Lesson 10

Calculating pH of Bases

Learning Goals

- ❑ I will be able to calculate the pH of solutions of bases.

$$pH = -\log[H^+]$$

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$$

$$K_w = [H^+][OH^-]$$

$$= 1.0\times10^{-14} \quad \text{(at 25°C)}$$

Example 1

Calculate the pH of a 0.080 mol/L barium hydroxide solution.

$Ba(OH)_2(aq) \rightarrow Ba^{2+}(aq) + 2\ OH^-(aq)$

$c_{Ba(OH)2} = 0.080$ mol/L

strong base; 100% dissociated

$[OH^-] = 2(0.080 \text{ mol/L})$
$= 0.16 \text{ mol/L}$

$K_w = [H^+][OH^-]$

$$[H^+] = \frac{K_w}{[OH^-]}$$
$$= \frac{1.0\times10^{-14}}{0.16}$$
$$= 6.25\times10^{-14} \text{ mol/L}$$

$$pH = -\log[H^+]$$
$$= -\log(6.25\times10^{-14})$$
$$= -(-13.2041...)$$
$$= 13.2041...$$

Therefore, the pH of the solution is 13.20.

Example 2

Calculate the pH of a 0.080 mol/L trimethylamine solution.

$(CH_3)_3N_{(aq)} + H_2O_{(l)} \rightleftharpoons (CH_3)_3NH^+_{(aq)} + OH^-_{(aq)}$ $K_b = 7.4 \times 10^{-5}$

$c_{(CH3)3N} = 0.080$ mol/L

weak base; equilibrium

	$(CH_3)_3N_{(aq)}$	$+ H_2O_{(l)} \rightleftharpoons$	$(CH_3)_3NH^+_{(aq)}$	$+ OH^-_{(aq)}$
INITIAL CONCENTRATION	0.080	—	0	~0
CHANGE IN CONCENTRATION	−x	—	+x	+x
EQUILIBRIUM CONCENTRATION	0.080−x	—	x	x

EQUILIBRIUM CONCENTRATION	$0.080-x$	x	x

$$K_b = \frac{[(CH_3)_3NH^+][OH^-]}{[(CH_3)_3N]}$$

$$7.4\times10^{-5} = \frac{(x)\,(x)}{0.080-x}$$

$$7.4\times10^{-5} = \frac{x^2}{0.080-x}$$

assume $0.080-x = 0.080$

$$7.4\times10^{-5} = \frac{x^2}{0.080}$$

$$\sqrt{(7.4\times10^{-5})(0.080)} = x$$

$$2.4331...\times10^{-3} = x$$

$$[OH^-]_{eq} = x \text{ mol/L}$$
$$= 2.4331...\times10^{-3} \text{ mol/L}$$

$$[H^+] = \frac{K_w}{[OH^-]}$$
$$= \frac{1.0\times10^{-14}}{2.4331...\times10^{-3}}$$
$$= 4.1099...\times10^{-12} \text{ mol/L}$$

$$pH = -\log[H^+]$$
$$= -\log(4.1099...\times10^{-12})$$
$$= -(-11.3861...)$$
$$= 11.3861...$$

Therefore, the pH of the solution is 11.39.

Success Criteria

- ❑ I can calculate the pH of solutions of bases.

PRACTICE:

- Worksheet: ‘PRACTICE: **CALCULATING pH OF BASES’**