Pearson

8.2: Solution Concentration

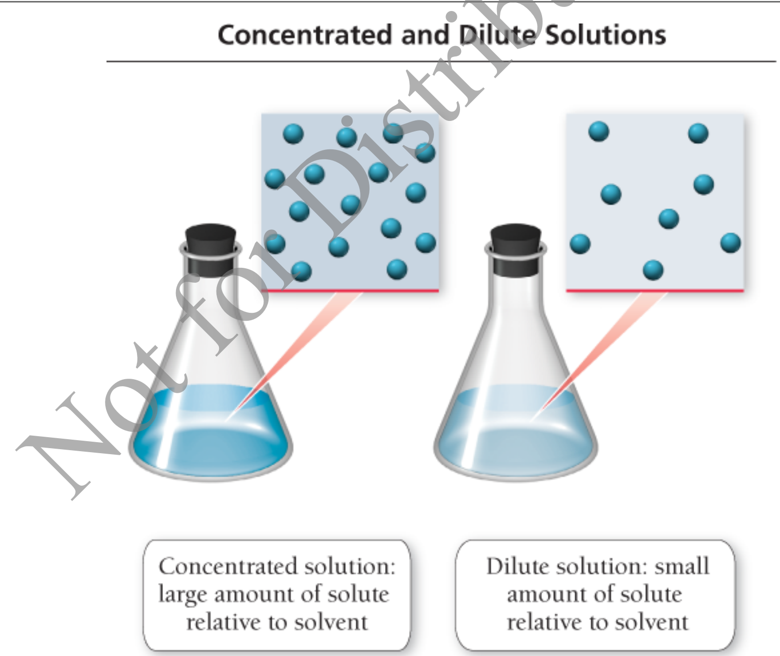
A homogeneous mixture of two substances—such as salt and water—is a **solution**. The major component of the mixture is the **solvent**, and the minor component is the **solute**. An **aqueous solution** is one in which water acts as the solvent. In this section, we examine how to quantify the concentration of a solution (the amount of solute relative to solvent), how to carry out calculations involving the concentration, and how to calculate the effects of diluting a solution.

Quantifying Solution Concentration

The amount of solute in a solution is variable. For example, we can add just a little salt to water to make a **dilute solution**, one that contains a small amount of solute relative to the solvent, or we can add a lot of salt to water to make a **concentrated solution**, one that contains a large amount of solute relative to the solvent (Figure 8.2). A common way to express solution concentration is **molarity (M)**, the amount of solute (in moles) divided by the volume of solution (in liters).

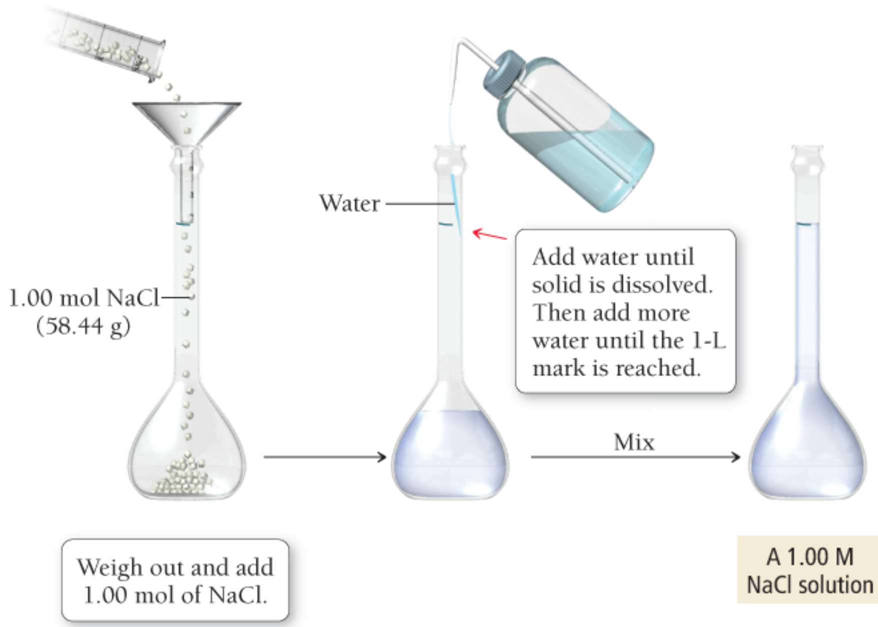
$$\text{molarity}\ \ (\text{M}) = \frac{\text{amount of solute}\ \ (\text{in mol})}{\text{volume of solution}\ \ (\text{in L})}$$

Figure 8.2 **Concentrated and Dilute Solutions**



Notice that molarity is a ratio of the amount of solute per liter of *solution*, not per liter of solvent. To make an aqueous solution of a specified molarity, we usually put the solute into a flask and then add water to reach the desired volume of solution. For example, to make 1 L of a 1 M NaCl solution, we add 1 mol of NaCl to a flask and then add enough water to make 1 L of solution (Figure 8.3).

