

Buffer

A solution of a weak acid and its salt (i.e. its conjugate base) OR a weak base and its salt (i.e. its conjugate acid)

What a buffer does: resists a change in pH

How a buffer works: since a buffer contains a pair of acid/base conjugates, it has both an acidic and a basic component that can react

- When a small amount of strong acid is added to a buffer, the base in the buffer "attacks" the H^+ (aka H_3O^+) that was added, thus neutralizing the H^+ and **pH does not change** significantly
- When a small amount of strong base is added to a buffer, the acid in the buffer "attacks" the OH^- that was added, thus neutralizing the OH^- and **pH does not change** significantly

How to solve buffer problems:

$$[H^+] = K_a \times \left(\frac{\text{mols weak acid}}{\text{mols conjugate base}} \right)$$

Buffer capacity: once all the moles of weak acid or all the moles of conjugate base are used in neutralization, the buffer is destroyed and the solution is no longer a buffer (because either the weak acid or conjugate base is missing), thus the addition of strong acid or base to the solution will cause pH to change.

- More concentrated buffers are able to neutralize more acid or base, thus having a higher capacity

Perfect Buffer: the most effective buffer, occurs when...

- mols acid = mols base
- $[H^+] = K_a$
- $pH = pK_a$
- When choosing a buffer, pick the an acid whose pK_a is close to the desired pH
 - Example: if you want a buffered solution with pH of 5.00, pick an acid whose pK_a is about 5.00 (meaning a K_a of 1.0×10^{-5})

Example

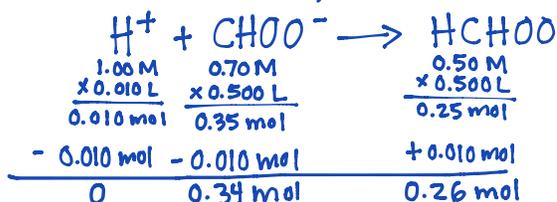
Calculate the pH of 0.500 L of a buffer solution composed of 0.50 M methanoic acid (HCOOH) and 0.70 M sodium methanoate (NaCOO⁻) before and after adding 10.0 mL of 1.00 M HCl. K_a of methanoic acid is 1.8×10^{-4} .

$$\text{Before: } [H^+] = K_a \times \left(\frac{\text{mols WA}}{\text{mols CB}} \right) = K_a \times \left(\frac{\text{mols HCOOH}}{\text{mols COO}^-} \right) = 1.8 \times 10^{-4} \times \left(\frac{0.50 \text{ M} \times 0.500 \text{ L}}{0.70 \text{ M} \times 0.500 \text{ L}} \right)$$

$$[H^+] = 1.3 \times 10^{-4} \text{ M}$$

$$pH = -\log[H^+] = -\log(1.3 \times 10^{-4}) = 3.89 = pH \text{ before HCl added}$$

After: HCl added to buffer, the base of the buffer "attacks" & neutralizes the HCl



$$\begin{array}{l}
 \left(\frac{\text{mols HCOOH}}{\text{mols COO}^-} \right) \times K_a = [H^+] \\
 \left(\frac{0.26 \text{ mol HCOOH}}{0.34 \text{ mol COO}^-} \right) \times (1.8 \times 10^{-4}) = 1.4 \times 10^{-4} \text{ M} \\
 pH = -\log(1.4 \times 10^{-4}) = 3.86 = pH \text{ after HCl added}
 \end{array}$$

* Note: v. small ΔpH
yay for buffers!!