Name ___________________ Chemistry ___/___/___

Strong Acids

Strong acids are acids that dissociate completely when dissolved in water. The six most important strong acids are listed below:

Hydrochloric acid	$HCl \rightarrow H^+ + Cl^-$
Hydrobromic acid	$HBr \rightarrow H^+ + Br^-$
Hydroiodic acid	$HI \rightarrow H^+ + I^-$
Perchloric acid	$HClO_4 \rightarrow H^+ + ClO_4^-$
Nitric acid	$HNO_3 \rightarrow H^+ + NO_3^-$
Sulfuric acid	$H_2SO_4 \rightarrow H^+ + HSO_4^-$

For sulfuric acid, only the first hydrogen is considered strong. The second hydrogen only dissociates slightly.

Strong Bases

Strong bases are bases that dissociate completely when dissolved in water. Only the hydroxide of Group 1 (IA) metals, calcium, strontium & barium dissociate to any appreciable degree.

Calculating the pH of a Strong Acid or a Strong Base

Strong acids and strong bases dissociate completely. Thus, for an acid, the concentration of hydrogen is the same as the original concentration of the acid. For a base, the concentration of hydroxide is equal to the concentration of the base. Once you know the $[H^+]$ or $[OH^-]$ you can determine the pH & pOH of a substance.

Example 1: Calculate the pH & pOH of a 0.025 M HCl
Since HCl is a **strong acid**, it dissociates completely. $HCl \rightarrow H^+ + Cl^-$. Thus, [HCl] is equal to $[H^+]$.
Thus, pH = -log[0.025]
pH = 1.6; pOH = 12.4

Example 2: Calculate the pH & pOH of 0.0038 M NaOH.
NaOH is a **strong base**, so it dissociates completely. $NaOH \rightarrow Na^+ + OH^-$ Thus [NaOH] is equal to $[OH^-]$.
* Remember, when a base dissociates, hydroxide (OH^-) ion is produced, so you calculate the pOH first.
pOH = -log[0.0038]
pOH = 2.4; pH = 11.6

Calculating Dissociation Constants of a Weak Acid

An **acid dissociation constant (K_a)** is *the ratio of the concentration of the dissociated form of an acid to the undissociated form*. Scientists calculate an acid's dissociation constant to determine how much of an acid is in the ionic form. Example: A 0.1000 M solution of acetic acid is only partially ionized. The $[H^+]$ in the solution is measured as 1.34×10^{-3}M. What is the acid dissociation constant? **Note: You must always create an ICE table when solving for Ka or Kb.**

Concentrations	$[CH_3COOH]$	$[H^+]$	$[CH_3COO^-]$
Initial	0.1000	0	0
Change	-x	+x	+x
Equilibrium	0.0987	1.34×10^{-3}	1.34×10^{-3}

$$K_a = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]}$$

$$K_a = \frac{(1.34 \times 10^{-3})(1.34 \times 10^{-3})}{0.0987}$$

$$K_a = \mathbf{1.82 \times 10^{-5}}$$

Ionization Constants of Weak Acids		
Monoprotic Acid	**Name**	**K_a**
HIO_3	iodic acid	1.69×10^{-1}
HNO_2	nitrous acid	7.1×10^{-4}
HF	hydrofluoric acid	6.8×10^{-4}
$HCHO_2$	formic acid	1.8×10^{-4}
$HC_3H_5O_3$	lactic acid	1.38×10^{-4}
$HC_7H_5O_2$	benzoic acid	6.28×10^{-5}
$HC_4H_7O_2$	butanoic acid	1.52×10^{-5}
HN_3	hydrazoic acid	1.8×10^{-5}
$HC_2H_3O_2$	acetic acid	1.8×10^{-5}
$HC_3H_5O_2$	propanoic acid	1.34×10^{-5}
HOCl	hypochlorous acid	3.0×10^{-8}
HCN	hydrocyanic acid	6.2×10^{-10}
HC_6H_5O	phenol	1.3×10^{-10}
HOI	hypoiodous acid	2.3×10^{-11}
H_2O_2	hydrogen peroxide	1.8×10^{-12}

