

Lecture Notes – Ch 15.5 & 15.6 Std 5

Ch 15.5 – pH of strong Acids

For the preclass today we did these problems:

Page 506, problem 53 in class and you will see problems like this on the quiz. Here are the answers:

53. a. $[H^+] = \text{anti-log} [-8.51] = 3.1 \times 10^{-9} \text{ M}$
pOH = $14 - 8.51 = 5.49$
 $[OH^-] = \text{anti-log} [-5.49] = 3.2 \times 10^{-6} \text{ M}$
- b. $[OH^-] = \text{anti-log} [-9.39] = 4.1 \times 10^{-10} \text{ M}$
pOH = $14 - 9.39 = 4.61$
 $[H^+] = \text{anti-log} [-4.61] = 2.5 \times 10^{-5} \text{ M}$
- c. $[H^+] = \text{anti-log} [-2.54] = 2.9 \times 10^{-3} \text{ M}$
pOH = $14 - 2.54 = 11.46$
 $[OH^-] = \text{anti-log} [-11.46] = 3.5 \times 10^{-12} \text{ M}$
- d. $[OH^-] = \text{anti-log} [-4.82] = 1.5 \times 10^{-5} \text{ M}$
pOH = $14 - 4.82 = 9.18$
 $[H^+] = \text{anti-log} [-9.18] = 6.6 \times 10^{-10} \text{ M}$

FOR STRONG ACIDS

Concentrations of $[H^+]$ ions = concentrations given of strong acids

Example:

$HCl \rightarrow H^+ + Cl^-$ Notice !!! No backwards arrow for the equilibrium. This means that the **HCl ionizes 100% in water.**

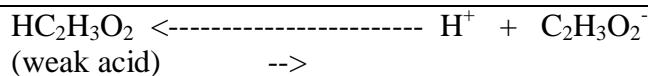
Next we did problem 57 on page 506, from the homework. As you see in each problem, the $[H^+]$ is always the same as the concentration of the strong acid. That is because all the HCl become H^+ and Cl^- ions. Here are the answers:

P. 506 problem 57

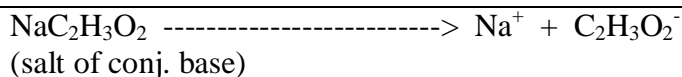
- a. $[H^+] = 1.04 \times 10^{-4} \text{ M}$ and to find the pH we take the $-\log [H^+] = 3.98$
- b. $[H^+] = 0.00301$ or $3.01 \times 10^{-3} \text{ M}$ and to find the pH we take the $-\log [H^+] = 2.5$
- c. $[H^+] = 5.41 \times 10^{-4} \text{ M}$ and to find the pH we take the $-\log [H^+] = 3.27$
- d. $[H^+] = 6.42 \times 10^{-2} \text{ M}$ and to find the pH we take the $-\log [H^+] = 1.19$

Ch 15.6 – Buffers

Buffers are created by **mixing a weak acid and the salt of its conjugate base.**



Notice the equilibrium favors the acid (left) very strongly. That is why it is a weak acid. Not many ions get formed because it doesn't ionize much.



Notice the salt ($\text{NaC}_2\text{H}_3\text{O}_2$) ionization is 100% , which means that all of the salt is ionized and none of the original $\text{NaC}_2\text{H}_3\text{O}_2$ is left. **This puts a lot of the Na^+ and $\text{C}_2\text{H}_3\text{O}_2^-$ ions into the water.**

After we put $\text{HC}_2\text{H}_3\text{O}_2$ and $\text{NaC}_2\text{H}_3\text{O}_2$ into the water, the ions and substances that can affect the pH are:



We now call this solution a buffer. This buffer will have a $\text{pH} = 4 (\pm 2)$ depending on the concentrations of the ions.

Let's see what happens when we add acids or bases to the buffer:

a. **ADD ACID**



The H^+ ions from the HNO_3 make the solution more acidic (lower pH), but they are quickly “grabbed” by the $\text{C}_2\text{H}_3\text{O}_2^-$ ions which act as “assassin” ions and hold on to the H^+ ions. This will form $\text{HC}_2\text{H}_3\text{O}_2$, which returns the pH to where it was. Because the H^+ isn't in the water any more, the $[\text{H}^+]$ returns to where it was before we put in the HNO_3 , which is $\text{pH} = 4$.

b. **ADD BASE**

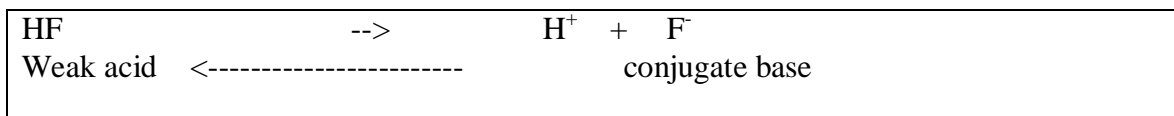


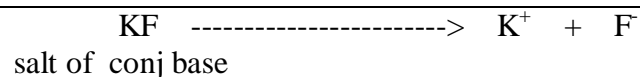
The OH^- ions act like “assassins” and “grab” the H^+ away from the $\text{HC}_2\text{H}_3\text{O}_2$ to form water. That takes the OH^- ions out of the water and makes more $\text{C}_2\text{H}_3\text{O}_2^-$. Adding the OH^- will at first make the solution more basic for a short time, but when this second reaction takes place the pH returns to $\text{pH} = 4$ because the OH^- ions are used up forming water.

Here are additional explanations of Buffers given in class just before Quiz 3.1

Buffer examples

1. Hydrofluoric acid (HF) a weak acid, and potassium fluoride (KF) the salt of the conjugate base.



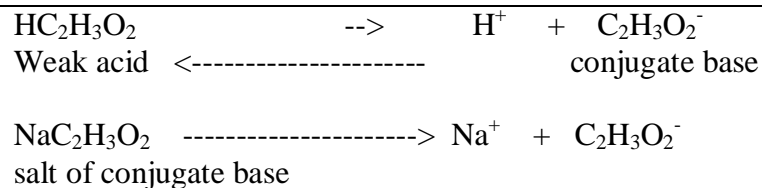


If we add OH^- (NaOH) - What takes out the OH^- so the pH doesn't change much?

If we add H^+ (HCl) – What takes out the H^+ so the pH doesn't change much?

Answers

2. Acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$), and sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$), the salt of the conjugate base.



If we add OH^- (NaOH) - What takes out the OH^- so the pH doesn't change much? If we add OH^- (NaOH) - What takes out the OH^- so the pH doesn't change much?

If we add H^+ (HCl) – What takes out the H^+ so the pH doesn't change much?

If we add H^+ (HCl) – What takes out the H^+ so the pH doesn't change much?

Answers

3. Can salts of strong acids act as buffers by themselves?

Examples: NaCl, KCl, CaCl_2 , MgCl_2 [Answers](#)

4. Here is an important **summary of buffer information**:

The a and b below explain how a buffer works to keep the pH relatively constant when acids or bases are added to their solutions.

- a. Buffers **resist change** in pH even if H^+ or OH^- ions are added from strong acids or bases.
- b. H^+ ions react with conjugate bases of weak acids. OH^- ions react with the H^+ ions from weak acid and take the hydrogen away from them.