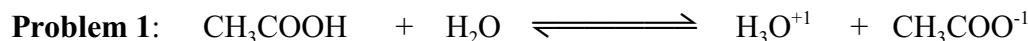


The Common Ion Effect

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



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, $\text{CH}_3\text{COONa}_{(\text{aq})}$, 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, \therefore decreasing the $[\text{H}_3\text{O}^+]$, 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^+ , 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 $0.100 \text{ mol dm}^{-3}$ nitrous acid, $\text{HNO}_2_{(\text{aq})}$, 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, $\text{HCl}_{(\text{aq})}$ is added to a 0.1 mol dm^{-3} solution of aqueous acetic acid solution,

$$K_a = 1.8 \times 10^{-5}. \quad (\text{ans: } [\text{CH}_3\text{COO}^-] = 1.8 \times 10^{-5})$$

Problem 4:

Calculate the $[\text{OH}^-]$, the pH, and the percent ionisation of an aqueous solution of 1.0 mol dm^{-3} aqueous solution of ammonia, $\text{NH}_3_{(\text{aq})}$, that also contains 0.10 mol dm^{-3} ammonium chloride,

$$K_{b(\text{NH}_3)} = 1.8 \times 10^{-5}. \quad (\text{Ans: } [\text{OH}^-] = 1.8 \times 10^{-4}, \text{ pH} = 10.26, \% \text{ ionization} = 1.8 \times 10^{-2} \%)$$

Problem 5:

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

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

c) What deductions can you make from these calculations.

(Ans: (a) % ionization = 2.68 %, (b) $7.2 \times 10^{-2} \%$, (c) \therefore the presence of a common ion greatly inhibits the %ionization of an acid and lowers the $[\text{H}_3\text{O}^+]$, hence increasing the pH of the solution, i.e makes it less acidic)

Problem 6:

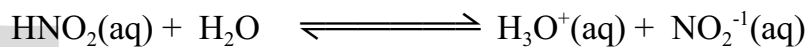
What is the pH of a solution which contains 0.10 mol dm^{-3} KClO and a 0.05 mol dm^{-3} HClO ,

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

K_a & THE COMMON ION EFFECT

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

- When 0.0250 mol sodium nitrite solid is added to a 500.0 cm³ solution of 0.100 mol dm⁻³ nitrous acid, the resulting pH is 3.00. Determine the K_a for nitrous acid, HNO₂(aq).



NOTE: The NO₂⁻¹ 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 NO₂⁻ ion

$$[\text{H}_3\text{O}^+]_{\text{E}} = 10^{-\text{pH}}$$

$$= 10^{-3.00}$$

$$= 1.00 \times 10^{-3} \text{ or } 0.00100$$

$$[\text{NO}_2^{-1}] = \text{mol/dm}^3 = 0.025 \text{ mol} / 0.5\text{dm}^3 = 0.05 \text{ mol dm}^{-3}$$

	[HNO ₂] (mol dm ⁻³)	[H ₃ O ⁺] (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

$$K_a = \frac{[\text{H}^+][\text{NO}_2^{-1}]}{[\text{HNO}_2]}$$

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

$$= 5.152 \times 10^{-4}$$

∴ The K_a is 5.15 × 10⁻⁴ mol dm⁻³