SOLVING ACID – BASE PROBLEMS

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"Don't fight the problem, it will win."

Aqueous acid-base reactions involve more components than gas equilibrium (problems will be more complicated), therefore a systematic approach to acid-base problems must be developed and obeyed.

1. **Identify the MAJOR SPECIES in the solution.**

   Major species: - those components present in large amounts
   - are written as they occur in solution

   Example: 1.0 M HCl

   Major Species: HCl
   H_2O

2. **Write reactions for substances that produce or consume H^+.**

   $HCl \rightleftharpoons H^+ + Cl^-$

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

3. **Determine the dominant source of H^+ in solution.**

   $HCl \longrightarrow H^+ + Cl^-$ $\quad 1.0\,M = [H^+] \quad K_a \ggg 1$

   $H_2O \rightleftharpoons H_3O^+ + OH^-$ $\quad 1\times10^{-7}M = [H^+] \quad K_w = 1\times10^{-14}$

*** *"Intelligent approximations are the key to solving acid-base problems (pg 33)."* ***

Calculating pH of WEAK ACIDS

**Remember weak acids only partially ionize – Have small K_a values. **

Ex: The hypochlorite ion (OCl^-) is a strong oxidizing agent often found in household bleaches and disinfectants. It is also the active ingredient that forms when swimming pool water is treated with chlorine. In addition to its oxidizing abilities, the hypochlorite ion has a relatively high affinity for protons and forms the weakly acidic hypochlorous acid ($HClO$, $K_a = 3.5 \times 10^{-8}$). Calculate the pH of a 0.100 M aqueous solution of hypochlorous acid.

① Major Species: $HClO$
H_2O

② Acid Dissociation: $HClO \rightleftharpoons H^+ + ClO^-$ $K_a = 3.5 \times 10^{-8}$ * will dominate production of H^+
$H_2O \rightleftharpoons H^+ + OH^-$ $K_w = 1.0 \times 10^{-14}$

③
← neglect contrib. from H_2O

	$HClO \rightleftharpoons$	H^+ +	ClO^-
I	0.100 M	0	0
C	−x	+x	+x
E	0.100 M − x	x	x

$$K_a = \frac{[H^+][ClO^-]}{[HClO]} = 3.5 \times 10^{-8}$$

$$= \frac{(x)^2}{(0.100 - x)} = 3.5 \times 10^{-8}$$

x is so small, can approx that $0.100 - x = 0.100$

$$\therefore \frac{x^2}{0.100} = 3.5 \times 10^{-8}$$

$$x = 5.9 \times 10^{-5}\,M$$

$\therefore [H^+] = 5.9 \times 10^{-5}\,M$

$pH = 4.23$

"5% Rule"

* x must be less than 5% of [HA] to drop "x"

Check: $\frac{5.9 \times 10^{-5}\,M}{0.100\,M} \times 100 = 0.059\%$

Solving Weak Acid Equilibrium Problems:

1. List the major species in solution.
2. Choose the species that can produce H^+, and write the balanced equations for the reactions producing H^+.
3. Using the values of K_a written, decide which equilibrium will dominate in producing H^+.
4. Write the equilibrium expression for the dominant equilibrium.
5. Construct an ICE table.
6. Substitute the equilibrium "concentrations" into the equilibrium expression.
7. Solve for x → assume $[HA]_o - x = [HA]_o$
   Use 5% rule to verify whether the approximation is valid. ✱
8. Calculate $[H^+]$ and pH.

Ex: Calculate the pH, [H+], [F-], and percent dissociation of a 1.00 M solution of HF ($K_a = 7.4 \times 10^{-4}$).

① Major Species: HF
H_2O

② Dissociation: $HF \rightleftharpoons H^+ + F^-$ $K_a = 7.4\times10^{-4}$ ← dominant
$H_2O \rightleftharpoons H^+ + OH^-$ $K_w = 1.0\times10^{-14}$

③

	$HF \rightleftharpoons$	$H^+ +$	F^-
I	1.00M	0	0
C	$-x$	$+x$	$+x$
E	1.00M $-x$	x	x
	9.73M	0.027M	0.027M

$$K_a = \frac{[H^+][F^-]}{[HF]} = 7.4\times10^{-4} = \frac{(x^2)}{(1.00M - \cancel{x})} = \frac{x^2}{(1.00M)}$$

$$2.7\times10^{-2} = x$$

5% Rule: $\frac{2.7\times10^{-2}M}{1.00M} = 2.7\%$

$pH = -\log(0.027M)$
$= 1.57$

$\%\ \text{Dissociation} = \frac{0.027M}{1.00M} = 2.7\%$

↑

$$\frac{\#\ \text{mol of HA Dissociated}}{\#\ \text{original mol/L of HA}} \times 100$$

