Sample Exercise 10.1 Converting Pressure Units

(a) Convert 0.357 atm to torr. **(b)** Convert 6.6×10^{-2} torr to atmospheres. **(c)** Convert 147.2 kPa to torr.

Solution

Analyze In each case we are given the pressure in one unit and asked to convert it to another unit. Our task, therefore, is to choose the appropriate conversion factors.

Plan We can use dimensional analysis to perform the desired conversions.

Solve

(a) To convert atmospheres to torr, we use the relationship 760 torr = 1 atm:

$$(0.357\ \cancel{\text{atm}})\left(\frac{760\ \text{torr}}{1\ \cancel{\text{atm}}}\right) = 271\ \text{torr}$$

Note that the units cancel in the required manner.

(b) We use the same relationship as in part (a). To get the appropriate units to cancel, we must use the conversion factor as follows:

$$(6.6 \times 10^{-2}\ \cancel{\text{torr}})\left(\frac{1\ \text{atm}}{760\ \cancel{\text{torr}}}\right) = 8.7 \times 10^{-5}\ \text{atm}$$

(c) The relationship 760 torr = 101.325 kPa allows us to write an appropriate conversion factor for this problem: 760 torr = 101.325

$$(147.2\ \cancel{\text{kPa}})\left(\frac{760\ \text{torr}}{101.325\ \cancel{\text{kPa}}}\right) = 1104\ \text{torr}$$

Sample Exercise 10.1 Converting Pressure Units

Continued

Check In each case, compare the magnitude of the answer with the starting value. The torr is a much smaller unit than the atmosphere (since there are 760 torr in 1 atm), so we expect the *numerical* answer to be larger than the starting quantity in **(a)** and smaller in **(b)**. In **(c)** notice that there are nearly 8 torr per kPa, so the numerical answer in torr should be about eight times larger than its value in kPa, consistent with our calculation.

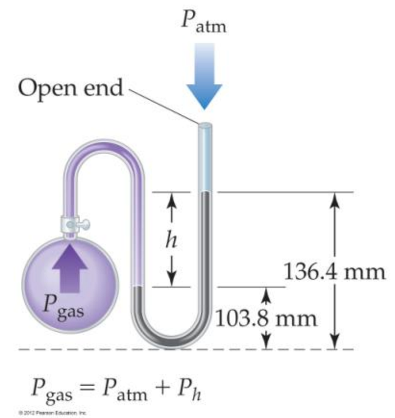
Practice Exercise

(a) In countries that use the metric system, atmospheric pressure in weather reports is given in kilopascals. Convert a pressure of 745 torr to kilopascals. **(b)** The pressure at the center of Hurricane Katrina was 902 mbar (millibars). There are 1000 mbar in 1 bar; convert this pressure to atmospheres.

Answers: **(a)** 99.3 kPa, **(b)** 0.890 atm

Sample Exercise 10.2 Using a Manometer to Measure Gas Pressure

On a certain day a laboratory barometer indicates that the atmospheric pressure is 764.7 torr. A sample of gas is placed in a flask attached to an open-end mercury manometer (**FIGURE 10.3**), and a meter stick is used to measure the height of the mercury in the two arms of the U tube. The height of the mercury in the open-end arm is 136.4 mm, and the height in the arm in contact with the gas in the flask is 103.8 mm. What is the pressure of the gas in the flask **(a)** in atmospheres, **(b)** in kilopascals?



Solution

Analyze We are given the atmospheric pressure (764.7 torr) and the mercury heights in the two arms of the manometer and asked to determine the gas pressure in the flask. Recall that millimeters of mercury is a pressure unit. We know that the gas pressure from the flask must be greater than atmospheric pressure because the mercury level in the arm on the flask side (103.8 mm) is lower than the level in the arm open to the atmosphere (136.4 mm). Therefore, the gas from the flask is pushing mercury from the arm in contact with the flask into the arm open to the atmosphere.

Plan We will use the difference in height between the two arms (*h* in Figure 10.3) to obtain the amount by which the pressure of the gas exceeds atmospheric pressure. Because an open-end mercury manometer is used, the height difference directly measures the pressure difference in mm Hg or torr between the gas and the atmosphere.

Sample Exercise 10.2 Using a Manometer to Measure Gas Pressure

Continued

Solve

(a) The pressure of the gas equals the atmospheric pressure plus *h*:

$$P_{gas} = P_{atm} + h$$
$$= 764.7 \text{ torr} + (136.4 \text{ torr} - 103.8 \text{ torr})$$
$$= 797.3 \text{ torr}$$

We convert the pressure of the gas to atmospheres:

$$P_{gas} = (797.3 \text{ torr})\left(\frac{1 \text{ atm}}{760 \text{ torr}}\right) = 1.049 \text{ atm}$$

(b) To calculate the pressure in kPa, we employ the conversion factor Between atmospheres and kPa:

$$1.049 \text{ atm}\left(\frac{101.3 \text{ kPa}}{1 \text{ atm}}\right) = 106.3 \text{ kPa}$$

Check The calculated pressure is a bit more than 1 atm, which is about 101 kPa. This makes sense because we anticipated that the pressure in the flask would be greater than the atmospheric pressure (764.7 torr = 1.01 atm) acting on the manometer.

