Chapter 8.8
Buffer Systems

Learning Goals: I will be able to…

1. **describe** the chemical characteristics of buffer solutions and **solve** related problems.

Buffers

- **Buffers** are mixtures of weak acid and its conjugate base or weak base and its conjugate acid.
- These conjugate pairs will allow a solution to resist changes in pH when small amounts of acids and/or bases are added.
- Usually buffers have common ions in them that act as a reservoir and help maintain a relatively constant pH.
- Buffer problems should be treated as common ion problems.

Buffers in Action

- Living organisms are very sensitive to pH changes as enzymes carry out their function optimally over a small pH range.
- Human blood plasma has a pH of about 7.4 maintained by the hydrogen carbonate-carbonate buffer system.
- Any change in pH of more than 0.2 induced by poisoning or disease is life-threatening.
- If blood was not buffered, the acid absorbed from a glass of orange juice would be fatal.

How Do Buffers Work?

- Le Chatelier's Principle explains the changes

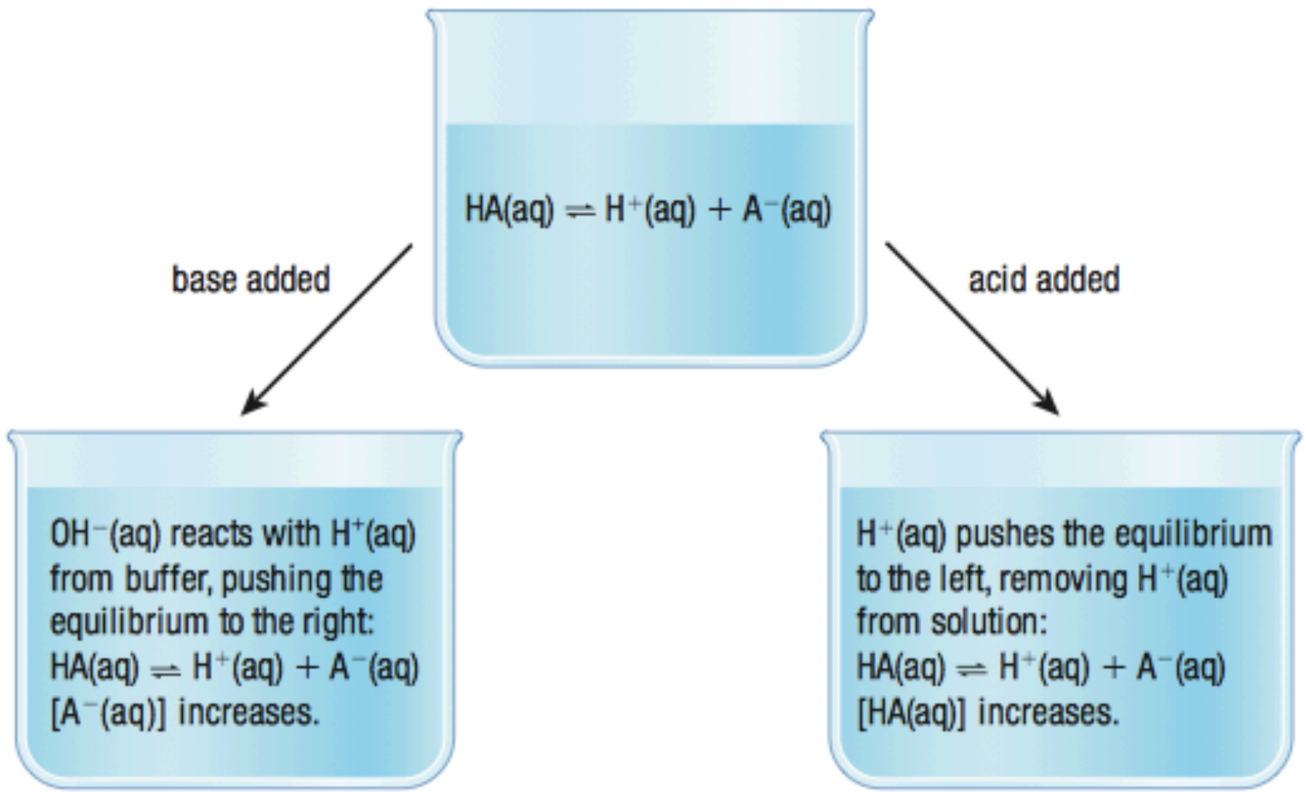

Figure 3 Le Châtelier's principle explains the changes that occur in an acidic buffer solution when a base or an acid is added.

Example: Acetic acid-Acetate ion Buffer

- When base is added:
  - $CH_3COOH_{(aq)} + OH^-_{(aq)} \rightleftharpoons CH_3COO^-_{(aq)} + H_2O_{(l)}$
- When acid is added:
  - $CH_3COO^-_{(aq)} + H_3O^+_{(aq)} \rightleftharpoons CH_3COOH_{(aq)} + H_2O_{(l)}$
- Therefore adding acid or base has **no effect** since $OH^-_{(aq)}$ or $H_3O^+_{(aq)}$ are removed by one component of buffer solution.
- Eventually if enough acid or base is added a considerable pH change will occur.
- The amount of acid or base that can be added is considered the **buffer capacity**.
  - It is determined by the concentrations of its conjugate acid-base pairs.
  - The highest buffer capacity is when at a point half-way to the equivalence point when pH = pK_a

Example 1

Calculate the pH of a buffer that contains 0.20 M acetic acid and 0.20 M sodium acetate.

$CH_3COOH_{(aq)} \rightleftharpoons CH_3COO^-{}_{(aq)} + H_3O^+{}_{(aq)}$

Calculating the pH of a Buffer

Henderson-Hasselbach Equation:

Can be used whenever both conjugate acid-base species are present and assumes approximation method can be used

$$pH = pK_a + \log\left(\frac{[\text{conjugate}]}{[\text{acid}]}\right)$$

$$pOH = pK_b + \log\left(\frac{[\text{conjugate}]}{[\text{base}]}\right)$$

Example 2

Calculate the pH of a buffer containing 0.33 M of ammonia and 0.33 M of ammonium chloride.

Using the Henderson – Hasselbach equation:

Did You Learn?

- Buffers contain a weak acid, $HA_{(aq)}$, and a salt of its conjugate base, $A^{-}_{(aq)}$, or a weak base, $B_{(aq)}$, and a salt of its conjugate acid, $BH^{+}_{(aq)}$.
- Buffers play an important role in many biological and industrial processes.
- When an acid or a base is added to a buffer, the system resists change in pH by removing $OH^{-}_{(aq)}$ or $H^{+}_{(aq)}$ ions from solution until its buffering capacity is exceeded.
- The buffering capacity of a solution containing $HA_{(aq)}$ and $A^{-}_{(aq)}$ depends on the ratios of $[HA_{(aq)}]$ to $[A^{-}_{(aq)}]$. Buffering is most efficient when the ratio is close to 1.

HOMEWORK

Required Reading:

p. 558 – 567

(remember to supplement your notes!)

Questions:

P. 565 #1-3

P. 567 #1-8