Pearson

7.4: Reaction Stoichiometry: How Much Carbon Dioxide?

Key Concept Video **Reaction Stoichiometry**

Now that we have examined how to write balanced chemical equations, we can return to the question posed in Section 7.1: How much carbon dioxide is produced from the world's combustion of fossil fuels (and how does that compare to the amount produced by volcanoes)? The balanced chemical equations for fossil fuel combustion reactions provide the relationship between the amount of fossil fuel burned and the amount of carbon dioxide emitted. In this discussion, we use octane (a component of gasoline) as a representative fossil fuel. The balanced equation for the combustion of octane is:

$$2\ C_8H_{18}(l) + 25\ O_2(g) \rightarrow 16\ CO_2(g) + 18\ H_2O\ (g)$$

The balanced equation shows that 16 CO_2 molecules are produced for every 2 molecules of octane burned. We can extend this numerical relationship between molecules to the amounts in moles as follows:

> The coefficients in a chemical equation specify the relative amounts in moles of each of the substances involved in the reaction.

In other words, from the equation, we know that 16 *moles* of CO_2 are produced for every 2 *moles* of octane burned. The numerical relationships between chemical amounts in a balanced chemical equation are called reaction **stoichiometry**. Stoichiometry allows us to predict the amounts of products that will form in a chemical reaction based on the amounts of reactants that react. Stoichiometry also allows us to determine the amount of reactants necessary to form a given amount of product. These calculations are central to chemistry, allowing chemists to plan and carry out chemical reactions to obtain products in the desired quantities.

> Stoichiometry is pronounced stoy-kee-AHM-e-tree.

Making Pizza: The Relationships among Ingredients

The concepts of stoichiometry are similar to those in a cooking recipe. Calculating the amount of carbon dioxide produced by the combustion of a given amount of a fossil fuel is analogous to calculating the number of pizzas that can be made from a given amount of cheese. For example, suppose we use the following pizza recipe:

$$1\ \text{crust} + 5\ \text{ounces tomato sauce} + 2\ \text{cups cheese} \rightarrow 1\ \text{pizza}$$

The recipe contains the numerical relationships between the pizza ingredients. It says that if we have 2

The recipe contains the numerical relationships between the pizza ingredients. It says that if we have 2 cups of cheese—and enough of everything else—we can make 1 pizza. We can write this relationship as a ratio between the cheese and the pizza:

$$2 \text{ cups cheese} : 1 \text{ pizza}$$

What if we have 6 cups of cheese? Assuming that we have enough of everything else, we can use this ratio as a conversion factor to calculate the number of pizzas:

$$6 \; \cancel{\text{cups cheese}} \times \frac{1 \text{ pizza}}{2 \; \cancel{\text{cups cheese}}} = 3 \text{ pizzas}$$

Six cups of cheese are sufficient to make 3 pizzas. The pizza recipe contains numerical ratios between other ingredients as well, including the following:

$$1 \text{ crust} : 1 \text{ pizza}$$
$$5 \text{ ounces tomato sauce} : 1 \text{ pizza}$$

Making Molecules: Mole-to-Mole Conversions

In a balanced chemical equation, we have a "recipe" for how reactants combine to form products. From our balanced equation for the combustion of octane, for example, we can write the following stoichiometric ratio:

$2 \text{ mol } C_8H_{18} : 16 \text{ mol } CO_2$

We can use this ratio to determine how many moles of CO_2 form when a given number of moles of C_8H_{18} burns. Suppose that we burn 22.0 moles of C_8H_{18}; how many moles of CO_2 form? We use the ratio from the balanced chemical equation in the same way that we used the ratio from the pizza recipe. The ratio acts as a conversion factor between the amount in moles of the reactant (C_8H_{18}) and the amount in moles of the product (CO_2).

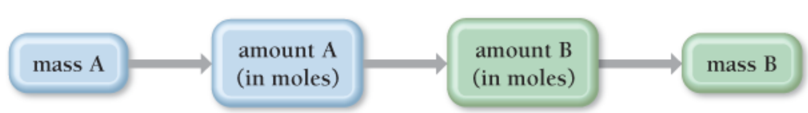
$$22.0 \; \cancel{\text{mol } C_8H_{18}} \times \frac{16 \text{ mol } CO_2}{2 \; \cancel{\text{mol } C_8H_{18}}} = 176 \text{ mol } CO_2$$

The combustion of 22.0 moles of C_8H_{18} adds 176 moles of CO_2 to the atmosphere.

Making Molecules: Mass-to-Mass Conversions

According to the U.S. Department of Energy, the world burned 3.3×10^{10} barrels of petroleum in 2013, the equivalent of approximately 3.7×10^{15} g of gasoline. We can estimate the mass of CO_2 emitted into the atmosphere from burning this much gasoline using the combustion of 3.7×10^{15} g octane as the representative reaction. This calculation is similar to the one we just completed, except that we are now given the *mass* of octane instead of the *amount* of octane in moles. Consequently, we must first convert the mass (in grams) to the amount (in moles).

