5

Gases

I. Measurements on Gases

A. Volume

1. Units

a. liter (L)

b. milliliter (mL)

c. cubic centimeter (cm^3)

d. cubic meter (m^3)

2. Equivalences

1 L = ______________________ mL.

1 L = ______________________ cm^3.

1 L = ______________________ m^3.

B. Amount

1. Unit/Symbol

a. moles (mol)/n

b. mass in grams/m

2. Equivalence

$$n = \frac{m}{MM}$$

C. Temperature

1. Unit

a. Kelvin (K)

b. Centigrade or Celsius (°C)

2. Equivalence

K = (°C) + 273.15

D. Pressure

1. Unit

a. millimeters of mercury (mm Hg)

b. atmosphere (atm)

c. bar

2. Equivalences

1 atm = ______________ mm Hg

1 atm = ______________ bar

II. Ideal Gas Law

A. Mathematical Expression

1. PV = nRT

P = ______________

V = ______________

n = ______________

R = ______________

T = ______________

2. n is always in moles.
T is always in Kelvin.

3. R – gas constant
If P is in atm and V is in L, then R is ______________. This is the most commonly used R.

B. Final and initial state problems

1. Steps to follow in solving these problems:

   a. List all variables (P, V, n, T) in two columns, one for the initial state, the other for the final state.

   b. Write two gas equations, one for the initial state, the other for the final state. Give subscript 1 for the initial state and 2 for the final state for the variables that vary. Those that remain constant do not get any subscript.

   c. Rewrite both gas equations so that subscripted variables are on the left and nonsubscripted variables are on the right. Equate left sides of both equations to each other.

   d. Substitute into the single equation obtained in (c) the variables given for both initial and final state, and solve for the desired variable.

2. Exercises

a. A sample of gas occupies 355 mL at 15 °C and 755 mm Hg pressure. What temperature will the gas have at the same pressure if its volume increases to 453 mL?

Solution:

(1) List all variables (P, V, n, T) in two columns, one for the initial state, the other for the final state.

P = 755 mm Hg	P = 755 mm Hg
V = 355 mL	V = 453 mL
T = 288 K	T = ?
n = constant	n = constant

As you can see, volume and temperature change, whereas pressure and amount remain constant.

(2) Our two gas equations are therefore

$$PV_1 = nRT_1 \qquad PV_2 = nRT_2$$

We did not give the subscripts 1 and 2 to P and n because P and n remain constant.

(3) Putting all constants (no subscripts) on one side and all variables (with subscripts) on the other, we get

$$\frac{V_1}{T_1} = \frac{nR}{P} \qquad \frac{V_2}{T_2} = \frac{nR}{P}$$

resulting in the single equation

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

(4) Substituting, we get

$$\frac{355\ \text{mL}}{288\ \text{K}} = \frac{453\ \text{mL}}{T_2}$$

$$T_2 = \frac{288\ \text{K} \times 453\ \text{mL}}{355\ \text{mL}} = 368\ \text{K}$$

A word about units: For initial and final state problems, do not worry about making units compatible with those of R. You only have to make sure that the units of the initial state and those of the final state remain the same. Temperature, however, should always be in Kelvin.

b. A sample of gas at 30.0°C has a pressure of 1.23 bar and a volume of 0.0247 m^3. If the volume of the gas is compressed to 0.00839 m^3 at the same temperature, what is its pressure at this volume? **(E1)** *c.* A sample of gas at 27.0°C has a volume of 2.08 L and a pressure of 750.0 mm Hg. If the gas is in a sealed container, what is its pressure in bar when the temperature (in °C) doubles? **(E2)**

C. One-state problems

1. One-state or single-state problems involve R and the variables P, V, n, and T. In these problems one measurement for each of three variables is given and you are asked to solve for the fourth variable. The units for the R that you use will determine the units that you will have to use for the variables. Remember T must always be in Kelvin.

2. Exercises:

a. How many grams of oxygen gas in a 10.0-L container will exert a pressure of 712 mm Hg at a temperature of 25.0°C?

