# ACID-BASE EQUILIBRIUM IN WATER

- Like other ELECTROLYTES, acids and bases IONIZE to some extent in water

- STRONG electrolytes ionize completely. Acids and bases that ionize completely in water are called STRONG ACIDS and STRONG BASES

- WEAK electrolytes ionize partially, remaining mostly non-ionized. Acids and bases that ionize only partially in solution are called WEAK ACIDS and WEAK BASES.

- Most acids and bases are WEAK!

Common strong acids

$HCl$

$HNO_3$

$H_2SO_4$ (only 1st proton)

$HBr$

$HI$

Common strong bases

$NaOH$, $KOH$ ] alkali metal hydroxides (Group IA)

$Ca(OH)_2$

# SIMPLE pH CALCULATIONS: STRONG ELECTROLYTES

- With strong acids and bases, the acid or base completely ionizes in water. So, we only have to worry about the effect of the acid or base on the water equilibrium itself.

- Since the equilibrium constant for the self-ionization of water is so small, the strong acid or base will overpower the hydronium (for acids) or hydroxide (for bases) produced by the water.

$$2H_2O \rightleftharpoons H_3O^+ + OH^-$$

Consider 0.025 M $HNO_3$

Assume all $H_3O^+$ comes from acid!

$$HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$$

So, $[H_3O^+]$ = original $[HNO_3]$ = 0.025 M

$$pH = -\log_{10}[H_3O^+] = -\log_{10}(0.025) = \boxed{1.60}$$

For logarithms, the places after the decimal are counted as significant. The places before are merely the exponent.

What would the pOH be?

$$pH + pOH = 14.00$$

$$pOH = 12.40$$

... but we usually don't care what pOH is.

# Consider 0.0125 M NaOH

$$2H_2O \rightleftharpoons H_3O^+ + OH^-$$

Assume all $OH^-$ comes from base

$$NaOH \rightarrow Na^+ + OH^-$$

$$[OH^-] = \text{original } [NaOH] = 0.0125\ M$$

$$pOH = -\log_{10}[OH^-] = -\log_{10}(0.0125) = 1.90$$

... now change to pH so we can compare this to the acid problem we just worked

$$pH + pOH = 14.00$$

$$pH + 1.90 = 14.00$$

$$\boxed{pH = 12.10}$$

Let's find the concentration of the hydronium ion, since that will equal the amount of hydroxide produced by the water. (This should be a lot smaller than the 0.0125 M hydroxide from the base!)

$$[H_3O^+] = 10^{-12.10} = 7.9 \times 10^{-13}\ M$$

... and it IS much smaller than 0.0125!

(A) What is the concentration of hydronium ion in an aqueous solution whose pH is 10.50? (B) What is the hydroxide ion concentration? (C) What molar concentration of sodium hydroxide solution would provide this pH?

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A) pH = 10.50   $[H_3O^+] = ?$   $[H_3O^+] = 10^{-pH}$

$[H_3O^+] = 10^{-10.50} = 3.16 \times 10^{-11}\ M = [H_3O^+]$

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B) $2H_2O \rightleftharpoons H_3O^+ + OH^-$ ; $K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$

$1.00 \times 10^{-14} = (3.16 \times 10^{-11})[OH^-]$

$3.16 \times 10^{-4}\ M = [OH^-]$

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C) Sodium hydroxide (NaOH) is a STRONG BASE which completely ionizes in water, so:

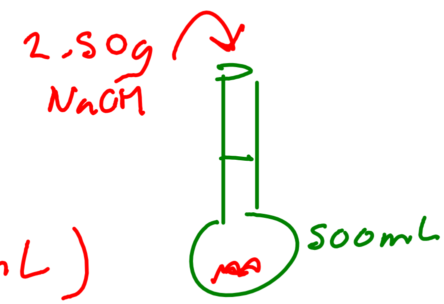
$NaOH \rightarrow Na^+ + OH^-$

1:1 ratio

$3.16 \times 10^{-4}\ M$ NaOH

What is the pH of a sodium hydroxide solution made from dissolving 2.50 g of sodium hydroxide in enough water to make 500.0 mL of solution?

$NaOH: 40.00\ g/mol$




Find the molarity of the sodium hydroxide.

$$M = \frac{\text{moles NaOH}}{L} \leftarrow 0.5000\ L\ (500.0\ mL)$$

$$2.50\ g\ NaOH \times \frac{mol\ NaOH}{40.00\ g\ NaOH} = 0.0625\ mol\ NaOH$$

$$M = \frac{0.0625\ mol\ NaOH}{0.5000\ L} = 0.125\ \underline{M}\ NaOH$$

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Sodium hydroxide is a strong base, so the hydroxide concentration equals the sodium hydroxide concentration:

$$\underline{NaOH} \rightarrow Na^+ + \underline{OH^-}$$

$$[OH^-] = 0.125\ M$$

$$[H_3O^+][OH^-] = 1.00 \times 10^{-14}$$

$$[H_3O^+](0.125) = 1.00 \times 10^{-14}$$

$$[H_3O^+] = 8.00 \times 10^{-14}$$

$$pH = 13.10$$

OR

$$-\log_{10}(0.125) = pOH = 0.90$$

$$pH + pOH = 14.00$$

$$pH + 0.90 = 14.00$$

$$pH = 13.10$$

For a WEAK ACID, equilibrium does not lie far to the right. The ionization equilibrium of the acid itself is important!

