

Ch. 16.3: Buffers Part 2

- HW # 16-4: p. 722 #13, 16, 28, 37, 38

**16.13 The pH of a bicarbonate-carbonic acid buffer is 8.00. Calculate the ratio of carbonic acid ( $\text{H}_2\text{CO}_3$ ) to bicarbonate ( $\text{HCO}_3^-$ ) ion.**



We can use the Henderson-Hasselbalch equation to calculate the ratio  $[\text{HCO}_3^-]/[\text{H}_2\text{CO}_3]$ . The Henderson-Hasselbalch equation is:

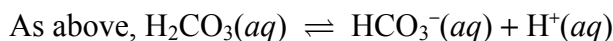
$$\text{pH} = \text{p}K_a + \log \frac{[\text{conjugate base}]}{[\text{acid}]}$$

$$K_{a_1} = 4.2 \times 10^{-7}; \text{p}K_{a_1} = 6.38$$

$$\text{pH} = 8.00 = \text{p}K_a + \log \frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]} = 6.38 + \log \frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]}$$

$$\log \frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]} = 1.62 \Rightarrow \frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]} = 41.7 \Rightarrow \frac{[\text{H}_2\text{CO}_3]}{[\text{HCO}_3^-]} = \mathbf{0.024}$$

**16.16 The pH of blood plasma is 7.40. Assuming the principal buffer system is  $\text{HCO}_3^-/\text{H}_2\text{CO}_3$ , calculate the ratio  $[\text{HCO}_3^-]/[\text{H}_2\text{CO}_3]$ . Is this buffer more effective against an added acid or an added base?**



$$\text{pH} = 7.40 = 6.38 + \log \frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]}$$

The [conjugate base]/[acid] ratio is:

$$\log \frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]} = 7.40 - 6.38 = 1.02 \Rightarrow \frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]} = 10^{1.02} = 1.0 \times 10^1$$

The buffer should be more effective against an added acid because ten times more base is present compared to acid. Note that a pH of 7.40 is only a two significant figure number; the final result should only have two significant figures.

**28 A solution is made by mixing exactly 500. mL of 0.167 M NaOH with exactly 500. mL of 0.100 M  $\text{CH}_3\text{COOH}$ . Calculate the equilibrium concentrations of  $\text{H}^+$ ,  $\text{CH}_3\text{COOH}$ ,  $\text{CH}_3\text{COO}^-$ ,  $\text{OH}^-$ , and  $\text{Na}^+$ .**

The resulting solution is not a buffer system. There is excess NaOH and the neutralization is well past the equivalence point.

$$\text{Moles NaOH} = 0.500 \text{ L} \times \frac{0.167 \text{ mol}}{1 \text{ L}} = 0.0835 \text{ mol}$$

$$\text{Moles CH}_3\text{COOH} = 0.500 \text{ L} \times \frac{0.100 \text{ mol}}{1 \text{ L}} = 0.0500 \text{ mol}$$

Note:  $\text{Na}^+$  is a spectator ion and has been removed for simplicity

$$\text{CH}_3\text{COOH}(aq) + \text{OH}^-(aq) \rightarrow \text{CH}_3\text{COO}^-(aq) + \text{H}_2\text{O}(l)$$

Initial (mol):	0.0500	0.0835	0
Change (mol):	-0.0500	-0.0500	+0.0500
Final (mol):	0	0.0335	0.0500

The volume of the resulting solution is 1.000 L (500. mL + 500. mL = 1000. mL = 1.000 L).

$$[\text{OH}^-] = \frac{0.0335 \text{ mol}}{1.00 \text{ L}} = \mathbf{0.0335 \text{ M}}$$

$$[\text{Na}^+] = \frac{(0.0335 + 0.0500) \text{ mol}}{1.00 \text{ L}} = \mathbf{0.0835 \text{ M}}$$

$$[\text{H}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{0.0335} = \mathbf{3.0 \times 10^{-13} \text{ M}}$$

$$[\text{CH}_3\text{COO}^-] = \frac{0.0500 \text{ mol}}{1.00 \text{ L}} = \mathbf{0.0500 \text{ M}}$$

$$\text{CH}_3\text{COO}^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{CH}_3\text{COOH}(aq) + \text{OH}^-(aq)$$

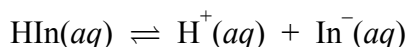
Initial (M):	0.0500	0	0.0335
Change (M):	-x	+x	+x
Equilibrium (M):	0.0500 - x	x	0.0335 + x

$$K_b = \frac{[\text{CH}_3\text{COOH}][\text{OH}^-]}{[\text{CH}_3\text{COO}^-]} = \frac{(x)(0.0335 + x)}{(0.0500 - x)} \approx \frac{(x)(0.0335)}{(0.0500)} = 5.6 \times 10^{-10}$$

$$x = [\text{CH}_3\text{COOH}] = \mathbf{8.4 \times 10^{-10} \text{ M}}$$

- 37  $K_a$  for indicator  $\text{HIn}$  is  $1.0 \times 10^{-6}$ . Its nonionized form is red, and its ionized form is yellow. What color is it in a solution of  $\text{pH} = 4.00$ ?

The weak acid equilibrium is



$$K_a = \frac{[\text{H}^+][\text{In}^-]}{[\text{HIn}]}, \text{ or } \frac{[\text{HIn}]}{[\text{In}^-]} = \frac{[\text{H}^+]}{K_a}$$

From the  $\text{pH}$ , we can calculate the  $\text{H}^+$  concentration.

$$[\text{H}^+] = 10^{-\text{pH}} = 10^{-4} = 1.0 \times 10^{-4} \text{ M}$$

$$\frac{[\text{HIn}]}{[\text{In}^-]} = \frac{[\text{H}^+]}{K_a} = \frac{1.0 \times 10^{-4}}{1.0 \times 10^{-6}} = \mathbf{100}$$

Since the concentration of  $\text{HIn}$  is 100 times greater than the concentration of  $\text{In}^-$ , the color of the solution will be that of  $\text{HIn}$ , the nonionized form. The color of the solution will be **red**.

**38** At what pH will HIn,  $K_a = 2.0 \times 10^{-6}$ , change from green (HIn) to red ( $\text{In}^-$ )?

The indicator color changes at this point. When  $[\text{HIn}] \approx [\text{In}^-]$  we can write:

$$\frac{[\text{In}^-]}{[\text{HIn}]} = \frac{K_a}{[\text{H}^+]} = 1, \text{ so } [\text{H}^+] = K_a = 2.0 \times 10^{-6}, \text{ and } \text{pH} = \mathbf{5.70}$$