Solution: Since R = 0.0821 L·atm/K·moles, V has to be in liters, P in atmospheres, T in Kelvin and n in moles. In this problem

V = 10.0 L

$$P = 712 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}} = 0.937 \text{ atm}$$

T = 25.0°C + 273.15 = 298.1 K

n = unknown variable asked for in grams

Substituting into the ideal gas equation, we get

$$PV = nRT$$

$$n = \frac{0.937 \text{ atm} \times 10.0 \text{ L}}{0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times 298.1 \text{ K}} = 0.383 \text{ mol } O_2$$

Since the amount of gas has to be expressed in grams of oxygen, we use the conversion factor

$$32.0 \text{ g} = 1 \text{ mol}$$

Thus

$$0.383 \text{ mol} \times \frac{32.0 \text{ g}}{1 \text{ mol}} = 12.3 \text{ g } O_2$$

b. At what temperature will 26.42 g of sulfur dioxide in a 250.0-mL container exert a pressure of 0.842 atm? **(E3)**

D. Other calculations involving the Ideal Gas Law

1. Molar mass calculations

a. We can often determine the molar mass of a gas by measuring its mass at a given volume, temperature, and pressure. To do this, we simply write m/MM instead of n in the ideal gas equation. We therefore get

$$PV = \frac{m}{MM}RT \qquad \text{instead of} \qquad PV = nRT$$

where MM is the molar mass, and m is the mass in grams. Thus

$$MM = \frac{mRT}{PV}$$

If you remember to keep your units straight, this type of problem becomes a simple "plug-in".

b. ***Exercises***

(1) When 4.93 g of carbon tetrachloride gas are in a 1.00-L container at 400.0 K, the gas exerts a pressure of 800.0 mm Hg. What is the molar mass of carbon tetrachloride?

Solution: Using the equation MM = mRT/PV, we get

$$MM = \frac{4.93\text{ g} \times 0.0821\dfrac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \times 400.0\text{ K}}{1.00\text{ L} \times 800.0\text{ mm} \times \dfrac{1\text{ atm}}{760\text{ mm}}} = 154\ \frac{\text{g}}{\text{mol}}$$

To check, we know that carbon tetrachloride is written as CCl_4. Thus, its molar mass must be 12.0 + 4(35.45) = 153.8. Remember that the compound with the smallest molar mass is hydrogen (2.02 g/mol). Therefore, if you arrive at a molar mass of less than 2.02 g/mol, you are doing something wrong!

(2) A gaseous compound with a mass of 3.216 g has a volume of 2236 mL at 27.0°C and 735 mm Hg. Calculate the approximate molar mass of the compound. **(E4)**

2. Density calculations

a. The equation PV = nRT can also be used to calculate the density of a gas at given conditions of temperature and pressure.
The variable to find is n.
Since density is the mass of gas that occupies one liter, then V = 1.00 L.
After finding n, change n to mass and you get mass/L.

b. Given the density of a gas at conditions of temperature and pressure, one can also determine the molar mass of the gas.
Take the mass portion of the density value and call that m.
The volume is 1.00 L.
You can then use the expression

$$PV = \frac{m}{MM} RT$$

to find MM.

c. Exercises

(1) Calculate the density of ammonia at 0°C and 1 atm pressure.
Solution: Here P = 1.00 atm, T = 273 K and V = 1.00 L. We find n using the ideal gas equation, PV = nRT:

$$n = \frac{1.000 \text{ atm} \times 1.000 \text{ L}}{0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times 273 \text{ K}} = 0.0446 \text{ mol}$$

We now use the molar mass of NH_3 (17.03 g/mol) to determine how many grams are equivalent to 0.0446 mol.

$$0.0446 \times 17.03 \frac{\text{g}}{\text{mol}} = 0.760 \frac{\text{g}}{\text{L}}$$

Note that the densities of gases are expressed as g/L, whereas those of liquids and solids are expressed in g/mL.

(2) Calculate the molar mass of a compound with a density of 0.630 g/L at 25.0°C and 730.0 mm Hg pressure. **(E5)**

III. Stoichiometry of Reactions Involving Gases

A. Volumes at specified temperature and pressure measurements

Since gases are often measured in liters or milliliters instead of grams, you need to convert that volume into either grams or moles in order to solve stoichiometric problems involving gases.