$$HA + H_2O \rightleftharpoons H_3O^+ + A^-$$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

Again, water's concentration will not change significantly, so it is folded into the ionization constant

acid ionization constant

(HA) = concentration of undissociated acid

For a WEAK BASE, equilibrium does not lie far to the right. The ionization equilibrium of the base itself is important!

$$B + H_2O \rightleftharpoons BH^+ + OH^-$$

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

base ionization constant

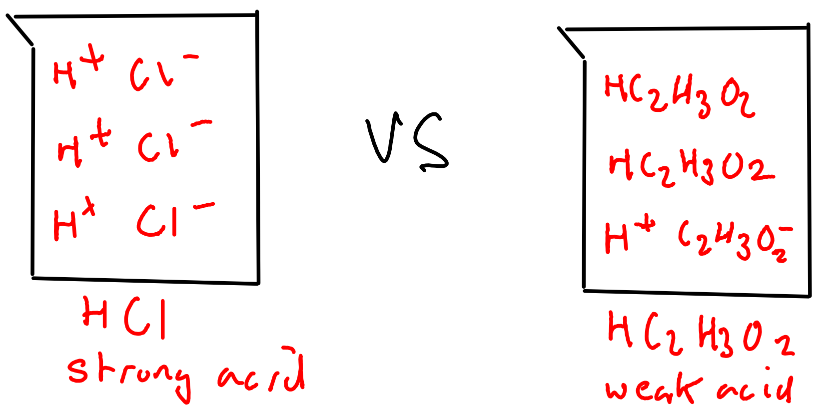
Values for Ka and Kb can often be found in data books / tables / or on the web.

In Ebbing, this data is in the appendices, on pages A-13 and A-14

# WEAK ELECTROLYTES

- In solutions of weak acids or bases, the UNDISSOCIATED form is present in significantly high concentration.

- The pH of a solution of weak acid will be HIGHER than the pH of a strong acid solution with the same nominal concentration!




Fewer molecules of the weak acid ionize, so the concnetration of hydrogen/hydronium ion is lower, meaning a higher pH!

- The pH of a solution of weak base will be LOWER than the pH of a strong base solution with the same nominal concentration!

Consider a 0.100M solution of the WEAK ACID $HNO_2$

$$HNO_2 + H_2O \rightleftharpoons H_3O^+ + NO_2^-$$

$$K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]} = 5.1 \times 10^{-4}$$

values for Ka are determined experimentally

What is the pH of the solution?

To find the pH, we need to determine the concnetration of hydronium, $[H_3O^+]$

... so we need to solve the equilibrium expression. But we don't know all of the concentrations AT EQUILIBRIUM to do so!

... but they ARE related!

We assume the amount of hydronium from the water is small enough to ignore

| SPECIES | INITIAL CONC | CHANGE | EQUILIBRIUM CONC |
|---|---|---|---|
| $[H_3O^+]$ | 0 | $+X$ | $X$ |
| $[NO_2^-]$ | 0 | $+X$ | $X$ |
| $[HNO_2]$ | 0.100 | $-X$ | $0.100 - x$ |

... this is similar to the problems from the equilibrium chapter!

$$5.1 \times 10^{-4} = \frac{[H_3O^+][NO_2^-]}{[HNO_2]}$$

$$5.1 \times 10^{-4} = \frac{(x)\ (x)}{(0.100 - x)}$$

$$5.1 \times 10^{-4} = \frac{x^2}{0.100 - x}$$

Quadratic equation!

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

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

Assume that $x << 0.100$

What's this?
For situations where the amount of dissociated acid or base is much, much smaller than the original amount, it's safe to assume that the amount of undissociated acid remains relatively constant.

In this case, 0.100 - x is essentially the same thing as 0.100 .

$$5.1 \times 10^{-4} = \frac{x^2}{0.100}$$

$$x^2 = 5.1 \times 10^{-5}$$

$$x = 7.14 \times 10^{-3} = [H_3O^+]$$

$$pH = 2.15$$

... if we'd used the quadratic equation, our answer would have been pH = 2.16

Compare:

- Weak acid $HNO_2$ : pH of 0.10 M solution = 2.15

- Strong acid: pH of 0.10 M solution = 1.00

The stronger the acid, the lower the pH of a solution of given concentration will be!