<u>Buffers</u>

- buffers resist large changes in pH resulting from the addition of acids or bases
- consist of **a weak acid and its conjugate weak base** or **a weak base and its conjugate weak acid** (usually close to equal concentrations of acid and base)
- often added as soluble salts

  ex. CH_3COOH and $\cancel{Na}CH_3COO$

  $NH_4\cancel{NO}_3$ and NH_3

  NaH_2PO_4 and Na_2HPO_4

buffer = solution where the solutes protect it against large changes in pH even when strong acids or bases are added

- buffers are Brønsted-Lowry equilibria that work by shifting to reduce the effect of adding H_3O^+ or OH^-
  - > consider the equation of a buffer:

$$CH_3COOH(aq) + H_2O(l) \rightleftarrows CH_3COO^-(aq) + H_3O^+(aq)$$

  - > if H_3O^+ is added, the $[H_3O^+]$ increases & the equilibrium shifts left:

$$CH_3COOH(aq) + H_2O(l) \rightleftarrows \mathbf{CH_3COO^-(aq)} + \uparrow H_3O^+(aq)$$

  - > if OH^- is added, it combines with H_3O^+ to form water and $[H_3O^+]$ decrease and the equilibrium shifts right:

$$\mathbf{CH_3COOH(aq)} + H_2O(l) \rightleftarrows CH_3COO^-(aq) + \downarrow H_3O^+(aq)$$

- Bronsted-Lowry acids and bases do not act effectively as buffers on their own. Both conjugate acid and bases must be present in approximately equal amounts to allow shifting in both directions.

$$\mathbf{CH_3COOH(aq)} + H_2O(l) \rightleftarrows \mathbf{CH_3COO^-(aq)} + H_3O^+(aq)$$

- the pH of a buffer depends on the ratio of conjugate acid and base
- the capacity of a buffer refers to the amount (moles) of acid or base a buffer can react with until a large change in pH will occur
- the actual number of moles of acid and base that are used to prepare the buffer determine its capacity
- Diluting a buffer has no effect on the pH because both conjugate acid and base are diluted equally.
- Buffers DO NOT maintain a constant pH, they simply resist large changes.
- Weak acids & their salts are better as buffers for $pH < 7$
- Weak bases & their salts are better as buffers for $pH > 7$
- Strong acids / bases can not be used to prepare buffers, they are too reactive

Buffers in Biological Systems

- buffers are important in biological systems to prevent systems from being overwhelmed by changes in acidity/basicity of the environment
- blood is an example of a buffered solution
- the principal acid and ion responsible for the buffering action is carbonic acid, H_2CO_3, and the bicarbonate ion, HCO_3^-

$$H_2O \quad + \quad H_2CO_3 \quad \leftrightarrows \quad H_3O^+ \quad + \quad HCO_3^-$$

- this maintains a blood pH near 7.35
- variations in blood pH are usually less than 0.1 of a pH unit; a change of 0.4 is likely to be fatal

1. A buffer solution is prepared using 1M NH_3 and 1M NH_4Cl.

a) write the equilibrium equation describing this buffer:

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

b) When a small amount of HCl is added, ↓ the $[OH^-]$ quickly __↓__creases (the pH goes __↓__)

c) As a result, the equilibrium shifts to the __right__ →

and the $[OH^-]$ gradually __↑__creases (the pH goes back __up__)

d) So as a result of adding HCl, there was a small net __↓__crease in the $[OH^-]$ (a small net __↓__crease in pH)

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

e) Using the same buffer system, a small amount of NaOH is added, the $[OH^-]$ quickly __↑__creases (the pH goes __↑__)

f) As a result, the equilibrium shifts to the __left__

and the $[OH^-]$ gradually __↓__creases (the pH goes back __↓__)

g) So, as a result of adding NaOH, there was a small net __↑__crease in the $[OH^-]$ (a small net __↑__crease in pH)

$CO_2 \longrightarrow H_2CO_3$

H

$SO_2 + H_2O \rightleftharpoons HSO_3 + OH$

$SO_2 \longrightarrow H_2SO_4$