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

$$CH_3COOH_{(aq)} \rightleftharpoons H^+_{(aq)} + CH_3COO^-_{(aq)}$$

Lots of weak acid

Salts completely dissociate.

$$NaCH_3CO_{2(aq)} \rightarrow Na^+_{(aq)} + CH_3COO^-_{(aq)}$$

Lots of conjugate base

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

Buffered Solutions

$$CH_3COOH_{(aq)} \rightleftharpoons H^+_{(aq)} + CH_3COO^-_{(aq)}$$

Lots of weak acid — Lots of conjugate base

$$K_a = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]}$$

$$[H^+] = K_a \times \frac{[CH_3COOH]}{[CH_3COO^-]}$$

$$-\log[H^+] = -\log K_a - 1\log\frac{[CH_3COOH]}{[CH_3COO^-]}$$

$$pH = pK_a + 1\log\frac{[CH_3COO^-]}{[CH_3COOH]}$$

Calculating the pH of Buffers

Two possible procedures:

1) The method shown in:
   "Ex1) The Common Ion Effect"

2) The Henderson-Hasselbalch equation
   - This method is a little easier.

Henderson – Hasselbalch Equation

$$pH = pK_a + 1\log\frac{[A^-]}{[HA]} \quad \textbf{or} \quad pOH = pK_b + 1\log\frac{[HB^+]}{[B]}$$

$pK_a = -\log_{10}K_a$

$pK_b = -\log_{10}K_b$

$[A^-]$ = molarity of conjugate base (concentration of salt)

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

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

$[HB^+]$ = molarity of conjugate acid (concentration of salt)

Ex1) Henderson – Hasselbalch

Ex1) Find the pH of a buffered solution created by mixing 0.15 mol NH_4NO_3 with 0.65 L of a 0.25 *M* NH_3 solution. Assume that the volume change is negligible. (K_b = 1.8 x 10^{-5})

Step 1. Find the molarity of NH_4NO_3

Ex1) Henderson – Hasselbalch (cont.)

Step 2. Find the pOH

$$\text{pOH} = \text{p}K_b + \log\frac{[NH_4^+]}{[NH_3]}$$

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

	NH_3	+ H^+	$\rightleftharpoons NH_4^+$
I*nitial*			
C*hange*			
E*quilibrium*			

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

Step 2. Find the pOH

$$\text{pOH} = \text{p}K_b + \log\frac{[NH_4^+]}{[NH_3]}$$

Step 3. Find pH

Buffered Solutions and pH

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

Ex) Reactions with Buffered Solutions

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