pH of Strong Acids

- The six strong acids are: HCl, HBr, HI, HNO_3, $HClO_4$, and H_2SO_4.
- Calculating the pH of strong acids:
  - Since strong acids 100% ionize, the $[H^+]$ = the initial [strong acid]
  - Example) 1.00×10^{-5} M HCl
    - $[HCl]_0 = 1.00 \times 10^{-5}$ M, thus $[H^+] = 1.00 \times 10^{-5}$ M
    - $pH = -\log(1.00 \times 10^{-5}) = 5.000$

pH of Strong Bases

- The eight strong bases are: LiOH, NaOH, KOH, RbOH, CsOH, $Ca(OH)_2$, $Sr(OH)_2$, and $Ba(OH)_2$. They are Group 1 and 2 hydroxides that form the "b" on the periodic table.
- Calculating the pH of strong bases:
  - Since strong bases 100% ionize, the $[OH^-]$ will be determined by the [strong base] and stoichiometry
  - Example 1) 1.00×10^{-4} M NaOH
    - $[NaOH]_0 = 1.00 \times 10^{-4}$ M, thus $[OH^-] = 1.00 \times 10^{-4}$ M
    - $pOH = -\log(1.00 \times 10^{-4}) = 4.000$
    - $pH = 14 - 4.000 = 10.000$
  - Example 2) 1.00×10^{-4} M $Ca(OH)_2$
    - $[Ca(OH)_2]_0 = 1.00 \times 10^{-4}$ M, thus $[OH^-] = 2(1.00 \times 10^{-4}) = 2.00 \times 10^{-4}$ M
    - $pOH = -\log(2.00 \times 10^{-4}) = 3.699$
    - $pH = 14 - 3.699 = 10.301$

pH of Weak Acids

- Since weak acids DO NOT completely ionize, the $[H^+]$ DOES NOT EQUAL the initial [weak acid]
- In fact, $[H^+] < [\text{weak acid}]_0$
- An equilibrium will be established between the weak acid molecules (i.e. unionized molecules, HA) and its ions (H^+ and A^-)
- The equilibrium $[H^+]$ is used to calculate pH
- Write the ionization equation of the weak acid in water, H_3O^+ will always be a product
- Use an ICE chart and Ka to determine equilibrium $[H_3O^+]$
- Example) Calculate the pH of a 2.00 M solution of nitrous acid, HNO_2. (Ka of HNO_2 = 6.0×10^{-4})

	$HNO_2(aq)$	+ $H_2O(l)$	$\rightleftharpoons$ $H_3O^+(aq)$	+ $NO_2^-(aq)$
I	2 M		0	0
C	-x		+x	+x
E	2M - x		x	x

$$K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]}$$

$$6.0 \times 10^{-4} = \frac{x^2}{2-x}$$

* Since "x" will be much smaller than 2, we can ignore "x".

$$6.0 \times 10^{-4} = \frac{x^2}{2}$$

$$12 \times 10^{-4} = x^2$$

$$x = 3.464 \times 10^{-2} \text{ M} = [H_3O^+]$$

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

$$pH = -\log(3.464 \times 10^{-2}) = \boxed{1.73}$$

pH of Weak Bases

- Same procedure as a weak acid with the following differences:
  - Write the ionization equation of the weak base in water, OH^- will always be a product
  - Use an ICE chart and Kb to determine equilibrium $[OH^-]$
  - Use the equilibrium $[OH^-]$ to calculate pOH
  - Use pOH to calculate pH
- Example) Calculate the pH for a 15.0 M solution of NH_3. (Kb of $NH_3 = 1.8 \times 10^{-5}$)

	$NH_3(aq)$	+ $H_2O(l)$ $\rightleftharpoons$	$NH_4^+(aq)$	+ $OH^-(aq)$
I	15 M		0	0
C	$-x$		$+x$	$+x$
E	$15M - x$		x	x

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

$$1.8 \times 10^{-5} = \frac{x^2}{15 - x}$$

* since "x" will be much smaller than 15, we can ignore "x".

$$1.8 \times 10^{-5} = \frac{x^2}{15}$$

$$2.7 \times 10^{-4} = x^2$$

$$x = 1.643 \times 10^{-2}\ M = [OH^-]$$

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

$$pOH = -\log (1.643 \times 10^{-2}) = 1.7843$$

$$pH = 14 - pOH$$

$$pH = 14 - 1.7843 = \boxed{12.22}$$

Percent Dissociation

- Strong acids and bases have 100% dissociation
- Weak acids and bases will have less than 100% dissociation
- Calculating percent dissociation:
  - Weak acid percent dissociation = $\frac{[H^+]}{[initial\ weak\ acid]}$ x 100
  - Weak base percent dissociation = $\frac{[OH^-]}{[initial\ weak\ base]}$ x 100

5% Dissociation Rule

- When the percent dissociation of a weak acid/base is less than 5%, the value of "– x" in a weak acid/base ICE chart can be ignored. Thus, the equilibrium concentration of the weak acid/base will be the same as its initial concentration. This will be the case on the AP exam. On the exam, make sure you note that you are ignoring "x" because "x" is much smaller than the initial concentration due to the low percent dissociation.
- How do you know when to ignore "x"?
  - Compare the initial weak acid/base concentration to 100 times Ka (or 100 Kb). If the initial concentration is much larger than 100 Ka (or 100 kb), then ignore "x".
  - On the AP exam, you may ignore "x", but make sure to state why you are allowed to ignore "x".

pH of Weak Acid Mixtures

Only the acid with largest Ka value will contribute an appreciable $[H^+]$. Determine the pH based on this acid and ignore any others.

Example) Calculate the pH of a solution that contains 1.00 M HCN (Ka = 6.2×10^{-10}) and 5.00 M HF (Ka = 7.2×10^{-4}).

