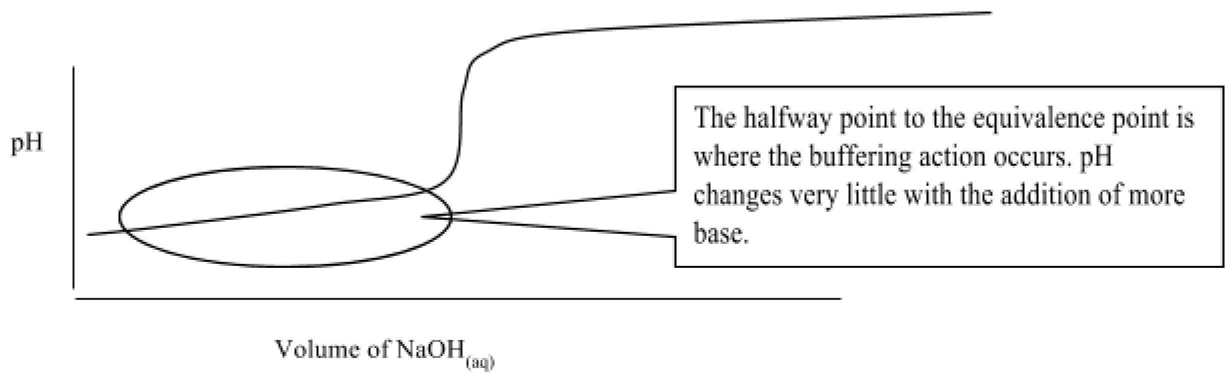
8.5 Buffers

- All pH curves involving a weak acid or weak base have at least one region where a **buffering action** occurs (a region on the curve where the **pH changes very little** despite the addition of a significant amount of acid or base.
- Buffer: a mixture of a conjugate acid-base pair that maintains a nearly constant pH when diluted or when a strong acid or base is added; an equal mixture of a weak acid and its conjugate base.



- Example of a common buffer is equal amounts of ethanoic acid ($\mathbf{HC_2H_3O_2}_{(aq)}$) and sodium ethanoate ($Na^+\mathbf{C_2H_3O_2^-}_{(aq)}$).

- When H^+ is added, the $C_2H_3O_2^-{}_{(aq)}$ forms $HC_2H_3O_{2\,(aq)}$.
- When OH^- is added the $HC_2H_3O_{2\,(aq)}$ is converted to $C_2H_3O_2^-{}_{(aq)}$ and water.
- This change will result in a small pH change. In effect, ethanoic acid removes OH^- from solution and ethanoate removes H^+ from solution and only a small pH change is observed.

The Capacity of a Buffer

- Buffering capacity is limited and with continued addition of acid or base, the buffer would eventually be used up and pH will jump dramatically.
- See figure 2 on page 616 for examples.

Buffers in Action

- Human blood has a pH of 7.4 and it must remain stable since many biological reactions occur at this pH. Some enzymes only work at specific pHs. In cells we find a buffer system comprised of $H_2PO_4^-{}_{(aq)}$ and $HPO_4^{2-}{}_{(aq)}$. In blood the buffer system is comprised of $H_2CO_{3(aq)}$ and $HCO_3^-{}_{(aq)}$.
- Interesting bit of trivia…if our blood was not buffered a glass of orange juice (citric acid) would be fatal.

Calculating pH in a Buffer

- A 1.0 L buffer is prepared that contains 0.20 mol/L acetic acid and 0.20 mol/L acetate at equilibrium.

(a) Calculate the pH of the buffer

(b) If 0.10 mol of $H^+_{(aq)}$ is added to the buffer without changing its volume, calculate the pH. (the volume change has no effect on overall calculation)

(c) Calculate and compare the change in pH expected if the same amount of $H^+_{(aq)}$ is added to water.

(a) $HC_2H_3O_{2\,(aq)} \leftrightarrow H^+_{(aq)} + C_2H_3O_2^-{}_{(aq)}$ $\qquad K_a = 1.8 \times 10^{-5}$

$$K_a = \frac{\left[H^+_{(aq)}\right]\left[C_2H_3O^-_{2(aq)}\right]}{\left[HC_2H_3O_{2(aq)}\right]}$$

rearrange

$$\left[H^+_{(aq)}\right] = \frac{K_a\left[HC_2H_3O_{2(aq)}\right]}{\left[C_2H_3O^-_{2(aq)}\right]}$$

$$\left[H^+_{(aq)}\right] = 1.8 \times 10^{-5}\, mol/L$$

$$pH = -\log\left[H^+_{(aq)}\right] = -\log\left(1.8 \times 10^{-5}\, mol/L\right) = 4.74$$

(b) The $H^+_{(aq)}$ will react with the acetate ions in this buffer.
$H^+_{(aq)} + C_2H_3O_2^-{}_{(aq)} \rightarrow HC_2H_3O_{2\,(aq)}$

In 1 L, $C_2H_3O_2^-{}_{(aq)}$ = 0.2 mol and $HC_2H_3O_{2\,(aq)}$ = 0.2 mol

By adding 0.1 mol of $H^+_{(aq)}$, 0.1 mol of $C_2H_3O_2^-{}_{(aq)}$ will be consumed and an additional 0.1 mol of $HC_2H_3O_{2\,(aq)}$ will be formed to make a total of 0.3 mol.

Therefore since there was not a volume change then we would have 0.1 mol/L of $C_2H_3O_2^-{}_{(aq)}$ and 0.3 mol/L of $HC_2H_3O_{2\,(aq)}$.

$$K_a = \frac{\left[H^+_{(aq)}\right]\left[C_2H_3O^-_{2(aq)}\right]}{\left[HC_2H_3O_{2(aq)}\right]} = 1.8\times10^{-5}$$
$$\left[H^+_{(aq)}\right] = \frac{K_a\left[HC_2H_3O_{2(aq)}\right]}{\left[C_2H_3O^-_{2(aq)}\right]}$$
$$\left[H^+_{(aq)}\right] = \frac{\left(1.8\times10^{-5}\right)(0.1)}{0.3} = 5.4\times10^{-5}\,mo/L$$
$$pH = -\log\left[H^+_{(aq)}\right] = -\log\left(5.4\times10^{-5}\,mol/L\right) = 4.27$$

a pH difference of 0.47

(c) In water there presence of 0.1 mol of $H^+_{(aq)}$ in 1 L of water will dramatically affect the pH.

$$pH = -\log\left[H^+_{(aq)}\right] = -\log\left(0.1mol/L\right) = 1.0$$

Pure water with a pH of 7.0 would drop to a pH of 1.0.

Homework

- Practice 1,2,3
- Questions 1,2,3,4,5,6,7,8,9