The general conceptual plan for calculations in which we are given the mass of a reactant or product in a chemical reaction and asked to find the mass of a different reactant or product takes this form:



where A and B are two different substances involved in the reaction. We use the molar mass of A to convert from the mass of A to the amount of A (in moles). We use the appropriate ratio from the balanced chemical equation to convert from the amount of A (in moles) to the amount of B (in moles). And finally, we use the molar mass of B to convert from the amount of B (in moles) to the mass of B.

To calculate the mass of CO_2 emitted upon the combustion of 3.7×10^{15} gof octane, we use the following conceptual plan:

Conceptual Plan

$$\text{g } C_8H_{18} \xrightarrow[\text{114.22 g } C_8H_{18}]{\text{1 mol } C_8H_{18}} \text{mol } C_8H_{18} \xrightarrow[\text{2 mol } C_8H_{18}]{\text{16 mol } CO_2} \text{mol } CO_2 \xrightarrow[\text{1 mol } CO_2]{\text{44.01 g } CO_2} \text{g } CO_2$$

Relationships Used

molar mass $C_8H_{18} = 114.22$ g/mol
2 mol C_8H_{18} : 16 mol CO_2 (from chemical equation)
molar mass $CO_2 = 44.01$ g/mol

Solution

We follow the conceptual plan to solve the problem, beginning with grams C_8H_{18} and canceling units to arrive at grams CO_2.

$$3.7 \times 10^{15} \text{ g } C_8H_{18} \times \frac{1 \cancel{\text{ mol } C_8H_{18}}}{114.22 \cancel{\text{ g } C_8H_{18}}} \times \frac{16 \cancel{\text{ mol } CO_2}}{2 \cancel{\text{ mol } C_8H_{18}}} \times \frac{44.01 \text{ g } CO_2}{1 \cancel{\text{ mol } CO_2}} = 1.1 \times 10^{16} \text{ g } CO_2$$

The world's petroleum combustion produces 1.1×10^{16} g CO_2 $(1.1 \times 10^{13}$ kg$)$ per year. In comparison, volcanoes produce about 2×10^{11} kg CO_2 per year. In other words, volcanoes emit only $\frac{2.0 \times 10^{11} \text{ kg}}{1.1 \times 10^{13} \text{ kg}} \times 100\% = 1.8\%$ as much CO_2 per year as petroleum combustion.* The argument that volcanoes emit more carbon dioxide than fossil fuel combustion is clearly erroneous. Examples 7.4 and 7.5 provide additional practice with stoichiometric calculations.

> The percentage of CO_2 emitted by volcanoes relative to all fossil fuels is even less than 1.8% because the combustion of coal and natural gas also emits CO_2.

Example 7.4 Stoichiometry

In photosynthesis, plants convert carbon dioxide and water into glucose $(C_6H_{12}O_6)$according to the reaction:

$$6\ CO_2(g) + 6\ H_2O(l) \xrightarrow{\text{sunlight}} 6\ O_2(g) + C_6H_{12}O_6(aq)$$

Suppose you determine that a particular plant consumes 37.8 g of CO_2 in one week. Assuming that there is more than enough water present to react with all of the CO_2, what mass of glucose (in grams) can the plant synthesize from the CO_2?

SORT The problem gives the mass of carbon dioxide and asks you to find the mass of glucose that the plant can produce.

GIVEN: 37.8 g CO_2

FIND: g $C_6H_{12}O_6$

STRATEGIZE The conceptual plan follows the general pattern of mass A → amount A (in moles) → amount B (in moles) → mass B. From the chemical equation, deduce the relationship between moles of carbon dioxide and moles of glucose. Use the molar masses to convert between grams and moles.

CONCEPTUAL PLAN

$$\text{g } CO_2 \xrightarrow{\frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2}} \text{mol } CO_2 \xrightarrow{\frac{1 \text{ mol } C_6H_{12}O_6}{6 \text{ mol } CO_2}} \text{mol } C_6H_{12}O_6 \xrightarrow{\frac{180.2 \text{ g } C_6H_{12}O_6}{1 \text{ mol } C_6H_{12}O_6}} \text{g } C_6H_{12}O_6$$

RELATIONSHIPS USED

molar mass $CO_2 = 44.01$ g/mol
6 mol CO_2 : 1 mol $C_6H_{12}O_6$ (from chemical equation)
molar mass $C_6H_{12}O_6 = 180.2$ g/mol

SOLVE Follow the conceptual plan to solve the problem. Begin with g CO_2 and use the conversion factors to arrive at g $C_6H_{12}O_6$

SOLUTION

$$37.8 \; \cancel{\text{g } CO_2} \times \frac{1 \; \cancel{\text{mol } CO_2}}{44.01 \; \cancel{\text{g } CO_2}} \times \frac{1 \; \cancel{\text{mol } C_6H_{12}O_6}}{6 \; \cancel{\text{mol } CO_2}} \times \frac{180.2 \text{ g } C_6H_{12}O_6}{1 \; \cancel{\text{mol } C_6H_{12}O_6}}$$
$$= 25.8 \text{ g } C_6H_{12}O_6$$

CHECK The units of the answer are correct. The magnitude of the answer (25.8 g) is less than the initial mass of CO_2 (37.8 g). This is reasonable because each carbon in CO_2 has two oxygen atoms associated with it, while in $C_6H_{12}O_6$ each carbon has only one oxygen atom associated with it and two hydrogen atoms, which are much lighter than oxygen. Therefore, the mass of glucose the plant produces should be less than the mass of carbon dioxide for this reaction.

