

5.4: Gas Stoichiometry

What you Should Know

Before beginning this section, you should know and understand:

- The concept of [Stoichiometric proportions](#) in reactions
- [Avogadro's Constant](#) and its uses in converting grams to moles
- [The Ideal Gas Law](#): $PV = nRT$
- [Dalton's Law](#) of Partial Pressure

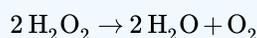
Chemical reactions between gaseous materials are quite similar to reactions between solids and liquids, except the Ideal Gas Law ($PV = nRT$) can now be included in the calculations. If a chemical reaction is reversible (such as the decay and formation of dinitrogen tetraoxide), then Dalton's Law of Partial Pressure may be used to determine the moles of reactants and products at which the reaction ceases (and subsequently, the temperature, pressure and volume of each gas can be determined as well).

Non-Reversible Reactions

A non-reversible reaction uses the reactants to form the products. The reaction goes in one direction; that is, using the product to recreate the reactants has greatly different requirements. One of the most common forms of non-reversible reactions is combustion (once an organic molecule has been converted to water and hydrogen gas, it is extremely difficult to reform). Other non-reversible reactions produce a state change, such as Hydrogen Peroxide (the gaseous material produces water, a liquid). To understand how the Ideal Gas Law applies to reactions, we shall use a nonreversible reaction as an example.

✓ Example 5.4.1: the decomposition of Hydrogen Peroxide to water and oxygen gas

If 4.000 grams of hydrogen peroxide is placed within a sealed 250 milliliter container at 500 K. What is the pressure of the oxygen gas produced in atmospheres?



Solution

First, we need to determine the moles of O_2 produced, just like any other stoichiometric problem.

$$(4g \text{ H}_2\text{O}_2) \times \left(\frac{1 \text{ mol H}_2\text{O}_2}{34.016 \text{ g H}_2\text{O}_2} \right) \left(\frac{1 \text{ mol O}_2}{2 \text{ mol H}_2\text{O}_2} \right) = 0.0588 \text{ mol O}_2$$

With the moles of oxygen determined, we can now use the Ideal Gas Law to determine the pressure.

$$PV = nRT$$

The volume (250 mL = 0.25 L) and temperature (500 K) are already given to us, and R (0.0820574 Latm mol⁻¹K⁻¹) is a constant.

$$\begin{aligned} P &= \frac{nRT}{V} \\ &= \frac{(0.0588 \text{ mol O}_2) \times (0.0820 \text{ L atm mol}^{-1} \text{ K}^{-1}) \times (500 \text{ K})}{0.25 \text{ L}} \\ &= 9.65 \text{ atm} \end{aligned}$$

By using the Ideal Gas Law for unit conversions, properties such as the pressure, volume, moles, and temperature of a gas involved in a reaction can be determined. However, a different approach is needed to solve reversible reactions.

For further clarification, when solving equations with gases, we must remember that gases behave differently under different conditions. For example, if we have a certain temperature or pressure, this can change the number of moles produced or the volume. This is unlike regular solids where we only had to account for the mass of the solids and solve for the mass of the product

by stoichiometry. In order to solve for the temperature, pressure, or volume of a gas using chemical reactions, we often need to have information on two out of three of these variables. So we need either the temperature and volume, temperature and pressure, or pressure and volume. The mass we can find using stoichiometric conversions we have learned before.

The reason why gases require additional information is because gases behave as ideal gases and ideal gases behave differently under different conditions. To account for these conditions, we use the ideal gas equation $PV=nRT$ where P is the pressure measured in atmosphere(atm), V is the volume measured in liters (L), n is the number of moles, R is the gas constant with a value of $.08206 \text{ L atm mol}^{-1} \text{ K}^{-1}$, and T is the temperature measured in kelvin (K).

✓ Example 5.4.2

Suppose we have the following combustion reaction (below). If we are given 2 moles of ethane at STP, how many liters of CO_2 are produced?



Solution

Step 1

First use stoichiometry to solve for the number of moles of CO_2 produced.

$$(2 \text{ mol C}_2\text{H}_6) \left(\frac{4 \text{ mol CO}_2}{2 \text{ mol C}_2\text{H}_6} \right) = 4 \text{ mol CO}_2$$

So 4 moles of Carbon Dioxide are produced if we react 2 moles of ethane gas.

Step 2

Now we simply need to manipulate the ideal gas equation to solve for the variable of interest. In this case we are solving for the number of liters.

Since we are told ethane is at STP, we know that the temperature is 273 K and the pressure is 1 atm.

So the variables we have are:

- $V = ?$
- $T = 273\text{K}$
- $P = 1 \text{ atm}$
- $n = 4 \text{ moles CO}_2$
- $R \text{ (gas constant)} = 0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1}$

Isolating the variable of interest from $PV = nRT$, we get

$$\begin{aligned} V &= \frac{nRT}{P} \\ &= \frac{4 \text{ mol} \times 0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1} \times 273 \text{ K}}{1 \text{ atm}} \\ &= 89.61 \text{ L} \end{aligned}$$

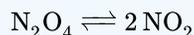
So we have a volume of 89.61 liters.

