

Problem 8-05 Solution of weak base – pH

What is pH value of the ammonia solution with concentration of $0.002 \text{ mol dm}^{-3}$ at the temperature of 20°C . The value of acidity constant of ion NH_4^+ is $5.56 \cdot 10^{-10}$ and the ionic product of water is $K_w = 6.8 \cdot 10^{-15}$ (both for standard state of infinite dilution, $c^{\text{st}} = 1 \text{ mol dm}^{-3}$). You can assume that the amount of hydrogen ions formed by water autoprotolysis is negligible small and all activity coefficients are equal to one.

[pH = 10.345]

Solution:

NH_4^+ acidity constant is the equilibrium constant of the reaction

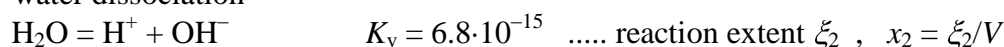


We take into consideration two simultaneous reaction:

- base NH_3 accepts proton



- water dissociation



Balance:

$$c_{\text{NH}_3} = c_0 - x_1, \quad c_0 = 0.002 \text{ mol dm}^{-3}$$

$$c_{\text{NH}_4^+} = x_1$$

$$c_{\text{OH}^-} = x_2$$

$$c_{\text{H}^+} = x_2 - x_1 \rightarrow 0 \quad \Rightarrow \quad x_1 \approx x_2 \equiv x$$

$$K_w \approx c_{\text{H}^+} \cdot c_{\text{OH}^-}$$

$$K = \frac{1}{K_{\text{NH}_4^+}} = \frac{c_{\text{NH}_4^+}}{c_{\text{NH}_3} \cdot c_{\text{H}^+}}$$

$$\frac{K_w}{K_{\text{NH}_4^+}} = \frac{c_{\text{NH}_4^+} \cdot c_{\text{OH}^-}}{c_{\text{NH}_3}} = \frac{x^2}{c_0 - x}$$

$$\frac{6.8 \cdot 10^{-15}}{5.56 \cdot 10^{-10}} = \frac{x^2}{0.002 - x}$$

$$1.223 \cdot 10^{-5} \cdot 0.002 - 1.223 \cdot 10^{-5} \cdot x = x^2$$

$$x = -6.115 \cdot 10^{-6} \pm (3.73945 \cdot 10^{-11} + 2.446 \cdot 10^{-8})^{1/2} =$$
$$= -6.115 \cdot 10^{-6} + 1.56516 \cdot 10^{-4} = 1.504 \cdot 10^{-4} \text{ mol dm}^{-3}$$

$$\text{pOH} = -\log 1.504 \cdot 10^{-4} = 3.82275$$

$$\text{pH} = -\log K_w - \text{pOH} = -\log (6.8 \cdot 10^{-15}) - 3.82275 = -(-14.16749) - 3.82275 = 10.3447$$

pH = 10.345

Note: Acidity constant and dissociation constant are connected by a relation:

