Pearson

10.10: Gases in Chemical Reactions: Stoichiometry Revisited

In Chapter 7, we discussed how to use the coefficients in chemical equations as conversion factors between number of moles of reactants and number of moles of products in a chemical reaction. We can use these conversion factors to determine, for example, the mass of product obtained in a chemical reaction based on a given mass of reactant or the mass of one reactant needed to react completely with a given mass of another reactant. The general conceptual plan for these kinds of calculations is:

mass A → amount A (in moles) → amount B (in moles) → mass B

where A and B are two different substances involved in the reaction and the conversion factor between amounts (in moles) of each comes from the stoichiometric coefficients in the balanced chemical equation.

In reactions involving *gaseous* reactant or products, we often specify the quantity of a gas in terms of its volume at a given temperature and pressure. As we have seen, stoichiometry involves relationships between amounts in moles. For stoichiometric calculations involving gases, we can use the ideal gas law to determine the amounts in moles from the volumes or to determine the volumes from the amounts in moles:

$$n = \frac{PV}{RT} \quad V = \frac{nRT}{P}$$

> The pressures here could also be partial pressures.

The general conceptual plan for these kinds of calculations is:

P, V, T of gas A → amount A (in moles) → amount B (in moles) → P, V, T of gas B

Examples 10.14 and 10.15 demonstrate this kind of calculation.

Example 10.14 Gases in Chemical Reactions

Methanol (CH_3OH) can be synthesized by the reaction:

$$CO(g) + 2\,H_2(g) \rightarrow CH_3OH(g)$$

What volume (in L) of hydrogen gas, at a temperature of 355 K and a pressure of 738 mmHg, do we need to synthesize 35.7 g of methanol?

SORT The problem gives the mass of methanol, the product of a chemical reaction. You are asked to find the required volume of one of the reactants (hydrogen gas) at a specified temperature and pressure.

GIVEN: $35.7\text{ g } CH_3OH$,
$T = 355\text{ K}$, $P = 738\text{ mmHg}$

FIND: V_{H_2}

STRATEGIZE Calculate the required volume of hydrogen gas from the number of moles of hydrogen gas, which you can obtain from the number of moles of methanol via the stoichiometry of the reaction.

gas, which you can obtain from the number of moles of methanol via the stoichiometry of the reaction.

First, find the number of moles of methanol from its mass by using the molar mass.

Then use the stoichiometric relationship from the balanced chemical equation to find the number of moles of hydrogen you need to form that quantity of methanol.

Finally, substitute the number of moles of hydrogen together with the pressure and temperature into the ideal gas law to find the volume of hydrogen.

CONCEPTUAL PLAN

$$\text{g } CH_3OH \xrightarrow{\frac{1 \text{ mol } CH_3OH}{32.04 \text{ g } CH_3OH}} \text{mol } CH_3OH$$

$$\text{mol } CH_3OH \xrightarrow{\frac{2 \text{ mol } H_2}{1 \text{ mol } CH_3OH}} \text{mol } H_2$$

$$n \text{ (mol } H_2), P, T \xrightarrow{PV = nRT} V_{H_2}$$

RELATIONSHIPS USED

$PV = nRT$ (ideal gas law)

$2 \text{ mol } H_2 : 1 \text{ mol } CH_3OH$ (from balanced chemical equation)

molar mass $CH_3OH = 32.04 \text{ g/mol}$

SOLVE Follow the conceptual plan to solve the problem. Begin by using the mass of methanol to determine the number of moles of methanol.

Next, convert the number of moles of methanol to moles of hydrogen.

Finally, use the ideal gas law to find the volume of hydrogen. Before substituting into the equation, you need to convert the pressure to atmospheres.

SOLUTION

$$35.7 \ \cancel{\text{g } CH_3OH} \times \frac{1 \text{ mol } CH_3OH}{32.04 \ \cancel{\text{g } CH_3OH}} = 1.1\underline{1}42 \text{ mol } CH_3OH$$

$$1.1\underline{1}42 \ \cancel{\text{mol } CH_3OH} \times \frac{2 \text{ mol } H_2}{1 \ \cancel{\text{mol } CH_3OH}} = 2.2\underline{2}84 \text{ mol } H_2$$

$$V_{H_2} = \frac{n_{H_2}RT}{P}$$

$$P = 738 \ \cancel{\text{mmHg}} \times \frac{1 \text{ atm}}{760 \ \cancel{\text{mmHg}}} = 0.97\underline{1}05 \text{ atm}$$

$$V_{H_2} = \frac{\left(2.2\underline{2}84 \ \cancel{\text{mol}}\right)\left(0.08206 \frac{\text{L} \cdot \cancel{\text{atm}}}{\cancel{\text{mol}} \cdot \cancel{\text{K}}}\right)\left(355 \ \cancel{\text{K}}\right)}{0.97\underline{1}05 \ \cancel{\text{atm}}} = 66.9 \text{ L}$$

CHECK The units of the answer are correct. The magnitude of the answer (66.9 L) seems reasonable. You are given slightly more than one molar mass of methanol, which is therefore slightly more than 1 mol of methanol. From the equation you can see that you need 2 mol hydrogen to make 1 mol methanol, so the answer must be slightly greater than 2 mol hydrogen. Under standard temperature and pressure, slightly more than 2 mol hydrogen occupies slightly more than $2 \times 22.4 \text{ L} = 44.8 \text{ L}$. At a temperature greater than standard temperature, the volume would be even greater; therefore, this answer is reasonable.

FOR PRACTICE 10.14 In the following reaction, 4.58 L of O_2 was formed at $P = 745$ mmHg and $T = 308$ K. How many grams of Ag_2O decomposed?

$$2\ Ag_2O\,(s) \rightarrow 4\ Ag\,(s) + O_2\,(g)$$

FOR MORE PRACTICE 10.14 In the reaction in For Practice 10.14, what mass of $Ag_2O\,(s)$ (in grams) is required to form 388 mL of oxygen gas at $P = 734$ mmHg and 25 °C?

Interactive Worked Example 10.14 **Gases in Chemical Reactions**

Molar Volume and Stoichiometry

In Section 10.5, we saw that, under standard temperature and pressure, 1 mol of an ideal gas occupies 22.4 L. Consequently, if a reaction occurs at or near standard temperature and pressure, we can use 1 mol = 22.4 L as a conversion factor in stoichiometric calculations, as we demonstrate in Example 10.15.

Example 10.15 Using Molar Volume in Gas Stoichiometric Calculations

How many grams of water form when 1.24 L of H_2 gas at STP completely reacts with O_2?

$$2\ H_2\,(g) + O_2\,(g) \rightarrow 2\ H_2O\,(g)$$

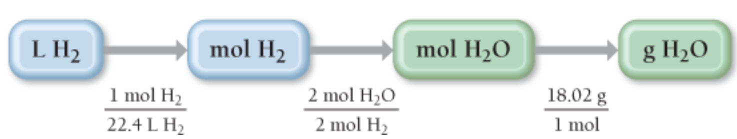
SORT You are given the volume of hydrogen gas (a reactant) at STP and asked to determine the mass of water that forms upon complete reaction.

GIVEN: 1.24 L H_2

FIND: g H_2O

STRATEGIZE Because the reaction occurs under standard temperature and pressure, you can convert directly from the volume (in L) of hydrogen gas to the amount in moles. Then use the stoichiometric relationship from the balanced equation to find the number of moles of water formed. Finally, use the molar mass of water to obtain the mass of water formed.

CONCEPTUAL PLAN



RELATIONSHIPS USED

1 mol = 22.4 L (at STP)

2 mol H_2 : 2 mol H_2O (from balanced equation)

molar mass H_2O = 18.02 g/mol

SOLVE Follow the conceptual plan to solve the problem.

$$1.24\ \cancel{\text{L H}_2} \times \frac{1\ \cancel{\text{mol H}_2}}{22.4\ \cancel{\text{L H}_2}} \times \frac{2\ \cancel{\text{mol H}_2\text{O}}}{1\ \cancel{\text{mol H}_2}} \times \frac{18.02\ \text{g H}_2\text{O}}{1\ \cancel{\text{mol H}_2\text{O}}} = 0.998\ \text{g H}_2\text{O}$$

CHECK The units of the answer are correct. The magnitude of the answer (0.998 g) is about 1/18 of the molar mass of water, roughly equivalent to the approximately 1/22 of a mole of hydrogen gas given, as you would expect for the 1:1 stoichiometric relationship between number of moles of hydrogen and number of moles of water.

FOR PRACTICE 10.15 How many liters of oxygen (at STP) are required to form 10.5 g of H_2O?

$$2\ \text{H}_2\,(g) + \text{O}_2\,(g) \rightarrow 2\ \text{H}_2\text{O}\,(g)$$

Conceptual Connection 10.7 **Pressure and Number of Moles**

Interactive