Guided Practice Activities

**Module**: States of Matter

**Section**: Gas Stoichiometry - Key
Writing Gas Law Reactions

Activity 1

The purpose of this activity is to check your understanding of chemical change involving gases.

1. Correctly write and balance the chemical reaction that occurs between hydrogen gas and oxygen gas to produce water (in gas phase).

   \[ 2H_2 (g) + O_2 (g) \rightarrow 2H_2O (g) \]

2. Draw a small particle model of this reaction showing the reactant gases present as a mixture in a container on the left hand side of the equation and the product gas shown on the right hand side of the equation. In your drawing assume that you have exact stoichiometric amounts of the reactants and products needed to completely react away both reactants.

   Because these molecules are rather small it is easy to include each atom in the drawing. However, simply drawing spheres for a gas molecule is acceptable in a small particle model.

   ![Particle Model](image)

   H₂ gas and O₂ gas  →  H₂O gas

3. Assuming that your reactants and products are at the same temperature and pressure. What is the relationship between the volume of the reactants and the volume of the products?

   There are more total moles of gas on the reactants side. If the temperature and pressure are constant for the reactant mixture and the product gas then the volume will be greater on the side with the larger number of gas moles, which is the reactant side. The volume of the reactant container is 1.5 times larger than the product container because the ratio of reactant gas moles to product gas moles is 3 to 2.

4. Comment on the relationship between the moles of the reactants and products and the volumes of the reactants and products. What gas law is associated with this relationship?

   A greater number of gas moles corresponds to a larger volume (if temperature and pressure are constant values). As explained above, the reactant mixture for this reaction has 1.5 times more moles of gas total than the product gas. Therefore, the reactant mixture has a
larger volume than the product gas. The gas law associated with this relationship is Avogadro’s Law: \( V_{1n_2} = V_{2n_1} \)

**Activity 2**

The purpose of this activity is to check your understanding of chemical change involving a gas.

1. Correctly write and balance the chemical reaction that occurs between butane gas and oxygen gas to produce carbon dioxide gas and water gas.

\[
2C_4H_{10} \text{ (g)} + 13O_2 \text{ (g)} \rightarrow 8CO_2 \text{ (g)} + 10H_2O \text{ (g)}
\]

2. Draw a small particle model of this reaction showing the reactant gases present as a mixture in a container on the left hand side of the equation and the product gases shown on the right hand side of the equation. In your drawing assume that you have exact stoichiometric amounts of the reactants and products needed to completely react away both reactants.

All gas molecules will be represented as spheres in this small particle model.

![Particle model](image)

3. Assuming that your reactants and products are at the same temperature and pressure. What is the relationship between the volume of the reactants and the volume of the products?

There are more total moles of gas on the products side. If the temperature and pressure are constant for the reactant mixture and the product mixture then the volume will be greater on the side with the larger number of gas moles, which is the products side. The ratio of product gas moles to reactant gas moles is 6 to 5 (simplified from 18:15) therefore the ratio of the product volume to the reactant volume is also 6 to 5. We could also say that the product volume is 1.2 times larger than the reactant volume.
4. Comment on the relationship between the moles of the reactants and products and the volumes of the reactants and products. What gas law is associated with this relationship?

A greater number of gas moles corresponds to a larger volume (if temperature and pressure are constant values). As explained above, the product mixture for this reaction has 1.2 times more moles of gas total than the reactant mixture. Therefore, the product mixture has a larger volume than the reactant mixture. The gas law associated with this relationship is Avogadro's Law: $V_1n_2 = V_2n_1$

**Gas Stoichiometry Problems**

**ACTIVITY 1**

The purpose of this activity is to check your quantitative skills associated with gas stoichiometry.

1. The chemical reaction that occurs in an automobile air bag is that solid sodium azide (NaN₃) is ignited and then quickly decomposes to form solid sodium and nitrogen gas according to the following balanced equation: $2\text{NaN}_3 \rightarrow 2\text{Na} + 3\text{N}_2$ Assuming you start with exactly 1 mole of sodium azide how many moles of nitrogen gas would be produced upon complete decomposition of the sodium azide?

   $$1\text{mol NaN}_3 \times \frac{3\text{mol N}_2}{2\text{mol NaN}_3} = 1.5\text{mol N}_2$$

2. Consider the gas producing chemical equation in the previous question. Assume you want to inflate the air bag to a volume of 35 mL with a pressure of 1.6 atm at 22 °C. How many grams of sodium azide must be present in the deflated bag tucked behind your steering wheel in order to inflate the bag to the desired volume and pressure?

