Key

COMMON ION

** Encountered with a solution containing a weak acid and a soluble salt containing the anion (conjugate base) of the weak acid.
**The acid and salt have the same anion. Ex: H_2S and K_2S

Ex: Calculate the $[H^+]$ in a solution containing 1.00 M HF ($K_a = 7.2 \times 10^{-4}$) and 0.50M sodium fluoride.

→ HF H_2O Na^+ F^-

① Major Species $HF \rightleftharpoons H^+ + F^-$ $K_a = 7.2 \times 10^{-4}$ ← dominant

$H_2O \rightleftharpoons H^+ + OH^-$ $K_w = 1.0 \times 10^{-14}$

$NaF \rightarrow Na^+ + F^-$ (Soluble Salt)

Could Have: $F^- + H_2O \rightarrow HF + OH^-$ Already included in equilibrium

② $HF \rightleftharpoons H^+ + F^-$ $K_a = \frac{[H^+][F^-]}{[HF]} = 7.2 \times 10^{-4}$

dissociation HF provides H^+

	HF	H^+	F^-
I	1.00M	0	0.50M
C	−x	+x	+x
E	1.00M−x	x	0.50M+x

shift right? (ie H^+ is zero)

$$7.2 \times 10^{-4} = \frac{(x)(0.50M + x)}{(1.00M - x)}$$

both insignificant (below 5%) have to check each one

When can ignore one... usually can ignore the other

$1.4 \times 10^{-3} = x$

$x = 1.4 \times 10^{-3} M$

$pH = 2.85$

"Common Ion Effect" = when a solution contains a salt having an ion common with one in the equilibrium, the position of the equilibrium is driven away from the side containing that ion.

" . . . the shift in an ionic equilibrium caused by the addition of a solute that provides an ion that takes part in the equilibrium (Ebbing)."

Ex: The degree if ionization of acetic acid, $HC_2H_3O_2$, in a 0.10 M aqueous solution at 25°C is 0.013. Ka at this temperature is 1.7×10^{-5}. Calculate the degree of ionization of $HC_2H_3O_2$ in a 0.10 M solution at 25°C to which HCl is added to make it 0.010 M HCl. How is the degree of ionization affected?

Major Species: $HC_2H_3O_2$, H_2O, HCl

$HC_2H_3O_2 \rightleftharpoons H^+ + C_2H_3O_2^-$ $K_a = 1.7 \times 10^{-5}$

$H_2O \rightleftharpoons H^+ + OH^-$ $K_w = 1.0 \times 10^{-14}$

$HCl \rightarrow H^+ + Cl^-$ $K_a >>> 1$

	$HCl \rightarrow$	H^+ +	Cl^-
I	0.010M	0	0
C	$-x$	$+x$	$+x$
E	0	0.010M	0.010M

$x = 0.010M$

Original Ionization = 0.013

Creation of H^+ affects eq. of $HC_2H_3O_2$

	$HC_2H_3O_2 \rightleftharpoons$	H^+ +	$C_2H_3O_2^-$
I	0.10 M	0.010M	0
C	$-x$	$+x$	$+x$
E	$0.10M - x$	$0.010 + x$	x
	$\approx 0.10M$	≈ 0.010	1.7×10^{-4}

$K_a = 1.7 \times 10^{-5} = \frac{(0.010 + x)(x)}{(0.10M - x)}$

$1.7 \times 10^{-5} = \frac{0.010(x)}{0.10M}$

$1.7 \times 10^{-4} M = x$

$x \rightarrow [H^+]$ produced $= 1.7 \times 10^{-4} M$

$\frac{1.7 \times 10^{-4}}{0.1} = 0.0017$ of Degree Ionization

Ex: A solution is prepared to be 0.10 acetic acid, and 0.20 M sodium acetate. What is the pH of this solution at 25°C? K_a for acetic acid at 25°C is 1.7×10^{-5}.

$HC_2H_3O_2$ ← dominant
H_2O
Na^+
$C_2H_3O_2^-$

	$HC_2H_3O_2 \rightleftharpoons$	$H^+ +$	$C_2H_3O_2^-$
I	0.10 m	0	0.20 m
C	−x	+x	+x
E	0.10 m − x	x	0.20 m + x

$$K_a = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

$$1.7 \times 10^{-5} = \frac{(x)(0.20m \cancel{+x})}{(0.10m \cancel{-x})}$$

$8.5 \times 10^{-6} M = x = [H^+]$

pH = 5.07

Without $C_2H_3O_2^-$ from $NaC_2H_3O_2$

	$HC_2H_3O_2 \rightleftharpoons$	$H^+ +$	$C_2H_3O_2^-$
I	0.10 m	0	0
C	−x	+x	+x
E	0.10 − x	x	x

$$K_a = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

$$1.7 \times 10^{-5} = \frac{(x^2)}{(0.10 \cancel{-x})}$$

$0.0013 = x$

pH = 2.89

Le Chat. Principle Shown!