Figure 8.3 **Preparing a 1.00 M NaCl Solution**



We *do not* combine 1 mol of NaCl with 1 L of water because the resulting solution would have a total volume exceeding 1 L and therefore a molarity of less than 1 M. To calculate molarity, we divide the amount of the solute in moles by the volume of the solution (solute *and* solvent) in liters, as shown in Example 8.1.

Example 8.1 Calculating Solution Concentration

If you dissolve 25.5 g KBr in enough water to make 1.75 L of solution, what is the molarity of the solution?

SORT You are given the mass of KBr and the volume of a solution and asked to find its molarity.

GIVEN: 25.5 g KBr, 1.75 L of solution

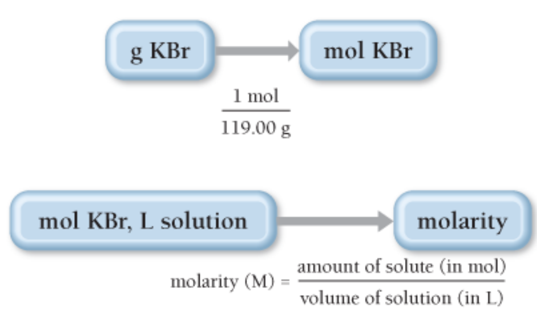
FIND: molarity (M)

STRATEGIZE When formulating the conceptual plan, think about the definition of molarity: the amount of solute *in moles* per *liter of solution*.

You are given the mass of KBr, so first use the molar mass of KBr to convert from g KBr to mol KBr.

Then use the number of moles of KBr and liters of solution to find the molarity.

CONCEPTUAL PLAN



RELATIONSHIPS USED molar mass of KBr = 119.00 g/mol

SOLVE Follow the conceptual plan. Begin with g KBr and convert to mol KBr; then use mol KBr and L solution to calculate molarity.

SOLUTION

$$25.5\ \cancel{\text{g-KBr}} \times \frac{1\ \text{mol KBr}}{119.00\ \cancel{\text{g-KBr}}} = 0.21\underline{4}29\ \text{mol KBr}$$

$$\begin{aligned} \text{molarity (M)} &= \frac{\text{amount of solute (in mol)}}{\text{volume of solution (in L)}} \\ &= \frac{0.21\underline{4}29\ \text{mol KBr}}{1.75\ \text{L solution}} \\ &= 0.122\ \text{M} \end{aligned}$$

CHECK The units of the answer (M) are correct. The magnitude is reasonable since common solutions range in concentration from 0 to about 18 M. Concentrations significantly above 18 M are suspect and should be double-checked.

FOR PRACTICE 8.1 Calculate the molarity of a solution made by adding 45.4 g of $NaNO_3$ to a flask and dissolving it with water to create a total volume of 2.50 L.

FOR MORE PRACTICE 8.1 What mass of KBr (in grams) do you need to make 250.0 mL of a 1.50 M KBr solution?

Interactive Worked Example 8.1 **Calculating Solution Concentration**

Using Molarity in Calculations

We can use the molarity of a solution as a conversion factor between moles of the solute and liters of the solution. For example, a 0.500 M NaCl solution contains 0.500 mol NaCl for every liter of solution.

$$\frac{0.500\ \text{mol NaCl}}{\text{L solution}} \quad \textit{converts} \quad \boxed{\text{L solution}} \longrightarrow \boxed{\text{mol NaCl}}$$

This conversion factor converts from L solution to mol NaCl. If we want to go the other way, we invert the conversion factor.

$$\frac{\text{L solution}}{0.500\ \text{mol NaCl}} \quad \textit{converts} \quad \boxed{\text{mol NaCl}} \longrightarrow \boxed{\text{L solution}}$$

Example 8.2 illustrates how to use molarity in this way.

Example 8.2 Using Molarity in Calculations

How many L of a 0.125 M NaOH solution contain 0.255 mol of NaOH?

SORT You are given the concentration of an NaOH solution. You are asked to find the volume of the solution that contains a given amount (in moles) of NaOH.