Practice Exercise

Convert a pressure of 0.975 atm into Pa and kPa.

Answers: 9.88×10^4 Pa and 98.8 kPa

Sample Exercise 10.3 Evaluating the Effects of Changes in *P, V, n,* and *T* on a Gas

Suppose we have a gas confined to a cylinder with a movable piston. (Sections 5.2, 5.3) Consider the following changes (assuming no leaks): **(a)** Heat the gas from 298 K to 360 K at constant pressure. **(b)** Reduce the volume from 1 L to 0.5 L at constant temperature. **(c)** Inject additional gas, keeping temperature and volume constant. Indicate how each change affects the average distance between molecules, the pressure of the gas, and the number of moles of gas in the cylinder.

Solution

Analyze We need to think how each change affects (1) the distance between molecules, (2) the pressure of the gas, and (3) the number of moles of gas in the cylinder.

Plan We will use the gas laws and the general properties of gases to analyze each situation.

Solve

(a) Heating the gas while maintaining constant pressure will cause the piston to move and the volume to increase (Charles's law). Thus, the distance between molecules will increase. At constant pressure, obviously, the pressure will not change. The total number of moles of gas remains the same.

(b) Compressing the gas into a smaller volume does not change the total number of gas molecules; thus, the total number of moles remains the same. The average distance between molecules, however, must decrease because of the smaller volume. The reduction in volume causes the pressure to increase (Boyle's law).

Sample Exercise 10.3 Evaluating the Effects of Changes in *P, V, n,* and *T* on a Gas

Continued

(c) Injecting more gas into the cylinder while keeping the volume and temperature constant results in more molecules and, thus, an increase in the number of moles of gas in the cylinder. The average distance between molecules must decrease because their number per unit volume increases. Avogadro's law tells us that the volume of the cylinder should have increased when we added more gas, but here the volume is fixed. Boyle's law comes to our aid: If the volume is low, then pressure is high. Therefore, we expect that the pressure will increase in the cylinder if we inject more gas, keeping volume and temperature constant.

Practice Exercise

Recall that density is mass per volume. (Section 1.4) What happens to the density of a gas as **(a)** the gas is heated in a constant-volume container; **(b)** the gas is compressed at constant temperature; **(c)** additional gas is added to a constant-volume container?

Answers: **(a)** no change, **(b)** increases, **(c)** increases

Sample Exercise 10.4 Using the Ideal-Gas Equation

Calcium carbonate, $CaCO_3(s)$, the principal compound in limestone, decomposes upon heating to $CaO(s)$ and $CO_2(g)$. A sample of $CaCO_3$ is decomposed, and the carbon dioxide is collected in a 250-mL flask. After decomposition is complete, the gas has a pressure of 1.3 atm at a temperature of 31 °C. How many moles of CO_2 gas were generated?

Solution

Analyze We are given the volume (250 mL), pressure (1.3 atm), and temperature (31 °C) of a sample of CO_2 gas and asked to calculate the number of moles of CO_2 in the sample.

Plan Because we are given *V*, *P*, and *T*, we can solve the ideal-gas equation for the unknown quantity, *n*.

Solve

In analyzing and solving gas law problems, it is helpful to tabulate the information given in the problems and then to convert the values to units that are consistent with those for R (0.08206 L-atm/mol-K). In this case the given values are

$$V = 250 \text{ mL} = 0.250 \text{ L}$$
$$P = 1.3 \text{ atm}$$
$$T = 31 \text{ °C} = (31 + 273) \text{ K} = 304 \text{ K}$$

Remember: *Absolute temperature must always be used when the ideal-gas equation is solved.*

We now rearrange the ideal-gas equation (Equation 10.5) to solve for *n*

$$PV = nRT \qquad [10.5]$$

$$n = \frac{PV}{RT}$$

$$n = \frac{(1.3 \text{ atm})(0.250 \text{ L})}{(0.08206 \text{ L-atm/mol-K})(304 \text{ K})} = 0.013 \text{ mol } CO_2$$

Sample Exercise 10.4 Using the Ideal-Gas Equation

Continued

Check Appropriate units cancel, thus ensuring that we have properly rearranged the ideal-gas equation and have converted to the correct units.

