problems in two steps. First a stoichiometry problem of the acid base reaction and then an equilibrium ICE diagram problem.

A buffered solution contains 0.50M acetic acid $HC_2H_3O_2$, $K_a$=1.8e-5 and 0.50M sodium acetate $NaC_2H_3O_2$. Calculate the pH of this solution.

Ex:

This equation is know as the Henderson-Hasselbalch equation.

$$pH = pK_a + \log\left(\frac{[A^-]}{[HA]}\right)$$

Calculate the change in pH that would occur when 0.010 mol of solid sodium hydroxide was added to 1.0 L of the buffer from the previous example.

Step 1- Stoichiometry

$$OH^- + HC_2H_3O_2 \iff H_2O + C_2H_3O_2^-$$

| | | | |
|---|---|---|---|
| 0.010 | 0.50 | | 0.50 |
| -0.010 | -0.010 | | +0.010 |
| 0 | 0.49 | | 0.51 |

Step 2-Equilibrium

$$HC_2H_3O_{2(aq)} \iff H^+_{(aq)} + C_2H_3O^-_{2(aq)}$$

| | | |
|---|---|---|
| 0.49 | 0 | 0.51 |
| -x | +x | +x |
| 0.49-x | x | 0.51+x |

$$K_a = \frac{[X][0.51+X]}{[0.49-X]} \quad K_a = \frac{[X][0.51]}{[0.49]} \quad [H^+] = 1.7 \cdot 10^{-5} \quad pH = 4.76$$

$$pH = pK_a + \log\left(\frac{[A^-]}{[HA]}\right) \qquad pH = -\log(1.8 \cdot 10^{-5}) + \log\left(\frac{[0.51]}{[0.49]}\right) = 4.76$$