5.12: Titrations - Measuring Moles by Volume of Solution

Masses are commonly measured with a laboratory balance that registers in grams. Masses of industrial chemicals are measured with much larger industrial scales that commonly give masses in kilograms or tons. In doing laboratory stoichiometric measurements with species in solution, it is often convenient to measure volumes of solution rather than masses of reactants. Solutions can be prepared that contain known numbers of moles per unit volume of solution. The volume of the reagent that must be added to another reagent to undergo a particular reaction can be measured with a device called a **buret**. A buret is shown in Figure 5.3. By measuring the volume of a solution of known concentration of solute required to react with another reactant, the number of moles of solute reacting can be calculated and stoichiometric calculations can be performed based upon the reaction. This procedure is commonly used in chemical analysis and is called **titration**.

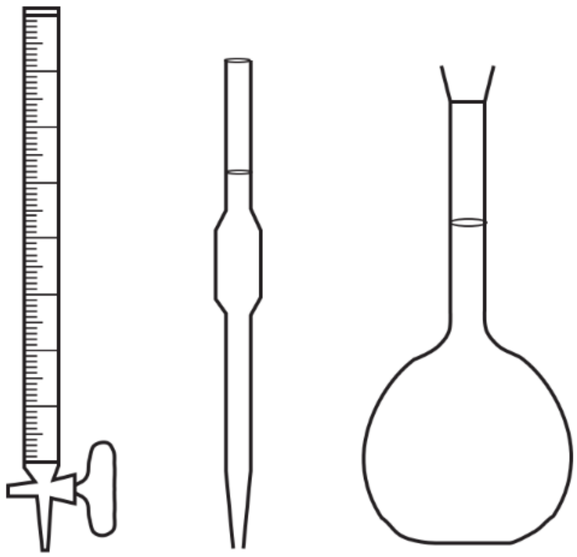

Figure 5.3. A buret consists of a narrow glass tube marked off in divisions of milliliters (mL) further subdivided into tenths of mL, usually with a total capacity of 50 mL. It enables accurate measurements of volumes of solution delivered through a stopcock estimated to the nearest 0.01 mL. A pipet delivers a fixed volume of solution. A volumetric flask contains a fixed volume of solution.

It is especially easy to relate volumes of solutions stoichiometrically when the solution concentrations are expressed as *molar concentration*, *M*. This concentration unit is defined as

$$\mathrm{M}=\frac{\text{moles of solute}}{\text{number of liters of solution}} \tag{5.12.1}$$

The number of moles of a substance, in this case the moles of solute, is related to the mass of the substance by

$$\text{Moles of solute}=\frac{\text{moles of solute}}{\text{molar mass of solute, g/mol}} \tag{5.12.2}$$

These two relationships can be combined to give the following useful equation:

$$\mathrm{M}=\frac{\text{mass of solute}}{(\text{molar mass of solute})\times(\text{number of liters of solution})} \tag{5.12.3}$$

A solution of known concentration that is added to a reaction mixture during the procedure of titration is a **standard solution**. One of the most common of these is a standard base solution of sodium hydroxide, NaOH. Typically, the concentration of sodium hydroxide in such a standard solution is 0.100 mol/L. Suppose that it is desired to make exactly 2 liters of a solution of 0.100mol/L sodium hydroxide. What mass of NaOH, molar mass 40.0 g/mol, is dissolved in this solution? To do this calculation, use Equation 5.12.3 rearranged to solve for mass of solute:

$$\text{Mass NaOH}=\mathrm{M}\times(\text{molar mass NaOH})\times(\text{liters NaOH}) \tag{5.12.4}$$

$$\text{Mass NaOH} = 0.100\ \text{mol/L} \times 40.0\ \text{g/mol} \times 2.00\ \text{L} = 8.00\ \text{g NaOH} \tag{5.12.5}$$

A common titration procedure is to use a standard solution of base to titrate an unknown solution of acid or to use standard acid to determine base. As an example consider an analysis for acid of a sample of water used to scrub exhaust gas from a hospital incinerator. The water is acidic because of the presence of hydrochloric acid produced by the scrubbing of HCl gas from the incinerator stack gas where the HCl was produced in the burning of polyvinyl chloride in the incinerator. Suppose that a sample of 100 mL of the scrubber water was taken for titration with a 0.125 mol/L standard NaOH and that the volume of standard NaOH consumed was 11.7 mL. What was the molar concentration of HCl in the stack gas scrubber water? To solve this problem it is necessary to know that the reaction between NaOH and HCl is,

$$\text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O} \tag{5.12.6}$$

a neutralization reaction in which water and a salt, NaCl are produced. Examination of the reaction shows that 1 mole of HCl reacts for each mole of NaOH. Equation 5.12.1 applies to both the standard NaOH solution and the HCl solution being titrated leading to the following equations:

$$M_{HCl} = \frac{moles_{HCl}}{liters_{HCl}} = \text{and M}_{NaOH} = \frac{moles_{NaOH}}{liters_{NaOH}} \tag{5.12.7}$$

