HOW DOES A BUFFER MAINTAIN PH?

A buffer is a special solution that stops massive changes in pH levels. Every buffer that is made has a certain buffer capacity, and buffer range. The buffer capacity is the amount of acid or base that can be added before the pH begins to change significantly. It can be also defined as the quantity of strong acid or base that must be added to change the pH of one liter of solution by one pH unit. The buffer range is the pH range where a buffer effectively neutralizes added acids and bases, while maintaining a relatively constant pH.

INTRODUCTION

The equation for pH also shows why pH does not change by much in buffers.

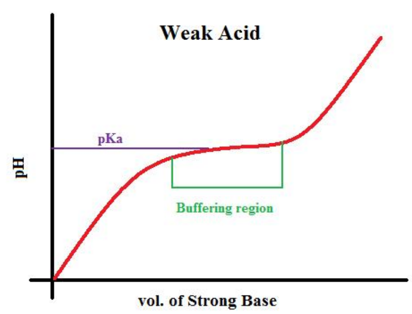
$$K_a = \frac{[H^+][A^-]}{[HA]} \tag{1}$$

$$pH = pK_a + \log \frac{[A^-]}{[HA]} \tag{2}$$

Where,

- A^- is the concentration of the conjugate base
- HA is the concentration of the acid

When the ratio between the conjugate base/ acid is equal to 1, the pH = pK_a. If the ratio between the two is 0.10, the pH drops by 1 unit from pK_a since log (0.10) = -1. If a ratio increases to a value of 10, then the pH increases by 1 unit since \log (10) = 1. The buffer capacity has a range of about 2. This means when a buffer is created, the pH can be changed by -1 by acid or +1 by base before the pH begins to change substantially. After the addition of base to raise the pH by 1 or more, most of the conjugate acid will have been depleted to try to maintain a certain pH, so the pH will be free to increase faster without the restraint of the conjugate acid. The same goes for the addition of acid, once the conjugate base has been used up, the pH will drop faster since most of the conjugate base has been used up.



Titration of a weak acid with a strong base

EXAMPLE 1

What is the effect on the pH of adding 0.006 mol HCL to 0.3L of a buffer solution that is 0.250M $HC_2H_3O_2$ and 0.560 M $NaC_2H_3O_2$? pK_a= 4.74

$$pH = 4.74 + \log \frac{0.560}{0.250} = 4.74 + 0.35 = 5.09 \tag{1}$$

$$C_2H_3O_2^- + H_3O^+ \rightleftharpoons HC_2H_3O_2 + H_2O \tag{2}$$

Calculate the starting amount of $C_2H_3O_2^-$

$$0.300\ L \times 0.560\ M = 0.168\ mol\ C_2H_3O_2^- \tag{3}$$

Calculate the starting amount of $HC_2H_3O_2$

$$0.300\ L \times 0.250\ M = 0.075\ mol\ HC_2H_3O_2 \tag{4}$$

	$C_2H_3O_2^-$	H_3O^-	$HC_2H_3O_2$
Original Buffer	0.168 mol		0.075 mol
Add		0.006 mol	
Change	-0.006 mol	-0.006 mol	+0.006 mol
Final Amount	0.162 mol		0.081 mol

Now calculate the new concentrations of $C_2H_3O_2^-$ and $HC_2H_3O_2$:

$$\frac{0.162\ mol}{0.300\ L} = 0.540\ M\ C_2H_3O_2^- \tag{5}$$

$$\frac{0.081\ mol}{0.300\ L} = 0.540\ M\ HC_2H_3O_2 \tag{6}$$

Using the new concentrations, we can calculate the new pH:

$$pH = 4.74 + \log \frac{0.540}{0.270} = 4.74 + 0.30 = 5.04 \tag{7}$$

Calculate the pH change:

$$pH_{final} - pH_{initial} = 5.04 - 5.09 = -0.05 \tag{8}$$

Therefore, the pH dropped by 0.05 pH units.

EXAMPLE 2

What is the effect on the pH of adding 0.006 mol NaOH to 0.3L of a buffer solution that is 0.250M $HC_2H_3O_2$ and 0.560 M $NaC_2H_3O_2$? pK_a= 4.74

$$pH = 4.74 + \log \frac{0.560}{0.250} = 4.74 + 0.35 = 5.09 \tag{9}$$

$$HC_2H_3O_2 + OH^- \rightleftharpoons C_2H_3O_2^- + H_2O \tag{10}$$

Calculate the starting amount of $HC_2H_3O_2$

$$0.300\ L \times 0.250\ M = 0.075\ mol\ HC_2H_3O_2 \tag{11}$$

Calculate the starting amount of $C_2H_3O_2^-$

$$0.300\ L \times 0.560\ M = 0.168\ mol\ C_2H_3O_2^- \tag{12}$$

	$HC_2H_3O_2$	OH^-	$C_2H_3O_2^-$
Original Buffer	0.075 mol		0.168 mol
Add		0.006 mol	
Change	-0.006 mol	-0.006 mol	+0.006 mol
Final Amount	0.069 mol		0.174 mol

Now calculate the new concentrations of $HC_2H_3O_2$ and $C_2H_3O_2$-:

$$\frac{0.069\ mol}{0.300\ L} = 0.230\ M\ HC_2H_3O_2 \tag{13}$$

$$\frac{0.174\ mol}{0.300\ L} = 0.580\ M\ C_2H_3O_2^- \tag{14}$$

Using the new concentrations, we can calculate the new pH:

$$pH = 4.74 + \log \frac{0.580}{0.230} = 4.74 + 0.40 = 5.14 \tag{15}$$

Calculate the pH change:

$$pH_{final} - pH_{initial} = 5.14 - 5.09 = +0.05 \tag{16}$$

Therefore, the pH increased by 0.05 pH units.

BUFFERS IN THE HUMAN BODY

Blood contains large amounts of carbonic acid, a weak acid, and bicarbonate, a base. Together they help maintain the bloods pH at 7.4. If blood pH falls below 6.8 or rises above 7.8, one can become sick or die. The bicarbonate neutralizes excess acids in the blood while the carbonic acid neutralizes excess bases.

Another example is when we consume antacids or milk of magnesia. After eating a meal with rich foods such as sea food, the stomach has to produce gastric acid to digest the food. Some of the acid can splash up the lower end of the esophagus causing a burning sensation. To relieve this burning, one would take an antacid, which when dissolved the bases buffer the excess acid by binding to them.

REFERENCES

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