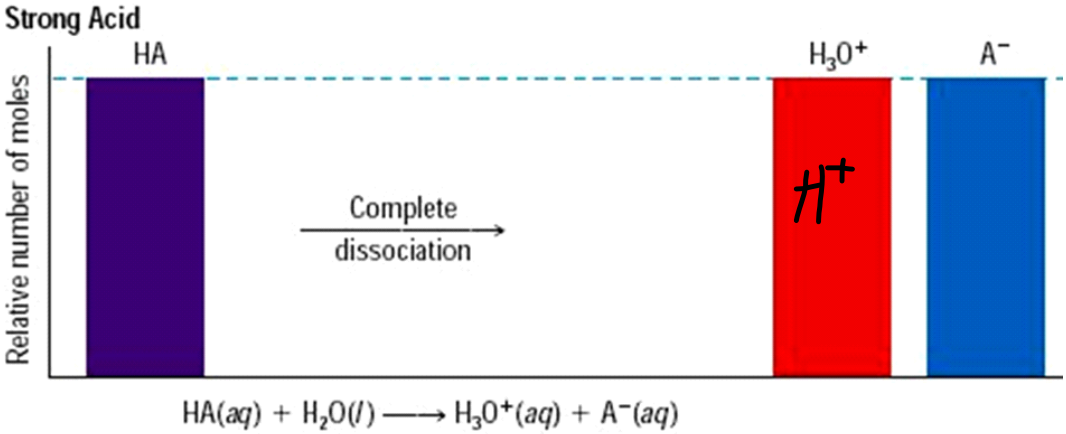
Strong Acids:

Strong acids are completely, 100% (quantitatively) ionized in aqueous solution. They have a high $[H^+]$ and low pH. There are only six strong acids.

$$HNO_3 \rightarrow H^+ + NO_3^-$$



1. Calculate the pH and pOH of a 0.100 M solution of hydrochloric acid.

$$HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$$
$$0.100M \quad 0.100M$$

$pH = -\log[H^+] = -\log[0.100] = 1.000$

$pOH = 14.00 - pH = 13.000$

2. Calculate the pH of a 5.0×10^{-2} M nitric acid solution.

$$HNO_{3(aq)} \rightarrow H^+_{(aq)} + NO_3^-_{(aq)}$$
$$5.0\times10^{-2}M \quad 5.0\times10^{-2}M$$

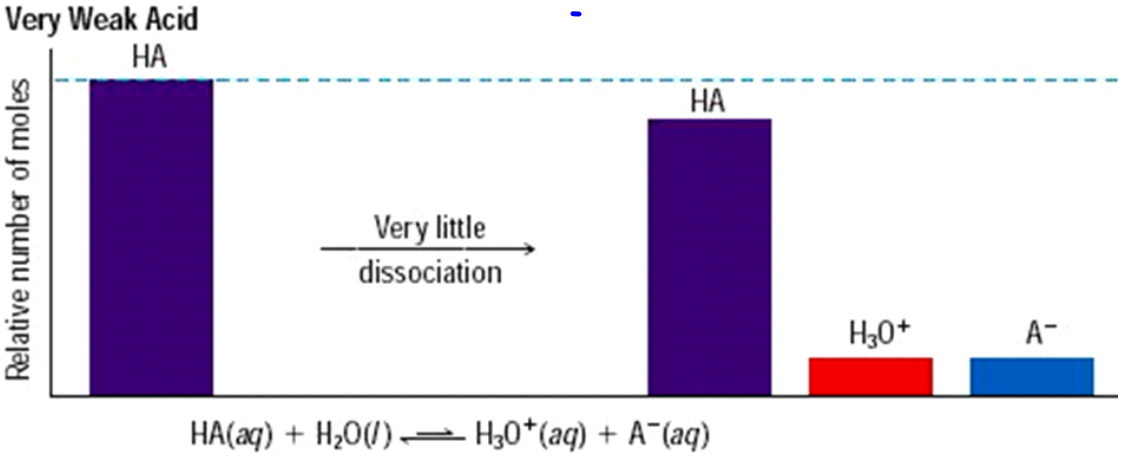
$pH = -\log[5.0\times10^{-2}M] = 1.30$

Weak Acids:

Weak acids are less than 100 % dissociated, all other acids outside of the top 6 are weak!

– Their general formula is $HA_{(aq)}$.