Practice Exercise

Tennis balls are usually filled with either air or N_2 gas to a pressure above atmospheric pressure to increase their bounce. If a tennis ball has a volume of 144 cm^3 and contains 0.33 g of N_2 gas, what is the pressure inside the ball at 24 °C?

Answer: 2.0 atm

Sample Exercise 10.5 Calculating the Effect of Temperature Changes on Pressure

The gas pressure in an aerosol can is 1.5 atm at 25 °C. Assuming that the gas obeys the ideal-gas equation, what is the pressure when the can is heated to 450 °C?

Solution

Analyze We are given the initial pressure (1.5 atm) and temperature (25 °C) of the gas and asked for the pressure at a higher temperature (450 °C).

Plan The volume and number of moles of gas do not change, so we must use a relationship connecting pressure and temperature. Converting temperature to the Kelvin scale and tabulating the given information, we have

	P	**T**
INITIAL	1.5 atm	298 K
FINAL	P_2	723 K

Solve To determine how *P* and *T* are related, we start with the ideal-gas equation and isolate the quantities that do not change (*n*, *V*, and *R*) on one side and the Variables (*P* and *T*) on the other side.

$$\frac{P}{T} = \frac{nR}{V} = \text{constant}$$

Because the quotient *P*/*T* is a constant, we can write

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

(where the subscripts 1 and 2 represent the initial and final states, respectively). Rearranging to solve for P_2 and substituting the given data give

$$P_2 = (1.5 \text{ atm})\left(\frac{723 \text{ K}}{298 \text{ K}}\right) = 3.6 \text{ atm}$$

Sample Exercise 10.5 Calculating the Effect of Temperature Changes on Pressure

Continued

Check This answer is intuitively reasonable—increasing the temperature of a gas increases its pressure.

Comment It is evident from this example why aerosol cans carry a warning not to incinerate.

Practice Exercise

The pressure in a natural-gas tank is maintained at 2.20 atm. On a day when the temperature is –15 °C, the volume of gas in the tank is 3.25×10^3 m^3. What is the volume of the same quantity of gas on a day when the temperature is 31 °C?

Answer: 3.83×10^3 m^3

Sample Exercise 10.6 Calculating the Effect of Changing *P* and *T* on Gas Volume

An inflated balloon has a volume of 6.0 L at sea level (1.0 atm) and is allowed to ascend until the pressure is 0.45 atm. During ascent, the temperature of the gas falls from 22 °C to –21 °C. Calculate the volume of the balloon at its final altitude.

Solution

Analyze We need to determine a new volume for a gas sample when both pressure and temperature change.

Plan Let's again proceed by converting temperatures to kelvins and tabulating our information.

	P	*V*	*T*
INITIAL	1.0 atm	6.0 L	295 K
FINAL	0.45 atm	V_2	252 K

Because *n* is constant, we can use Equation 10.8.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \qquad [10.8]$$

Solve Rearranging Equation 10.8 to solve for V_2 gives

$$V_2 = V_1 \times \frac{P_1}{P_2} \times \frac{T_2}{T_1} = (6.0\ \mathrm{L})\left(\frac{1.0\ \mathrm{atm}}{0.45\ \mathrm{atm}}\right)\left(\frac{252\ \mathrm{K}}{295\ \mathrm{K}}\right) = 11\ \mathrm{L}$$

Sample Exercise 10.6 Calculating the Effect of Changing *P* and *T* on Gas Volume

Continued

Check The result appears reasonable. Notice that the calculation involves multiplying the initial volume by a ratio of pressures and a ratio of temperatures. Intuitively, we expect decreasing pressure to cause the volume to increase. Similarly, decreasing temperature should cause the volume to decrease. Because the pressure difference is more dramatic than the temperature difference, we expect the effect of the pressure change to predominate in determining the final volume, as it does.

Practice Exercise

A 0.50-mol sample of oxygen gas is confined at 0 °C and 1.0 atm in a cylinder with a movable piston. The piston compresses the gas so that the final volume is half the initial volume and the final pressure is 2.2 atm. What is the final temperature of the gas in degrees Celsius?

Answer: 27 °C

Sample Exercise 10.7 Calculating Gas Density

What is the density of carbon tetrachloride vapor at 714 torr and 125 °C?

Solution

Analyze We are asked to calculate the density of a gas given its name, its pressure, and its temperature. From the name we can write the chemical formula of the substance and determine its molar mass.

