BUFFER SOLUTIONS

Buffers

a solution that is able to withstand changes in pH (so that the pH is almost constant) upon addition of small amounts of acid or base - based upon the common ion effect

Common Ion Effect

shift of an ionic equilibrium upon addition of a solute which contains an ion that participates in the equilibrium

EX 12. What is the pH of 1.0 M HF? ($K_a = 6.6 \times 10^{-4}$) (same as EX 9.)

$$HF(aq) \quad + \quad H_2O(l) \quad \Leftrightarrow \quad H_3O^+(aq) \quad + \quad F^-(aq)$$

EX 13. Consider the addition of enough NaF to the solution in EX 12 to make the solution 1 M in both HF and NaF. Since

$$NaF(s) \rightarrow Na^+(aq) + F^-(aq)$$

the concentration table changes:

$$HF(aq) \quad + \quad H_2O(l) \quad \Leftrightarrow \quad H_3O^+(aq) \quad + \quad F^-(aq)$$

Note: The addition of F^- has shifted the equilibrium to the left where HF is less dissociated and a lower concentration of $[H_3O^+]$ is present. This result is perfectly in accord with Le Châtelier's Principle. Since K_a is a constant, if $[F^-]$ is increased then 1) $[H_3O^+]$ must decrease and 2) [HF] increase in order to maintain the same value for K_a.

simpler way to do problem

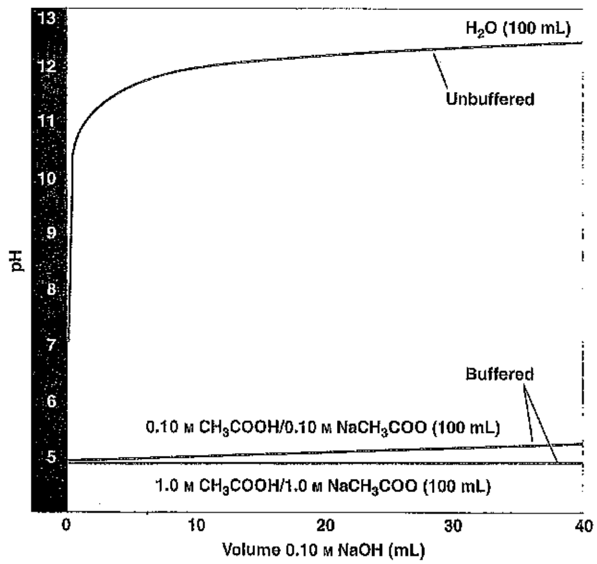
Buffers in Practice

types

buffering action

buffer concentration

FIG IV. Adding Base to Buffered and Unbuffered Solutions



Henderson-Hasselbalch Equation (p. 373)

Working with Buffer Solutions

based on a **weak acid** (HA) and its **conjugate base** (B:)

$$HA(aq) + H_2O(l) <=> H_3O^+(aq) + A^-(aq) \quad K_a = \frac{[H_3O^+][A^-]}{[HA]} \text{ or } pH = pK_a + \log_{10}\frac{[A^-]}{[HA]}$$

$$\text{Henderson – Hasselbalch Equation:} \quad pH = pK_a + \log_{10}\frac{[A^-]_0}{[HA]_0}$$

based on a **weak base** (B:) and its **conjugate acid** (BH^+)

$$B:(aq) + H_2O(l) <=> OH^-(aq) + BH^+(aq) \quad K_b = \frac{[OH^-][BH^+]}{[B:]} \quad \text{use } pK_a \text{ for } BH^+$$

$$\text{Henderson – Hasselbalch Equation:} \quad pH = pK_a + \log_{10}\frac{[B:]_0}{[BH^+]_0}$$

Method I: Addition of Strong Base to a Weak Acid (or Strong Acid to a Weak Base)

EX 14. How many moles of potassium hydroxide would have to be added to 125 mL of a 0.494 M hydrocyanic acid ($K_a = 4.0 \times 10^{-10}$) solution in order to prepare a buffer with a pH of 9.010?

EX 15. An aqueous solution contains 0.372 M hydrofluoric acid ($K_a = 7.2 \times 10^{-4}$). How many mL of 0.376 M sodium hydroxide would have to be added to 125 mL of this solution in order to prepare a buffer with a pH of 3.360?

EX 16. An aqueous solution contains 0.331 M methylamine (CH_3NH_2). How many mL of 0.293 M hydoiodic acid would have to be added to 225 mL of this solution in order to prepare a buffer with a pH of 11.100?

Method II: Addition of Weak Acid and its Conjugate Base (or Weak Base and its Conjugate Acid)

EX 17. You want to prepare 500 mL of a solution buffered at a pH of 4.50 with a buffer concentration of 0.40 M. The buffer is to be made from 1.00 M $C_6H_5CO_2H$ ($K_a = 6.3 \times 10^{-5}$) and its sodium salt $NaC_6H_5CO_2$. What volume of the acid and its conjugate base would you need?