Marks 5

- 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).

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 ~ 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,

moles of benzoate = volume×concentration = 0.375×0.0100 = 3.75×10^{-3} mol

225 mL of a 0.0200 of benzoic acid contains,

moles of benzoic acid = 0.225×0.0200 = 4.50×10^{-3} mol

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

$$[\text{base}] = \frac{\text{number of moles}}{\text{volume}} = \frac{3.75\times10^{-3}}{0.600} = 6.25\times10^{-3}\ \text{M}$$

$$[\text{acid}] = \frac{4.50\times10^{-3}}{0.600} = 7.50\times10^{-3}\ \text{M}$$

As pK_a = -log$_{10}K_a$,

pK_a = -log$_{10}$(6.3 × 10^{-5}) = 4.2

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

$$\text{pH} = \text{pK}_a + \log_{10}\left(\frac{[\text{base}]}{[\text{acid}]}\right) = 4.2 + \log_{10}\left(\frac{6.25\times10^{-3}}{7.50\times10^{-3}}\right) = 4.1$$

Answer: **pH = 4.1**