The manufacture of soap requires a number of chemistry techniques. One necessary piece of information is the saponification number. This is the amount of base needed to hydrolyze a certain amount of fat to produce the free fatty acids that are an essential part of the final product. The fat is heated with a known amount of base (usually \(\text{NaOH}\) or \(\text{KOH}\)). After hydrolysis is complete, the leftover base is titrated to determine how much was needed to hydrolyze the fat sample.

**Titration Calculations**

At the equivalence point in a neutralization, the moles of acid are equal to the moles of base.

\[
\text{moles acid} = \text{moles base}
\]

Recall that the molarity \(\text{M} = \frac{\text{moles solute}}{\text{L}}\) of a solution is defined as the moles of the solute divided by the liters of solution. \(\text{M} = \frac{\text{moles solute}}{\text{L}}\). So the moles of solute are therefore equal to the molarity of a solution multiplied by the volume in liters.

\[
\text{moles solute} = \text{M} \times \text{L}
\]

We can then set the moles of acid equal to the moles of base.

\[
\text{M}_A \times V_A = \text{M}_B \times V_B
\]

\(\text{M}_A\) is the molarity of the acid, while \(\text{M}_B\) is the molarity of the base. \(V_A\) and \(V_B\) are the volumes of the acid and base, respectively.

Suppose that a titration is performed and \(20.70 \: \text{mL}\) of \(0.500 \: \text{M} \: \text{NaOH}\) is required to reach the end point when titrated against \(15.00 \: \text{mL}\) of \(\text{HCl}\) of unknown concentration. The above equation can be used to solve for the molarity of the acid.

\[
\text{M}_A = \frac{\text{M}_B \times V_B}{V_A} = \frac{0.500 \: \text{M} \times 20.70 \: \text{mL}}{15.00 \: \text{mL}} = 0.690 \: \text{M}
\]

The higher molarity of the acid compared to the base in this case means that a smaller volume of the acid is required to reach the equivalence point.

The above equation works only for neutralizations in which there is a 1:1 ratio between the acid and the base. The example below demonstrates the technique to solve a titration problem for a titration of sulfuric acid with sodium hydroxide.

Example

In a titration of sulfuric acid against sodium hydroxide, \(32.20 \: \text{mL}\) of \(0.250 \: \text{M} \: \text{NaOH}\) is required to neutralize \(26.60 \: \text{mL}\) of \(\text{H}_2\text{SO}_4\). Calculate the molarity of the sulfuric acid.

\[
\text{H}_2\text{SO}_4(aq) + 2\text{NaOH}(aq) \rightarrow \text{Na}_2\text{SO}_4(aq) + 2\text{H}_2\text{O}
\]
**Solution**

*Step 1: List the known values and plan the problem.*

- Molarity $\ce{NaOH} = 0.250 \text{ M}$
- Volume $\ce{NaOH} = 32.20 \text{ mL}$
- Volume $\ce{H_2SO_4} = 26.60 \text{ mL}$

**Unknown**

- Molarity $\ce{H_2SO_4} =$ ?

First determine the moles of $\ce{NaOH}$ in the reaction. From the mole ratio, calculate the moles of $\ce{H_2SO_4}$ that reacted. Finally, divide the moles of $\ce{H_2SO_4}$ by its volume to get the molarity.

*Step 2: Solve.*

\[\begin{align} &\text{mol} \ce{NaOH} = \text{M} \times \text{L} = 0.250 \times 0.03220 = 8.05 \times 10^{-3} \text{ mol} \ce{NaOH} \\
&8.05 \times 10^{-3} \text{ mol} \ce{NaOH} \times \frac{1 \text{ mol} \ce{H_2SO_4}}{2 \text{ mol} \ce{NaOH}} = 4.03 \times 10^{-3} \text{ mol} \ce{H_2SO_4} \\
&\frac{4.03 \times 10^{-3} \text{ mol} \ce{H_2SO_4}}{0.02660 \text{ L}} = 0.151 \text{ M} \ce{H_2SO_4} \end{align}\]

*Step 3: Think about your result.*

The volume of $\ce{H_2SO_4}$ required is smaller than the volume of $\ce{NaOH}$ because of the two hydrogen ions contributed by each molecule.

**Summary**

- The process of calculating concentration from titration data is described and illustrated.

**Contributors and Attributions**

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