

## Chapter 16: Buffer Calculations

### Example 1 - Buffer Method 1

What is the pH of a buffer made by adding 4.68g of sodium benzoate ( $\text{NaC}_7\text{H}_5\text{O}_2$ ) to 250.0mL of 0.15M benzoic acid solution?  $K_a = 6.5 \times 10^{-5}$

Need concentration of the conjugate base benzoate ion ( $\text{C}_7\text{H}_5\text{O}_2^-$ ):

$$(4.68\text{g NaC}_7\text{H}_5\text{O}_2) \left( \frac{1\text{mol NaC}_7\text{H}_5\text{O}_2}{144.1\text{g NaC}_7\text{H}_5\text{O}_2} \right) \left( \frac{1\text{mol C}_7\text{H}_5\text{O}_2^-}{1\text{mol NaC}_7\text{H}_5\text{O}_2} \right) = 0.325\text{mol C}_7\text{H}_5\text{O}_2^-$$

$$[\text{C}_7\text{H}_5\text{O}_2^-] = \frac{0.0325\text{mol C}_7\text{H}_5\text{O}_2^-}{0.25\text{L}} = 0.13\text{M C}_7\text{H}_5\text{O}_2^-$$

	$\text{HC}_7\text{H}_5\text{O}_2$	+	$\text{H}_2\text{O}$	$\rightleftharpoons$	$\text{H}_3\text{O}^+$	+	$\text{C}_7\text{H}_5\text{O}_2^-$
I					0		
C					+x		
E					x		

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{C}_7\text{H}_5\text{O}_2^-]}{[\text{HC}_7\text{H}_5\text{O}_2]}$$

CHECK: Henderson - Hasselbalch Equation

If assume x is small compared to initial concentrations of acid and conjugate base, one can use the initial values of the acid and conjugate base for equilibrium concentrations.

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

$$[\text{H}^+] = \frac{K_a[\text{HA}]}{[\text{A}^-]}$$

$$-\log[\text{H}^+] = -\log K_a - \log\left(\frac{[\text{HA}]}{[\text{A}^-]}\right)$$

$$\text{pH} = \text{p}K_a + \log\left(\frac{[\text{A}^-]}{[\text{HA}]}\right)$$

$$\text{pH} = \text{p}K_a + \log\left(\frac{[\text{base}]}{[\text{acid}]}\right)$$

$$\text{pH} =$$

### Example 2 - Buffer Method 2

What is the pH of a buffer made by adding 125mL of 0.14M HC<sub>7</sub>H<sub>5</sub>O<sub>2</sub> solution with 165mL of 0.16M NaC<sub>7</sub>H<sub>5</sub>O<sub>2</sub> solution?  $K_a = 6.5 \times 10^{-5}$

Need concentrations of the conjugate acid and base in the mixture.

	HC <sub>7</sub> H <sub>5</sub> O <sub>2</sub>	+	H <sub>2</sub> O	⇌	H <sub>3</sub> O <sup>+</sup>	+	C <sub>7</sub> H <sub>5</sub> O <sub>2</sub> <sup>-</sup>
I					0		
C					+x		
E					x		

$$K_a = \frac{[H_3O^+][C_7H_5O_2^-]}{[HC_7H_5O_2]}$$

CHECK:  $pH = pK_a + \log\left(\frac{[base]}{[acid]}\right) =$

### Example 3 – Buffer Range

At what pH does buffer work best? Best buffer system when  $pH = pK_a$

What is the range of a buffer? Buffer works within  $\pm 1$  of  $pK_a$  value so buffer  $pH = pK_a \pm 1$

What is the buffer range of benzoic acid/benzoate buffer?

$$pK_a = -\log(6.5 \times 10^{-5}) = 4.19 \quad \text{Buffer range} = 4.19 \pm 1 \quad \text{so} \quad 3.19 \text{ to } 5.19$$

What is the pH for best buffer capacity and what is the buffering range for each of the following?

