

13. Buffers (IV.19)

a) Purpose

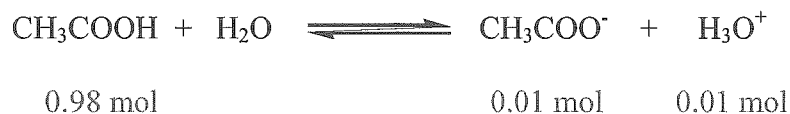
prevents a significant change in pH when acid or base is added

No Buffer
0.1 mol HCl added to water
pH changes 6 units! (7 to 1)

With Buffer
0.1 mol HCl added to water
containing "buffer system"
pH changes only 0.08 units!

b) Acidic Buffer Systems

i) 1 mol CH₃COOH in 1 L water: (*buffers against base only*)

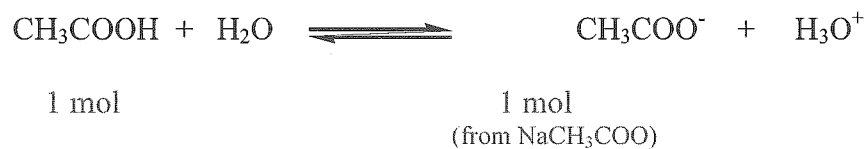


① If add OH⁻, the H₃O⁺ present will "mop up" (neutralize) the OH⁻ added. Thus, no net increase in pH.

The equilibrium will also shift right to replace the H₃O⁺ used to neutralize the OH⁻.

② But, if add H₃O⁺, the equilibrium will shift left but will quickly stop as soon as CH₃COO⁻ is used up. Thus, the pH will decrease.

ii) 1 mol CH₃COOH and 1 mol NaCH₃COO in 1 L water: (*buffers against acid and base*)



① If add OH⁻, system prevent pH increase as described above.

② If add H₃O⁺, there is now plenty of CH₃COO⁻ which can react with the H₃O⁺, shift the equilibrium to the left, and prevent a decrease in pH.

iii) Why is it called an “acidic” buffer system?

$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]}$$

$$1.8 \times 10^{-5} = \frac{[1.0\text{M}][\text{H}_3\text{O}^+]}{[1.0\text{M}]} \quad 1.8 \times 10^{-5} = [\text{H}_3\text{O}^+] \quad \text{pH} = 4.74$$

This is an acidic buffer system because it will maintain pH around 4.74

iv) Example: If we add 0.1 mol HCl to 1L water that contains the above buffer system, what is the change in pH?



1.1 mol

0.9 mol

0.1 mol of CH_3COO^- reacts with the 0.1 mol H_3O^+ , shifting equilibrium to the left, thus decreasing CH_3COO^- by 0.1 mol and increasing CH_3COOH by 0.1 mol.

$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]} \quad 1.8 \times 10^{-5} = \frac{[0.9\text{M}][\text{H}_3\text{O}^+]}{[1.1\text{M}]}$$

$$2.2 \times 10^{-5} = [\text{H}_3\text{O}^+] \quad \text{pH} = 4.66 \quad \text{pH change} = 4.74 - 4.66 = 0.08$$

v) An acidic buffer is made of a weak acid (CH_3COOH) and a salt containing its conjugate base (NaCH_3COO)

c) Basic Buffer Systems

i) 1 mol NH_3 and 1 mol NH_4Cl in 1 L water:



1 mol

1 mol
(from NH_4Cl)

① If add OH^- , excess NH_4^+ mops it up

② If add H_3O^+ , OH^- mops it up and excess NH_3 shifts right to replace OH^-

ii) Why is it called a "basic" buffer system?

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$

$$1.79 \times 10^{-5} = \frac{[1.0\text{M}][\text{OH}^-]}{[1.0\text{M}]} \quad 1.79 \times 10^{-5} = [\text{OH}^-] \quad \text{pOH} = 4.74 \quad \text{pH} = 9.25$$

This is a basic buffer system because it will maintain pH around 9.25

iii) A basic buffer system is made of a weak base (NH_3) and a salt containing its conjugate acid (NH_4Cl).

d) Weak Acid/Base Titrations are Buffers

i) titrations of a weak acid with a strong base sets up a buffer equilibrium system



Halfway to the equivalence point ($V_{\text{eq}}/2$) is when we have a buffer because $[\text{CH}_3\text{COOH}] = [\text{CH}_3\text{COO}^-]$

ii) titrations of a weak base with a strong acid also sets up a buffer system



Halfway to the equivalence point ($V_{\text{eq}}/2$) is when we have a buffer because $[\text{NH}_3] = [\text{NH}_4^+]$

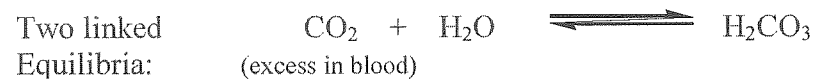
e) Application

i) The Haemoglobin enzyme in blood works at an optimum pH of 7.4

pH > 7.4 O_2 not released

pH < 7.4 O_2 not picked up

ii) Blood has a buffer system to prevent drastic pH changes



134 - 143 (IV, 19-20)