FOR PRACTICE 7.4 Magnesium hydroxide, the active ingredient in milk of magnesia, neutralizes stomach acid, primarily HCl, according to the reaction:

$$Mg(OH)_2(aq) + 2\,HCl(aq) \rightarrow 2\,H_2O(l) + MgCl_2\,(aq)$$

What mass of HCl, in grams, is neutralized by a dose of milk of magnesia containing 3.26 g $Mg(OH)_2$?

Interactive Worked Example 7.4 **Stoichiometry**

Example 7.5 Stoichiometry

Sulfuric acid (H_2SO_4)is a component of acid rain that forms when SO_2, a pollutant, reacts with oxygen and water according to the simplified reaction:

$$2\,SO_2(g) + O_2(g) + 2\,H_2O\,(l) \rightarrow 2\,H_2SO_4\,(aq)$$

The generation of the electricity used by a medium-sized home produces about 25 kg of SO_2 per year. Assuming that there is more than enough O_2 and H_2O, what mass of H_2SO_4, in kilograms, can form from this much SO_2?

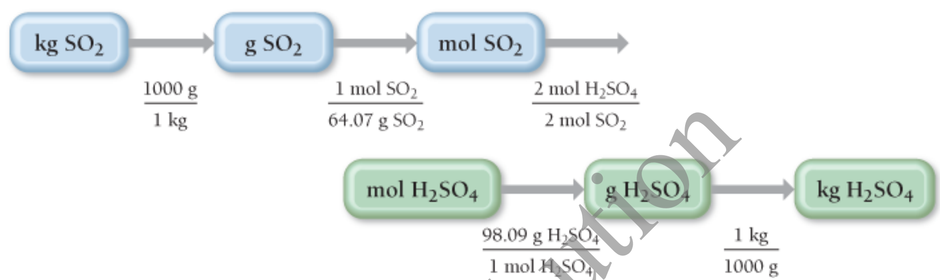
SORT The problem gives the mass of sulfur dioxide and asks you to find the mass of sulfuric acid.

GIVEN: 25 kg SO_2

FIND: kg H_2SO_4

STRATEGIZE The conceptual plan follows the standard format of mass → amount (in moles) → amount (in moles) → mass. Because the original quantity of SO_2 is given in kilograms, you must first convert to grams. You can deduce the relationship between moles of sulfur dioxide and moles of sulfuric acid from the balanced chemical equation. Because the final quantity is requested in kilograms, you convert to kilograms at the end.

CONCEPTUAL PLAN



RELATIONSHIPS USED

1 kg = 1000 g | 2 mol SO_2 : 2 mol H_2SO_4 (from chemical equation)

molar mass SO_2 = 64.07 g/mol | molar mass H_2SO_4 = 98.09 g/mol

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SOLVE Follow the conceptual plan to solve the problem. Begin with the given amount of SO_2 in kilograms and use the conversion factors to arrive at kg H_2SO_4

SOLUTION

$$25\ \cancel{\text{kg}}\ \cancel{SO_2} \times \frac{1000\ \cancel{\text{g}}}{1\ \cancel{\text{kg}}} \times \frac{1\ \cancel{\text{mol } SO_2}}{64.07\ \cancel{\text{g}}\ \cancel{SO_2}} \times \frac{2\ \cancel{\text{mol } H_2SO_4}}{2\ \cancel{\text{mol } SO_2}} \times \frac{98.09\ \cancel{\text{g}}\ H_2SO_4}{1\ \cancel{\text{mol } H_2SO_4}} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 38 \text{ kg } H_2SO_4$$

CHECK The units of the final answer are correct. The magnitude of the final answer (38 kg H_2SO_4) is larger than the given amount of SO_2 (25 kg). This is reasonable because in the reaction each SO_2 molecule "gains weight" by reacting with O_2 and H_2O.

FOR PRACTICE 7.5 Another component of acid rain is nitric acid, which forms when NO_2, also a pollutant, reacts with oxygen and water according to the simplified equation:

$$4\, NO_2(g) + O_2(g) + 2\, H_2O(l) \rightarrow 4\, HNO_3(aq)$$

The generation of the electricity used by a medium-sized home produces about 16 kg of NO_2 per year. Assuming that there is adequate O_2 and H_2O, what mass of HNO_3, in kilograms, can form from this amount of NO_2 pollutant?

Conceptual Connection 7.4 Stoichiometry

Conceptual Connection 7.5 Stoichiometry II

Interactive Interac

*Gerlach, T. M., Present-day CO_2 emissions from volcanoes; *Eos, Transactions, American Geophysical Union*, Vol. 72, No. 23, June 4, 1991, pp. 249 and 254–255.