Marks

- Codeine, a cough suppressant extracted from crude opium, is a weak base with a $pK_b = 5.79$. What is the pH of a 0.020 M solution of codeine? **2**

As codeine, is a weak base and so $[OH^-]$ must be calculated. for example using a reaction table:

	codeine	H_2O	$\rightleftharpoons$	codeineH^+	OH^-
initial	**0.020**	**large**		**0**	**0**
change	**$-x$**	**negligible**		**$+x$**	**$+x$**
final	**$0.020 - x$**	**large**		**x**	**x**

The equilibrium constant K_b is given by:

$$K_b = \frac{[\text{codeineH}^+][\text{OH}^-]}{[\text{codeine}]} = \frac{x^2}{(0.020 - x)}$$

As $pK_b = -\log_{10}K_b = 5.79$, $K_b = 10^{-5.79}$. Hence,

$$\frac{x^2}{(0.020 - x)} = 10^{-5.79}$$

As K_b is very small, $0.020 - x \sim 0.020$ and hence:

$$x^2 = 0.020 \times 10^{-5.79} \quad \text{or} \quad x = 1.8 \times 10^{-4}\ \text{M} = [\text{OH}^-(\text{aq})]$$

Hence, the pOH is given by:

$$\text{pOH} = -\log_{10}[\text{OH}^-] = -\log_{10}[1.8 \times 10^{-4}] = 3.74$$

Finally, pH + pOH = 14.00 so

$$\text{pH} = 14.0 - 3.74 = 10.26$$

Answer: **10.26**

- A buffer solution is formed with 0.250 M CH_3COOH and 0.350 M CH_3COONa. What is the pH of this buffer solution? (K_a of acetic acid $= 1.8 \times 10^{-5}$ M.) **3**

The pH of a buffer solution is given by the Henderson-Hasselbalch equation:

$$\text{pH} = \text{p}K_a + \log\left(\frac{\text{base}}{\text{acid}}\right)$$

As $pK_a = -\log_{10}K_a = -\log_{10}(1.8\times10^{-5}) = 4.74$. With [base] = $[CH_3COONa]$ = 0.350 M and [acid] = $[CH_3COOH]$ = 0.250 M,

ANSWER CONTINUES ON THE NEXT PAGE

$$\mathbf{pH = 4.74 + \log\left(\frac{0.350}{0.250}\right) = 4.89}$$

As the buffer contains a higher concentration of base than acid, the pH > pK_a.

Answer: **4.89**

Calculate the pH of the solution formed when 6.3×10^{-2} mol of NaOH is added to 1.0 L of the buffer solution.

In a 1.0 L solution of the buffer, there is 0.250 mol of CH_3COOH and 0.350 mol of CH_3COO^-.

The OH^- will react with the CH_3COOH to produce CH_3COO^-. The concentration of the former will therefore decrease whilst the concentration of the latter will increase. After the OH^- is added:

number of moles of CH_3COOH = $(0.250 - 6.3 \times 10^{-2})$ M = 0.187 mol
number of moles of CH_3COO^- = $(0.350 + 6.3 \times 10^{-2})$ M = 0.413 mol

As the volume of solution does not change, these are also the new acid and base concentrations. Hence, the buffer now has:

$$\mathbf{pH = 4.74 + \log\left(\frac{0.413}{0.187}\right) = 5.09}$$

As base has been added, there is an increase in the pH. As it is being added to a buffer system, this change is small. Addition of this quantity of base to water would increase the pH by 1.20 units.

Answer: **5.09**