Calculate the pH of a 0.010 M solution of aspirin at 25 °C. The pK_a of aspirin is 3.5 at this temperature.

Marks 7

As aspirin is a weak acid, $[H_3O^+]$ must be calculated using a reaction table:

	$C_9H_8O_4$	H_2O	$\rightleftharpoons$	H_3O^+	$C_9H_7O_4^-$
initial	**0.010**	**large**		**0**	**0**
change	**$-x$**	**negligible**		**$+x$**	**$+x$**
final	**$0.010 - x$**	**large**		**x**	**x**

The equilibrium constant K_a is given by:

$$K_a = \frac{[H_3O^+][C_9H_7O_4^-]}{[C_9H_7O_4]} = \frac{x^2}{0.010 - x}$$

As pK_a = $-\log_{10}K_a$, $K_a = 10^{-3.5}$ and is very small, $0.010 - x \sim 0.010$ and hence:

$x^2 = 0.010 \times 10^{-3.5}$ **or** $x = 1.8 \times 10^{-3}$ M $= [H_3O^+]$

Hence, the pH is given by:

$$pH = -\log_{10}[H_3O^+] = -\log_{10}(1.8 \times 10^{-3}) = 2.8$$

pH = **2.8**

Aspirin, $C_9H_8O_4$ is not very soluble. "Soluble aspirin" can be made by reacting aspirin with sodium hydroxide. Write the chemical equation for this reaction.

$$C_9H_8O_4(s) + OH^-(aq) \rightarrow C_9H_7O_4^-(aq) + H_2O(l)$$

Is a solution of "soluble aspirin" acidic or basic? Briefly explain your answer.

Basic. The $C_9H_7O_4^-$(aq) ion reacts with water (*i.e.* undergoes hydrolysis) to generate a small amount of OH^- ions. The $C_9H_7O_4^-$(aq) ion is a weak base, so the following equilibrium reaction lies very much in favour of the reactants.

$$C_9H_7O_4^-(aq) + H_2O(l) \rightleftharpoons C_9H_8O_4(aq) + OH^-(aq)$$