Exercises

1. Consider the reaction

$$MnO_2\,(s) + 2\,Cl^-\,(aq) + 4\,H^+\,(aq) \rightarrow Mn^{2+}\,(aq) + Cl_2\,(g) + 2\,H_2O$$

How many grams of MnO_2 are required to produce 1.200 L of Cl_2 gas at 1.00 atm and 200°C?

Solution: Here you want to convert 1.200 L of Cl_2 (at 1.00 atm and 200°C) to grams of MnO_2. For this reaction, the equivalences are

$$1\text{ mol } MnO_2 = 1\text{ mol } Cl_2 = 86.94\text{ g } MnO_2 = 70.90\text{ g } Cl_2$$

Using the ideal gas law you can determine the number of moles of Cl_2.

$$n = \frac{1\text{ atm} \times 1.200\text{ L}}{0.0821\ \dfrac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \times 473\text{ K}} = 0.0309\text{ mol } Cl_2$$

Now you can use a conversion factor (from the equivalences above) to change moles of Cl_2 to grams of MnO_2.

$$0.0309\text{ mol } Cl_2 \times \frac{86.94\text{ g } MnO_2}{1\text{ mol } Cl_2} = 2.69\text{ g } MnO_2$$

2. When sodium metal reacts with water, sodium ions, hydroxide ions, and hydrogen gas are produced.

a. Write a balanced equation for the reaction.

b. How many liters of hydrogen gas will be collected at 25.00°C and 1.00 atm pressure if 2.15 g of sodium are used? **(E6)**

B. Volumes measured when P and T remain constant

Law of Combining Volumes: ____

____

This law can be applied to the reaction

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

in the following way:

$$2 \text{ mol } H_2 = 1 \text{ mol } O_2 = 2 \text{ mol } H_2O = 4.04 \text{ g } H_2 = 32.0 \text{ g } O_2 = 36.0 \text{ g } H_2O$$

or NOT *and*

$$2 \text{ L } H_2 = 1 \text{ L } O_2 = 2 \text{ L } H_2O$$

You may NOT write 36.0 g H_2O = 2 L H_2O. The volume equivalences may be written only if the volumes are measured *at the same temperature and pressure.* The temperature and pressure do not need to be specified, but the problem has to state that the conditions remain the same.

Exercises

1. Consider the reaction

$$2\,NO\,(g) + O_2\,(g) \rightarrow 2\,NO_2\,(g)$$

What volume of NO_2 will be produced by 1.38 L of O_2 when both volumes are measured at the same temperature and pressure?

Solution:

$$2 \text{ L NO} = 1 \text{ L } O_2 = 2 \text{ L } NO_2$$

$$1.38 \text{ L } O_2 \times \frac{2 \text{ L } NO_2}{1 \text{ L } O_2} = 2.76 \text{ L } NO_2$$

2. Consider the reaction

$$2\ Bi_2S_3\ (s) + 9\ O_2\ (g) \rightarrow 2\ Bi_2O_3\ (s) + 6\ SO_2\ (g)$$

The reaction is carried out and kept at 1.00 atm pressure and 175°C throughout. How many liters of O_2 are needed to produce 3.87 L of SO_2? How many grams of Bi_2S_3 are needed? **(E7)**

IV. Gas Mixtures

A. Dalton's Law of Partial Pressures

1. Statement: ____

____

2. Mathematical relationship

$$P_{total} = P_A + P_B + \ldots$$

3. Exercises:

a. A 200.0-mL sample of O_2 is collected over water at 27°C and 748 mm Hg. The vapor pressure of water at 27°C is 28 mm Hg. What mass of O_2 is in the sample?

Solution: To arrive at the mass of O_2, we must first determine the pressure exerted by the oxygen gas alone. To do this we use Dalton's Law of Partial Pressures.

$$\begin{aligned} P_{total} &= P_A + P_B \\ &= P_{oxygen} + P_{water} \\ 748 &= P_{oxygen} + 28 \\ P_{oxygen} &= 720 \text{ mm Hg} \end{aligned}$$

Now, we can calculate the mass of O_2 by using the ideal gas law. The volume of the sample is used as the volume of oxygen.

$$n = \frac{720 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm}} \times 0.200 \text{ L}}{0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times 300 \text{ K}} = 0.00769 \text{ mol}$$

Since 1 mole of O_2 = 32.0 g,

$$0.00769 \text{ mol} \times \frac{32.0 \text{ g}}{1 \text{ mol}} = 0.246 \text{ g } O_2$$

b. Consider the reaction

$$2\,Al\,(s) + 6\,H^+\,(aq) \rightarrow 2\,Al^{3+}\,(aq) + 3\,H_2\,(g)$$

The hydrogen gas is collected over water at 30.0°C and the total pressure is 740.0 mm Hg (VP_{H_2O} = 32 mm Hg).