Ionization Equation: $HA_{(aq)} \leftrightarrows H^{+}_{(aq)} + A^{-}_{(aq)}$



Percent Ionization / Percent Reaction

p = [H+] x 100%
% [HA]

[H⁺] = hydrogen ion concentration,
[HA]= acid concentration

1. A 0.10 M solution of acetic acid is 1.3 % dissociated. Calculate the pH.

$CH_3COOH_{(aq)} \rightleftharpoons H^+_{(aq)} + CH_3COO^-_{(aq)}$

0.10M 1.3% 0.0013M

$\frac{1.3\%}{100} = \frac{[H^+]}{0.10M}$ $[H^+] = 0.013 \cdot 0.10M = 0.0013M$

$pH = -\log[0.0013] = 2.89$

2. The pH of a 0.10 M methanoic acid solution is 2.38. Calculate the percent ionization.

$HCOOH_{(aq)} \rightleftharpoons H^+_{(aq)} + COOH^-_{(aq)}$

0.10M 4.2×10^{-3}

$P = \frac{[H^+]}{[HA]} \times 100$ $[H^+] = 10^{-2.38}$ $[H^+] = 4.2 \times 10^{-3} \frac{mol}{L}$

$P = \frac{4.2 \times 10^{-3}}{0.10} \times 100$

$P = 4.2\%$

Acid Ionization Constant

Weak acids have a small Ka value; strong acids have a larger Ka due to the higher [H⁺].

*Always use Ka for weak acids!

$K_a = \frac{[H^+][A^-]}{[HA]}$ $K_a = \frac{[H^+]^2}{[HA]}$

$\sqrt{[H^+]^2} = \sqrt{K_a \cdot [HA]}$

$[H^+] = \sqrt{K_a \cdot [HA]}$

Diprotic and triprotic acids lose their hydrogens one at a time. Each ionization has a different ionization constant.

$H_3PO_4 \rightleftharpoons H^+ + H_2PO_4^-$

$H_2PO_4^- \rightleftharpoons H^+ + HPO_4^{2-}$

diff K_a

1. What is the pH of a 0.100 M solution of acetic acid?

$CH_3COOH_{(aq)} \rightleftharpoons H^+ + CH_3COO^-_{(aq)}$

0.100M 1.34×10^{-3}

$[H^+] = \sqrt{K_a \cdot [HA]} = \sqrt{1.8 \times 10^{-5} \frac{mol}{L} \cdot 0.100 \frac{mol}{L}} = 1.34 \times \frac{mol}{L}$

$pH = -\log[1.34 \times 10^{-3}] = 2.872$

2. The acid dissociation constant for benzoic acid is 6.3×10^{-5}. What is the pH of a 0.300M solution of benzoic acid.

$C_6H_5COOH_{(aq)} \rightleftharpoons H^+_{(aq)} + C_6H_5COO^-_{(aq)}$

0.300M

$[H^+] = \sqrt{K_a \cdot [HA]} = \sqrt{6.35 \times 10^{-5} \frac{mol}{L} \cdot 0.300 \frac{mol}{L}}$

$[H^+] = 4.35 \times 10^{-3} \frac{mol}{L}$

$pH = -\log[4.35 \times 10^{-3}] = 2.362$

3. 1.35g of hydrogen cyanide gas is dissolved in 500. mL. What is the pH of the final solution.

$HCN_{(aq)} \rightleftharpoons H^+_{(aq)} + CN^-_{(aq)}$

~~0.0999M~~ ~~$7.87 \times 10^{-6} \frac{mol}{L}$~~

① moles of HCN

$n = \frac{m}{M} = \frac{1.35g}{27.03g/mol} = 0.0499 mol$

② Conc. of HCN

$c = \frac{n}{V} = \frac{0.0499 mol}{0.500L} = 0.0999M$

③ $[H^+]$ using Ka

$[H^+] = \sqrt{K_a [HA]} = \sqrt{6.2 \times 10^{-10} \frac{mol}{L} \cdot 0.0999 \frac{mol}{L}}$

$[H^+] = 7.87 \times 10^{-6} \frac{mol}{L}$

$pH = -\log[7.87 \times 10^{-6}]$

$pH = 5.104$

STRONG BASES

Strong bases are all highly soluble ionic hydroxides. Example: NaOH, LiOH, KOH… Group 1 & 2.

$NaOH_{(s)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$ 100% dissociation

5.0×10^{-2} 5.0×10^{-2}

Example:

1. Calculate the pH of a 5.0×10^{-2} mol/L NaOH solution.

$pOH = -\log[5.0 \times 10^{-2}] = 1.30$

$pH = 14.00 - 1.30 = 12.70$