Buffers

- pH of a bufffer 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 Ka or Kb equilibrium
  - Ka or Kb 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. Ka = 1.8 x10^{-5}
- 1) Neutralization equation
- $HC_2H_3O_2$ <==> H^+ + $C_2H_3O_2$ + Cl^-
- Reaction runs backward consumes H+ from strong acid and makes wk acid molecules.

Buffers: Neutralization

$$HC_2H_3O_2 \Longleftrightarrow H^+ + C_2H_3O_2^- + Cl^-$$

- 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 $HC_2H_3O_2$
- 0.016 moles/***202*** mL = 0.079 M $C_2H_3O_2^-$

- Now the problem is wk acid diassociation 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.
  - $HC_2H_3O_2 <==> H^+ + C_2H_3O_2^-$
- I 0.119 0 0.079
- C 0.119-x x 0.079 + x
- E

Buffers: ICE table

$HC_2H_3O_2 <==> H^+ + C_2H_3O_2^-$

- I 0.119 0 0.079
- C 0.119-x x 0.079 + x
- E 0.119 2.71 x10^{-5} 0.079

$[H^+][C_2H_3O_2^-]/[HC_2H_3O_2]$

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