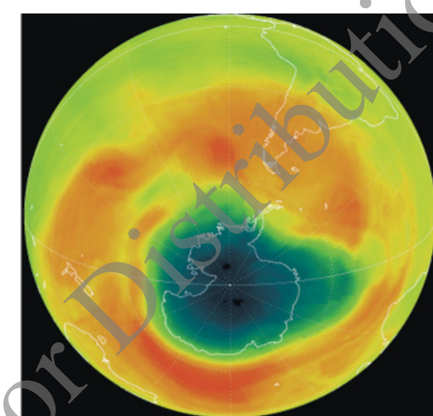
4.10: Composition of Compounds

A chemical formula, in combination with the molar masses of its constituent elements, indicates the relative quantities of each element in a compound, which is extremely useful information. For example, about 35 years ago, scientists began to suspect that synthetic compounds known as chlorofluorocarbons (CFCs) were destroying ozone (O_3) in Earth's upper atmosphere. Upper atmospheric ozone is important because it acts as a shield, protecting life on Earth from the sun's harmful ultraviolet light.

CFCs are chemically inert compounds used primarily as refrigerants and industrial solvents. Over time CFCs have accumulated in the atmosphere. In the upper atmosphere, sunlight breaks bonds within CFCs, releasing chlorine atoms. The chlorine atoms then react with ozone, converting the ozone into O_2. So the harmful part of CFCs is the chlorine atoms that they carry. How can we determine the mass of chlorine in a given mass of a CFC?



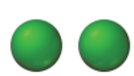
The chlorine in chlorofluorocarbons caused the ozone hole over Antarctica. The dark blue color indicates depressed ozone levels.

One way to express how much of an element is in a given compound is to use the element's mass percent composition for that compound. The **mass percent composition** or **mass percent** of an element is that element's percentage of the compound's total mass. We calculate the mass percent of element X in a compound from the chemical formula as follows:

$$\text{mass percent of element X} = \frac{\text{mass of element X in 1 mol of compound}}{\text{mass of 1 mol of the compound}} \times 100\%$$

Suppose, for example, that we want to calculate the mass percent composition of Cl in the chlorofluorocarbon CCl_2F_2. The mass percent Cl is given by:

CCl_2F_2



$$\text{mass percent Cl} = \frac{2 \times \text{molar mass Cl}}{\text{molar mass } CCl_2F_2} \times 100\%$$

We multiply the molar mass of Cl by two because the chemical formula has a subscript of 2 for Cl, indicating that 1 mol of CCl_2F_2 contains 2 mol of Cl atoms. We calculate the molar mass of CCl_2F_2 as follows:

$$\begin{aligned}\text{molar mass} &= 12.01\ \text{g/mol} + 2(35.45\ \text{g/mol}) + 2\,(19.00\ \text{g/mol})\\ &= 120.91\ \text{g/mol}\end{aligned}$$

So the mass percent of Cl in CCl_2F_2 is:

$$\begin{aligned}\text{mass percent Cl} &= \frac{2 \times \text{molar mass Cl}}{\text{molar mass CCl}_2\text{F}_2} \times 100\%\\ &= \frac{2 \times 35.45\ \text{g}/\cancel{\text{mol}}}{120.91\ \text{g}/\cancel{\text{mol}}} \times 100\%\\ &= 58.64\%\end{aligned}$$

Example 4.11 Mass Percent Composition

Calculate the mass percent of Cl in Freon-112 $(C_2Cl_4F_2)$ a CFC refrigerant.

SORT You are given the molecular formula of freon-112 and asked to find the mass percent of Cl.

GIVEN: $C_2Cl_4F_2$

FIND: mass percent Cl

STRATEGIZE The molecular formula tells you that there are 4 mol of Cl in each mole of Freon-112. Find the mass percent composition from the chemical formula by using the equation that defines mass percent. The conceptual plan shows how to use the mass of Cl in 1 mol of $C_2Cl_4F_2$ and the molar mass of $C_2Cl_4F_2$ to find the mass percent of Cl.

CONCEPTUAL PLAN

$$\text{mass \% Cl} = \frac{4 \times \text{molar mass Cl}}{\text{molar mass C}_2\text{Cl}_4\text{F}_2} \times 100\%$$

RELATIONSHIPS USED

$$\text{mass percent of element X} = \frac{\text{mass of element X in 1 mol of compound}}{\text{mass of 1 mol of compound}} \times 100\%$$

SOLVE Calculate the necessary parts of the equation and substitute the values into the equation to find mass percent Cl.

