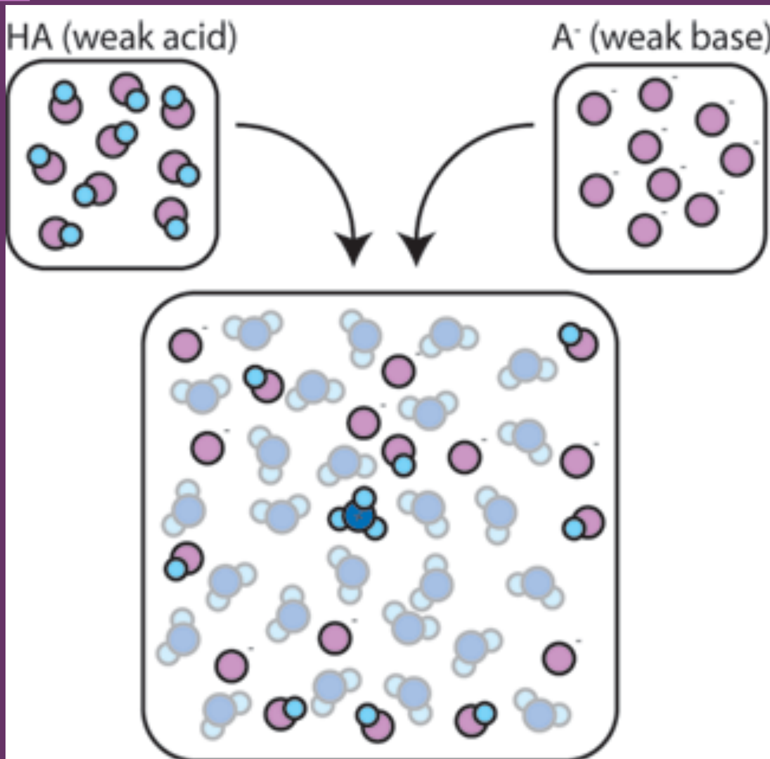

Buffers

Buffer

- any substance or mixture of compounds that, added to a solution, is **capable of neutralizing both acids and bases** without appreciably changing the original acidity or alkalinity of the solution.

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- A buffer is created when both a weak acid and its conjugate base are present in solution (or a weak base and its conjugate acid)
- i.e. CH_3COOH and sodium acetate

Equilibria in Acidic Buffer Solutions

Acidic Buffer

- Weak acid and its conjugate base
- Both in relatively high concentrations

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- to write the equilibrium equation for an acidic buffer, write the IONIZATION of the WEAK ACID in water

Example 1

- A buffer solution is prepared that contains 1M CH_3COOH and 1M $NaCH_3COO$
- Step 1: Determine the Ionization Equation

$$CH_3COOH + H_2O \rightarrow CH_3COO^- + H_3O^+$$

- Step 2: Determine the effect on the equilibrium Equation
  - increase the H_3O^+ / decrease the H_3O^+

Focus on the H_3O^+ in Acid Buffers

- When acid or base is added to the buffer it changes the amount of H_3O^+ present in the solution
- The equilibrium will shift either left or right to counteract the change in H_3O^+

Example 2

- A small amount of HCl is added to an unbuffered solution, and the pH changes from 7.0 to 5.0. If the same amount of HCl is added to a buffer solution with the same volume as the unbuffered solution, what would the resultant change in pH be?
- Draw a graph to demonstrate ☺

- the pH goes DOWN when acid is added

Equilibria in Basic Buffer Solutions

Basic Buffer

- Weak base and its conjugate acid (in a salt)

Example 1

- A buffer solution is prepared containing 1M NH_3 and 1M NH_4Cl.
- Step 1: Determine the equilibrium equation

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

- Step 2: Determine the effect on the system

Focus on OH- in Basic Buffers

- When acid or base is added to the buffer it changes the amount of OH- present in the solution
- The equilibrium will shift either left or right to counteract the change in OH-

Example 2

- A small amount of NaOH is added to an unbuffered solution, and the pH changes from 7.0 to 9.0. If the same amount of sodium hydroxide is added to a buffer solution with the same volume as the unbuffered solution, how would the net pH of the solution respond?
- Draw a graph ☺

Adding an Acid to an Acidic Buffer

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- A small amount of HCl is added to a buffer solution containing 1M acetic acid CH_3COOH and 1M sodium acetate $NaCH_3COO$.

- The equilibrium equation is:

$$CH_3COOH + H_2O \Leftrightarrow H_3O^+ + CH_3COO^-$$

What Happens?

1. The addition of HCl immediately increases the H_3O+ and decreases the pH

2. To counteract the increase in H_3O+, the equilibrium shifts to the left

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3. As the equilibrium shifts to the left, CH_3COO^- reacts with H_3O^+. Because there initial $[CH_3COO^-]$ is high, there is enough of it to react with much of the excess H_3O^+
4. The H_3O^+ goes down close to, but not as low as, its original value
5. In the overall process, the final H_3O^+ is slightly higher and the pH is slightly lower than it was before the acid was added

Adding a Base to an Basic Buffer

Example 4

- A small amount of NaOH solution is added to a buffer solution consisting of 1M NH_3 and 1M NH_4Cl.

- The buffer equilibrium equation is:

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- When NaOH is added to the buffer, the [OH-] immediately _______________, and the pH immediately ________________.

- This would cause a shift to the _____________, and the [OH-] would gradually _____________ during the shift. The pH of the solution would gradually _____________ during this shift.

- Draw a graph showing what happens to the [OH-] from before the time NaOH is added until the shift is complete and a new equilibrium mixture exists

Adding an Base to an Acidic Buffer

A small amount of NaOH is added to an acetic acid / acetate buffer

Adding an Acid to a Basic Buffer

A small amount of HCl is added to an ammonia / ammonium buffer

Buffer Limitations

Moles Matter

- Once all the buffer is used up, buffering will no longer occur and the effect of the added acid/base will be noticeable
- It all depends on the amount of moles present in solution!
- i.e. if you add 6moles of HCl to a 2mole basic buffer, the buffer will only handle 2moles of HCl.

Buffer Uses

- Stabilizing pH in hot tubs and swimming
- Maintaining pH for pharmaceuticals
- Controlling pH in wines and foods
- Calibrating pH meters

Blood Buffers

- CO_2 is a product of metabolism in human cells.
- The CO_2 dissolves in blood plasma (water) to form H_2CO_3:

$$\mathbf{CO_2(g) + H_2O(l) \Leftrightarrow H_2CO_3(aq)}$$

- H_2CO_3 ionizes to form H_3O+ and HCO_3-:

$H_2CO3(aq) + H_2O(l) \Leftrightarrow H_3O+(aq) + HCO_3-(aq)$

- This process is summarized in the body as:

$\mathbf{CO_2(g) + 2H_2O(l) \Leftrightarrow H_3O+(aq) + HCO_3-(aq)}$

- In this case, $CO_2(g)$ is the weak base, HCO_3-(aq) is the conjugate acid.
- This buffer helps to keep the blood pH level close to 7.35
- During periods of high physical exertion, CO_2 is controlled by breathing and [HCO_3-] is controlled by the kidneys