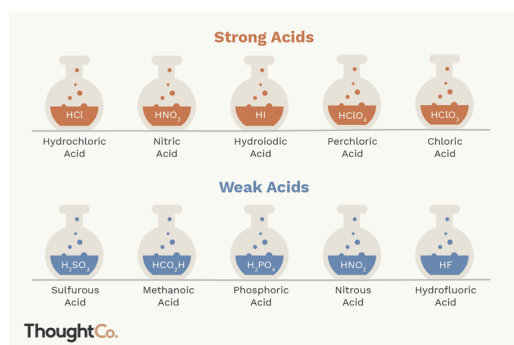


14.5 - pH of Weak Acid Solutions



Weak Acid Calculations

A weak acid dissolving and ionizing in water harkens back to equilibrium calculations with elements of the RICE method.

To solve these, we need to be aware of the major species in solution, and their K_a values.

Note: while water is a major species in any aqueous solution, the contribution of H^+ ions only becomes significant if the K_a value of the weak acid is extremely low (like E^{-12}).

1. HF Example

What are the **major species**, and what is the pH of a 1.00 M HF solution? $K_a = 7.2 E^{-4}$.

Water, and HF are the major species (low K_a).

First: the equilibrium reaction: $HF \rightleftharpoons H^+ + F^-$

Next, the equilibrium expression from which we can launch mathematics:

$$K_a = \frac{[H^+][F^-]}{[HF]}$$

Letting x equal the amount of dissociated HF, and recognizing that the concentration of HF vs. the low K_a , we can approximate the following:

$$K_a = \frac{[H^+][F^-]}{[HF]} \Rightarrow \frac{(x)(x)}{1.00 - x} \approx \frac{(x)(x)}{1.00}$$

$$7.2E-4 \approx x^2 \quad \Rightarrow \quad x \approx \sqrt{7.2E-4} = 2.7E-2 M$$

HF Answer

Finally, calculate pH:

$$pH = -\log[H^+] = -\log 2.7E-2 = \boxed{1.57}$$

HF Example Discussion

What if the concentration of HF had been much less, or if the K_a had been much greater?

In that case, we would not have been able to approximate that $1.00 - x$ was close enough to 1.00.

In such a case, we would have had to proceed through until we could use the quadratic formula (or graphing) to determine the concentration x .

A rule of thumb: if the error produced by these approximations are less than 5%, use them.

For our example:

$$\frac{x}{[HF]_0} = \frac{2.7E-2}{1.00} \cdot 100\% = 2.7\%$$

2 & 3: Weak Acid Mixtures

2. What are the **major species**, and calculate the pH of a solution containing 1.00 M HCN, and 5.00 M HNO₂.

3. Calculate the concentration of the cyanide ion at equilibrium.

Major species: HCN, HNO₂, H₂O.

First: K_a comparison from Table 14.1:

HCN: $K_a = 6.2 E^{-10}$; HNO₂: $K_a = 4.0 E^{-4}$.

HCN won't play a significant part in this problem.

2. pH Solution

Using the Equilibrium Expression for HNO_2 to solve for x , we get:

$$K_a = 4.0E-4 = \frac{[H^+][NO_2^-]}{[HNO_2]} \Rightarrow \frac{(x)(x)}{5.00-x} \approx \frac{(x)(x)}{5.00}$$

$$2.0E-3 \approx x^2$$

$$x \approx \sqrt{2.0E-3} = 4.5E-2 M$$

And pH:

$$pH = -\log[H^+] = -\log 4.5E-2 = \boxed{1.35}$$

3. $[\text{CN}^-]$ Solution

This gets interesting.

HCN has a low K_a , so contributes very little $[\text{H}^+]$.

Hence: $[\text{HCN}]_0 \approx [\text{HCN}] \approx 1.00 M$

However, it is the *only* source of CN^- ions, and its equilibrium expression must account for ALL H^+ IONS, even though the vast majority of them came from the dissociation of HNO_2 .

