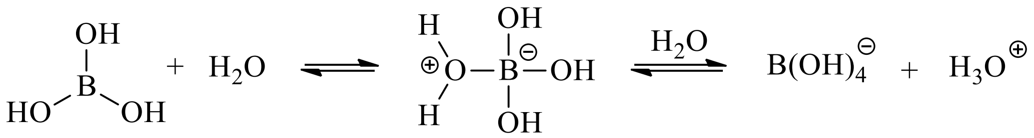
Marks
9

- Boric acid, $B(OH)_3$, is a weak acid (pK_a = 9.24) that is used as a mild antiseptic and eye wash. Unusually, the Lewis acidity of the compound accounts for its Brønsted acidity. By using an appropriate chemical equation, show how this compound acts as a Brønsted acid in aqueous solution.

The boron atom in $B(OH)_3$ is electron deficient. It acts as a Lewis acid by readily accepting the lone pair from the oxygen in a water molecule to go from sp^2 to sp^3 hybridisation.



Solution A consists of a 0.050 M aqueous solution of boric acid at 25 °C. Calculate the pH of Solution A.

As boric is a weak acid, $[H_3O^+]$ must be calculated using a reaction table (acid = $B(OH)_3$ and base = $B(OH)_2^-$)

	acid	**H_2O**	$\rightleftharpoons$	**H_3O^+**	**base**
initial	**0.050**	**large**		**0**	**0**
change	**$-x$**	**negligible**		**$+x$**	**$+x$**
final	**$0.050 - x$**	**large**		**x**	**x**

The equilibrium constant K_a is given by: $K_a = \dfrac{[H_3O^+][\text{base}]}{[\text{acid}]} = \dfrac{x^2}{0.050 - x}$

As pK_a = $-\log_{10}K_a$, $K_a = 10^{-9.24}$ and is very small, $0.050 - x \sim 0.050$ and hence:

$x^2 = 0.050 \times 10^{-9.24}$ **or** $x = 5.36 \times 10^{-6}$ **M** $= [H_3O^+]$

Hence, the pH is given by:

pH $= -\log_{10}[H_3O^+] = -\log_{10}(5.36 \times 10^{-6}) = 5.27$

pH = **5.27**

At 25 °C, 1.00 L of Solution B consists of 10.18 g of $NaB(OH)_4$ dissolved in water. Calculate the pH of Solution B.

The molar mass of $NaB(OH)_3$ is:

molar mass = [22.99 (Na) + 10.81 (B) + 4 (16.00 (O) + 1.008 (H))] g mol^{-1}
= 101.83 g mol^{-1}

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A mass of 10.18 g therefore corresponds to:

number of moles = mass / molar mass = 10.18 g / (101.83 g mol^{-1}) = 0.100 mol

A 1.00 L solution contains this amount has a concentration of 0.100 M.

As it is a weak base, [OH$^-$] must be calculated by considering the equilibrium:

	base	H_2O	$\rightleftharpoons$	acid	OH^-
initial	0.100	large		0	0
change	$-y$	negligible		$+y$	$+y$
final	$0.10 - y$	large		y	y

The equilibrium constant K_b is given by:

$$K_b = \frac{[\text{acid}][\text{OH}^-]}{[\text{base}]} = \frac{y^2}{(0.100 - y)}$$

For an acid and its conjugate base:

$$\text{p}K_a + \text{p}K_b = 14.00$$

$$\text{p}K_b = 14.00 - 9.24 = 4.76$$

As pK_b = 4.76, $K_b = 10^{-4.76}$. K_b is very small so $0.100 - y \sim 0.100$ and hence:

$$y^2 = 0.100 \times 10^{-4.76} \text{ or } y = 0.00132 \text{ M} = [\text{OH}^-]$$

Hence, the pOH is given by:

$$\text{pOH} = -\log_{10}[\text{OH}^-] = \log_{10}[0.00132] = 2.88$$

Finally, pH + pOH = 14.00 so

$$\text{pH} = 14.00 - 2.88 = 11.12$$

pH = **11.12**

Using both Solutions A and B, calculate the volumes (mL) required to prepare a 1.0 L solution with a pH = 8.50.

The ratio of acid to conjugate base needed can be calculated using the Henderson-Hasselbalch equation, pH = pK_a + log$\frac{[\text{base}]}{[\text{acid}]}$:

$$8.50 = 9.24 + \log\frac{[\text{base}]}{[\text{acid}]} \quad \text{so} \quad \frac{[\text{base}]}{[\text{acid}]} = 10^{-0.74} = 0.182$$

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A volume V_a of the acid and V_b of base are added together to give a solution with a total volume of 1.0 L so:

$$V_a + V_b = 1.0 \text{ L}$$

Using $c_1V_1 = c_2V_2$, this mixing reduces the concentration of both:

acid: $(0.050 \text{ M}) \times V_a = c_{acid} \times (1.0 \text{ L})$ **so** $V_a = 20. \times c_{acid}$
base: $(0.100 \text{ M}) \times V_b = c_{base} \times (1.0 \text{ L})$ **so** $V_b = 10.0 \times c_{base}$

Using the concentration ratio from the Henderson-Hasselbalch equation above, the ratio of the volumes needed is therefore:

$$V_b / V_a = (10.0 / 20.) \times c_{base} / c_{acid} = (10. / 20.) \times 0.182 = 0.0910$$

or

$$V_b = 0.0910 \times V_a$$

From above, $V_a + V_b = 1.0$ L so:

$$V_a + (0.0910 \times V_a) = 1.0 \text{ L}$$
$$1.0091\ V_a = 1.0 \text{ L}$$
$$V_a = 0.917 \text{ L}$$

Hence, $V_b = 0.083$ L.