

Buffers

- pH of a buffer is determined by the ratio of $[A^-]/[HA]$
- Buffer capacity is determined by molarity $[HA]$ and $[A^-]$. Higher molarity means bigger reservoir of HA and A^- to counteract change.

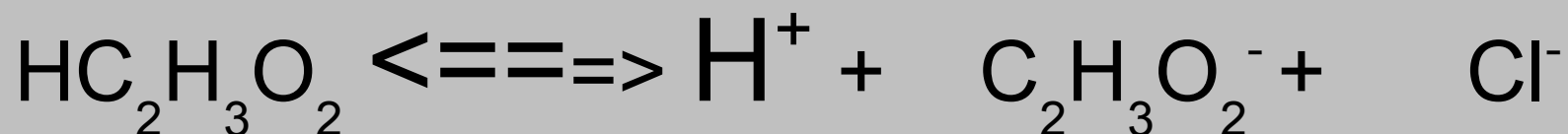
Buffers: Problem solving

- Approach buffer problems by breaking them down into steps.
- 1) What is the change to the buffer by adding an acid or base
 - a.) neutralization
 - b.) dilution
- 2) How will the weak acid or base respond to the new concentrations to establish K_a or K_b equilibrium
 - K_a or K_b problem with common ion

Buffers: Example

- What would the pH be at equilibrium of a buffer system consisting of 0.1 M acetic acid, 0.1 M sodium acetate if 2 ml of 2M HCl is added to 200 mL of this buffer solution. $K_a = 1.8 \times 10^{-5}$
- 1) Neutralization equation
- $\text{HC}_2\text{H}_3\text{O}_2 \rightleftharpoons \text{H}^+ + \text{C}_2\text{H}_3\text{O}_2^- + \text{Cl}^-$
- Reaction runs backward consumes H^+ from strong acid and makes wk acid molecules.

Buffers: Neutralization



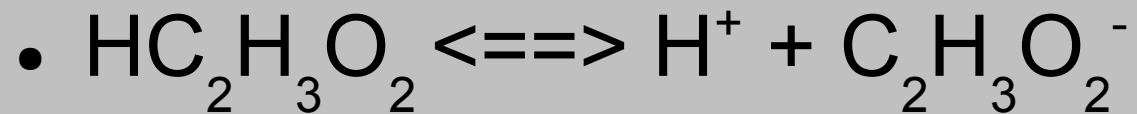
- I 0.02 moles 0.004 mol 0.02 mol ---
- C 0.02 +0.004 -0.004 mol 0.02 -0.004 --
- E 0.024 0 0.016
- Adding strong acid reacted with the conj. Base of the wk acid and more weak acid molecule made. Free H⁺ goes to zero.

Buffers: Dilution

- 0.024 moles/**202** mL = 0.119 M HC₂H₃O₂
- 0.016 moles/**202** mL = 0.079 M C₂H₃O₂⁻
- Now the problem is wk acid dissociation with a non-zero amount of the common ion.
- ICE table

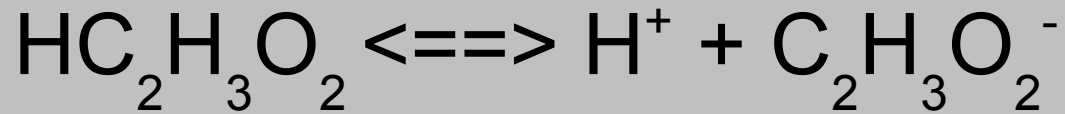
Buffers: ICE table

- Now that we know the actual concentrations after neutralization and dilution. Allow the weak acid to dissociate.

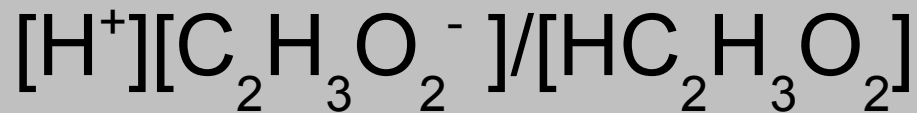


- I 0.119 0 0.079
- C 0.119-x x 0.079 + x
- E

Buffers: ICE table



- I 0.119 0 0.079
- C 0.119-x x 0.079 + x
- E 0.119 2.71 x 10⁻⁵ 0.079



- $\frac{(x)(0.079 + x)}{(0.119-x)} = 1.8 \times 10^{-5}$
- X = 2.71 x 10⁻⁵ pH = 4.57