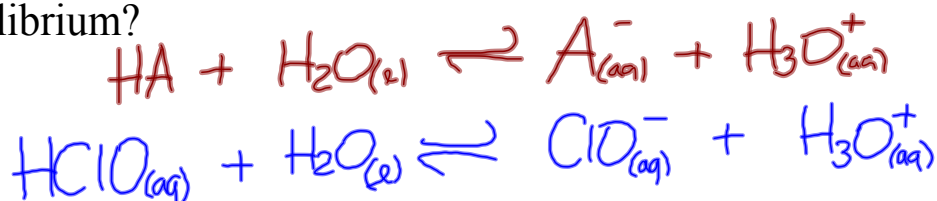


#22, 23 p. 610

What is the pH of a 0.150 mol/L hypochlorous acid solution at equilibrium?



0.150 mol/L

$$K_a = \frac{[\text{ClO}^-_{(aq)}][\text{H}_3\text{O}^+_{(aq)}]}{[\text{HClO}_{(aq)}]}, \quad [\text{ClO}^-_{(aq)}] = [\text{H}_3\text{O}^+_{(aq)}]$$

$$K_a = \frac{[\text{H}_3\text{O}^+_{(aq)}]^2}{[\text{HClO}_{(aq)}]}$$

$$2.9 \times 10^{-8} = \frac{[\text{H}_3\text{O}^+_{(aq)}]^2}{[0.150]}$$

$$[\text{H}_3\text{O}^+_{(aq)}] = \sqrt{(2.9 \times 10^{-8})(0.150)}$$

$$\underline{[\text{H}_3\text{O}^+_{(aq)}] = 6.59 \times 10^{-5} \text{ M}}$$

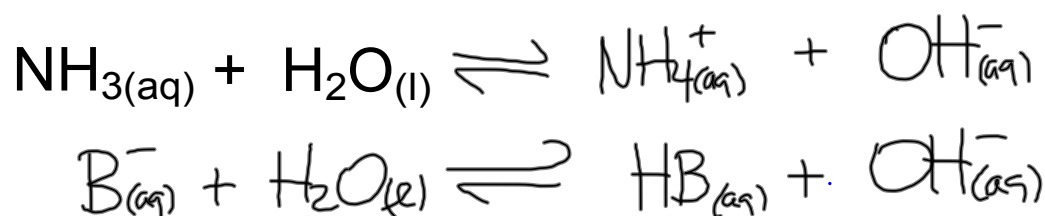
$$\text{pH} = -\log [\text{H}_3\text{O}^+_{(aq)}]$$

$$\text{pH} = -\log [6.59 \times 10^{-5}]$$

$$\boxed{\text{pH} = 4.181}$$

Weak Bases

Weak bases react with water to form the hydroxide ion and conjugate acid of the base.



$$K = \frac{[\text{NH}_4^+(\text{aq})][\text{OH}^-(\text{aq})]}{[\text{NH}_3(\text{aq})][\text{H}_2\text{O}(\text{l})]}$$

$$K_b = \frac{[\text{NH}_4^+(\text{aq})][\text{OH}^-(\text{aq})]}{[\text{NH}_3(\text{aq})]}$$

base dissociation constant

Calculate the pH of a 0.221 mol/L solution of $\text{NH}_3(\text{aq})$ at equilibrium.



$$K_b = \frac{[\text{NH}_4^+(\text{aq})][\text{OH}^-(\text{aq})]}{[\text{NH}_3(\text{aq})]}, \quad [\text{NH}_4^+(\text{aq})] = [\text{OH}^-(\text{aq})]$$

$$K_b = \frac{[\text{OH}^-(\text{aq})]^2}{[\text{NH}_3(\text{aq})]}$$

$$K_a K_b = K_w$$

$$K_b = \frac{K_w}{K_a}$$

$$\frac{K_w}{K_a} = \frac{[\text{OH}^-(\text{aq})]^2}{[\text{NH}_3(\text{aq})]}$$

$$\frac{1.0 \times 10^{-14}}{5.8 \times 10^{-10}} = \frac{[\text{OH}^-(\text{aq})]^2}{0.221}$$

$$\sqrt{(1.72 \times 10^{-5})(0.221)} = [\text{OH}^-(\text{aq})]$$

$$\underline{[\text{OH}^-(\text{aq})] = 0.00195 \text{ M}}$$

$$p\text{OH} = -\log [\text{OH}^-(\text{aq})]$$

$$p\text{OH} = -\log [0.00195]$$

$$\underline{p\text{OH} = 2.710}$$

$$p\text{H} + p\text{OH} = 14.000$$

$$p\text{H} = 14.000 - 2.710$$

$$\boxed{p\text{H} = 11.290}$$

Worksheet