Reversible Reactions in Gases

A reversible reaction is a chemical reaction in which reactants produces a product, which then decays back to the reactants. This continues until the products and reactants are in equilibrium. In other words, the final state of the gas includes both the reactants and the products. For example, Reactant A combines with Reactant B to form Product AB, which then breaks apart into A and B, until an equilibrium of the three is reached. In a reaction between gases, determining gas properties such as partial pressure and moles can be quite difficult. For this example, we consider the thermal decomposition of Dinitrogen Tetraoxide into Nitrogen Dioxide.

✓ Example 5.4.3

For this example, we shall use Dinitrogen Tetraoxide, which decomposes to form Nitrogen Dioxide.



2 atm of dinitrogen tetraoxide is added to a 500 mL container at 273 K. After several minutes, the total pressure of N_2O_4 and 2NO_2 at equilibrium is found to be 3.2 atm. Find the partial pressures of both gases.

Solution

The simplest way of solving this problem is to begin with an **ICE table**.

	N_2O_4	2NO_2	Description of Each Letter
Initial	2atm	0atm	The initial amounts of reactants and products
Change	-X	+2X	The unknown change, X, each one multiplied by their stoichiometric factor
Equilibrium	2-Xatm	2x atm	The initial + the change

With this data, a simple equation can be derived to determine the value of X.

$$P_{total} = (2 - X) + 2X = 3.2 \text{ atm} \quad (5.4.2)$$

$$X = 1.2 \text{ atm} \quad (5.4.3)$$

$$P_{\text{NO}_2} = 2x = 2.4 \text{ atm} \quad (5.4.4)$$

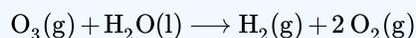
$$P_{\text{N}_2\text{O}_4} = 2 - x = 0.8 \text{ atm} \quad (5.4.5)$$

Law of Combining Volumes

This law of combining volumes was first discovered by the famous scientist Gay-Lussac who noticed this relationship. He determined that if certain gases that are products and reactions in a chemical reaction are measured at the same conditions, temperature and pressure, then the volume of gas consumed/produced is equal to the ratio between the gases or the ratio of the coefficients.

✓ Example 5.4.4

If ozone, hydrogen, and oxygen were all measured at 35°C and at 753 mmHg, then how many liters of ozone was consumed if you had 5 liters of oxygen gas?



Solution

Step 1

Identify what we are looking for and if any relationships can be spotted. In this case, we can see that there are three gases all at the same temperature and pressure, which follows Gay-Lussac's Law of Combining Volumes. We can now proceed to use his law.

Step 2

We simply change the coefficients to volumetric ratios. So for every 1 Liter of Ozone gas we have, we produce 1 Liter of H_2 gas and 2 Liter of O_2 gas.

We are given 5 liters of Oxygen gas and want to solve for the amount of liters of ozone consumed. We simply use the 2:1 stoichiometry of the reaction.

$$5L O_2 \left(\frac{1 L O_3}{2 L O_2} \right) = 2.5 L O_3 \quad (5.4.6)$$

Problems

1. A 450 mL container of oxygen gas is at STP. Hydrogen gas is pumped into the container, producing water. What is the least amount of mL of Hydrogen gas needed in order to react the oxygen to completion?



2. This reaction occurred at 427 Kelvin, with 37 g of CH_4 and an excess of oxygen. The carbon dioxide produced was captured in a 30L sealed container. What is the pressure of the carbon dioxide within the container?



Solution

1. 900 mL H_2 : To solve this question you simply use Gay-Lussac's law of combining volume because both gases are at the same temperature at pressure.

$$450 \text{ mL } O_2 \left(\frac{2 \text{ mL } H_2}{1 \text{ mL } O_2} \right) = 900 \text{ mL } H_2 \quad (5.4.9)$$

2. 2.7 atm: For a question like this, you want to first determine the number of moles of the compound in interest first using stoichiometry and then using the ideal gas law to solve for the variable of interest.

$$37 \text{ g } CH_4 \left(\frac{1 \text{ mol } CH_4}{16 \text{ g } CH_4} \right) \left(\frac{1 \text{ mol } CO_2}{1 \text{ mol } CH_4} \right) = 2.3 \text{ mol } CO_2 \quad (5.4.10)$$

Now use the ideal gas law to solve for the pressure of CO_2 .

- $PV=nRT$
- $P = ?$
- $V = 30L$
- $n = 2.3 \text{ mol } CO_2$
- $T = 427 \text{ K}$
- $P = nRT/V$

$$P = (2.3 \text{ mol} * .08206 \text{ L atm mol}^{-1} \text{ K}^{-1} * 427 \text{ K}) / (30 \text{ L})$$

$$P = 2.7 \text{ atm}$$

References

1. Petrucci, Harwood, Herring, Madura. General Chemistry: Principles and Modern Applications. Prentice Hall, New Jersey 2007.

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