The Common-Ion Effect with Acids and Bases

If you have something in common with someone else that means that the two of you have something that is the same about you. For example, if you and your friend both like cookie dough ice cream more than chocolate ice cream then you have something in common.

The common ion effect is similar to having something in common with a friend. If two substances both contain nitrate (NO_3^-) then they have that ion "in common." The common ion effect is based on LeChatelier's Principle. One way it occurs is when there is a weak acid or base in a solution along with a second soluble substance that has the same ion as the weak acid or base.

Critical Thinking Questions

1. Consider a solution of acetic acid ($HC_2H_3O_2$). Why will sodium acetate ($NaC_2H_3O_2$) cause the "common ion effect," but sodium chloride (NaCl) will not?

2. Let's say we have a solution of the weak base ammonia (NH_3) $K_b = 1.8x10^{-5}$

   a) Complete the equilibrium equation below:

   $$NH_3 + H_2O \rightleftarrows$$

   b) Will NH_4Cl cause a "common ion effect?" will KNO_3? EXPLAIN.

Let's consider the calculation of the pH of a 0.075 M solution of the weak acid acetic acid, $HC_2H_3O_2$, with a $K_a = 1.7x10^{-5}$. The problem can be written like this:

$$HC_2H_3O_2 + H_2O \rightleftarrows C_2H_3O_2^- + H_3O^+$$

$$1.7x10^{-5} = \frac{[x][x]}{[0.075 - x]}$$

$$1.7x10^{-5} = \frac{x^2}{0.075}$$

$$x = 0.0011 = [H^+]$$

$$pH = -\log(0.0011) = \mathbf{2.95}$$

Questions like the previous question we have done before. We assume that the acetic acid is in pure water, so the initial amounts of the products are zero. The problem becomes a little different if the water has a common ion dissolved in it. For example…

3. Calculate the pH of a solution that is 0.075 M acetic acid **AND** 0.033 M sodium acetate. The following table may be helpful: *(note: given the size of the K_a value, x is negligible compared to a* ***non-zero*** *amount)*

	$HC_2H_3O_2$	+	H_2O	⇄	$C_2H_3O_2^-$	+	H_3O^+
Initial:	0.075 M				0.033 M		0 M
Change:	-x				+x		+x
Equilibrium:	0.075 – x ≈ 0.075				0.033 + x ≈ 0.033		x

4. When comparing the example on the bottom of the other side of this sheet with question 3 above, the initial concentration of all of the products was not zero. What effect did this have on the pH of the solution?

5. Calculate the pH of a solution that is both 0.052 M NH_3 ($K_b = 1.8x10^{-5}$) **AND** 0.028 M NH_4Cl.

	NH_3	+	H_2O	⇄	NH_4^+	+	OH^-
Initial:							
Change:							
Equilibrium:							

6. Calculate the pH of a solution that is 0.15 M acetic acid **AND** 0.023 M calcium acetate, $Ca(C_2H_3O_2)_2$.

	$HC_2H_3O_2$	+	H_2O	⇄	$C_2H_3O_2^-$	+	H_3O^+
Initial:							
Change:							
Equilibrium:							