(1) What is the partial pressure of the hydrogen gas?

(2) How many grams of H_2 were produced if 2.00 L of wet gas were collected?

(3) How much aluminum metal was used to produce the hydrogen gas? **(E8)**

B. Mole fraction and partial pressure

1. Mole fraction

For gas A, it is the fraction of the total number of moles that is accounted for by gas A.

$$X_A = \text{mole fraction of gas A} = \frac{n_A}{n_{total}}$$

$$n_A = \text{number of moles of gas A}$$

$$n_{total} = \text{number of moles of all gases combined}$$

2. Mathematical relationship between the mole fraction and partial pressure

$$P_A = X_A P_{total}$$

3. Exercises

a. What pressure does a mixture made up of 4.00 g of NO_2 gas and 2.00 g of oxygen gas exert on a 2.00-L container at 27.00°C? What is the partial pressure of the oxygen?

Solution: To obtain the total pressure, we use the ideal gas law, taking the total number of moles as n. Thus we can write

$$P_{total} V = n_{total} RT$$

We calculate n_{total} by calculating n_{NO_2} and n_{O_2}

$$n_{O_2} = 2.00 \text{ g} \times \frac{1 \text{ mol}}{32.0 \text{ g}} = 0.0625 \text{ mol } O_2$$

$$n_{NO_2} = 4.00 \text{ g} \times \frac{1 \text{ mol}}{46.0 \text{ g}} = 0.0870 \text{ mol } NO_2$$

Adding we get

$$n_{total} = n_{O_2} + n_{NO_2} = 0.0625 \text{ mol } O_2 + 0.0870 \text{ mol } NO_2 = 0.1495 \text{ mol}$$

Now we substitute n_{total} into the ideal gas law to obtain P_{total}.

$$P_{total} = \frac{0.1495 \text{ mol} \times 0.0821 \dfrac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times 300.15 \text{ K}}{2.00 \text{ L}} = 1.84 \text{ atm}$$

Using the relationship

$$P_A = X_A P_{total}$$

where

$$X_A = \frac{n_A}{n_{total}}$$

and $A = O_2$, we get

$$X_{O_2} = \frac{0.0625}{0.0625 + 0.0870} = 0.418$$

and

$$P_{O_2} = 0.418 \times 1.84 = 0.769 \text{ atm}$$

b. Nitrogen gas with a volume of 200.0 mL, a pressure of 750.0 mm Hg, and a temperature of 27.00°C is mixed with oxygen gas and transferred to a 750.0-mL container at 27.00°C. The total pressure of the gas is found to be 680.0 mm Hg, at a temperature of 27.00°C.

(1) Calculate $n_{nitrogen}$.

(2) Calculate n_{total}.

(3) Calculate the partial pressure of each gas. **(E9)**

V. Kinetic Theory of Gases

A. The kinetic theory of gases is based on the idea that ______________________

____.

B. Assumptions of the molecular model

1. ____

____

2. ____

____

3. ____

____

4. ____

____

C. Expression for Pressure

Mathematical relationship

$$P = \frac{Nmu^2}{3V}$$

where the ratio N/V expresses ______________________________

and the product mu^2 is a measure of ______________________________.

D. Average Kinetic Energy of Translational Motion

1. Symbol: E_t

2. Mathematical expression for E_t

$$E_t = \frac{3RT}{2N_A}$$

3. This means that

a. At a given temperature, molecules of different gases must have __________

____.

b. E_t is directly proportional to ______________________________

E. Average speed of gas particles

1. Symbol used for average speed: ______________________________
2. Equation relating average speed to temperature for any gas.

$$u = \left(\frac{3RT}{MM}\right)^{\frac{1}{2}}$$

3. The average speed of a gas particle is

a. directly proportional to ______________________________

____

Equation:

b. inversely proportional to ______________________________

____

Equation:

4. Exercises

a. Calculate the average speed of a neon atom at 27°C.