SOLUTION

$$\begin{aligned}4 \times \text{molar mass Cl} &= 4(35.45\ \text{g/mol}) = 141.8\ \text{g/mol}\\ \text{molar mass C}_2\text{Cl}_4\text{F}_2 &= 2(12.01\ \text{g/mol}) + 4(35.45\ \text{g/mol}) + 2\,(19.00\ \text{g/mol})\\ &= 24.02\ \text{g/mol} + 141.8\ \text{g/mol} + 38.00\ \text{g/mol} = 203.8\ \text{g/mol}\end{aligned}$$

$$\text{mass \% Cl} = \frac{4 \times \text{molar mass Cl}}{\text{molar mass C}_2\text{Cl}_4\text{F}_2} \times 100\% = \frac{141.8\ \text{g}/\cancel{\text{mol}}}{203.8\ \text{g}/\cancel{\text{mol}}} \times 100\% = 69.58\%$$

CHECK The units of the answer (%) are correct. The magnitude is reasonable because (1) it is between 0 and 100% and (2) chlorine is the heaviest atom in the molecule and there are four of them.

FOR PRACTICE 4.11 Acetic acid ($C_2H_4O_2$)is the active ingredient in vinegar. Calculate the mass percent composition of oxygen in acetic acid.

FOR MORE PRACTICE 4.11 Calculate the mass percent composition of sodium in sodium oxide.

Conceptual Connection 4.9 **Chemical Formula and Mass Percent Composition**

Mass Percent Composition as a Conversion Factor

The mass percent composition of an element in a compound is a conversion factor between the mass of the element and the mass of the compound. For example, we saw that the mass percent composition of Cl in CCl_2F_2 is 58.64%. Since percent means *per hundred*, there are 58.64 grams Cl *per hundred* grams CCl_2F_2, which we express as the ratio:

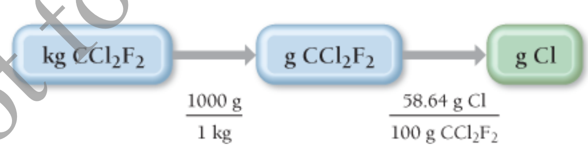
$$58.64 \text{ g Cl} : 100 \text{ g } CCl_2F_2$$

or, in fractional form:

$$\frac{58.64 \text{ g Cl}}{100 \text{ g } CCl_2F_2} \text{ or } \frac{100 \text{ g } CCl_2F_2}{58.64 \text{ g Cl}}$$

These ratios function as conversion factors between grams of Cl and grams of CCl_2F_2. For example, to calculate the mass of Cl in 1.00 kg CCl_2F_2, we use the following conceptual plan:

Conceptual Plan



Notice that the mass percent composition acts as a conversion factor between grams of the compound and grams of the constituent element. To calculate grams Cl, we follow the conceptual plan.

Solution

$$1.00 \cancel{\text{kg}}\ \cancel{CCl_2F_2} \times \frac{1000\ \cancel{\text{g}}}{1\ \cancel{\text{kg}}} \times \frac{58.64 \text{ g Cl}}{100\ \cancel{\text{g}}\ \cancel{CCl_2F_2}} = 5.86 \times 10^2 \text{ g Cl}$$

Example 4.12 Using Mass Percent Composition as a Conversion Factor

The U.S. Food and Drug Administration (FDA) recommends that a person consume less than 2.4 g of sodium per day. What mass of sodium chloride (in grams) can you consume and still be within the FDA guidelines? Sodium chloride is 39% sodium by mass.

SORT You are given a mass of sodium and the mass percent of sodium in sodium chloride. You are asked to find the mass of NaCl that contains the given mass of sodium.

GIVEN: 2.4 g Na

FIND: g NaCl

STRATEGIZE Convert between mass of a constituent element and mass of a compound by using mass percent composition as a conversion factor.

CONCEPTUAL PLAN

$$\text{g Na} \xrightarrow{\frac{100\ \text{g NaCl}}{39\ \text{g Na}}} \text{g NaCl}$$

RELATIONSHIPS USED

39 g Na: 100 g NaCl

SOLVE Follow the conceptual plan to solve the problem.