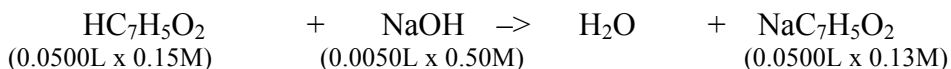
Buffer components	pH best buffering capacity	Buffer range
HC <sub>2</sub> H <sub>3</sub> O <sub>2</sub> / C <sub>2</sub> H <sub>3</sub> O <sub>2</sub> <sup>-</sup>		
HClO / ClO <sup>-</sup>		
HCO <sub>3</sub> <sup>-</sup> / CO <sub>3</sub> <sup>2-</sup>		
HCHO <sub>2</sub> / CHO <sub>2</sub> <sup>-</sup>		
NH <sub>4</sub> <sup>+</sup> / NH <sub>3</sub>		

Example 4 –Addition of base to buffer

What is the new pH if 5.0 mL of 0.50M NaOH is added to 50.0mL of benzoic acid-benzoate buffer from problem 1?

$$[\text{HC}_7\text{H}_5\text{O}_2] = 0.15\text{M} \quad [\text{C}_7\text{H}_5\text{O}_2^-] = 0.13\text{M}$$

First determine how much buffer reacts since NaOH is a strong base, NaOH will react 100% with the acid component of the buffer. MUST WORK STOICHIOMETRY IN MOLES.



Initial mol

Reaction

Final mol

Note: Must divide by the total volume before entering new values into ICE chart.

Total volume = 50.0 mL + 5.0 mL = 55.0 mL or 0.0550L

$$[\text{HC}_7\text{H}_5\text{O}_2]=$$

$$[\text{C}_7\text{H}_5\text{O}_2^-]=$$

	HC <sub>7</sub> H <sub>5</sub> O <sub>2</sub>	+	H <sub>2</sub> O	⇌	H <sub>3</sub> O <sup>+</sup>	+	C <sub>7</sub> H <sub>5</sub> O <sub>2</sub> <sup>-</sup>
I					0		
C					+x		
E					x		

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{C}_7\text{H}_5\text{O}_2^-]}{[\text{HC}_7\text{H}_5\text{O}_2]}$$

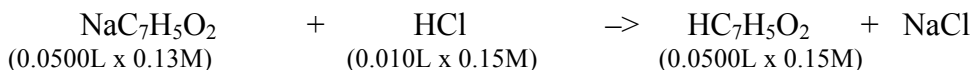
$$\text{CHECK: } \text{pH} = \text{p}K_a + \log\left(\frac{[\text{base}]}{[\text{acid}]}\right) =$$

Example 5 –Addition of acid to buffer

What is the new pH if 10.0 mL of 0.15M HCl was added to 50.0mL of benzoic acid-benzoate buffer from problem 1?

$$[\text{HC}_7\text{H}_5\text{O}_2] = 0.15\text{M} \quad [\text{C}_7\text{H}_5\text{O}_2^-] = 0.13\text{M}$$

First determine how much buffer reacts since HCl is a strong acid, HCl will react 100% with the base component of the buffer. MUST WORK STOICHIOMETRY IN MOLES.



Initial mol

Reaction

Final mol

Note: Must divide by the total volume before entering new values into ICE chart.

Total volume =

	HC <sub>7</sub> H <sub>5</sub> O <sub>2</sub>	+	H <sub>2</sub> O	⇌	H <sub>3</sub> O <sup>+</sup>	+	C <sub>7</sub> H <sub>5</sub> O <sub>2</sub> <sup>-</sup>
I					0		
C					+x		
E					x		

$$K_a = \frac{[H_3O^+][C_7H_5O_2^-]}{[HC_7H_5O_2]}$$

CHECK:  $pH = pK_a + \log\left(\frac{[base]}{[acid]}\right) =$

Example 6 – Depleting buffer

How much of the 0.15M HCl would have to be added to 50.0 mL of the benzoic acid/benzoate buffer to deplete it completely?

$$[HC_7H_5O_2] = 0.15M \quad [C_7H_5O_2^-] = 0.13M$$

Since an acid added to a buffer will act with the conjugate base, determine how much acid will be required to react with 50.0 mL of 0.13M C<sub>7</sub>H<sub>5</sub>O<sub>2</sub><sup>-</sup>.