*only the strongest acid affects pH

∴ HF determines pH b/c Ka of HF > Ka of HCN

	$HF_{(aq)}$	+ $H_2O_{(l)}$ ⇌	$H_3O^+_{(aq)}$	+ $F^-_{(aq)}$
I	5 M		0	0
C	-x		+x	+x
E	5M - x		x	x

$$Ka = \frac{[H_3O^+][F^-]}{[HF]}$$

$$7.2 \times 10^{-4} = \frac{x^2}{5 - x}$$

ignore "x" b/c "x" is much smaller than initial [HF]

$$7.2 \times 10^{-4} = \frac{x^2}{5}$$

$$3.6 \times 10^{-3} = x^2$$

$$x = 0.06\ M = [H_3O^+]$$

$$pH = -\log(0.06) = \boxed{1.22}$$

Acids Donate ONE proton at a time

- Monoprotic – acid donating one H^+ (ex. $HC_2H_3O_2$)
- Diprotic – acid donating two H^+'s (ex. $H_2C_2O_4$)
- Polyprotic – acid donating many H^+'s (ex. H_3PO_4)
- Acids donate ONE proton at a time
- Acids with more than one ionizable hydrogen will ionize in steps. Each dissociation has its own Ka value. The first dissociation (Ka_1) will be the greatest and subsequent dissociations will have much smaller Ka's. As each H^+ is removed, the remaining acid gets weaker and therefore has a smaller Ka. As the negative charge on the acid increases it becomes more difficult to remove the positively charged proton.
- Since Ka_2 and Ka_3 are so much smaller than Ka_1, the $[H^+]$ obtained from Ka_2 and Ka_3 are so small that they can be ignored. Thus, the pH of a polyprotic acid depends only on the first dissociation and Ka_1.

Example: Calculate the pH of a 5.0 M H_3PO_4 solution.

$$H_3PO_4(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + H_2PO_4^-(aq) \qquad Ka_1 = 7.5 \times 10^{-3}$$
$$H_2PO_4^-(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + HPO_4^{2-}(aq) \qquad Ka_2 = 6.2 \times 10^{-8}$$
$$HPO_4^{2-}(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + PO_4^{3-}(aq) \qquad Ka_3 = 4.8 \times 10^{-13}$$

	$H_3PO_{4(aq)}$	+ $H_2O_{(l)}$ ⇌	$H_3O^+_{(aq)}$	+ $H_2PO_4^-{}_{(aq)}$
I	5 M		0	0
C	-x		+x	+x
E	5M - x		x	x

$$Ka = \frac{[H_3O^+][H_2PO_4^-]}{[H_3PO_4]}$$

$$7.5 \times 10^{-3} = \frac{x^2}{5 - x}$$

ignore "x" b/c "x" is much smaller than initial $[H_3PO_4]$.

$$7.5 \times 10^{-3} = \frac{x^2}{5}$$

$$3.75 \times 10^{-2} = x^2$$

$$x = 0.1936\ M = [H_3O^+]$$

$$pH = -\log(0.1936) = \boxed{0.71}$$

Acid Base Properties of Salts

Salts are produced when an acid and base react. Salts are not always neutral. Some hydrolyze with water to produce aqueous solutions with pH's other than 7.00. To determine if a salt is neutral, acidic or basic you must think about from which acid and base the salt formed.

- Strong acid + strong base = NEUTRAL salt
  - Example) $HNO_3(aq) + NaOH(aq) \rightarrow H_2O(l) + NaNO_3(aq)$
    Strong Strong Water Neutral Salt
    - Na^+ is neutral because it came from a strong base
    - NO_3^- is neutral because it came from a strong acid

- Strong acid + weak base = ACIDIC salt
  - Example) $HNO_3(aq) + NH_3(aq) \rightarrow NH_4NO_3(aq)$
    Strong Weak Acidic Salt
    - NO_3^- is neutral because it came from a strong acid
    - NH_4^+ is acidic because it came from a weak base and hydrolyzes water according to the equation:
    - $NH_4^+(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + NH_3(aq)$

- Weak acid + strong base = BASIC salt
  - Example) $HC_2H_3O_2(aq) + NaOH \rightarrow H_2O(l) + NaC_2H_3O_2(aq)$
    Weak Strong Basic Salt
    - Na^+ is neutral because it came from a strong base
    - $C_2H_3O_2^-$ is basic because it came from a weak acid and hydrolyzes water according to the equation:
    - $C_2H_3O_2^-(aq) + H_2O(l) \leftrightarrow HC_2H_3O_2(aq) + OH^-(aq)$

- Weak acid + weak base = ??? salt
  - pH depends on the Ka value of the cation and Kb value of the anion
  - If Ka is larger, the solution is acidic
  - If Kb is larger, the solution is basic

Practice

1. Calculate the pH of a 0.100 M aqueous solution of hypochlorous acid, HOCl. (Ka of HOCl = 3.5×10^{-8}).
2. Calculate the pH of a 1.5 M solution of methylamine, CH_3NH_2 (Kb of CH_3NH_2 = 4.38×10^{-4}).
3. Calculate the percent dissociation of 0.100 M acetic acid, $HC_2H_3O_2$ (Ka = 1.8×10^{-5}).
4. Calculate the percent dissociation of 3.1×10^{-4} M carbonic acid, H_2CO_3 (Ka = 4.4×10^{-7}).
5. If a 0.375M weak acid (HA) dissociates 2.5% when put in water, what is the acid's Ka value?
6. Lactic acid ($HC_3H_5O_3$) is a waste product that accumulates in muscle tissue during exercise, leading to pain and a feeling of fatigue. In a 0.100 M aqueous solution, lactic acid is 3.7% dissociated. Calculate the value of Ka for this acid.