More Stuff about Aqueous Equilibria CH17, Part 1

Common Ion Effect

- assume sln. of acetic acid @ equilibrium

$$CH_3COOH\ (aq) \rightleftharpoons H^+\ (aq) + CH_3COOH^-\ (aq)$$

- What happens if we add solid sodium acetate?

$$NaCH_3COOH\ (aq)\ \cancel{\rightleftharpoons} \rightarrow Na^+\ (aq) + CH_3COOH^-\ (aq)$$

- The presence of acetate from sodium acetate in an acetic acid sln. will shift the equilibrium, according to Le Chatlier's Principle

    The sln. is less acidic.

- assume sln. of ammonia happily @ equilibrium

$$NH_3\ (aq) \rightleftharpoons NH_4^+\ (aq) + OH^-\ (aq)$$

- What happens if add solid ammonium chloride?

$$NH_4Cl\ (aq) \rightarrow NH_4^+\ (aq) + Cl^-\ (aq)$$

- presence of ammonium from amonium chloride in an ammonia sln. will shift equilibrim left.

    The sln. is less basic.

Common Ion Effect

- The shift in equilibrim that occurs by adding a compound that has an ion in common with the dissolved substance.
- The presence of a common ~~acid~~ ion supresses the ionization of a weak acid or base
- (It can also affect solubility, as we will see later)

BUFFERS!

pH of 1.0 L of 0.30 M acetic acid sln. before and after adding 0.30 mol sodium acetate?

Before : $K_a = 1.8 \times 10^{-5} = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]} = \frac{x^2}{0.30 - x}$

ignore x in denominator → $x = 2.3 \times 10^{-3}$ M $= [H^+]$, pH $= 2.63$

After :

	$[CH_3COOH]$ (M)	H^+ (M)	$[CH_3COO^-]$ (M)
I	0.30	0	0.30
C	$-x$	$+x$	$+x$
E	$0.30 - x$	x	$0.30 + x$

$K_a = 1.8 \times 10^{-5} = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]} = \frac{x(0.30 + x)}{0.30 - x}$

Assume that x is small compared to 0.30.

$$1.8 \times 10^{-5} = \frac{x\,(0.30)}{(0.30)}$$

$x = 1.8 \times 10^{-5} = [H^+]$

pH $= 4.78$

- Sln. is less acidic after adding the common ion acetate
- The acid dissociates less because of the common ion

Acid-Base Homeostasis

- If you burn 2,500 calories / day, you are producing 1 mole of CO_2 per hour.

$$CO_2\,(aq) + H_2O \rightleftharpoons \underset{\text{carbonic acid}}{H_2CO_3\,(aq)}$$

- If your blood volume is 4 liters, this is 0.25 M acid added to your bloodstream per hour!

Acid – Base Buffers

- A buffer is a mixture of a weak acid and its conjugate base.
- Adding a small amount of H^+ or OH^- reacts only slightly with one component of the buffer, so the pH doesn't change that much

$$HA \rightleftharpoons H^+ (aq) + A^- (aq)$$

weak acid, conjugate base

insignificant change in pH

① Add OH^-, reacts with HA

$$HA + OH^- \rightarrow A^- + H_2O \quad \text{neutralization}$$

② Add H^+, reacts with A^-

$$A^- + H^+ \rightarrow HA \quad \text{remove } H^+ \text{ from sln.}$$

Henderson – Hasselbalch Equation

$$k_a = \frac{[H^+][A^-]}{[HA]}$$

- take $-\log$ of both sides : $-\log k_a = -\log [H^+] + -\log \frac{[A^-]}{[HA]}$

- Rearrange : $-\log [H^+] = -\log k_a + \log \frac{[A^-]}{[HA]}$

- Simplify : $\cancel{pH = pk_a + \log \frac{[A^-]}{[HA]}}$

$$\boxed{pH = pk_a + \log \frac{[A^-]}{[HA]}}$$

$\cancel{pH \neq pk_a}$

What is it good for?

① To calculate the pH of sln. when [HA], $[A^-]$, and pk_a are known.

② To calculate the pk_a of any weak acid when [HA], $[A^-]$, and pH are known

③ To calculate the ratio of HA to A^- needed to construct a buffer system at a specific pH.