4.16: Concentrations as Conversion Factors

Learning Objective

- Apply concentration units as conversion factors.

Concentration can be a conversion factor between the amount of solute and the amount of solution or solvent (depending on the definition of the concentration unit). As such, concentrations can be useful in a variety of stoichiometry problems. In many cases, it is best to use the original definition of the concentration unit; it is that definition that provides the conversion factor.

A simple example of using a concentration unit as a conversion factor is one in which we use the definition of the concentration unit and rearrange; we can do the calculation again as a unit conversion, rather than as a definition. For example, suppose we ask how many moles of solute are present in 0.108 L of a 0.887 M NaCl solution. Because 0.887 M means 0.887 mol/L, we can use this second expression for the concentration as a conversion factor:

$$0.108\ L\ \cancel{NaCl} \times \frac{0.887\, mol\, NaCl}{L\ \cancel{NaCl}} = 0.0958\, mol\, NaCl$$

(There is an understood 1 in the denominator of the conversion factor.) If we used the definition approach, we get the same answer, but now we are using conversion factor skills. Like any other conversion factor that relates two different types of units, the reciprocal of the concentration can be also used as a conversion factor.

Example 4.16.1

Using concentration as a conversion factor, how many liters of 2.35 M $CuSO_4$ are needed to obtain 4.88 mol of $CuSO_4$?

Solution

This is a one-step conversion, but the concentration must be written as the reciprocal for the units to work out:

$$4.88\ mol\, \cancel{CuSO_4} \times \frac{1\, L}{2.35\ \cancel{mol}} = 2.08\, L\, of\ solution$$

Exercise 4.16.1

Using concentration as a conversion factor, how many liters of 0.0444 M CH_2O are needed to obtain 0.0773 mol of CH_2O?

Answer:

1.74 L

Of course, once quantities in moles are available, another conversion can give the mass of the substance, using molar mass as a conversion factor.

Example 4.16.2

What mass of solute is present in 0.765 L of 1.93 M NaOH?

Solution

This is a two-step conversion, first using concentration as a conversion factor to determine the number of moles and then the molar mass of NaOH (40.0 g/mol) to convert to mass:

$$0.765\ \cancel{L} \times \frac{1.93\ mol\ \cancel{NaOH}}{L\ \cancel{solution}} \times \frac{40.0 g\, NaOH}{1\ mol\ \cancel{NaOH}} = 59.1\, g\, NaOH$$

Exercise 4.16.2

What mass of solute is present in 1.08 L of 0.0578 M H_2SO_4?

Answer

6.12 g

More complex stoichiometry problems using balanced chemical reactions can also use concentrations as conversion factors. For example, suppose the following equation represents a chemical reaction:

$$2\,\mathrm{AgNO_3(aq)} + \mathrm{CaCl_2(aq)} \rightarrow 2\,\mathrm{AgCl(s)} + \mathrm{Ca(NO_3)_2(aq)}$$

If we wanted to know what volume of 0.555 M $CaCl_2$ would react with 1.25 mol of $AgNO_3$, we first use the balanced chemical equation to determine the number of moles of $CaCl_2$ that would react and then use concentration to convert to liters of solution:

$$1.25\;mol\;\cancel{AgNO_3} \times \frac{1\;mol\;\cancel{CaCl_2}}{2\;mol\;\cancel{AgNO_3}} \times \frac{1\,L\,solution}{0.555\;mol\;\cancel{CaCl_2}} = 1.13\,L\,CaCl_2$$

This can be extended by starting with the mass of one reactant, instead of moles of a reactant.

Example 4.16.3

What volume of 0.0995 M $Al(NO_3)_3$ will react with 3.66 g of Ag according to the following chemical equation?

$$3\,\mathrm{Ag(s)} + \mathrm{Al(NO_3)_3(aq)} \rightarrow 3\,\mathrm{AgNO_3} + \mathrm{Al(s)}$$

Solution

Here, we first must convert the mass of Ag to moles before using the balanced chemical equation and then the definition of molarity as a conversion factor:

$$3.66\;g\;\cancel{Ag} \times \frac{1\;mol\;\cancel{Ag}}{107.97\;g\;\cancel{Ag}} \times \frac{1\;mol\;\cancel{Al(NO_3)_3}}{3\;mol\;\cancel{Ag}} \times \frac{1\,L\,solution}{0.0995\;mol\;\cancel{Al(NO_3)_3}} = 0.114\,L$$

The strikeouts show how the units cancel.

Exercise 4.16.3

What volume of 0.512 M NaOH will react with 17.9 g of $H_2C_2O_4(s)$ according to the following chemical equation?

$$\mathrm{H_2C_2O_4(s)} + 2\,\mathrm{NaOH(aq)} \rightarrow \mathrm{Na_2C_2O_4(aq)} + 2\,\mathrm{H_2O}(\ell)$$

Answer:

0.777 L

We can extend our skills even further by recognizing that we can relate quantities of one solution to quantities of another solution. Knowing the volume and concentration of a solution containing one reactant, we can determine how much of another solution of another reactant will be needed using the balanced chemical equation.

