

- The K_a of benzoic acid is 6.3×10^{-5} M at 25 °C.

Calculate the pH of a 0.0100 M aqueous solution of sodium benzoate (C_6H_5COONa).

Marks
5

As benzoic acid is a weak acid, its conjugate base, $C_6H_5COO^-$, is a weak base and so $[OH^-]$ must be calculated using the reaction table:

	$C_6H_5COO^-$	H_2O	\rightleftharpoons	OH^-	C_6H_5COOH
initial	0.0100	large		0	0
change	-x	negligible		+x	+x
final	0.0100 - x	large		x	x

The equilibrium constant K_b is given by:

$$K_b = \frac{[OH^-][C_6H_5COOH]}{[C_6H_5COO^-]} = \frac{x^2}{0.0100 - x}$$

For an acid and its conjugate base in aqueous solution, $K_a \times K_b = K_w = 10^{-14}$. Hence,

$$K_b = \frac{10^{-14}}{6.3 \times 10^{-5}} = 1.6 \times 10^{-10}$$

As K_b is very small, $0.0100 - x \sim 0.0100$ and hence:

$$x^2 = 0.0100 \times (1.6 \times 10^{-10}) \quad \text{or} \quad x = 1.3 \times 10^{-6} \text{ M} = [OH^-(aq)]$$

Hence, the pOH is given by:

$$pOH = -\log_{10}[OH^-] = -\log_{10}[1.3 \times 10^{-6}] = 5.9$$

Finally, $pH + pOH = 14.0$ so

$$pH = 14.0 - 5.9 = 8.1$$

Answer: **pH = 8.1**

Answer: **pH = 4.1**

ANSWER CONTINUES ON THE NEXT PAGE

A buffer solution is prepared by adding 375 mL of this 0.0100 M aqueous solution of sodium benzoate to 225 mL of 0.0200 M aqueous benzoic acid. Calculate the pH of the buffer solution.

375 mL of a 0.0100 of benzoate contains,

$$\text{moles of benzoate} = \text{volume} \times \text{concentration} = 0.375 \times 0.0100 = 3.75 \times 10^{-3} \text{ mol}$$

225 mL of a 0.0200 of benzoic acid contains,

$$\text{moles of benzoic acid} = 0.225 \times 0.0200 = 4.50 \times 10^{-3} \text{ mol}$$

The mixture has a volume of (375 + 225) = 600 mL so the concentrations of benzoate (base) and benzoic acid (acid) are:

$$[\text{base}] = \frac{\text{number of moles}}{\text{volume}} = \frac{3.75 \times 10^{-3}}{0.600} = 6.25 \times 10^{-3} \text{ M}$$

$$[\text{acid}] = \frac{4.50 \times 10^{-3}}{0.600} = 7.50 \times 10^{-3} \text{ M}$$

As $pK_a = -\log_{10}K_a$,

$$pK_a = -\log_{10}(6.3 \times 10^{-5}) = 4.2$$

The pH of the buffer can be calculated using the Henderson-Hasselbalch equation,

$$\text{pH} = pK_a + \log_{10} \left(\frac{[\text{base}]}{[\text{acid}]} \right) = 4.2 + \log_{10} \left(\frac{6.25 \times 10^{-3}}{7.50 \times 10^{-3}} \right) = 4.1$$

Answer: pH = 4.1