Solution: This is a simple plug-in problem. Just remember to use 8.31×10^3 g · m²/s² · mol · K for R instead of 0.0821 L·atm/mol·K. Make sure that your units cancel!

To solve this particular problem, we take the formula

$$u = \left(\frac{3RT}{MM}\right)^{\frac{1}{2}}$$

and substitute 300 K for T and 20.18 g/mol for MM (molar mass for neon). We get

$$u = \left(\frac{3 \times 8.31 \times 10^3 \dfrac{g \cdot m^2}{s^2 \cdot mol \cdot K} \times 300\ K}{20.18 \dfrac{g}{mol}}\right)^{\frac{1}{2}} = 609 \frac{m}{s}$$

b. At what temperature will an HCl molecule have an average speed of 555 m/s? **(E10)**

c. The average speed of a helium atom at 27°C is 1.37×10^3 m/s.

(1) What is the average speed of a gaseous chlorine molecule at 27°C?

Solution: We use the relation

$$\frac{u_{He}}{u_{Cl_2}} = \left(\frac{MM_{Cl_2}}{MM_{He}}\right)^{\frac{1}{2}}$$

Substituting, we get

$$\frac{1.37 \times 10^3 \ \frac{m}{s}}{u_{Cl_2}} = \left(\frac{70.90}{4.003}\right)^{\frac{1}{2}}$$

$$u_{Cl_2} = \frac{1.37 \times 10^3 \ \frac{m}{s}}{4.21} = 325 \ \frac{m}{s}$$

(2) What is the average speed of a gaseous helium molecule at 270.0°C?

Solution: This time, we use the relation between speed and temperature for the same molecule. We designate u_2 as the speed at 270.0°C and u_1 as the speed at 27°C.

$$\frac{u_2}{1.37 \times 10^3} = \left(\frac{543}{300}\right)^{\frac{1}{2}}$$

$$u_2 = 1.84 \times 10^3 \ \frac{m}{s}$$

d. The average speed of a gaseous HBr molecule at 100°C is 3.39×10^4 m/s.

(1) What is its speed at 25°C?

(2) At what temperature could the HBr molecule have an average speed of 7.50×10^5 m/s? **(E11)**

F. Effusion

1. Definition: ____

____

2. Graham's Law

a. Statement: ____

____

b. Equations relating the rate of effusion to different variables

(1) To speed of molecules:

(2) To molar mass of the molecules:

(3) To time:

3. Exercises

a. An unknown gas goes through a small opening in 65.00 minutes while an equal mass of hydrogen gas goes through the same opening in 9.75 minutes. Calculate the molar mass of the unknown gas.

Solution: Since the variables in this problem are time and mass, we choose the equation that relates them, namely,

$$\frac{\text{time}_A}{\text{time}_B} = \left(\frac{MM_A}{MM_B}\right)^{\frac{1}{2}}$$

If we call the unknown gas A and designate hydrogen as B, then we can write

$$\frac{65.00}{9.75} = \left(\frac{MM_A}{2.02}\right)^{\frac{1}{2}}$$

Squaring both sides we get

$$MM_A = \left(\frac{65.00}{9.75}\right)^2 \times 2.02 = 89.8\ \frac{\text{g}}{\text{mol}}$$

b. What is the molar mass of a gas that diffuses one fifth as fast as helium? **(E12)**

VI. Real Gases

A. Conditions for deviating from the ideal gas law

1. ______________________________

____

2. ______________________________

____

B. Factors that the ideal gas law neglects

1. ______________________________

____

2. ______________________________

____

SELF-TEST

A. Multiple choice:

1. A sealed, rigid flask contains nitrogen gas. The flask is cooled from room temperature to −50°C. Which of the following statements is true?

a. The number of moles of nitrogen decreases.
b. The volume of nitrogen increases.
c. The pressure of nitrogen decreases.
d. The pressure of nitrogen increases.
e. The volume of nitrogen decreases.

2. The ideal gas law predicts that

— the volume of a gas goes to zero at absolute zero temperature.
— density increases with pressure.
— density increases with temperature.
— the product, PV/T, for a fixed amount of gas is constant.

