The Common Ion Effect

Consider a weak acid and a salt containing one ion common to both.

Problem 1: $CH_3COOH + H_2O \rightleftharpoons H_3O^{+1} + CH_3COO^{-1}$

If we were to take a 0.1 mol dm^{-3} CH_3COOH solution, $K_a = 1.8 \times 10^{-5}$.

Calculate its pH:

(answer: pH = 2.9)

Now, suppose we dissolve 0.1 mol dm^{-3} sodium acetate, $CH_3COONa_{(aq)}$, in that same solution, with your current knowledge of equilibrium, predict what would be the pH of the new solution. Justify your answer.

(Answer: reverse reaction will be favoured according to LeChatelier's Principle, .: decreasing the [H_3O^{+1}], hence pH will increase)

Now, determine the pH of this solution.

(answer: pH = 4.8)

Hence in conclusion, one can say that the presence of a common ion greatly inhibits the dissociation of an acid and lowers the concentrations of the H_3O^{+1}, thus increasing the pH of the solution, making the solution less acidic.

Problem 2: **See Page 2: K_a & The Common Ion Effect** on how to solve these problems.

When 0.025 mol sodium nitrite$_{(s)}$, is added to a 500 cm^3 solution of 0.100 mol dm^{-3} nitrous acid,$HNO_{2\ (aq)}$, the resulting pH is 3.00. Determine the K_a for nitrous acid.

Problem 3:

What will be the resulting acetate ion concentration when 0.10 mol dm^{-3} hydrochloric acid solution, $HCl_{(aq)}$ is added to a 0.1 mol dm^{-3} solution of aqueous acetic acid solution,

$K_a = 1.8 \times 10^{-5}$. (ans: [CH_3COO^{-1}] = 1.8 x10^{-5})

Problem 4:

Calculate the [OH^{-1}], the pH, and the percent ionisation of an aqueous solution of 1.0 mol dm^{-3} aqueous solution of ammonia, $NH_{3(aq)}$, that also contains 0.10 mol dm^{-3} ammonium chloride,

$K_{b(NH_3)} = 1.8 \times 10^{-5}$. (Ans: [$OH^{-1}$] = 1.8 x10^{-4}, pH = 10.26, % ionization = 1.8 x 10^{-2} %)

Problem 5:

a) Calculate the percent ionization of a 1.0 mol dm^{-3} $HF_{(aq)}$, $K_{a(HF)} = 7.2 \times 10^{-4}$.

b) Calculate the percent ionisation of a 1.0 mol dm^{-3} $HF_{(aq)}$ that also contains 1.0 mol dm^{-3} of sodium fluoride, NF. $K_{a(HF)} = 7.2 \times 10^{-4}$.

c) What deductions can you make from these calculations.

(Ans: (a) % ionization = 2.68 %, (b) 7.2 x 10^{-2} %, (c) .: the presence of a common ion greatly inhibits the %ionization of an acid and lowers the [H_3O^{+1}], hence increasing the pH of the solution, i.e makes it less acidic)

Problem 6:

What is the pH of a solution which contains 0.10 mol dm^{-3} KClO and a 0.05 mol dm^{-3} HClO,

$K_{a(HClO)} = 2.9 \times 10^{-8}$ (Ans: pH = 7.84)

K_a & THE COMMON ION EFFECT

We are given an acid and a salt containing one ion common to both.

1. When 0.0250 mol sodium nitrite solid is added to a 500.0 cm^3 solution of 0.100 mol dm^{-3} nitrous acid, the resulting pH is 3.00. Determine the K_a for nitrous acid, $HNO_2(aq)$.

$$HNO_2(aq) + H_2O \rightleftharpoons H_3O^+(aq) + NO_2^{-1}(aq)$$

$$NaNO_2(aq) \longrightarrow Na^+(aq) + NO_2^{-1}(aq)$$

NOTE: The NO_2^{-1} ion is common to both the salt and the acid.
The Na^+ ion is only a spectator ion.
Le Chatelier's Principle comes into effect.
The reaction will proceed to the left since there is an excess of NO_2^- ion

$[H_3O^+]_E = 10^{-pH}$
$= 10^{-3.00}$
$= 1.00 * 10^{-3}$ or 0.00100
$[NO_2^{-1}] =$ mol/dm^3 = 0.025 mol / 0.5dm^3 = 0.05 mol dm^{-3}

	$[HNO_2]$ (mol dm^{-3})	**$[H_3O^+]$ (mol dm^{-3})**	**$[NO_2^{-1}]$ (mol dm^{-3})**
Initial	0.100	-----	0.0500
Change	- 0.00100	+0.00100	+0.00100
Equil^m	0.099	0.00100	0.0510

$$K_a = \frac{[H^+][NO_2^{-1}]}{[HNO_2]}$$

$$= \frac{(0.00100)(0.0510)}{(0.099)}$$

$= 5.152 * 10^{-4}$

∴ The K_a is $5.15 * 10^{-4}$ mol dm^{-3}