RSC | Advancing the Chemical Sciences

Reckitt Benckiser

Determining the pK_a of aspirin

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. Aspirin and 95% ethanol are both harmful.

Principle

Aspirin is a weak acid. It partially ionises in water:

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

and 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:

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

Taking logarithms, the following relationships are derived

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

When the acid is 'half-neutralised', [A$^-$] = [HA], and pH = pK_a.
By titrating a solution of the aspirin 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.

Aspirin and sodium hydroxide react in a 1:1 mole ratio:

$$\mathrm{CH_3OCOC_6H_4COOH} + \mathrm{NaOH} \rightarrow \mathrm{CH_3OCOC_6H_4COONa} + \mathrm{H_2O}$$

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
- Aspirin − 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 0.36 g of aspirin 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 aspirin used (relative molecular mass of aspirin = 180) 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. From the graph, estimate the pH at the half-way point of the titration. This gives a value for the pK_a of aspirin.
5. Calculate the K_a of aspirin.

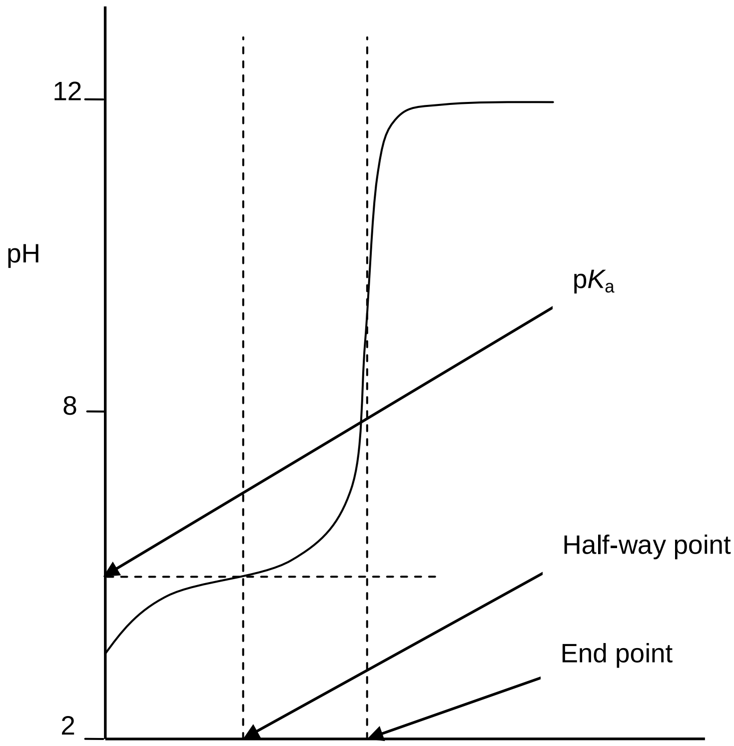

Figure Illustrative titration graph for aspirin against sodium hydroxide solution.