The number of true statements above is

a. 0 **b.** 1 **c.** 2 **d.** 3 **e.** 4

3. The gas constant, R, can have the units

— bar-m^3/mol-K
— bar-L/mol-K
— mm Hg-cm^3/mol-K
— mol-K/mm Hg-mL

The number of false statements above is
a. 0 **b.** 1 **c.** 2 **d.** 3 **e.** 4

4. In an ideal gas, the collisions of the molecules with the walls of the container account for
a. the velocity of the molecules.
b. the observed pressure.
c. the number of moles.
d. the observed temperature.
e. none of these.

5. For a fixed amount of gas at a fixed pressure, changing the temperature from 30°C to 60°C causes
a. the gas volume to decrease.
b. the gas volume to double.
c. the gas volume to increase but not to double.
d. the gas volume to decrease to half its original volume.
e. no change in the gas volume.

6. Attractive forces between gas molecules are most important at
a. low pressures and low temperatures.
b. low pressures and high temperatures.
c. high pressures and high temperatures.
d. high pressures and low temperatures.

B. Answer the questions below, using **LT** (for *is less than*), **GT** (for *is greater than*), **EQ** (for *is equal to*), or **MI** (for *more information required*) in the blanks provided.

______ **1.** At 100°C and 1 atm, the velocity of a molecule of hydrogen gas is (1) the velocity of a molecule of oxygen gas.

______ **2.** At 100°C, the average translational kinetic energy of a molecule of hydrogen gas is (2) the average translational kinetic energy of a molecule of oxygen gas.

______ **3.** At constant temperature and volume, the pressure exerted by ten moles of hydrogen gas is (3) the pressure exerted by ten moles of oxygen gas.

______ **4.** At the same temperature and volume, the pressure exerted by 10.00 g of hydrogen gas is __(4)__ the pressure exerted by 10.00 g of oxygen gas.

______ **5.** At 50°C, the velocity of a molecule of chlorine gas is __(5)__ the velocity of the same molecule of chlorine gas at 100°C.

______ **6.** At a constant volume, the pressure of 2.00 g of SO_2 (g) ($\mathcal{M}$ = 64.0 g/mol) at 400 K is __(6)__ the pressure of 2.00 g of O_2 (g) ($\mathcal{M}$ = 32.0 g/mol) at 200 K.

C. True or False:

Consider 2 flasks with the same volume, temperature and pressure. One flask contains nitrogen gas, the other flask has helium gas.

______ **1.** Both flasks have the same number of moles.

______ **2.** The nitrogen gas in one flask has a lower density than the helium gas in the other flask.

______ **3.** If a pinhole is created in both flasks, nitrogen gas would effuse faster than the hydrogen gas.

______ **4.** The average translational energy of the gases in both flasks is the same.

______ **5.** One mole of CO_2 gas is introduced to each flask. In the resulting mixture of gases, the mole fraction of nitrogen is equal to the mole fraction of helium.

D. Problems:

Consider ammonia gas (NH_3), which is produced by reacting nitrogen gas and hydrogen gas.

1. Write a balanced equation for the reaction.

2. A one-liter flask contains 0.20 mol of nitrogen at room temperature and 1.5 atm pressure. Another 1.0-L flask contains 0.60 mol of hydrogen at the same room temperature. What is the pressure of the hydrogen gas?

3. The same gases in Problem 2 are combined and put into a 3.0-L flask. The gas mixture exerts a total pressure of 6.0 atm. What temperature does the gas mixture have?

4. The ammonia produced by reacting 0.200 mol of nitrogen and 0.600 mol of hydrogen is measured to be 5.00 L at 27°C and one atm. How many moles of nitrogen were used up in the reaction?

5. Using the data in Problem 4, calculate the partial pressure of each gas in the reaction vessel if all the gases remain in the vessel. The temperature of the vessel is 27°C and the pressure in the vessel is 3.0 atm.

6. Calculate the density of ammonia at 27°C and 1.0 atm pressure.

7. How much faster does the hydrogen diffuse than the nitrogen?

8. Calculate the average speed of an ammonia molecule at 77°C in miles/hour (mph), if oxygen has an average speed of 482 m/s at 25°C.