SOLUTION

$$2.4\ \cancel{\text{g Na}} \times \frac{100\ \text{g NaCl}}{39\ \cancel{\text{g Na}}} = 6.2\ \text{g NaCl}$$

You can consume 6.2 g NaCl and still be within the FDA guidelines.

CHECK The units of the answer are correct. The magnitude seems reasonable because it is larger than the given amount of sodium, as expected, because sodium is only one of the elements in NaCl.

12.5 packets of salt contain 6.2 g of NaCl.

FOR PRACTICE 4.12 What mass (in grams) of iron(III) oxide contains 58.7 g of iron? Iron(III) oxide is 69.94% iron by mass.

FOR MORE PRACTICE 4.12 If someone consumes 22 g of sodium chloride per day, what mass (in grams) of sodium does that person consume? Sodium chloride is 39% sodium by mass.

Conversion Factors from Chemical Formulas

Conversion Factors from Chemical Formulas

Mass percent composition is one way to understand how much chlorine is in a particular chlorofluorocarbon or, more generally, how much of a constituent element is present in a given mass of any compound. However, we can also approach this type of problem in a different way. Chemical formulas contain within them inherent relationships between atoms (or moles of atoms) and molecules (or moles of molecules). For example, the formula for CCl_2F_2 tells us that 1 mol of CCl_2F_2 contains 2 mol of Cl atoms. We write the ratio as:

$$1 \text{ mol } CCl_2F_2 : 2 \text{ mol Cl}$$

With ratios such as these—which come from the chemical formula—we can directly determine the amounts of the constituent elements present in a given amount of a compound without having to calculate mass percent composition. For example, we calculate the number of moles of Cl in 38.5 mol of CCl_2F_2 as follows:

Conceptual Plan

$$\boxed{\text{mol } CCl_2F_2} \xrightarrow[\frac{2 \text{ mol Cl}}{1 \text{ mol } CCl_2F_2}]{} \boxed{\text{mol Cl}}$$

Solution

$$38.5 \ \cancel{\text{mol } CCl_2F_2} \times \frac{2 \text{ mol Cl}}{1 \ \cancel{\text{mol } CCl_2F_2}} = 77.0 \text{ mol Cl}$$

As we have seen, however, we often want to know not the *amount in moles* of an element in a certain number of moles of compound, but the *mass in grams* (or other units) of a constituent element in a given *mass* of the compound. Suppose we want to know the mass (in grams) of Cl contained in 25.0 grams CCl_2F_2. The relationship inherent in the chemical formula (2 mol Cl : 1 mol CCl_2F_2) applies to the amount in moles, not to mass. Therefore, we first convert the mass of CCl_2F_2 to moles CCl_2F_2. *Then* we use the conversion factor from the chemical formula to convert to moles Cl. Finally, we use the molar mass of Cl to convert to grams Cl.

Conceptual Plan

$$\boxed{\text{g } CCl_2F_2} \xrightarrow[\frac{1 \text{ mol } CCl_2F_2}{120.91 \text{ g } CCl_2F_2}]{} \boxed{\text{mol } CCl_2F_2} \xrightarrow[\frac{2 \text{ mol Cl}}{1 \text{ mol } CCl_2F_2}]{} \boxed{\text{mol Cl}} \xrightarrow[\frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}}]{} \boxed{\text{g Cl}}$$

Solution

$$25.0 \ \cancel{\text{g } CCl_2F_2} \times \frac{1 \ \cancel{\text{mol } CCl_2F_2}}{120.91 \ \cancel{\text{g } CCl_2F_2}} \times \frac{2 \ \cancel{\text{mol Cl}}}{1 \ \cancel{\text{mol } CCl_2F_2}} \times \frac{35.45 \text{ g Cl}}{1 \ \cancel{\text{mol Cl}}} = 14.7 \text{ g Cl}$$

Notice that we must convert from g CCl_2F_2 to mol CCl_2F_2 *before* we can use the chemical formula as a conversion factor. *Always remember that the chemical formula gives us a relationship between the amounts (in moles) of substances, not between the masses (in grams) of them.*

The general form for solving problems where we are asked to find the mass of an element present in a given mass of a compound is:

$$\text{mass compound} \rightarrow \text{moles compound} \rightarrow \text{moles element} \rightarrow \text{mass element}$$

We use the atomic or molar mass to convert between mass and moles, and we use relationships inherent in the chemical to convert between moles and moles.