Ex: 0.050 mol of weak acid (HA) is dissolved in enough water to make 1.0 L of solution. The pH of this solution is found to be 3.50. What is the value of K_a for this weak acid?

$$HA \rightleftharpoons H^+ + A^- \qquad K_a = \frac{[H^+][A^-]}{[HA]} = \frac{(3.2\times10^{-4}M)^2}{(0.050\,M)} = \underline{\underline{2.0\times10^{-6}}}$$

	HA	H^+	A^-
I	0.050M	0	0
C	$-x$	$+x$	$+x$
E	$0.050-x$ ≈ 0.050	x $3.2\times10^{-4}M$	x $3.2\times10^{-4}M$

$pH = 3.50 \;\therefore\; [H^+] = x = \underline{\underline{3.2\times10^{-4}M}}$

Ex: Calculate the pH of a solution that contains 1.00 M HCN ($K_a = 6.2 \times 10^{-10}$) and 5.00 M HNO_2 ($K_a = 4.0 \times 10^{-4}$). Also calculate the concentration of cyanide ion (CN^-) in this solution at equilibrium.

① Major Species:

$HCN \rightleftharpoons H^+ + CN^- \quad K_a = 6.2\times10^{-10}$

$HNO_2 \rightleftharpoons H^+ + NO_2^- \quad K_a = 4.0\times10^{-4}$ * dominant

$H_2O \rightleftharpoons H^+ + OH^- \quad K_w = 1.0\times10^{-14}$

②

	HNO_2	$\rightleftharpoons H^+$	$+ NO_2^-$
I	5.00M	0	0
C	$-x$	$+x$	$+x$
E	5.00M $-x$	x	x

$K = \frac{[H^+][NO_2^-]}{[HNO_2]} = 4.0\times10^{-4} = \frac{(x)(x)}{(5.00M \cancel{-x})}$

$= \frac{x^2}{5.00M}$

$4.5\times10^{-2}M = x$

$\therefore x = [H^+] = 4.5\times10^{-2}M$

$pH = 1.35$

5% Rule: $\frac{4.5\times10^{-2}M}{5.00M} \times 100 = 0.89\%$

From HNO_2

$[CN^-]$:

	HCN	$\rightleftharpoons H^+$	$+ CN^-$
I	1.00M	$4.5\times10^{-2}M$	0
C	$-x$	$+x$	$+x$
E	1.00M $-x$	$4.5\times10^{-2}+x$	x
	$\approx 1.00M$	$\approx 4.5\times10^{-2}$	

$K_a = \frac{[H^+][CN^-]}{[HCN]} = 6.2\times10^{-10}$

$[CN^-] = \frac{6.2\times10^{-10}[HCN]}{[H^+]}$

$= \frac{6.2\times10^{-10}(1.00M \cancel{-x})}{(4.5\times10^{-2}M \cancel{+x})}$

$= 1.4\times10^{-8}M = [CN^-]$

Key

AP Chemistry Quiz
Section 1- 4, Equilibrium Chapter 2

1. Calculate the pH of a solution of 2.0×10^{-3} M $Ca(OH)_2$.

2. Calculate $[H^+]$ and pH in a solution of 0.025 M formic acid, HCO_2H, if $K_a = 1.8 \times 10^{-4}$. Note: formic acid contains only 1 ionizable hydrogen.

(1) $2.0 \times 10^{-3} m\ Ca(OH)_2$

$[OH^-] = 4.0 \times 10^{-3}$

$pOH = 2.40$

$14 - 2.40 \Rightarrow 11.60 = pH$

$[OH^-] = 4.0 \times 10^{-3}$ $[H^+] = 2.5 \times 10^{-12}$

$pH = 11.60$

(2) $HCO_2H \rightleftharpoons H^+ + CO_2H^-$

	HCO_2H	H^+	CO_2H^-
I	0.025m	0	0
C	$-x$	$+x$	$+x$
E	$0.025-x$	x	x

$$K_a = \frac{[H^+][CO_2H^-]}{[HCO_2H]}$$

$$1.8 \times 10^{-4} = \frac{(x)(x)}{(0.025-x)}$$

$$1.8 \times 10^{-4} = \frac{x^2}{(0.025-\cancel{x})}$$

$$\sqrt{4.5 \times 10^{-6}} = \sqrt{x^2}$$

$0.0021\ m = x$

$$5\% = \frac{0.0021}{0.025} \times 100 = 8.4\%$$

$$1.8 \times 10^{-4} = \frac{x^2}{0.025-x)}$$

$4.5 \times 10^{-6} - 1.8 \times 10^{-4} - x^2 = 0$

$x = 0.0020M \rightarrow pH = 2.70$