

## Acids and Bases 14.10

### Buffered Solutions

### The Common Ion Effect is Used to Create Buffered Solutions

#### Buffers resist changes in pH

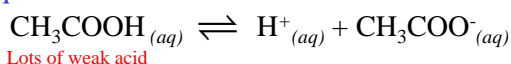
- When small quantities of strong acids or strong bases are added to a buffered solution, the changes in pH are small.

#### Buffer

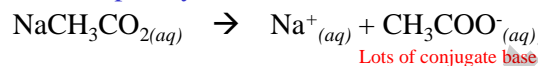
- A weak acid and its salt (extra conjugate base).  
or
- A weak base and its salt (extra conjugate acid).

### Buffered Solutions

Equilibrium lies far to the left for a weak acid

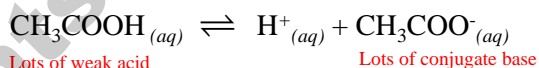


Salts completely dissociate.



The system has lots of acid to react with strong bases and lots of conjugate base to react with strong acids.

### Buffered Solutions



$$K_a = \frac{[\text{H}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]}$$

$$[\text{H}^+] = K_a \times \frac{[\text{CH}_3\text{COOH}]}{[\text{CH}_3\text{COO}^-]}$$

$$-\log[\text{H}^+] = -\log K_a - \log \frac{[\text{CH}_3\text{COOH}]}{[\text{CH}_3\text{COO}^-]}$$

$$\text{pH} = \text{p}K_a + \log \frac{[\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]}$$

### Calculating the pH of Buffers

Two possible procedures:

- The method shown in:  
"Ex1) The Common Ion Effect"
- The Henderson-Hasselbalch equation
  - This method is a little easier.

### Henderson – Hasselbalch Equation

$$\text{pH} = \text{p}K_a + \log \frac{[\text{A}^-]}{[\text{HA}]} \quad \text{or} \quad \text{pOH} = \text{p}K_b + \log \frac{[\text{HB}^+]}{[\text{B}]}$$

$$\text{p}K_a = -\log_{10} K_a$$

$$\text{p}K_b = -\log_{10} K_b$$

[A<sup>-</sup>] = molarity of conjugate base (concentration of salt)

[HA] = molarity of weak acid (initial molarity)

[B] = molarity of weak base (initial molarity)

[HB<sup>+</sup>] = molarity of conjugate acid (concentration of salt)

**Ex1) Henderson – Hasselbalch**

Ex1) Find the pH of a buffered solution created by mixing 0.15 mol  $\text{NH}_4\text{NO}_3$  with 0.65 L of a 0.25 M  $\text{NH}_3$  solution. Assume that the volume change is negligible. ( $K_b = 1.8 \times 10^{-5}$ )

**Step 1. Find the molarity of  $\text{NH}_4\text{NO}_3$**

**Ex1) Henderson – Hasselbalch (cont.)**

**Step 2. Find the pOH**

$$\text{pOH} = \text{p}K_b + \log \frac{[\text{NH}_4^+]}{[\text{NH}_3]}$$

**Step 3. Find pH**

**Ex2) Henderson – Hasselbalch  
Adding a Strong Acid**

Ex2) If 0.02 moles of HCl were added to 1.0 L of the buffered solution from example 1, what would be the new pH? Assume that the volume does not change.

**Step 1. ICE table with moles of each species**

	$\text{NH}_3$	+	$\text{H}^+$	$\rightleftharpoons$	$\text{NH}_4^+$
<b>Initial</b>					
<b>Change</b>					
<b>Equilibrium</b>					

**Ex2) Henderson – Hasselbalch  
Adding a Strong Acid (cont.)**

**Step 2. Find the pOH**

$$\text{pOH} = \text{p}K_b + \log \frac{[\text{NH}_4^+]}{[\text{NH}_3]}$$

**Step 3. Find pH**

**Buffered Solutions and pH**

- $\text{pH} = \text{p}K_a$  when  $[\text{A}^-] = [\text{HA}]$ , as  $\log(1) = 0$
- Buffers made from very weak acids and their salts have high pH values.
  - $\text{pH} = \text{p}K_a = -\log(1 \times 10^{-10}) = 10.0$
  - The basic anion from the salt increases  $[\text{OH}^-]$ .
- Buffers made from stronger weak acids and their salts have lower pH values:
  - $\text{pH} = \text{p}K_a = -\log(1 \times 10^{-4}) = 4.0$
  - The stronger acid increases  $[\text{H}^+]$ .

**Ex) Reactions with Buffered Solutions**

Ex) Hydrobromic acid is added to a buffered solution containing acetic acid and its salt.