   The only gas in the reaction is the product gas – nitrogen. So we must determine how many moles of nitrogen we need to produce in order to inflate the air bag to the desired conditions. We can use the ideal gas law for this part of the calculation. Then, we will use stoichiometry to work backward to determine the amount of sodium azide reactant needed to produce the necessary amount of nitrogen gas.
Find moles $N_3$:

\[
PV = nRT
\]

\[
P = 1.6 \text{ atm}
\]

\[
V = 35 \text{ mL} = 0.035 \text{ L}
\]

\[
R = 0.08206 \ \text{ L} \cdot \text{ atm} \cdot \text{ mol}^{-1} \cdot \text{ K}^{-1}
\]

\[
T = 22^\circ \text{C} = 295.15 \text{ K}
\]

\[
n = ?
\]

\[
n = \frac{PV}{RT}
\]

\[
n = \frac{(1.6 \text{ atm})(0.035 \text{ L})}{(0.08206 \ \text{ L} \cdot \text{ atm} \cdot \text{ mol}^{-1} \cdot \text{ K}^{-1})(295.15 \text{ K})}
\]

\[
n = 2.3 \times 10^{-3} \text{ mol } N_3
\]

Convert moles $N_3$ to moles $NaN_3$:

\[
2.3 \times 10^{-3} \text{ mol } N_3 \times \frac{2 \text{ mol } NaN_3}{3 \text{ mol } N_3} = 1.5 \times 10^{-3} \text{ mol } NaN_3
\]

Find grams $NaN_3$:

\[
\text{mass} = n \times \text{molarmass}
\]

molarmass $NaN_3 = 65 \ \text{ g/mol}$

\[
\text{mass} = (1.5 \times 10^{-3} \text{ mol } NaN_3)(65 \ \text{ g/mol})
\]

\[
\text{mass} = 0.10 \text{ grams } NaN_3
\]

3. Heating a 4.75 g sample of an ore containing a metal sulfide in the presence of excess oxygen, produces 1.20 L of dry $SO_2$ gas measured at 37 °C and 800 torr. Calculate the moles of $SO_2$ formed. What mass of S must have been present in the sample of the ore? What was the mass percentage of sulfur in the ore?

First, we can use the given conditions and the ideal gas law to calculate the number of $SO_2$ gas moles formed. Then, we will use stoichiometry to determine the mass of sulfur in the products and therefore its percentage in the ore.
Find moles $\text{SO}_2$:

$$PV = nRT$$
$$P = 800\text{Torr}$$
$$V = 1.20\text{L}$$
$$R = 62.364 \frac{\text{L-Torr}}{\text{mol-K}}$$
$$T = 37^\circ\text{C} = 310.15\text{K}$$

$$n = \frac{PV}{RT} = \frac{(800\text{Torr})(1.20\text{L})}{(62.364 \frac{\text{L-Torr}}{\text{mol-K}})(310.15\text{K})} = 4.96 \times 10^{-2}\text{mol} \text{SO}_2$$

**Mass of Sulfur**:

$$\text{moles SO}_2 \xrightarrow{\text{mole ratio}} \text{moles S} \xrightarrow{\text{atomic mass } S} \text{grams S}$$

$$4.96 \times 10^{-2}\text{mol} \text{SO}_2 \times \frac{1\text{mol } S}{1\text{mol} \text{SO}_2} \times 32.06\text{ g } S \times \frac{1\text{mol } S}{1\text{mol } S} = 1.59\text{ g } S$$

**Percentage of Sulfur in Ore**:

$$\text{mass}\% S = 100 \times \frac{\text{mass of sulfur}}{\text{mass of ore}}$$

$$\text{mass}\% S = 100 \times \frac{1.59\text{ g}}{4.75\text{ g}} = 33.5\%$$

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**ACTIVITY 2**

The purpose of this activity is check your understanding of gas stoichiometry when there is a limiting reactant to consider.

1. If 2.00 L of nitrogen and 4.00 L of hydrogen were allowed to react, how many liters of ammonia gas could form. Assume that all gases are at the same temperature and pressure and that the limiting reactant is used up. What would be the final volume of the mixture?

   First, we need to write the balanced chemical equation:

   $$\text{N}_2 (g) + 3\text{H}_2 (g) \rightarrow 2\text{NH}_3 (g)$$

   Notice for every one mole of nitrogen gas used, the reaction requires three moles of hydrogen gas. The reaction uses up hydrogen gas at a much faster rate. But we need to assess the given amounts of each reactant gas using their stoichiometric coefficients before
determining the limiting reactant. Volume and molar amounts are proportional at a given temperature and pressure. So when assessing the limiting reactant for gas reactants we can simply their volumes alone! Usually, we have to convert to moles first, but again volume and molar amounts are proportional at a given temperature and pressure.

Then, once we determine the limiting reactant, we can calculate the volumes of product formed AND of remaining reactant gas. They will both contribute the overall volume once the limiting reactant has been used up.

