

- The active ingredient in aspirin is the monoprotic acid, acetylsalicylic acid ($\text{HC}_9\text{H}_7\text{O}_4$) that has a K_a of 3.3×10^{-4} M at 25°C . What is the pH of a solution obtained when a tablet containing 200 mg of acetylsalicylic acid is dissolved in 125 mL of water?

The molar mass is $(8 \times 1.008 \text{ (H)}) + (9 \times 12.01) + (4 \times 16.00 \text{ (O)}) \text{ g mol}^{-1} = 180.154 \text{ g mol}^{-1}$ so 200 mg contains:

$$n(\text{HC}_9\text{H}_7\text{O}_4) = \frac{\text{mass}}{\text{molar mass}} = \frac{(200 \times 10^{-3} \text{ g})}{(180.154 \text{ g mol}^{-1})} = 1.11 \times 10^{-3} \text{ mol}$$

When this amount is dissolved in 125 mL, the concentration is:

$$n[\text{HC}_9\text{H}_7\text{O}_4] = \frac{\text{number of moles}}{\text{volume}} = \frac{(1.11 \times 10^{-3} \text{ mol})}{(125/1000 \text{ L})} = 8.88 \times 10^{-3} \text{ M}$$

The reaction table is:

	$\text{HC}_9\text{H}_7\text{O}_4$	H_2O	\rightleftharpoons	H_3O^+	$\text{C}_9\text{H}_7\text{O}_4^-$
initial	8.88×10^{-3}	large		0	0
change	$-x$	negligible		$+x$	$+x$
final	$8.88 \times 10^{-3} - x$	large		x	x

The equilibrium constant K_a is given by:

$$K_a = \frac{[\text{H}_3\text{O}^+(\text{aq})][\text{C}_9\text{H}_7\text{O}_4^-(\text{aq})]}{[\text{HC}_9\text{H}_7\text{O}_4(\text{aq})]} = \frac{x^2}{(8.88 \times 10^{-3} - x)} = 3.3 \times 10^{-4}$$

K_a is not sufficiently small in comparison to the initial concentration of acid that any approximation to this equation can be made. Hence, the quadratic expression must be solved:

$$x^2 = (3.3 \times 10^{-4} \times 8.88 \times 10^{-3}) - (3.3 \times 10^{-4})x$$

$$x^2 + (3.3 \times 10^{-4})x - (2.9 \times 10^{-6})$$

With $a = 1$, $b = +3.3 \times 10^{-4}$ and $c = -2.9 \times 10^{-6}$, the roots are:

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} = \frac{-(3.3 \times 10^{-4}) \pm \sqrt{(3.3 \times 10^{-4})^2 - (4 \times 1 \times -2.9 \times 10^{-6})}}{(2 \times 1)}$$

Only the positive root has physical significance so $x = 1.55 \times 10^{-3}$

$$[\text{H}_3\text{O}^+(\text{aq})] = x = 1.55 \times 10^{-3} \text{ M, pH} = -\log_{10}([\text{H}_3\text{O}^+(\text{aq})]) = -\log_{10}(1.55 \times 10^{-3}) = 2.81$$

Answer: pH = 2.81

ANSWER CONTIUNES ON THE NEXT PAGE

- A standard test for the presence of chloride ion in water involves the appearance of a precipitate of AgCl upon addition of 1 mL of AgNO₃ (0.03 M) to 100 mL of the water sample. What is the minimum concentration of Cl⁻ detectable by this method? $K_{sp}(\text{AgCl}) = 1.8 \times 10^{-10} \text{ M}^2$.

The number of moles of Ag⁺(aq) in 1 mL of 0.03 M AgNO₃ is:

$$n(\text{Ag}^+(\text{aq})) = \text{concentration} \times \text{volume} = (0.03 \text{ M}) \times (1 \times 10^{-3} \text{ L}) = 3 \times 10^{-5} \text{ mol}$$

In the test, this amount is present in (100 + 1) = 101 mL so its concentration is:

$$[\text{Ag}^+(\text{aq})] = \frac{\text{number of moles}}{\text{volume}} = \frac{(3 \times 10^{-5} \text{ mol})}{(101 \times 10^{-3} \text{ L})} = 3 \times 10^{-4} \text{ M}$$

The solubility equilibrium and product for AgCl are given by:



$$K_{sp} = [\text{Ag}^+(\text{aq})][\text{Cl}^-(\text{aq})] = 1.8 \times 10^{-10}$$

Hence,

$$[\text{Cl}^-(\text{aq})] = \frac{K_{sp}}{[\text{Ag}^+(\text{aq})]} = \frac{(1.8 \times 10^{-10})}{(3 \times 10^{-4})} = 6 \times 10^{-7} \text{ M}$$

If [Cl⁻(aq)] is less than this value, AgCl will not precipitate.

$$\text{Answer: } [\text{Cl}^-(\text{aq})] = 6 \times 10^{-7} \text{ M}$$