Determining the pK_a of 2-hydroxybenzoic acid

Student worksheet

Health and safety note

Wear eye protection and ensure no naked flames. 0.10 mol dm^{-3} sodium hydroxide solution is an irritant. 95% ethanol is highly flammable. 2-hydroxybenzoic acid and 95% ethanol are both harmful.

Principle

2-hydroxybenzoic acid is a weak acid

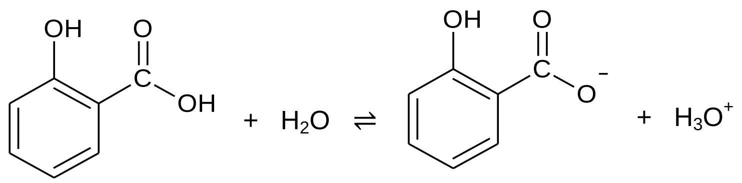

Figure 1 2-hydroxybenzoic acid ionises in aqueous solution. It is a weak acid.

This may be represented:

$$\mathrm{HA + H_2O \leftrightarrow A^- + H_3O^+}$$

or, more simply

$$\mathrm{HA \leftrightarrow A^- + H^+}$$

Its acid dissociation constant, K_a, is given by:

$$K_\mathrm{a} = \frac{[\mathrm{H_3O^+}][\mathrm{A^-}]}{[\mathrm{HA}]}$$

often written simply as:

$$K_\mathrm{a} = \frac{[\mathrm{H^+}][\mathrm{A^-}]}{[\mathrm{HA}]}$$

The pK_a value is given by:

$$\mathrm{p}K_\mathrm{a} = -\log_{10} K_\mathrm{a}$$

Taking logarithms, the following relationships are derived

$$\mathrm{pH} = \mathrm{p}K_\mathrm{a} - \log\frac{[\mathrm{HA}]}{[\mathrm{A^-}]}$$

or

$$\mathrm{pH} = \mathrm{p}K_\mathrm{a} + \log\frac{[\mathrm{A^-}]}{[\mathrm{HA}]}$$

When the acid is 'half-neutralised',

[A$^-$] = [HA], and pH = pK_a.

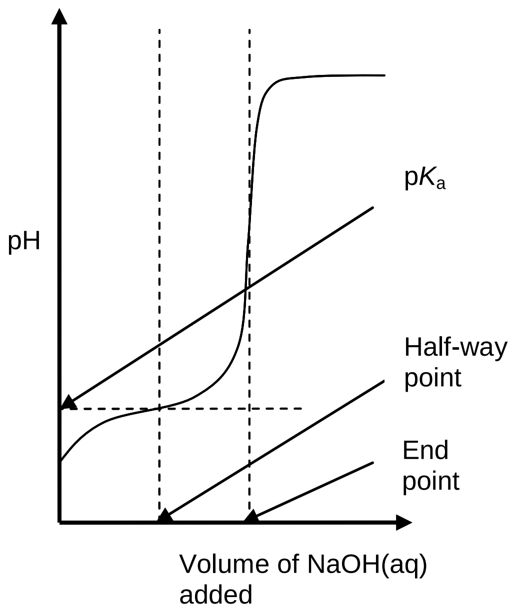

Figure 2 A graph of pH for the titration of a weak acid with sodium hydroxide solution.

By titrating a solution of the 2-hydroxybenzoic acid against a strong alkali, such as sodium hydroxide solution, the pH at the half-way point can be determined and this gives the pK_a of aspirin.

2-hydroxybenzoic acid and sodium hydroxide react in a 1:1 mole ratio:

$$\mathrm{HOCOC_6H_4COOH} + \mathrm{NaOH} \rightarrow \mathrm{HOCOC_6H_4COONa} + \mathrm{H_2O}$$

Equipment and materials

- Balance
- 50 cm^3 burette
- 250 cm^3 beaker
- Glass stirring rod
- 10 cm^3 and 100 cm^3 measuring cylinders
- Spatula
- pH probe and pH meter
- 2-hydroxybenzoic acid − Harmful
- 95% ethanol − Highly flammable, Harmful
- 0.10 mol dm^{-3} sodium hydroxide solution − Irritant

Method

1. Fill a burette with 0.10 mol dm^{-3} sodium hydroxide solution.
2. Weigh 0.28 g of 2-hydroxybenzoic acid into a 250 cm^3 beaker. Add 10 cm^3 of 95% ethanol, Stir with a glass rod and when the solid has dissolved add 90 cm^3 of deionised water. Stir the mixture until it is homogeneous.
3. Place a pH probe in the solution and connect it to a pH meter.

   Note: The pH probe should have been calibrated using suitable buffer solutions.
4. Add 2 cm^3 quantities of sodium hydroxide solution from the burette to the beaker, stirring well between additions and recording the pH.
5. Near the end-point the pH begins to rise rapidly. So after you have added 18 cm^3 of sodium hydroxide solution begin adding it in 0.5 cm^3 portions. After about 22 cm^3 start adding in 2 cm^3 portions again. Continue until total of 36 cm^3 has been added.

Processing data

1. Plot a graph of pH against volume of 0.10 mol dm^{-3} sodium hydroxide solution added.
2. From the graph, calculate the end-point of the titration.
3. Check this against the expected value by calculating the number of moles of 2-hydroxybenzoic acid used (relative molecular mass of 2-hydroxybenzoic acid = 138) and, therefore, the volume of 0.10 mol dm^{-3} sodium hydroxide solution needed to react with it in a 1:1 mole ratio.
4. At the half-way point to the end point, $[HOC_6H_4COO^-] = [H^+]$.

   From the graph, estimate the pH at the half-way point of the titration and, therefore, a value for the pK_a of 2-hydroxybenzoic acid.
5. Calculate K_a of 2-hydroxybenzoic acid.