When exactly enough NaOH has been added to react with all the HCl present, the reaction is complete with no excess of either HCl or NaOH. In a titration this **end point** is normally shown by the change of color of a dye called an **indicator** dissolved in the solution being titrated. At the endpoint moles HCl= moles NaOH and the two equations above can be solved to give,

$$\text{M}_{\text{HCl}} \times \text{liters}_{\text{HCl}} = \text{M}_{\text{NaOH}} \times \text{liters}_{\text{NaOH}} \tag{5.12.8}$$

which can be used to give the molar concentration of HCl:

$$\text{M}_{\text{HCl}} = \frac{\text{M}_{\text{NaOH}} \times \text{liters}_{\text{NaOH}}}{\text{liters}_{\text{HCl}}} \tag{5.12.9}$$

Converting the volumes given from mL to liters and substituting into this equation gives the molar concentration of HCl in the incinerator scrubber water:

$$\text{M}_{\text{HCl}} = \frac{0.125\text{mol/L} \times 0.0117\text{L}}{0.100\ \text{L}} = 0.0146\text{mol/L} \tag{5.12.10}$$

Determining Percentage Composition by Titration

A useful application of titration, or **titrimetric analysis** as it is called, is to determine the percentage of a substance in a solid sample that will react with the titrant. To see how this is done, consider a sample consisting of basic lime, $Ca(OH)_2$, and dirt with a mass of 1.26 g. Using titration with a standard acid solution it is possible to determine the mass of basic $Ca(OH)_2$ in the sample and from that calculate the percentage of $Ca(OH)_2$ in the sample. Assume that the solid sample is placed in water and titrated with 0.112 mol/L standard HCl, a volume of 42.2 mL(0.0422 L) of the acid being required to reach the end point. The $Ca(OH)_2$ reacts with the HCl

$$\text{Ca(OH)}_2 + 2\,\text{HCl} \rightarrow \text{CaCl}_2 + 2\,\text{H}_2\text{O} \tag{5.12.11}$$

whereas the dirt does not react. Examination of this reaction shows that at the end point the mole ratio

$$\frac{1\ \text{mol Ca(OH)}_2}{2\ \text{mol HCl}} \tag{5.12.12}$$

applies. At the end point, the number of moles of HCl can be calculated from

$$\text{Mol}_{\text{HCl}} = \text{liters}_{\text{HCl}} \times \text{M}_{\text{HCl}} \tag{5.12.13}$$

and, since the molar mass of $Ca(OH)_2$ is 74.1 (given atomic masses 40.1, 16.0, and 1.0 for Ca, O, and H, respectively), the mass of $Ca(OH)_2$ is given by

$$\text{Mass}_{\text{Ca(OH)}_2} = \text{moles}_{\text{Ca(OH)}_2} \times \text{molar mass}_{\text{Ca(OH)}_2} \tag{5.12.14}$$

With this information it is now possible to calculate the mass of $Ca(OH)_2$:

$$\text{Mass}_{\text{Ca(OH)}_2} = \text{mol}_{\text{Ca(OH)}_2} \times \frac{74.1 \text{ g Ca(OH)}_2}{1 \text{ mol Ca(OH)}_2} \tag{5.12.15}$$

$$\text{Mass}_{\text{Ca(OH)}_2} = \underbrace{\text{Liters}_{\text{HCl}} \times \text{M}_{\text{HCl}}}_{\text{Moles HCl reacting}} \times \underbrace{\frac{1 \text{ mol Ca(OH)}_2}{2 \text{ mol HCl}}}_{\text{Converts from moles HCl to moles Ca(OH)}_2} \times \underbrace{\frac{74.1 \text{ g Ca(OH)}_2}{1 \text{ mol Ca(OH)}_2}}_{\text{Gives mass Ca(OH)}_2 \text{ from moles Ca(OH)}_2} \tag{5.12.16}$$

$$\text{Mass}_{\text{Ca(OH)}_2} = 0.0422\text{L HCl} \times \frac{0.112 \text{ mol HCl}}{1 \text{ L HCl}} \times \frac{1 \text{ mol Ca(OH)}_2}{2 \text{ mol HCl}} \times \frac{74.1 \text{ g Ca(OH)}_2}{1 \text{ mol Ca(OH)}_2} \tag{5.12.17}$$

$$\text{Mass Ca(OH)}_2 = 0.175\ g \tag{5.12.18}$$

$$\text{Percent}_{\text{Ca(OH)}_2} = \frac{\text{mass Ca(OH)}_2}{\text{mass sample}} \times \frac{0.175g}{1.26g} \times 100 = 13.9\% \tag{5.12.19}$$

Exercise

Exercise: A 0.638 g sample consisting of oxalic acid, H_2C_2O4, and sodium oxalate, $Na_2C_2O_4$ was dissolved and titrated with 0.116 mol/L sodium hydroxide, of which 47.6mL (0.0476 L) was required. Each molecule of $H_2C_2O_4$ releases 2 H^+ ions. Calculate the percentage of oxalic acid in the sample.

Answer

38.9%