GIVEN: 0.125 M NaOH solution, 0.255 mol NaOH

FIND: volume of NaOH solution (in L)

STRATEGIZE The conceptual plan begins with mol NaOH and shows the conversion to L of solution using molarity as a conversion factor.

CONCEPTUAL PLAN

$$\text{mol NaOH} \xrightarrow{\dfrac{1\text{ L solution}}{0.125\text{ mol NaOH}}} \text{L solution}$$

RELATIONSHIPS USED

$$0.125\text{ M NaOH} = \frac{0.125\text{ mol NaOH}}{1\text{ L solution}}$$

SOLVE Follow the conceptual plan. Begin with mol NaOH and convert to L solution.

SOLUTION

$$0.255\ \cancel{\text{mol NaOH}} \times \frac{1\text{ L solution}}{0.125\ \cancel{\text{mol NaOH}}} = 2.04\text{ L solution}$$

CHECK The units of the answer (L) are correct. The magnitude is reasonable because the solution contains 0.125 mol per L. Therefore, roughly 2 L contains the given amount of moles (0.255 mol).

FOR PRACTICE 8.2 How many grams of sucrose ($C_{12}H_{22}O_{11}$) are in 1.55 L of 0.758 M sucrose solution?

FOR MORE PRACTICE 8.2 How many mL of a 0.155 M KCl solution contains 2.55 g KCl?

Interactive Worked Example 8.2 **Using Molarity in Calculations**

Conceptual Connection 8.1 **Solutions**

Interactive

When diluting acids, always add the concentrated acid to the water. Never add water to concentrated acid solutions, as the heat generated may cause the concentrated acid to splatter and burn your skin.

Solution Dilution

To save space in storerooms, laboratories often store solutions in concentrated forms called **stock solutions**. For example, hydrochloric acid is frequently stored as a 12 M stock solution. However, many lab procedures call for much less concentrated hydrochloric acid solutions, so we must dilute the stock solution to the required concentration. How do we know how much of the stock solution to use? When we dilute a solution by adding more solvent, the number of moles of solute do not change; the same number of moles are simply contained in a greater volume, which changes the concentration. Therefore, the most direct way to solve dilution problems is to use the following dilution equation:

[8.1]

$$M_1V_1 = M_2V_2$$

where M_1 and V_1 are the molarity and volume of the initial concentrated solution, and M_2 and V_2 are the molarity and volume of the final diluted solution. This equation works because the molarity multiplied by the volume gives the number of moles of solute, which does not change upon dilution.

$$\begin{aligned} M_1V_1 &= M_2V_2 \\ \text{mol}_1 &= \text{mol}_2 \end{aligned}$$

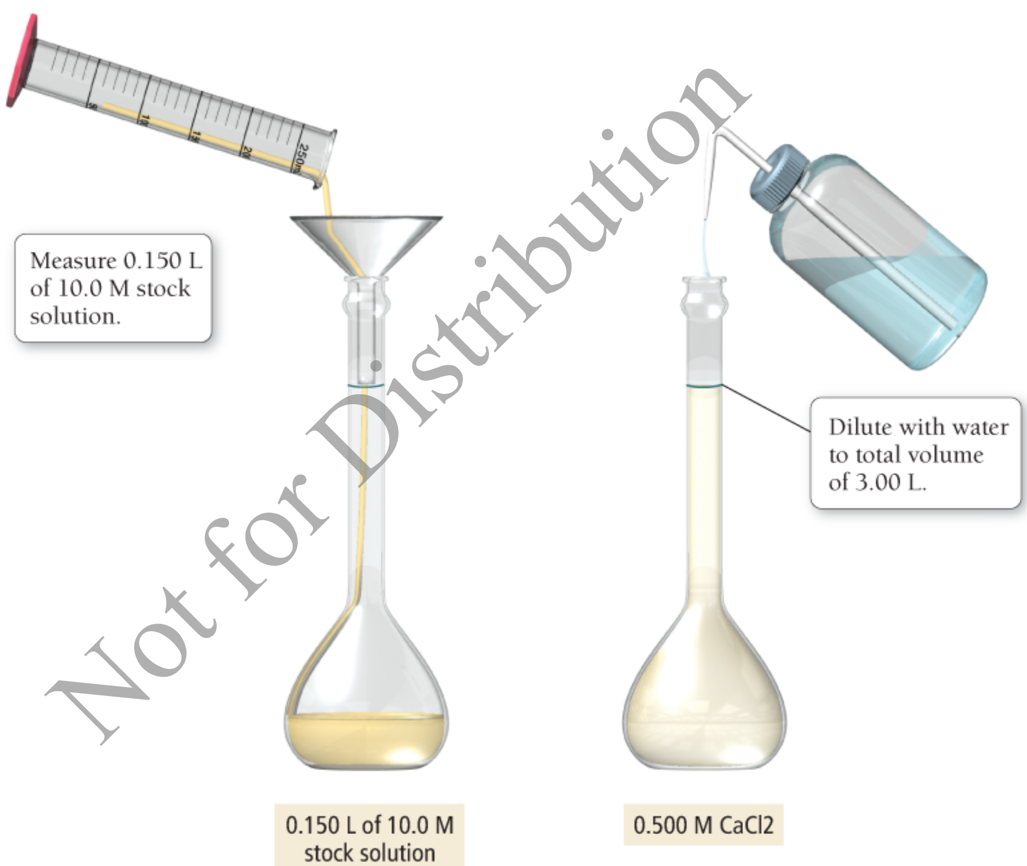
For example, suppose a procedure for spherification (see Section 8.1) calls for 3.00 L of a 0.500 M $CaCl_2$ solution. How can we prepare this solution from a 10.0 M stock solution? We solve Equation 8.1 for V_1, the volume of the stock solution required for the dilution, and then substitute in the correct values to calculate it.

