Acids and Bases

#6

Calculate the pH of a solution of 5.258 g NH_4Cl in 750 mL of solution.

Solution

NH_4Cl:

$$5.258\ \text{g} = \frac{5.258\ \text{g}}{53.5 \frac{\text{g}}{\text{mol}}} = 0.0983\ \text{mol}$$

$$\frac{0.0983\ \text{mol}}{0.750\ \text{L}} = 0.131\ \frac{\text{mol}}{\text{L}}$$

NH_4Cl:

= salt, completely dissociated

$NH_4Cl\ (aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$

NH_4^+ = weak acid (stronger than water)

Cl^- = weak base, weaker than water

$NH_4^+(aq) \rightarrow NH_3(aq) + H^+(aq)$ $\qquad K_{a_{NH_4^+}} = \frac{[NH_3] \times [H^+]}{[NH_4^+]} = 5.6 \times 10^{10}$

$\frac{\text{mol}}{\text{L}}$	NH_4^+	NH_3	H^+
Start	0.131	0	0*
Δ	– x	+ x	+ x
Equilibrium	0.131 – x	x	x

(*) The hydrogen ions delivered by the ionization of water are neglected. The weak acid NH_4^+ ($K_a = 5.6 \times 10^{-10}$) is much stronger than the weak acid H_2O ($K_a = 1.0 \times 10^{-14}$).

$$K_{a_{NH_4^+}} = \frac{x^2}{0.131 - x} = 5.6 \times 10^{-10}$$

$$\Rightarrow x^2 + 5.6 \times 10^{-10} x - 7.3 \times 10^{-11} = 0$$

$$\Rightarrow x = 8.5 \times 10^{-6}$$

$\frac{\text{mol}}{\text{L}}$	NH_4^+	NH_3	H^+
Equilibrium	0.131	8.5×10^{-6}	8.5×10^{-6}

$\Rightarrow$ **pH = 5.07**