Example 4.13 Chemical Formulas as Conversion Factors

Hydrogen may replace gasoline as a fuel in the future. Most major automobile companies are developing vehicles that run on hydrogen. These cars are environmentally friendly because their only emission is water vapor. One way to obtain hydrogen for fuel is to use an emission-free energy source such as wind power to form elemental hydrogen from water. What mass of hydrogen (in grams) is contained in 1.00 gallon of water? (The density of water is 1.00 g/mL.)

SORT You are given a volume of water and asked to find the mass of hydrogen it contains. You are also given the density of water.

GIVEN: 1.00 gal H_2O

$d_{H_2O} = 1.00\ \text{g/mL}$

FIND: g H

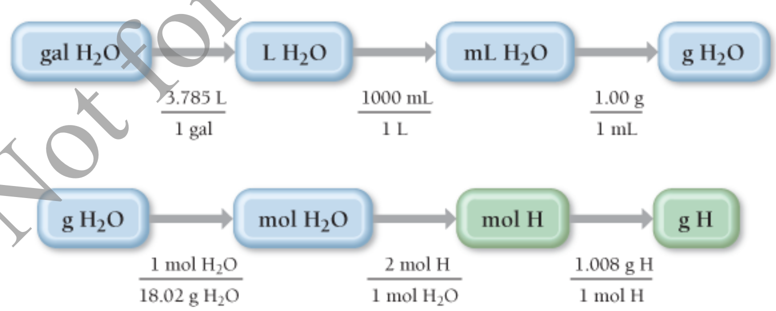
STRATEGIZE The first part of the conceptual plan shows how to convert the units of volume from gallons to liters and then to milliliters. It also shows how you can then use the density to convert mL to g.

The second part of the conceptual plan is the basic sequence of

$$\text{mass} \rightarrow \text{moles} \rightarrow \text{moles} \rightarrow \text{mass.}$$

Convert between moles and mass using the appropriate molar masses, and convert from mol H_2O to mol H using the conversion factor derived from the molecular formula.

CONCEPTUAL PLAN



RELATIONSHIPS USED

$$3.785\ \text{L} = 1\ \text{gal}$$
$$1\ \text{mL} = 10^{-3}\ \text{L}$$
$$1.00\ \text{g}\ H_2O = 1\ \text{mL}\ H_2O\ \ (\text{density of } H_2O)$$
$$\text{molar mass } H_2O = 2(1.008) + 16.00 = 18.02\ \text{g/mol}$$
$$2\ \text{mol H}: 1\ \text{mol}\ H_2O$$
$$1.008\ \text{g H} = 1\ \text{mol H}$$

SOLVE Follow the conceptual plan to solve the problem.

SOLUTION

$$1.00 \ \cancel{\text{gal } H_2O} \times \frac{3.785 \ \cancel{\text{L}}}{1 \ \cancel{\text{gal}}} \times \frac{1 \ \cancel{\text{mL}}}{10^{-3} \ \cancel{\text{L}}} \times \frac{1.0 \text{ g}}{\cancel{\text{mL}}} = 3.7\underline{8}5 \times 10^3 \text{ g } H_2O$$

$$3.7\underline{8}5 \times 10^3 \ \cancel{\text{g } H_2O} \times \frac{1 \ \cancel{\text{mol } H_2O}}{18.02 \ \cancel{\text{g } H_2O}} \times \frac{2 \ \cancel{\text{mol H}}}{1 \ \cancel{\text{mol } H_2O}} \times \frac{1.008 \text{ g H}}{1 \ \cancel{\text{mol H}}} = 4.23 \times 10^2 \text{ g H}$$

CHECK The units of the answer (g H) are correct. Since a gal of water is about 3.8 L, its mass is about 3.8 kg. H is a light atom, so its mass should be significantly less than 3.8 kg, as it is in the answer.

FOR PRACTICE 4.13 Determine the mass of oxygen in a 7.2-g sample of $Al_2(SO_4)_3$

FOR MORE PRACTICE 4.13 Butane (C_4H_{10}) is the liquid fuel in lighters. How many grams of carbon are present within a lighter containing 7.25 mL of butane? (The density of liquid butane is 0.601 g/mL.)

Interactive Worked Example 4.13 **Chemical Formulas as Conversion Factors**

Conceptual Connection 4.10 **Chemical Formulas and Elemental Composition**

Interactive