AP14

Acid-Base Equilibria Flow Chart

Strong Acids (easy calcs)

$HClO_4$	perchloric acid
H_2SO_4	sulfuric acid*
HNO_3	nitric acid
HI	hydroiodic acid
HBr	hydrobromic acid
HCl	hydrochloric acid
H_3O^+	hydronium ions

*only 1st H^+ is strong

$[HA]_{initial}=[H^+]_{equilibrium}$

100% ionize

Weak Acids (complicated calculations)

Any acid that is not on strong acid list

$[HA]_{initial} >> [H^+]_{equilibrium}$ K_a= small values $1.00 \times 10^{-14} = K_a \times K_b$

The acid concentration given in the problem is the $[HA]_{initial}$

$$\%\ \text{Ionization} = \frac{[H^+]_{@equilibrium}}{[HA]_{@initial}}$$

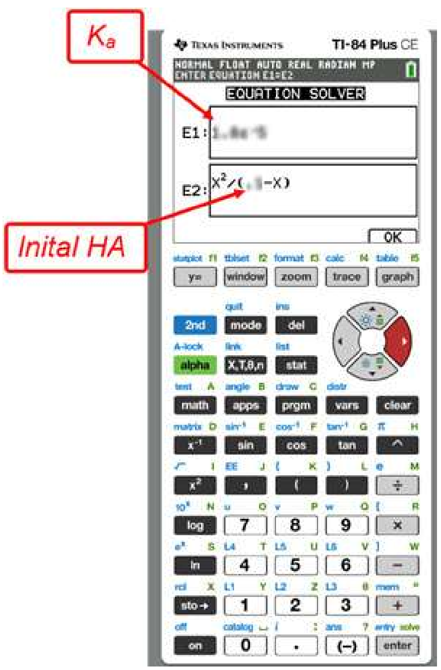
% Ionization increases with dilution and increased pH.

Use the RICE to understand the equilibrium.
Use the Equilibrium line with equilibrium expression.

	HA	⇌	H^+	+	A^-
Initial	HA				
Change	$-x$		$+x$		$+x$
Equil	HA-x		$+x$		$+x$

$$K_a = \frac{[H^+][A^-]}{[HA]} \text{ @ equilibrium}$$

x is small so $[HA-x] \approx [HA]$ $[H^+] \approx \sqrt{K_a \times [HA]_{start}}$.



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

Always carry pH to the hundredths place.

Acid-Base Equilibria Flow Chart

Strong Bases dissociate to produce OH^-

$BOH \rightarrow B^+ + OH^-$

Metal OH , alkali metal hydroxides
Metal $(OH)_2$, alkaline earth metal hydroxides
- Calcium hydroxide
- Barium hydroxide
- Strontium hydroxide

$[\text{Metal }(OH)_n]_{initial} = n \times [OH^-]_{equilibrium}$

Soluble hydroxides 100% ionize

Weak Bases
React with water to produce OH^-

NH_3 & many organic compounds such as morphine and caffeine; conjugates of weak acids

$[B]_{initial} >> [OH^-]_{equilibrium}$

Lower concentrations produce higher % ionization

K_b= small values $1.00 \times 10^{-14} = K_a \times K_b$

Use the RICE to understand the equilibrium.
Use the Equilibrium line with equilibrium expression

	$B + H_2O$	$\rightleftharpoons$	HB^+	+	OH^-
Initial	B		0		0
Change	$-x$		$+x$		$+x$
Equil	$B-x$		$+x$		$+x$

$$K_b = \frac{[HB^+][OH^-]}{[B]} \text{ @ equilibrium}$$

x so $B-x \approx B$ $\left[OH^-\right] \approx \sqrt{K_b \times [B]}$

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

$pH = 14.00 - pOH$

Always carry pOH and pH to the hundredths place.