Example 4.16.4

A student takes a precisely measured sample, called an *aliquot*, of 10.00 mL of a solution of $FeCl_3$. The student carefully adds 0.1074 M $Na_2C_2O_4$ until all the Fe^{3+}(aq) has precipitated as $Fe_2(C_2O_4)_3(s)$. Using a precisely measured tube called a burette, the student finds that 9.04 mL of the $Na_2C_2O_4$ solution was added to completely precipitate the Fe^{3+}(aq). What was the concentration of the $FeCl_3$ in the original solution? (A precisely measured experiment like this, which is meant to determine the amount of a substance in a sample, is called a *titration*.) The balanced chemical equation is as follows:

$$2\,FeCl_3(aq) + 3\,Na_2C_2O_4(aq) \rightarrow Fe_2(C_2O_4)_3(s) + 6\,NaCl(aq)$$

Solution

First we need to determine the number of moles of $Na_2C_2O_4$ that reacted. We will convert the volume to liters and then use the concentration of the solution as a conversion factor:

$$9.04\ m\cancel{L} \times \frac{1\ \cancel{L}}{1000\ m\cancel{L}} \times \frac{0.1074 mol\ Na_2C_2O_4}{\cancel{L}} = 0.000971\ mol\ Na_2C_2O_4$$

Now we will use the balanced chemical equation to determine the number of moles of Fe^{3+}(aq) that were present in the initial aliquot:

$$0.000971\ mol\ \cancel{Na_2C_2O_4} \times \frac{2 mol\ FeCl_3}{3\ mol \cancel{Na_2C_2O_4}} = 0.000647 mol\ FeCl_3$$

Then we determine the concentration of $FeCl_3$ in the original solution. Converting 10.00 mL into liters (0.01000 L), we use the definition of molarity directly:

$$M = \frac{mol}{L} = \frac{0.000647 mol\ FeCl_3}{0.01000 L} = 0.0647 M\ FeCl_3$$

Exercise 4.16.4

A student titrates 25.00 mL of H_3PO_4 with 0.0987 M KOH. She uses 54.06 mL to complete the chemical reaction. What is the concentration of H_3PO_4?

$$H_3PO_4(aq) + 3\,KOH(aq) \rightarrow K_3PO_4(aq) + 3\,H_2O$$

Answer:

0.0711 M

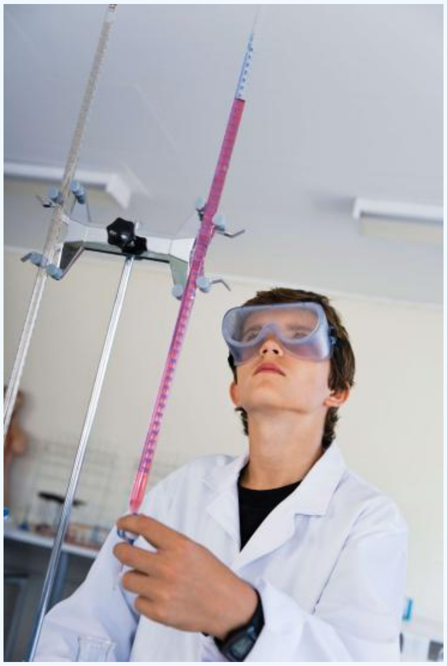

Figure 4.16.1 Titration © Thinkstock. When a student performs a titration, a measured amount of one solution is added to another reactant.

We have used molarity exclusively as the concentration of interest, but that will not always be the case. The next example shows a different concentration unit being used.

Example 4.16.5

H_2O_2 is used to determine the amount of Mn according to this balanced chemical equation:

$$2\,MnO_4^-(aq) + 5\,H_2O_2(aq) + 6\,H^+(aq) \rightarrow 2\,Mn_2^+(aq) + 5\,O_2(g) + 8\,H_2O(\ell)$$

What mass of 3.00% m/m H_2O_2 solution is needed to react with 0.355 mol of MnO_4^-(aq)?

Solution

Because we are given an initial amount in moles, all we need to do is use the balanced chemical equation to determine the number of moles of H_2O_2 and then convert to find the mass of H_2O_2. Knowing that the H_2O_2 solution is 3.00% by mass, we can determine the mass of solution needed:

$$0.355\ mol\ \cancel{MnO_4^-} \times \frac{5\ mol\ \cancel{H_2O_2}}{2\ mol \cancel{MnO_4^-}} \times \frac{34.02\ g\ \cancel{H_2O_2}}{mol\ \cancel{H_2O_2}} \times \frac{100 g\ solution}{3\ g\ \cancel{H_2O_2}} = 1006 g\ sol$$

The first conversion factor comes from the balanced chemical equation, the second conversion factor is the molar mass of H_2O_2, and the third conversion factor comes from the definition of percentage concentration by mass.

Exercise 4.16.5

Use the balanced chemical reaction for MnO_4^- and H_2O_2 to determine what mass of O_2 is produced if 258 g of 3.00% m/m H_2O_2 is reacted with MnO_4^-.

Answer

7.28 g

Summary

Know how to apply concentration units as conversion factors.