Therefore:

$$K_a = 6.2E-10 = \frac{[H^+][CN^-]}{[HCN]} \Rightarrow \frac{4.5E-2 \cdot [CN^-]}{1.00}$$

$$[CN^-] = \frac{6.2E-10 \cdot 1.00}{4.5E-2} = \boxed{1.4E-8 M}$$

 $[\text{CN}^-]$ Discussion

If it were only a 1.00 M solution of HCN , where $[\text{CN}^-]$ equaled the square root of K_a , we would get the usual:

$$K_a = 6.2E-10 = \frac{[H^+][CN^-]}{[HCN]} \approx \frac{x \cdot x}{1.00}$$

$$x = \sqrt{6.2E-10} = 2.5E-5 M$$

That the actual concentration of CN^- ions is so much less ($1.4E-8 M$) demands a look at the equilibrium environment: in the acid mixture, the abundance of hydrogen ions bombarding the cyanide ions forces that reaction back to the left, forming more HCN than if HCN were the lone acid, and there was more space for cyanide ions to exist.

Cray!

Percent Dissociation

It is often useful to specify the amount of weak acid that has been dissociated in achieving equilibrium.

Percent dissociation is given by:

$$\% \text{Dissociation} = \frac{\text{Amount Dissociated (M)}}{\text{Initial Concentration (M)}} \cdot 100\%$$

Earlier, we used HF in an example. For that problem, the percent dissociation is:

$$\% \text{Dissociation} = \frac{2.7E-2 M}{1.00 M} \cdot 100\% = \boxed{2.7\%}$$

4 & 5. Calculating % Dissociation

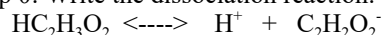
Calculate the percent dissociation of acetic acid in the following solutions:

4. 1.00 M $\text{HC}_2\text{H}_3\text{O}_2$

5. 0.100 M $\text{HC}_2\text{H}_3\text{O}_2$

Major Species: H_2O , $\text{HC}_2\text{H}_3\text{O}_2$.

Step 0: Write the dissociation reaction:



Strategy: First, use Table 14.1 to get K_a of acetic acid, then using the equilibrium expression solve for x (the low K_a value will allow the 5% rule to apply).

4. % Dissociation Solution

4. 1.00 M $\text{HC}_2\text{H}_3\text{O}_2$

First, solve for moles of dissociation:

$$K_a = 1.8E-5 = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} \Rightarrow \frac{(x)(x)}{1.00-x} \approx \frac{(x)(x)}{1.00}$$

$$1.8E-5 \approx x^2$$

$$x \approx \sqrt{1.8E-5} = 4.2E-3 M$$

Finally, percent dissociation:

$$\% \text{Dissociation} = \frac{4.2E-3 M}{1.00 M} \cdot 100\% = \boxed{0.42\%}$$

5. % Dissociation Solution5. 0.100 M HC₂H₃O₂

Moles dissociated:

$$K_a = 1.8E-5 = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} \Rightarrow \frac{(x)(x)}{0.100-x} \approx \frac{(x)(x)}{0.100}$$

$$1.8E-6 \approx x^2$$

$$x \approx \sqrt{1.8E-6} = 1.3E-3 M$$

Finally, percent dissociation:

$$\% \text{Dissociation} = \frac{1.3E-3 M}{0.10 M} \cdot 100\% = \boxed{1.3\%}$$

It is worth noting that as the concentration of a weak acid decreases, its % dissociation increases.

6. K_a From % DissociationLactic acid (HC₃H₅O₃) accumulates in muscle tissue during exertion. In a 0.100 M solution it is 3.7 % dissociated.What is the K_a of lactic acid?**K_a From % Dissociation Answer**Lactic acid is monoprotic, so its dissociation reaction is: HC₃H₅O₃ <----> H⁺ + C₃H₅O₃⁻

Using the % dissociation formula, we derive the hydrogen ion concentration, x:

$$\% \text{Dissociation} = \frac{\text{Amount Dissociated (M)}}{\text{Initial Concentration (M)}} \cdot 100\%$$

$$3.7\% = \frac{x}{0.100 M} \cdot 100\% \Rightarrow x = \frac{3.7}{100} \cdot 0.10 M = 3.7E-3 M$$

Finally, applying the equilibrium expression:

$$K_a = \frac{[H^+][C_3H_5O_3^-]}{[HC_3H_5O_3]} \Rightarrow \frac{(x)(x)}{0.100-x} \approx \frac{x^2}{0.100}$$

$$K_a = \frac{(3.7E-3)^2}{0.100} = \boxed{1.4E-4}$$

Homework

Preview 14.6 - 14.8

14.5 Problems in your Booklet

Due: Next Class