ACID-BASE TITRATIONS

An acid-base titration involves reacting an acid with a base in order to accurately determine the concentration of one of the solutions. The other solution must be a standard solution, meaning that its concentration is known accurately.

STANDARD SOLUTIONS

The standard solution can be either a ‘primary standard’ or a ‘secondary standard’, although a primary standard is preferred because the concentration is known more accurately. A primary standard is prepared by dissolving a known mass in a known volume of water. The concentration of a secondary standard is determined by titrating it against a primary standard.

The amount of a primary standard must be calculated from its mass, and so its molar mass must be known and must not change.

A primary standard should therefore:

- Be readily available in a pure form.
- Have a relatively high molar mass to minimise weighing errors.
- Be stable as both the solid and the solution. It must be stored without reacting with the atmosphere.
- Be completely soluble.
- React according to a known chemical equation so that mole ratios can be used.

METHOD:

The following procedure should be followed in an acid-base titration:

1. A **pipette** is used to accurately measure the volume of the sample. This amount of the sample is then known as an **aliquot**. The pipette should be rinsed with the sample itself so that the concentration of the sample does not change.

2. The sample is put into a **conical flask**. The conical flask must be rinsed with distilled water so that the number of moles of the sample does not change.

3. The standard solution is put into a **burette**. The burette must be rinsed with the standard solution so that the concentration of the standard solution does not change.

4. Two or three drops of indicator solution may be added to the conical flask so that there will be an observable colour change at the end point of the titration. The **equivalence point** of the titration is the point where chemically equivalent amounts of the two solutions, according to the mole ratios shown by the equation, are present. The **end point** of the titration occurs when a permanent colour change is first observed, and the titration is stopped.

5. Several titrations are performed, until there are three concordant results. The average of these volumes is known as the average **titre**.

CALCULATIONS

The following calculations must be completed after the titration:

1. A balanced chemical equation is always required.

2. The amount, in mole, of the standard solution must be calculated. The concentration is known accurately, and the volume is usually read from the burette. These values are substituted into the formula $n(s\tan dard\,solution) = C \times V$.

3. Use mole ratios to compare the amount of the standard solution with the amount of the sample. The balanced chemical equation must be used for this.

4. If necessary, use ratios to allow for dilution. Find the amount, in mole, of the unknown in the dilute solution. And therefore also the amount, in mole, of the unknown in the initial concentrated sample.

5. Calculate the concentration of the initial sample by substituting into the formula $C(sample) = \frac{n}{V}$. Alternatively, calculate the mass of the initial sample by substituting into the formula $mass(sample) = n \times M$.

EXAMPLE 1
A titration was performed to determine the concentration of a sample of sulfuric acid. An average titre of 22.18 mL of 0.116 M potassium hydroxide was required to react with 20.00 mL of the sulfuric acid. What was the concentration of the sulfuric acid?

Solution

1. Write a balanced chemical equation for the reaction between sulfuric acid and potassium hydroxide.

$$H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$$

2. Calculate the amount of potassium hydroxide.

$$\begin{aligned} n(KOH) &= C \times V \\ &= 0.116 \times 0.02218 = 0.02573\, mol \end{aligned}$$

3. Use mole ratios to determine the amount of sulfuric acid.

$$\begin{aligned} n(H_2SO_4) &= \frac{1}{2} \times n(KOH) \\ &= \frac{1}{2} \times 0.02573 = 0.01286 mol \end{aligned}$$

4. Calculate the concentration of the sample.

$$C(H_2SO_4) = \frac{n}{V} = \frac{0.01286}{0.02000} = 0.6432M$$

5. Give the answer with three significant figures (because the least number of significant figures in the question is three).

$$C(H_2SO_4) = 0.643M$$

EXAMPLE 2

A titration was performed to determine the concentration of acetic acid in commercial vinegar. A 20.00 ml sample of the commercial vinegar was put into a volumetric flask and the volume was made up to 250.00 ml with distilled water. A 20.00 ml sample of the dilute vinegar was put into a conical flask, and titrated against 0.110 M sodium hydroxide solution. An average titre of 22.40 ml was obtained. What was the concentration of the commercial vinegar?

Solution

1. Write a balanced chemical equation for the reaction between acetic acid and sodium hydroxide.

$$CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COO.Na(aq) + H_2O(l)$$

2. Calculate the amount of sodium hydroxide.

$$\begin{aligned} n(NaOH) &= C \times V \\ &= 0.110 \times 0.02240 = 2.46 \times 10^{-3}\ mol \end{aligned}$$

3. Use mole ratios to determine the amount of acetic acid in the dilute sample.

$$\begin{aligned} n(CH_3COOH) &= n(NaOH) \\ &= 2.46 \times 10^{-3}\ mole\ in\ 20.00\, ml\ of\ the\ dilute sample \end{aligned}$$

4. Use ratios to find the amount, in mole, of acetic acid in 250.00 ml of the dilute sample. And therefore in 20.00 ml of the commercial vinegar.

$$\begin{array}{lll} 2.46 \times 10^{-3}\ mol & in & 20.00\, ml \\ x\, mol & in & 250.00\, ml \end{array}$$

$$\begin{array}{ll} Cross\ multiply: & 20.00 \times x = 2.46 \times 10^{-3} \times 250.00 \\ & \therefore x = 0.0308\, mol\ \ in\, 250.00\, ml\ of\ the\ dilute\, sample \end{array}$$

Since only distilled water was added to 20.00 ml of the commercial vinegar

$$\begin{aligned} & n(CH_3COOH)\ in\ 20.00\, ml\ of\ concentrated\ vinegar \\ & \quad = n(CH_3COOH)\ in\ 250.00\, ml\ of\ dilute\ vinegar \\ & \quad = 0.0308\, mol \end{aligned}$$

5. Calculate the concentration of the commercial vinegar.

$$C(CH_3COOH) = \frac{n}{V} = \frac{0.0308}{0.02000} = 1.54\, M$$