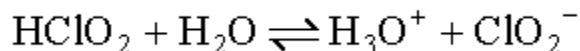


Weak Acid Calculation - When the Assumption Fails

Since the acid dissociates less than 5%, Δ is negligible. A good rule of thumb, if the value of the equilibrium constant is about 1×10^{-5} or less and the WA or WB concentration ≥ 0.01 M, the assumption is valid.

Calculate the equilibrium concentrations of H_3O^+ , HClO_2 , and ClO_2^- for a 0.10 M HClO_2 solution ($K_a = 1.1 \times 10^{-2}$)

1. Set up the reaction:



2. Set up the chart:

	$[\text{HClO}_2]$	$[\text{H}_3\text{O}^+]$	$[\text{ClO}_2^-]$
Initial:	0.10 M	0	0
Equilibrium:	$0.10 \text{ M} - \Delta$	Δ	Δ

$$\Delta = x$$

3. Write the K_a expression:

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{ClO}_2^-]}{[\text{HClO}_2]}$$

4. Substitute:

$$1.1 \times 10^{-2} = \frac{(\Delta)(\Delta)}{(0.10 - \Delta)} = \frac{\Delta^2}{(0.10 - \Delta)}$$

5. Assume that Δ is negligible when subtracted.

$$\frac{\Delta^2}{(0.10 - \Delta)} \rightarrow \frac{\Delta^2}{(0.10)}$$

6. Solve for Δ .

$$\frac{\Delta^2}{(0.10)} = 1.1 \times 10^{-2}$$

$$\Delta^2 = (0.10)(1.1 \times 10^{-2})$$

$$\Delta = \sqrt{(0.10)(1.1 \times 10^{-2})}$$

$$\Delta = 0.033 \text{ M}$$

7. Check the assumption.

$$\frac{\Delta}{\text{initial}} \times 100 = \frac{0.033 \text{ M}}{0.10 \text{ M}} \times 100 = \boxed{33\% > 5\%}$$

The assumption fails so we have to go back to the expression before the assumption:

$$1.1 \times 10^{-2} = \frac{\Delta^2}{(0.10 - \Delta)}$$

We have to now solve for Δ by using the quadratic formula. In order to use the quadratic formula, our equation has to be in the standard form:

$$ax^2 + bx + c = 0$$

So we have to rearrange our equation to fit the standard form:

$$1.1 \times 10^{-2} = 0.011 = \frac{\Delta^2}{(0.10 - \Delta)} \quad \text{Multiply both sides by } (0.10 - \Delta)$$

$$0.011(0.10 - \Delta) = \Delta^2 \quad \text{Distribute } 0.011 \text{ to both terms inside the parentheses}$$

$$0.0011 - 0.011\Delta = \Delta^2 \quad \text{Move all the terms on the left side of the equation to the right by adding the opposite of each term}$$

$$\Delta^2 + 0.011\Delta - 0.0011 = 0$$

$$a = 1 \quad b = 0.011 \quad c = -0.0011$$

The quadratic formula is:

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

Substitute in the values we found for a, b, and c and solve for Δ :

$$a = 1 \quad b = 0.011 \quad c = -0.0011$$

$$\Delta = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} = \frac{-0.011 + \sqrt{(0.011)^2 - 4(1)(-0.0011)}}{2(1)} = \boxed{0.028 \text{ M}}$$

According to our chart, $\Delta = [\text{H}_3\text{O}^+] = [\text{ClO}_2^-]$ and $[\text{HClO}_2] = 0.10 - \Delta$:

$$[\text{H}_3\text{O}^+] = [\text{ClO}_2^-] = \Delta = 0.028 \text{ M}$$

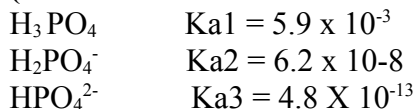
$$[\text{HClO}_2] = 0.10 - 0.028 = 0.072 \text{ M}$$

ex. What are the concentrations of all species present in a 0.50 M acetic acid solution?

The K_a is 1.7×10^{-5}

(p.8)

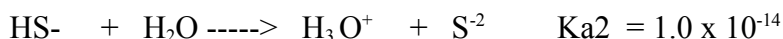
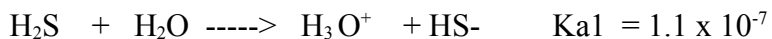
POLYPROTIC ACIDS: A weak acid that supplies two or more hydronium ions per molecule. (Refer to the dissociation WS for the dissociation reactions of these.)



The dissociation constants decrease with each successive step. It is easier to remove an H^+ from an uncharged species and becomes increasingly more difficult as the negative charge on the ions gets larger.

ex. Calculate the equilibrium concentrations of H_2S , H_3O^+ , HS^- , and S^{2-} present in a 0.10 M solution of H_2S .

(p.10)



Because $K_{a1} \gg K_{a2}$, we may assume nearly all of the H_3O^+ is derived from the 1st dissociation and that very little of HS^- formed in the 1st step will undergo further dissociation.