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:

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

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

Example

Calculate the pH of 0.500 L (V) of a buffer solution composed of 0.50 M methanoic acid (HCHOO) (WA) and 0.70 M sodium methanoate (NaCHOO) (CB) before and after adding 10.0 mL of 1.00 M HCl. Ka of methanoic acid is 1.8×10^{-4}.

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

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

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

* Note: v. small ΔpH

yay for Buffers!!

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

	H^+	+ $CHOO^-$	→ $HCHOO$
	1.00 M × 0.010 L = 0.010 mol	0.70 M × 0.500 L = 0.35 mol	0.50 M × 0.500 L = 0.25 mol
	− 0.010 mol	− 0.010 mol	+ 0.010 mol
	0	0.34 mol	0.26 mol

$\left(\frac{mols\ HCHOO}{mols\ CHOO^-}\right) \times Ka = [H^+]$

$\left(\frac{0.26\ mol\ HCHOO}{0.34\ mol\ CHOO^-}\right) \times (1.8 \times 10^{-4}) = 1.4 \times 10^{-4}\ M$

$pH = -\log(1.4 \times 10^{-4}) = \boxed{3.86 = pH\ after\ HCl\ added}$