

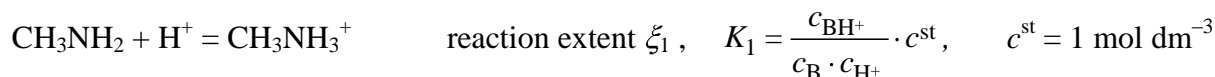
Problem 8-06 Solution of weak base – acidity and dissociation constants

The measurement of pH of methylamine solution of the concentration $0.002 \text{ mol dm}^{-3}$ at the temperature of 30.1°C yielded the value $\text{pH} = 10.87$. Calculate the acidity constant of CH_3NH_3^+ and the dissociation constant of methylamine (both for standard state of infinite dilution, $c^{\text{st}} = 1 \text{ mol dm}^{-3}$). Ionic product of water at 30.1°C is $K_w = 1.484 \cdot 10^{-14}$ and all the activity coefficients can be taken as equal to one.

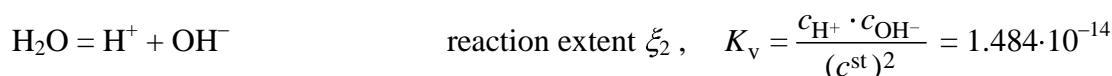
$$[K_{\text{BH}^+} = 1.104 \cdot 10^{-11}, K_{\text{dis}} = 1.344 \cdot 10^{-3}]$$

Řešení:

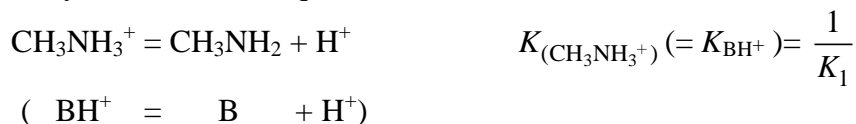
Base is a substance accepting a proton: (notation: $\text{CH}_3\text{NH}_2 \equiv \text{B}$):



At the same time occurs the dissociation of water:



Acidity constant is the equilibrium constant of the reaction



Balance:

$$c_{\text{B}} = c_0 - x_1, \quad c_0 = 0.002 \text{ mol dm}^{-3}, \quad x_1 = \xi_1/V$$

$$c_{\text{BH}^+} = x_1$$

$$c_{\text{H}^+} = x_2 - x_1 = 10^{-\text{pH}} = 10^{-10.87} = 1.349 \cdot 10^{-11} \text{ mol dm}^{-3}$$

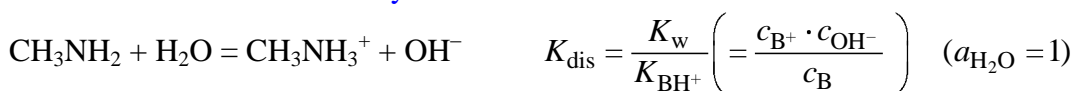
$$c_{\text{OH}^-} = x_2 = \frac{K_v}{c_{\text{H}^+}} = \frac{1.484 \cdot 10^{-14}}{1.349 \cdot 10^{-11}} = 0.0011 \text{ mol dm}^{-3}$$

$$x_1 = x_2 - c_{\text{H}^+} = 0.0011 - 1.349 \cdot 10^{-11} \doteq 0.0011 \text{ mol dm}^{-3}$$

Acidity constant of BH^+

$$K_{\text{BH}^+} = \frac{1}{K_1} = \frac{c_{\text{B}} \cdot c_{\text{H}^+}}{c_{\text{BH}^+}} = \frac{(0.002 - 0.0011) \cdot 1.349 \cdot 10^{-11}}{0.0011} = 1.104 \cdot 10^{-11}$$

Dissociation constant of methylamine:



$$K_{\text{dis}} = \frac{K_w}{K_{\text{BH}^+}} = \frac{1.484 \cdot 10^{-14}}{1.104 \cdot 10^{-11}} = 1.344 \cdot 10^{-3}$$