Calculating pH Using Dissociation Constants

Likewise, if you know the concentration of the acid and the acid dissociation constant, you can determine the pH.
Example: Calculate the pH of a 0.0100 M lactic acid solution. $K_a = 1.38 \times 10^{-4}$

$$HC_3H_5O_3 \rightleftharpoons H^{1+} + C_3H_5O_3^{1-}$$

$$K_a = \frac{[H^{1+}][C_3H_5O_3^{1-}]}{[HC_3H_5O_3]}$$

$$\frac{1.38 \times 10^{-4}}{1} = \frac{[x][x]}{0.0100}$$

$1.38 \times 10^{-6} = x^2$

$x = 0.00117$ therefore $[H^+] = 0.00117$

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

$pH = -\log[0.00117]$

pH = 2.39

Calculating Dissociation Constants of a Weak Base

A **base dissociation constant (K_b)** is *the ratio of the concentration of the dissociated form of a base to the undissociated form*. Scientists calculate a base's dissociation constant to determine how much of a base is in the ionic form. **Example**: A 0.1000 M solution of weak base is only partially ionized. The $[OH^-]$ in the solution is measured as 1.34×10^{-3}M. What is the base dissociation constant?

Concentrations	[XOH]	$[X^+]$	$[OH^-]$
Initial	0.1000	0	0
Change	-x	+x	+x
Equilibrium	0.0987	1.34×10^{-3}	1.34×10^{-3}

$$K_b = \frac{[X^+][OH^-]}{[XOH]}$$

$$K_b = \frac{(1.34 \times 10^{-3})(1.34 \times 10^{-3})}{0.0987}$$

$K_b =$ **1.82×10^{-5}**

Calculating pH Using Dissociation Constants

Likewise, if you know the concentration of the base and the base dissociation constant, you can determine the pH.
Example: Calculate the pH of a 0.010 M ammonia solution.

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

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

$$\frac{1.8 \times 10^{-5}}{1} = \frac{[x][x]}{0.010}$$

$1.8 \times 10^{-7} = x^2$

$x = 4.2 \times 10^{-4}$ therefore $[OH^-] = 4.2 \times 10^{-4}$

$pOH = -\log[OH^-]$

$pOH = -\log[4.2 \times 10^{-4}]$

pOH = 3.4

pH = 10.6

Ionization Constants of Weak Bases

Weak Base	Name	K_b
$(CH_3)_2NH$	dimethylamine	9.6×10^{-4}
CH_3NH_2	methylamine	4.4×10^{-4}
$CH_3CH_2NH_2$	ethylamine	5.6×10^{-4}
$(CH_3)_3N$	trimethylamine	7.4×10^{-5}
NH_3	ammonia	1.8×10^{-5}
N_2H_4	hyzadrine	9.6×10^{-7}
C_5H_5N	pyridine	1.7×10^{-9}
$C_6H_5NH_2$	aniline	3.8×10^{-10}

Homework:

1. List the six strong acids.

2. List six strong bases.

3. What makes an acid or a base strong?

4. Calculate the pH & pOH of a 0.0045 M NaOH solution.

5. Calculate the pH & pOH of a 0.000088 M HI solution.

6. Calculate the pH & pOH of a 0.20 M KOH solution.

7. Calculate the pH & pOH of a 0.085 M $HClO_4$ solution.

8. A 0.200 M solution of a weak acid has a $[H^+]$ of 9.86×10^{-4}M. What is the pH of this solution? What is the K_a of the solution?

9. A 0.500 M solution of a weak acid has a $[H^+]$ of 3.4×10^{-6}M. What is the pH of this solution? What is the K_a of the solution?

10. A 0.050 M solution of a weak acid has a $[H^+]$ of 2.9×10^{-3}M. What is the pH of this solution? What is the K_a of the solution?

11. Calculate the pH of a 0.50 M acetic acid solution.

12. Calculate the pH of a 0.10 M propanoic acid solution.

13. Calculate the pH of a 2.3 x 10^{-3} M lactic acid solution.

14. A 0.100 M solution of a weak base has a $[OH^-]$ of 4.86 x 10^{-3}M. What is the pH of this solution? What is the K_b of the solution?

15. A 0.00780 M solution of a weak base has a $[OH^-]$ of 2.4 x 10^{-7}M. What is the pH of this solution? What is the K_b of the solution?

16. A 0.050 M solution of a weak base has a $[OH^-]$ of 9.2 x 10^{-6}M. What is the pH of this solution? What is the K_b of the solution?

17. Calculate the pH of a 0.5 M trimethylamine solution.

18. Calculate the pH of a 0.1 M aniline solution.

19. Calculate the pH of a 2.3 x 10^{-3} M dimethylamine solution.