The Common Ion Effect

The addition of a common ion to a weak acid solution causes a stress on the equilibrium. According to Le Châtelier's Principle, the system will shift to relieve the effects of the stress. For the ionization of a weak acid,

$$HA + H_2O \leftrightarrows H_3O^+ + A^-$$

the stress caused by the addition of a common ion causes the equilibrium to shift left, decreasing $[H_3O^+]$, and therefore increasing pH.

Example: Determine the pH of the following solutions.

a) 0.100 M acetic acid
b) 0.100 M acetic acid with 0.300 M sodium acetate

Solution:

a)

	$HC_2H_3O_2$	$\leftrightarrows$	H^+	+	$C_2H_3O_2^-$
[initial]	0.100		0		0
Δ	− x		+ x		+ x
[equilibrium]	0.100 − x		x		x

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

$$1.8 \cdot 10^{-5} = \frac{(x)(x)}{0.100 - x}$$

$$1.8 \cdot 10^{-5} = \frac{x^2}{0.100}$$

$$1.8 \cdot 10^{-6} = x^2$$

$$x = 0.0013$$

$[H_3O^+]$ = 0.0013 mol/L pH = 2.9

b) $NaC_2H_3O_2 \rightarrow Na^+ + C_2H_3O2^-$

	$NaC_2H_3O_2$	Na^+	$C_2H_3O2^-$
[initial]	0.300	0	0
final]	0	0.300	0.300

	$HC_2H_3O_2$	$\leftrightarrows$	H^+	+	$C_2H_3O_2^-$
[initial]	0.100		0		0.300
Δ	– x		+ x		+ x
[equilibrium]	0.100 – x		x		0.300 + x

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

$$1.8 \cdot 10^{-5} = \frac{(x)(0.300 + x)}{0.100 - x}$$

$$1.8 \cdot 10^{-5}(.100 - x) = 0.300x + x2$$

$$1.8 \cdot 10^{-6} - 1.8 \cdot 10^{-5}x = 0.300x + x^2$$

$$0 = x^2 + 0.300018x - 1.8 \cdot 10^{-6}$$

$$x = \frac{-0.300018 \pm \sqrt{(0.300018)^2 - (4)(1)(1.8 \cdot 10^{-6})}}{(2)(1)}$$

$$x = \frac{-0.300018 \pm \sqrt{0.090018}}{2}$$

$$x = \frac{-3.00018 \pm 0.300029999}{2}$$

$$x = -0.30 \quad \text{or} \quad 6.0 \cdot 10^{-6}$$

$[H_3O^+] = 6.0 \cdot 10^{-6}$ mol/L pH = 5.2