$$\begin{aligned} M_1V_1 &= M_2V_2 \\ V_1 &= \frac{M_2V_2}{M_1} \\ &= \frac{0.500\ \cancel{\text{mol/L}} \times 3.00\ \text{L}}{10.0\ \cancel{\text{mol/L}}} \\ &= 0.150\ \text{L} \end{aligned}$$

We make the solution by diluting 0.150 L of the stock solution to a total volume of 3.00 L (V_2). The resulting solution is 0.500 M $CaCl_2$ (Figure 8.4).

Figure 8.4 Preparing 3.00 L of 0.500 M $CaCl_2$ from a 10.0 M Stock Solution

Diluting a Solution



$$M_1V_1 = M_2V_2$$

$$\frac{10.0\ \text{mol}}{\cancel{\text{L}}} \times 0.150\ \cancel{\text{L}} = \frac{0.500\ \text{mol}}{\cancel{\text{L}}} \times 3.00\ \cancel{\text{L}}$$

$$1.50\ \text{mol} = 1.50\ \text{mol}$$

Example 8.3 Solution Dilution

To what volume should you dilute 0.200 L of a 15.0 M NaOH solution to obtain a 3.00 M NaOH solution?

SORT You are given the initial volume, initial concentration, and final concentration of a solution, and you need to determine the final volume.

GIVEN:

$$V_1 = 0.200 \text{ L}$$
$$M_1 = 15.0 \text{ M}$$
$$M_2 = 3.00 \text{ M}$$

FIND: V_2

STRATEGIZE Equation 8.1 relates the initial and final volumes and concentrations for solution dilution problems. You are asked to find V_2. The other quantities $(V_1, M_1,$ and $M_2)$ are all given in the problem.

CONCEPTUAL PLAN

$$V_1, M_1, M_2 \xrightarrow[M_1V_1 = M_2V_2]{} V_2$$

RELATIONSHIPS USED

$$M_1V_1 = M_2V_2$$

SOLVE Begin with the solution dilution equation and solve it for V_2. Substitute in the required quantities and calculate V_2.

Make the solution by diluting 0.200 L of the stock solution to a total volume of 1.00 L (V_2). The resulting solution has a concentration of 3.00 M.

SOLUTION

$$M_1V_1 = M_2V_2$$
$$V_2 = \frac{M_1V_1}{M_2}$$
$$= \frac{15.0\ \cancel{\text{mol/L}} \times 0.200 \text{ L}}{3.00\ \cancel{\text{mol/L}}}$$
$$= 1.00 \text{ L}$$

CHECK The final units (L) are correct. The magnitude of the answer is reasonable because the solution is diluted from 15.0 M to 3.00 M, a factor of five. Therefore, the volume should increase by a factor of five.

FOR PRACTICE 8.3 To what volume (in mL) should you dilute 100.0 mL of a 5.00 M $CaCl_2$ solution to obtain a 0.750 M $CaCl_2$ solution?

FOR MORE PRACTICE 8.3 What volume of a 6.00 M $NaNO_3$ solution should you use to make 0.525 L of a 1.20 M $NaNO_3$ solution?

Conceptual Connection 8.2 Solution Dilution

Interactive