Also, because a pressure and temperature aren’t given, we can assume the gases are at STP and use the molar volume of an ideal gas at STP to convert between volume and moles (although, we’ll find that this value cancels out in each of our stoichiometric calculations!).

\[
\text{Find the L.R.:} \\
\text{Volume } \propto \text{ moles} \\
\text{Divide volume by stoich. coefficients:} \\
2.00L \text{ } N_2 = 2.00 \\
\quad 1 \text{ mol } N_2 \\
4.00L \text{ } H_2 = 1.33 \\
\quad 3 \text{ mol } H_2 \\
1.33 < 2.00 \\
\therefore H_2 \text{ is L.R.} \\
\]

\[
\text{Liters of } NH_3 \text{ formed:} \\
\text{Assume STP!} \\
4.00L \text{ } H_2 \times \frac{1 \text{ mol } H_2}{22.4L \text{ } H_2} \times \frac{2 \text{ mol } NH_3}{3 \text{ mol } H_2} \times \frac{22.4L \text{ } NH_3}{1 \text{ mol } NH_3} = 2.67L \text{ } NH_3 \\
\]

\[
\text{Liters of } N_2 \text{ leftover:} \\
4.00L \text{ } H_2 \times \frac{1 \text{ mol } H_2}{22.4L \text{ } H_2} \times \frac{1 \text{ mol } N_2}{3 \text{ mol } H_2} \times \frac{22.4L \text{ } N_2}{1 \text{ mol } N_2} = 1.33L \text{ } N_2 \text{ USED} \\
2.00L \text{ } N_2 - 1.33L \text{ } N_2 = 0.67L \text{ } N_2 \text{ leftover} \\
\]
\[
\text{Total Final Volume:} \\
\text{Volume } NH_3 \text{ formed + Volume } N_2 \text{ leftover} = 2.67L \text{ } NH_3 + 0.67L \text{ } N_2 = 3.34L \text{ gas total} \\
\]

2. Assume that you burn 10.0 L of ammonia in 15.0 L of oxygen at 500 °C. What volume of nitric oxide, NO, gas can form? What volume of steam (water vapor) is formed? Assume all gases are at the same temperature and pressure and that all the limiting reactant is used up.

First, we need to write the balanced chemical equation:

\[
4NH_3 (g) + 5O_2 (g) \rightarrow 4NO (g) + 6H_2O (g) \\
\]
Volume and molar amounts are proportional at a given temperature and pressure. So when assessing the limiting reactant for gas reactants we can simply their volumes alone! Usually, we have to convert to moles first, but again volume and molar amounts are proportional at a given temperature and pressure.

Then, once we determine the limiting reactant, we can calculate the volumes of product gases that are formed. Unfortunately, the gases are not at STP and we cannot use the molar volume of an ideal gas at STP to convert between volume and moles. We can solve for the molar volume of an ideal gas at 500°C and some pressure and use this number for our calculations. A pressure is not given, so let’s select a simple pressure for our calculations – 1 atm. Alternatively, we could use the ideal gas law to convert between volume and moles for each compound. Choose whichever method makes more sense to you!

\[ \text{Find the L.R.:} \]
\[ \text{Volume } \alpha \text{ moles} \]
\[ \frac{10.00L \text{ } NH_3}{4 \text{ mol } NH_3} = 2.5 \]
\[ \frac{15.0L \text{ } O_2}{5 \text{ mol } O_2} = 3 \]
\[ 2.5 < 3 \therefore \text{NH}_3 \text{ is L.R.} \]

\[ \text{Adjusted Molar Volume of Ideal Gases:} \]
\[ PV = nRT \]
\[ \text{Assume } P = 1 \text{ atm} \]
\[ T = 500° C = 773.15K \]
\[ R = 0.08206 \text{ Latm/mol K} \]
\[ \frac{V}{n} = ? \]
\[ \frac{V}{n} = \frac{RT}{P} = \frac{(0.08206 \text{ Latm/mol K})(773.15K)}{(1 \text{ atm})} = 63.445 \frac{L}{\text{mol}} \]

\[ \text{Liters of NO formed:} \]
\[ 10.00L \text{ } NH_3 \times \frac{1 \text{ mol } NH_3}{63.445L \text{ } NH_3} \times \frac{4 \text{ mol NO}}{4 \text{ mol } NH_3} \times \frac{63.445L \text{ NO}}{1 \text{ mol NO}} = 10.00L \text{ NO} \]

\[ \text{Liters of H}_2\text{O formed:} \]
\[ 10.00L \text{ } NH_3 \times \frac{1 \text{ mol } NH_3}{63.445L \text{ } NH_3} \times \frac{6 \text{ mol } H_2O}{4 \text{ mol } NH_3} \times \frac{63.445L \text{ } H_2O}{1 \text{ mol } H_2O} = 15.00L \text{ H}_2\text{O} \]
Notice, that despite all of our hard work, we didn’t actually need to calculate a new molar volume for the gases at this temperature and pressure. In both stoichiometric calculations, the adjusted molar volume cancelled out. This is again because volume and moles are proportional for a given set of conditions! Still it was good practice!