

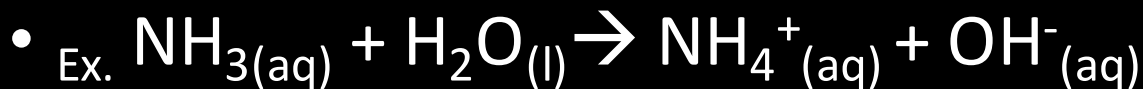
Weak acids and weak
bases calculations

At this point...

- We know how to do the following:
 - Write equilibrium expression
 - Understand strong acids and bases in terms of dissociation.
 - Calculate $[H^+]$ $[H_3O^+]$ or $[OH^-]$ from a strong acid/base dissociation with K_w .
 - Calculate the pH from the above concentrations.

- In addition to the K_w (water's ionization constant), acids and bases also have their own.
 - K_a is the equilibrium constant for acid (refer to table).
 - The higher the K_a , the more dissociation will occur, the stronger the acid.
- /5
- HCl has a K_a of 1.3×10^6 while boric acid has a K_a of 5.8×10^{-10}

- Bases also have their own equilibrium constant, which is K_b
- The higher the K_b value, the stronger the base, or weaker the acid.
- K_b uses the hydroxide ion in the calculation.



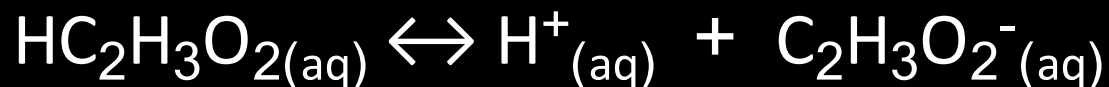
- $$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$

Weak acid and base calculations

- The only difference between weak acid and base calculations is that there is NO complete dissociation.
- This results in having the reactants in the calculation.
- Check out the example...

Ex. Calculate the hydrogen ion concentration in a 0.10 M acetic acid solution, $\text{HC}_2\text{H}_3\text{O}_2$. K_a for acetic acid, a weak acid, is 1.8×10^{-5} .

1. Write out the dissociation equation



Use K_a to solve for $[\text{H}^+]$ or $[\text{OH}^-]$

Ex. Calculate the hydrogen ion concentration in a 0.10 M acetic acid solution, $\text{HC}_2\text{H}_3\text{O}_2$. K_a for acetic acid, a weak acid, is 1.8×10^{-5} .

2. Use your K_a and write the equilibrium expression

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

We **know** the concentration of acetic acid from the question, but we **do not** know the concentration of the products because acetic acid does not fully dissociate.

Ex. Calculate the hydrogen ion concentration in a 0.10 M acetic acid solution, $\text{HC}_2\text{H}_3\text{O}_2$. K_a for acetic acid, a weak acid, is 1.8×10^{-5} .

3. Use algebra for the unknowns (products)

$$1.8 \times 10^{-5} = \frac{[x][x]}{[0.10\text{M}]}$$

$$x = 1.3 \times 10^{-3} \text{ M}$$

The 'x' is the concentration of $[\text{H}^+]$

Ex 2. Calculate the hydroxide ion concentration, $[OH^-]$, in a 0.025 M solution of analine, $C_6H_5NH_2$, a weak base with $K_b = 4.3 \times 10^{-10}$



$$\begin{aligned} K_b &= \frac{[C_6H_5NH^+][OH^-]}{[C_6H_5NH_2]} \\ 4.3 \times 10^{-10} &= \frac{(x)(x)}{0.025} \\ x &= 3.3 \times 10^{-6} \text{ mol/L} \end{aligned}$$

$$x = [OH^-]$$

Steps:

1. Write the balance equation out.
2. Use K_a or K_b equation to solve for $[H^+]$ or $[OH^-]$

Extension to these problems?

- Calculate the pH, pOH, [OH⁻] or [H⁺].
- From last question:

$$x = 3.3 \times 10^{-6} \text{ mol/L}$$

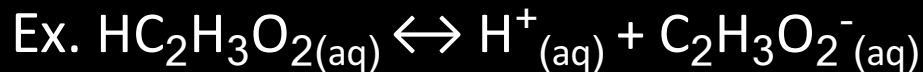
- [OH⁻] is $3.3 \times 10^{-6} \text{ mol/L}$
- So, we can calculate [H⁺] from the Kw.
 - $K_w = [H^+] (3.3 \times 10^{-6} \text{ M})$
 - Solve for [H⁺]
- Once you have [H⁺], you can calculate pH and everything else!

Extension to these problems?

Calculating % dissociated

- This is basically calculating how much of the acid was dissociated compared to the original.
- Formula:

$$\% \text{ dissociated} = \frac{[\text{dissociated ion}]}{[\text{starting acid/base}]} \times 100$$



$$\% \text{ dissociated} = \frac{[\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]} \times 100$$

Extension to these problems?

Ex. Calculate the $[\text{H}_3\text{O}^+]$ of a 0.38 mol/L weak acid that is dissociated 0.12%

- We know that this is a weak acid question.
- We know the total concentration of the original acid.
- We know 0.12% of the total concentration was dissociated.

Extension to these problems?

Ex. Calculate the $[\text{H}_3\text{O}^+]$ of a 0.38 mol/L weak acid that is dissociated 0.12%

From the 0.12%, we can calculate $[\text{H}_3\text{O}^+]$

$$0.12 \% = \frac{[\text{H}_3\text{O}^+]}{[0.38]} \times 100$$

$$[\text{H}_3\text{O}^+] = 4.56 \times 10^{-4} \text{M}$$

Extension to these problems?

Ex 2. If $[\text{H}_3\text{O}^+] = 4.5 \times 10^{-6} \text{ mol/L}$ in a 0.45 mol/L solution of the weak acid HX, calculate percent dissociation.

- We know that this is a weak acid question.
- We know the total concentration of the original acid.
- We know the concentration of the ion dissociated.

$$\begin{aligned} \text{\% dissociated} &= \frac{[4.5 \times 10^{-6}]}{[0.45\text{M}]} \times 100 \\ &= 0.0010\% \end{aligned}$$