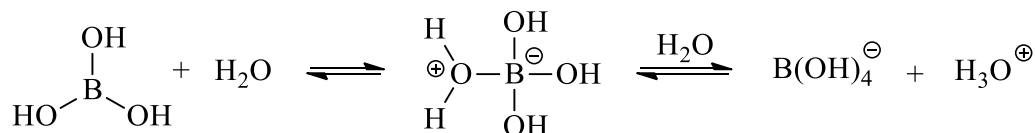


Marks
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- Boric acid, B(OH)_3 , is a weak acid ($\text{p}K_a = 9.24$) that is used as a mild antiseptic and eye wash. Unusually, the Lewis acidity of the compound accounts for its Brønsted acidity. By using an appropriate chemical equation, show how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

The boron atom in B(OH)_3 is electron deficient: it has 6 rather than 8 electrons in its valence shell. 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 0.60 M aqueous solution of boric acid at 25 °C. Calculate the pH of Solution A.

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

	acid	H_2O	\rightleftharpoons	H_3O^+	base
initial	0.60	large		0	0
change	-x	negligible		+x	+x
final	0.60 - x	large		x	x

The equilibrium constant K_a is given by: $K_a = \frac{[\text{H}_3\text{O}^+][\text{base}]}{[\text{acid}]} = \frac{x^2}{0.60-x}$

As $\text{p}K_a = -\log_{10}K_a$, $K_a = 10^{-9.24}$ and is very small, $0.60 - x \sim 0.60$ and hence:

$$x^2 = 0.60 \times 10^{-9.24} \quad \text{or} \quad x = 1.86 \times 10^{-5} \text{ M} = [\text{H}_3\text{O}^+]$$

Hence, the pH is given by:

$$\text{pH} = -\log_{10}[\text{H}_3\text{O}^+] = -\log_{10}(1.86 \times 10^{-5}) = 4.73$$

$$\text{pH} = 4.73$$

At 25 °C, 1.00 L of Solution B consists of 112 g of NaB(OH)_4 dissolved in water. Calculate the pH of Solution B.

The molar mass of NaB(OH)_3 is:

$$\begin{aligned} \text{molar mass} &= [22.99 (\text{Na}) + 10.81 (\text{B}) + 4 (16.00 (\text{O}) + 1.008 (\text{H}))] \text{ g mol}^{-1} \\ &= 101.83 \text{ g mol}^{-1} \end{aligned}$$

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A mass of 112 g therefore corresponds to:

$$\text{number of moles} = \text{mass} / \text{molar mass} = 112 \text{ g} / (101.83 \text{ g mol}^{-1}) = 1.10 \text{ mol}$$

A 1.00 L solution contains this amount has a concentration of 1.10 M.

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

	base	H_2O	\rightleftharpoons	acid	OH^-
initial	1.10	large		0	0
change	-y	negligible		+y	+y
final	$1.10 - y$	large		y	y

The equilibrium constant K_b is given by:

$$K_b = \frac{[\text{acid}][\text{OH}^-]}{[\text{base}]} = \frac{y^2}{(1.10 - y)}$$

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 $\text{p}K_b = 4.76$, $K_b = 10^{-4.76}$. K_b is very small so $1.00 - y \sim 1.00$ and hence:

$$y^2 = 1.10 \times 10^{-4.76} \text{ or } y = 0.00437 \text{ M} = [\text{OH}^-]$$

Hence, the pOH is given by:

$$\text{pOH} = -\log_{10}[\text{OH}^-] = \log_{10}[0.00437] = 2.36$$

Finally, $\text{pH} + \text{pOH} = 14.00$ so

$$\text{pH} = 14.00 - 2.36 = 11.64$$

$$\text{pH} = 11.64$$

Using both Solutions A and B, calculate the volumes (mL) required to prepare a 1.0 L solution with a $\text{pH} = 9.24$.

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

$$9.24 = 9.24 + \log \frac{[\text{base}]}{[\text{acid}]} \text{ so } \frac{[\text{base}]}{[\text{acid}]} = 1.00 \text{ or } [\text{base}] = [\text{acid}]$$

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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 1.00 L so:

$$V_a + V_b = 1.00 \text{ L}$$

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

$$\begin{aligned} \text{acid: } (0.60 \text{ M}) \times V_a &= c_{\text{acid}} \times (1.0 \text{ L}) \quad \text{so } V_a = 1.67 \times c_{\text{acid}} \\ \text{base: } (1.10 \text{ M}) \times V_b &= c_{\text{base}} \times (1.0 \text{ L}) \quad \text{so } V_b = 0.909 \times c_{\text{base}} \end{aligned}$$

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

$$V_b / V_a = (0.909 / 1.67) \times c_{\text{base}} / c_{\text{acid}} = (0.909 / 1.67) \times 1.00 = 0.545$$

or

$$V_b = 0.545 \times V_a$$

From above, $V_a + V_b = 1.00 \text{ L}$ so:

$$\begin{aligned} V_a + (0.545 \times V_a) &= 1.00 \text{ L} \\ 1.545 V_a &= 1.00 \text{ L} \\ V_a &= 0.647 \text{ L} \end{aligned}$$

Hence, $V_b = (1.00 - 0.647) \text{ L} = 0.35 \text{ L}$.

Answer: **650 mL of solution A and 350 mL of solution B**