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 Kw.
  - Calculate the pH from the above concentrations.

- In addition to the Kw (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.
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- 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.

- Ex. $NH_{3(aq)} + H_2O_{(l)} \rightarrow NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$

- $K_b = \dfrac{[NH_4^+][OH^-]}{[NH_3]}$

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, $HC_2H_3O_2$. K_a for acetic acid, a weak acid, is 1.8×10^{-5}.

1. Write out the dissociation equation

$$HC_2H_3O_{2(aq)} \leftrightarrow H^+_{(aq)} + C_2H_3O_2^-{}_{(aq)}$$

Use Ka to solve for [H+] or [OH-]

Ex. Calculate the hydrogen ion concentration in a 0.10 M acetic acid solution, $HC_2H_3O_2$. K_a for acetic acid, a weak acid, is 1.8×10^{-5}.

2. Use your K_a and write the equilibrium expression

$$Ka = \frac{[H^+]\,[C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

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, $HC_2H_3O_2$. K_a for acetic acid, a weak acid, is 1.8×10^{-5}.

3. Use algebra for the unknowns (products)

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

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

The 'x' is the concentration of $[H^+]$

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

- 1. $C_6H_5NH_2 + H_2O \leftrightarrow C_6H_5NH^+ + OH^-$

$$K_b = \frac{[C_6H_5NH^+]\,[OH^-]}{[C_6H_5NH_2]}$$

$$4.3 \times 10^{-10} = \frac{(x)\,(x)}{0.025}$$

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

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:

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

- [OH-] is 3.3×10^{-6} mol/L
- So, we can calculate [H+] from the Kw.
  - Kw = [H+] (3.3×10^{-6}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:

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

Ex. $HC_2H_3O_{2(aq)} \leftrightarrow H^+_{(aq)} + C_2H_3O_2^-{}_{(aq)}$

$$\% \text{ dissociated} = \frac{[C_2H_3O_2^-]}{[HC_2H_3O_2]} \times 100$$

Extension to these problems?

Ex. Calculate the $[H_3O^+]$ 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 [H_3O^+] of a 0.38 mol/L weak acid that is dissociated 0.12%

From the 0.12%, we can calculate [H_3O^+]

$$0.12\ \% = \frac{[H_3O^+]}{[0.38]} \times 100$$

[H_3O^+] = 4.56×10^{-4}M

Extension to these problems?

Ex 2. If $[H_3O^+] = 4.5 \times 10^{-6}$ 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.

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

$$= 0.0010\%$$