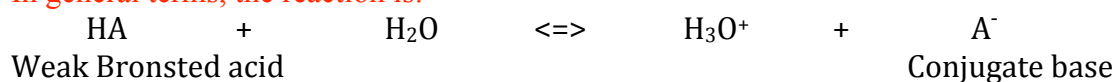


Chapter 15 Weak Acids & Bases

Equilibrium Expression for Acids

In general terms, the reaction is:



$$K_c = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{H}_2\text{O}][\text{HA}]}$$

$$K_c[\text{H}_2\text{O}] = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

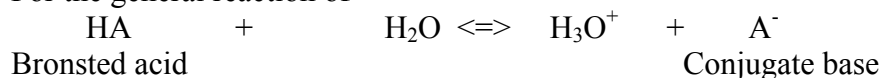
Sometimes reaction is shortened to: $\text{HA} \rightleftharpoons \text{H}^+ + \text{A}^-$

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

Better way to classify acid strength is by K_a .

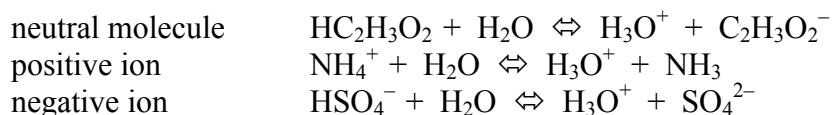
| | |
|-------------------------------|---------------|
| $K_a < 10^{-3}$ | Weak acid |
| $K_a = 1 \text{ to } 10^{-3}$ | Moderate acid |
| $K_a > 1$ | Strong acid |

For the general reaction of



the "A" denotes any acid; that is, particle with ability to donate a proton, so "A" could be:

Examples:



The stronger the acid, the weaker is the conjugate base and vice-versa.

Intersection
 $x = 0.0040 \text{ M}$

$$[\text{H}_3\text{O}^+] = [\text{F}^-] = x = 0.0040 \text{ M}$$

$$[\text{HF}] = 0.050 - x = 0.050 \text{ M} - 0.0040 \text{ M} = 0.046 \text{ M}$$

$$\text{pH} = -\log([\text{H}_3\text{O}^+]) = -\log(x) = -\log(0.0040) = 2.40$$

$$\text{pOH} = 14.00 - \text{pH} = 14.00 - 2.40 = 11.60$$

$$\% \text{ ionization} = \frac{[\text{H}_3\text{O}^+]}{[\text{HF}]_{\text{INITIAL}}} = \frac{x}{[\text{HF}]} = \frac{0.0040 \text{ M}}{0.050 \text{ M}} \times 100 = 8.0\%$$

Calculations - Weak Base Equilibria

Calculate the pH of a 0.015M NH_3 solution. $K_b = 1.8 \times 10^{-5}$
 $\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$

| | | | | | | | |
|---|---------------|---|----------------------|----------------------|-----------------|---|---------------|
| | NH_3 | + | H_2O | \rightleftharpoons | NH_4^+ | + | OH^- |
| I | 0.015 M | | | | 0 | | ~ 0 |
| C | -x | | | | +x | | +x |
| E | 0.015 - x | | | | x | | x |

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

$$1.8 \times 10^{-5} = \frac{x^2}{(0.015-x)}$$

$$0 = \frac{x^2}{(0.015-x)} - 1.8 \times 10^{-5} \quad \text{Window } x=0 \text{ to } 0.015$$

Intersection
 $x = 5.1 \times 10^{-4} \text{ M}$

$$[\text{OH}^-] = x = 5.1 \times 10^{-4} \text{ M}$$

$$\text{pOH} = -\log([\text{OH}^-]) = -\log(x) = -\log(5.1 \times 10^{-4}) = 3.29$$

$$\text{pH} = 14.00 - \text{pOH} = 14.00 - 3.29 = 10.71$$

Calculations - Weak Acid/Base Equilibria

pH of a salt solution

Calculate the pH of a 0.15 M KCN solution.

| | |
|----------------|-----------------|
| K ⁺ | CN ⁻ |
| A | B |

CN⁻ is the conjugate base of the weak acid HCN. $K_a = 6.2 \times 10^{-10}$

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{6.2 \times 10^{-10}} = 1.6 \times 10^{-5}$$

| | | |
|---|--------------------------------------|-----------------------|
| | CN ⁻ + H ₂ O ⇌ | HCN + OH ⁻ |
| I | 0.15 M | 0 ~ 0 |
| C | - x | + x + x |
| E | 0.15 - x | x x |

$$K_b = \frac{[\text{HCN}][\text{OH}^-]}{[\text{CN}^-]}$$

$$1.6 \times 10^{-5} = \frac{x^2}{(0.15 - x)}$$

$$0 = \frac{x^2}{(0.15 - x)} - 1.6 \times 10^{-5} \quad \text{Window } x = 0 \text{ to } 0.15$$

Intersection
 $x = 0.0015 \text{ M}$

$$[\text{OH}^-] = x = 0.0015 \text{ M}$$

$$\text{pOH} = -\log([\text{OH}^-]) = -\log(x) = -\log(0.0015) = 2.81$$

$$\text{pH} = 14.00 - \text{pOH} = 14.00 - 2.81 = 11.19$$