	NaC <sub>7</sub> H <sub>5</sub> O <sub>2</sub>	+	HCl	→	HC <sub>7</sub> H <sub>5</sub> O <sub>2</sub>	+	NaCl
	(0.0500L x 0.13M)		(? L x 0.15M)		(0.0500L x 0.15M)		
Initial mol	0.0065 mol		0.0065 mol		0.0075 mol		
Reaction	<u>- 0.0065 mol</u>		<u>- 0.0065 mol</u>		<u>mol</u>		
Final mol	0 mol		0 mol		mol		

Buffer would be depleted when \_\_\_\_\_ mL of 0.15M HCl has been added.

At this point only acid HC<sub>7</sub>H<sub>5</sub>O<sub>2</sub> remains. Note: can NOT use Henderson-Hasselbalch equation since NO longer a buffer.

$$[HC_7H_5O_2] = \frac{\text{mol}}{L} =$$

$$\text{Volume} = 50. \text{ mL} + \text{ \_\_\_\_\_\_ mL} =$$

	HC <sub>7</sub> H <sub>5</sub> O <sub>2</sub>	+	H <sub>2</sub> O	⇌	H <sub>3</sub> O <sup>+</sup>	+	C <sub>7</sub> H <sub>5</sub> O <sub>2</sub> <sup>-</sup>
I					0		
C					+x		
E					x		

$$K_a = \frac{[H_3O^+][C_7H_5O_2^-]}{[HC_7H_5O_2]}$$

Likewise, similar calculation can be done for depletion of acid component of buffer.

#### Example 7 – Determining acid/base ratio needed for buffer pH

When need to know either

(A) how much of one of the buffer components to add to produce a certain pH or

(B) what ratio of conjugate base to acid is needed to produce a certain pH,

use Henderson-Hasselbalch equation.

How many grams of sodium benzoate (NaC<sub>7</sub>H<sub>5</sub>O<sub>2</sub>) must be added to 100. mL of 0.13M benzoic acid (HC<sub>7</sub>H<sub>5</sub>O<sub>2</sub>) to produce a buffer with a pH of 4.30?

$$K_a = 6.5 \times 10^{-5} \quad pK_a = -\log K_a = -\log(6.5 \times 10^{-5}) = 4.19$$

$$pH = pK_a + \log\left(\frac{[\text{base}]}{[\text{acid}]}\right) \quad \text{Calculate concentration of base needed to produce pH of 4.30.}$$

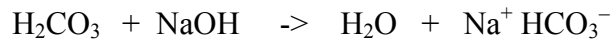
Then what mass of base NaC<sub>7</sub>H<sub>5</sub>O<sub>2</sub> is needed to produce desired concentration in 100. mL volume?

To prepare 100. mL of buffer at pH 4.30, measure out \_\_\_\_\_ g NaC<sub>7</sub>H<sub>5</sub>O<sub>2</sub> and add 0.13M acid HC<sub>7</sub>H<sub>5</sub>O<sub>2</sub> to total volume of 100. mL.

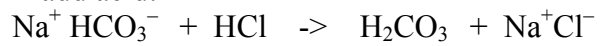
## Blood Buffer System

$\text{H}_2\text{CO}_3$  conjugate acid (formed when blood gas  $\text{CO}_2 + \text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{CO}_3$ )  
 $\text{Na}^+ \text{HCO}_3^-$  conjugate base

If add base:



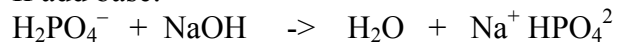
If add acid:



## Cell Buffer System

$\text{H}_2\text{PO}_4^-$  conjugate acid  
 $\text{HPO}_4^{2-}$  conjugate base

If add base:



If add acid:

