Dilution with water does not alter the numbers of moles of solute present.

To save time and space in the laboratory, routinely used solutions are often purchased or prepared in concentrated form (called **stock solutions**). Water is then added to achieve the molarity desired for a particular solution. This process is called **dilution.** For example, the common acids are purchased as concentrated solutions and diluted as needed. A typical dilution calculation involves determining how much water must be added to an amount of stock solution to achieve a solution of the desired concentration. The key to doing these calculations is to remember that:

$$\text{Moles of solute after dilution} = \text{moles of solute before dilution}$$

because only water (no solute) is added to accomplish the dilution.

For example, suppose we need to prepare 500. mL of 1.00 M acetic acid ($HC_2H_3O_2$) from a 17.4 M stock solution of acetic acid. What volume of the stock solution is required? The first step is to determine the number of moles of acetic acid in the final solution by multiplying the volume by the molarity (remembering that the volume must be changed to liters):

$$500.\ \cancel{\text{mL solution}} \times \frac{1\ \cancel{\text{L solution}}}{1000\ \cancel{\text{mL solution}}} \times \frac{1.00\ \text{mol } HC_2H_3O_2}{\cancel{\text{L solution}}} = 0.500\ \text{mol } HC_2H_3O_2$$

Thus we need to use a volume of 17.4 M acetic acid that contains 0.500 mol $HC_2H_3O_2$. That is,

$$V \times \frac{17.4\ \text{mol } HC_2H_3O_2}{\text{L solution}} = 0.500\ \text{mol } HC_2H_3O_2$$

Solving for V gives

$$V = \frac{0.500\ \cancel{\text{mol } HC_2H_3O_2}}{\dfrac{17.4\ \cancel{\text{mol } HC_2H_3O_2}}{\text{L solution}}} = 0.0287\ \text{L or } 28.7\ \text{mL solution}$$

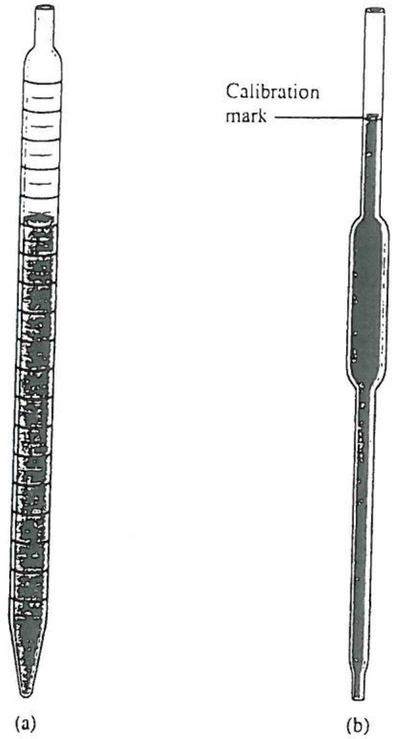

Figure 4.6

(a) A measuring pipet is graduated and can be used to measure various volumes of liquid accurately. (b) A volumetric (transfer) pipet is designed to measure one volume accurately. When filled to the mark, it delivers the volume indicated on the pipet.

Thus, to make 500 mL of a 1.00 M acetic acid solution, we can take 28.7 mL of 17.4 M acetic acid and dilute it to a total volume of 500 mL with distilled water.

A dilution procedure typically involves two types of glassware: a pipet and a volumetric flask. A **pipet** is a device for accurately measuring and transferring a given volume of solution. There are two common types of pipets: *volumetric* (or *transfer*) *pipets* and *measuring pipets*, as shown in Fig. 4.6. Volumetric pipets come in specific sizes, such as 5 mL, 10 mL, 25 mL, and so on. Measuring pipets are used to measure volumes for which a volumetric pipet is not available. For example, we would use a measuring pipet as shown in Fig. 4.7 to deliver 28.7 mL of

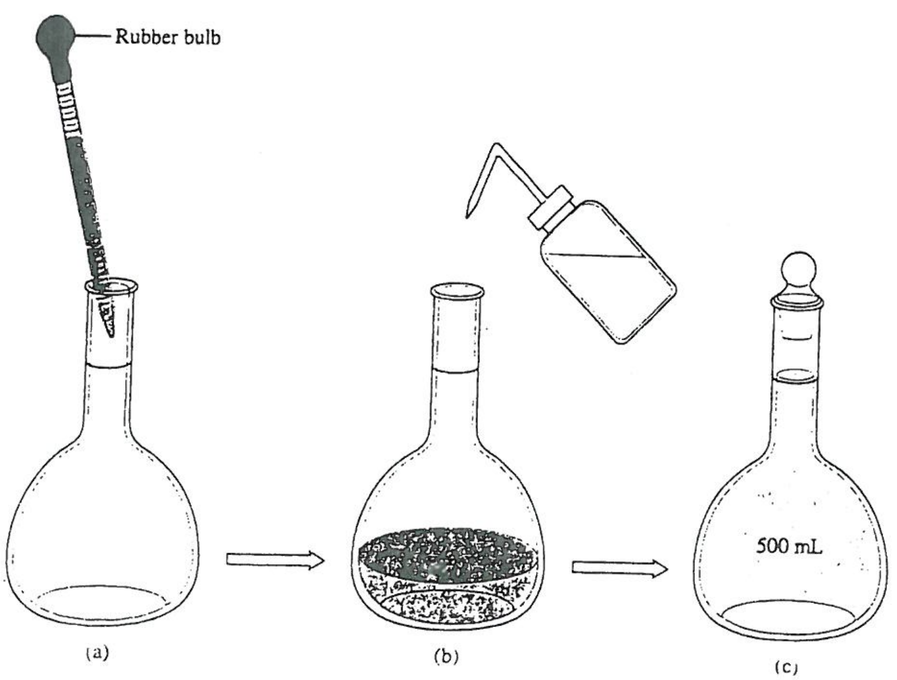

Figure 4.7

(a) A measuring pipet is used to transfer 28.7 mL of 17.4 M acetic acid solution to a volumetric flask. (b) Water is added to the flask to the calibration mark. (c) The resulting solution is 1 M acetic acid.

17.4 M acetic acid into a 500-mL volumetric flask and then add water to the mark to perform the dilution described above.

Sample Exercise 4.9

What volume of 16 M sulfuric acid must be used to prepare 1.5 L of a 0.10 M H_2SO_4 solution?

Solution

We must first determine the moles of H_2SO_4 in 1.5 L of 0.10 M H_2SO_4:

$$1.5 \cancel{\text{ L solution}} \times \frac{0.10 \text{ mol } H_2SO_4}{\cancel{\text{L solution}}} = 0.15 \text{ mol } H_2SO_4$$

Next, we must find the volume of 16 M H_2SO_4 that contains 0.15 mol H_2SO_4:

$$V \times \frac{16 \text{ mol } H_2SO_4}{\text{L solution}} = 0.15 \text{ mol } H_2SO_4$$

Solving for V gives

$$V = \frac{0.15 \cancel{\text{ mol } H_2SO_4}}{\dfrac{16 \cancel{\text{ mol } H_2SO_4}}{1 \text{ L solution}}} = 9.4 \times 10^{-3} \text{ L or } 9.4 \text{ mL solution}$$

Thus, to make 1.5 L of 0.10 M H_2SO_4 using 16 M H_2SO_4, we must take 9.4 mL of the concentrated acid and dilute it with water to 1.5 L. The correct way to do this is to add the 9.4 mL of acid to about 1 L of distilled water and then dilute to 1.5 L by adding more water.

In diluting an acid, "Do what you oughta, always add acid to water."

As noted above, the central idea in performing the calculations associated with dilutions is to recognize that the moles of solute are not changed by the dilution. Another way to express this condition is by the following equation:

$$M_1V_1 = M_2V_2$$

where M_1 and V_1 represent the molarity and volume of the original solution (before dilution) and M_2 and V_2 represent the molarity and volume of the diluted solution. This equation makes sense because

$$\begin{aligned} M_1 \times V_1 &= \text{mol solute before dilution} \\ &= \text{mol solute after dilution} = M_2 \times V_2 \end{aligned}$$

Repeat Sample Exercise 4.9 using the equation $M_1V_1 = M_2V_2$. Note that in doing so

$$M_1 = 16\ M \qquad M_2 = 0.10\ M \qquad V_2 = 1.5 \text{ L}$$

and V_1 is the unknown quantity sought. The equation $M_1V_1 = M_2V_2$ always holds for a dilution. This equation will be easy for you to remember if you understand where it comes from.