Plan We can use Equation 10.10 to calculate the density. Before we can do that, however, we must convert the given quantities to the appropriate units, degrees Celsius to kelvins and pressure to atmospheres. We must also calculate the molar mass of CCl_4.

$$d = \frac{n\mathcal{M}}{V} = \frac{P\mathcal{M}}{RT} \qquad [10.10]$$

Solve The absolute temperature is 125 + 273 = 398 K. The pressure is (714 torr) (1 atm/760 torr) = 0.939. The molar mass of CCl_4 is 12.01 + (4) (35.45) = 153.8 g/mol. Therefore,

$$d = \frac{(0.939 \text{ atm})(153.8 \text{ g/mol})}{(0.08206 \text{ L-atm/mol-K})(398 \text{ K})} = 4.42 \text{ g/L}$$

Check If we divide molar mass (g/mol) by density (g/L), we end up with L/mol. The numerical value is roughly 154/4.4 = 35 . That is in the right ballpark for the molar volume of a gas heated to 125 °C at near atmospheric pressure, so our answer is reasonable.

Sample Exercise 10.7 Calculating Gas Density

Continued

Practice Exercise

The mean molar mass of the atmosphere at the surface of Titan, Saturn's largest moon, is 28.6 g/mo. The surface temperature is 95 K, and the pressure is 1.6 atm. Assuming ideal behavior, calculate the density of Titan's atmosphere.

Answer: 5.9 g/L

Sample Exercise 10.8 Calculating the Molar Mass of a Gas

A large evacuated flask initially has a mass of 134.567 g. When the flask is filled with a gas of unknown molar mass to a pressure of 735 torr at 31 °C, its mass is 137.456 g. When the flask is evacuated again and then filled with water at 31 °C, its mass is 1067.9 g. (The density of water at this temperature is 0.997 g/mL.) Assuming the ideal-gas equation applies, calculate the molar mass of the gas.

Solution

Analyze We are given the temperature (31 °C) and pressure (735 torr) for a gas, together with information to determine its volume and mass, and we are asked to calculate its molar mass.

Plan We need to use the mass information given to calculate the volume of the container and the mass of the gas in it. From this we calculate the gas density and then apply Equation 10.11 to calculate the molar mass of the gas.

$$\mathcal{M} = \frac{dRT}{P} \qquad [10.11]$$

Solve The gas mass is the difference between the mass of the flask filled with gas and the mass of the evacuated flask:

$$137.456\text{ g} - 134.567\text{ g} = 2.889\text{ g}$$

The gas volume equals the volume of water the flask can hold, calculated from the mass and density of the water. The mass of the water is the difference between the masses of the full and evacuated flask:

$$1067.9\text{ g} - 134.567\text{ g} = 933.3\text{ g}$$

Rearranging the equation for density ($d = m/V$), we have

$$V = \frac{m}{d} = \frac{(933.3\text{ g})}{(0.997\text{ g/mL})} = 936\text{ mL} = 0.936\text{ L}$$

Sample Exercise 10.8 Calculating the Molar Mass of a Gas

Continued

Knowing the mass of the gas (2.889 g) and its volume (0.936 L), we can calculate the density of the gas:

2.889 g/0.936 L = 3.09 g/L

After converting pressure to atmospheres and temperature to kelvins, we can use Equation 10.11 to calculate the molar mass:

$$\mathcal{M} = \frac{dRT}{P} \qquad [10.11]$$

$$\begin{aligned}\mathcal{M} &= \frac{dRT}{P} \\ &= \frac{(3.09\ \text{g/L})(0.08206\ \text{L-atm/mol-K})(304\ \text{K})}{(0.09671)\ \text{atm}} \\ &= 79.7\ \text{g/mol}\end{aligned}$$

Check The units work out appropriately, and the value of molar mass obtained is reasonable for a substance that is gaseous near room temperature.

Practice Exercise

Calculate the average molar mass of dry air if it has a density of 1.17 g/L at 21 °C and 740.0 torr.

Answer: 29.0 g/mol

Sample Exercise 10.9 Relating a Gas Volume to the Amount of Another Substance in a Reaction

Automobile air bags are inflated by nitrogen gas generated by the rapid decomposition of sodium azide, NaN_3:

$$2\ NaN_3(s) \rightarrow 2\ Na(s) + 3\ N_2(g)$$

If an air bag has a volume of 36 L and is to be filled with nitrogen gas at 1.15 atm and 26.0 °C, how many grams of NaN_3 must be decomposed?

Solution

Analyze This is a multistep problem. We are given the volume, pressure, and temperature of the N_2 gas and the chemical equation for the reaction by which the N_2 is generated. We must use this information to calculate the number of grams of NaN_3 needed to obtain the necessary N_2.

Plan We need to use the gas data (*P*, *V*, and *T*) and the ideal-gas equation to calculate the number of moles of N_2 gas that should be formed for the air bag to operate correctly. We can then use the balanced equation to determine the number of moles of NaN_3 needed. Finally, we can convert moles of NaN_3 to grams.

Gas data → mol N_2 → mol NaN_3 → g NaN_3

Sample Exercise 10.9 Relating a Gas Volume to the Amount of Another Substance in a Reaction

Continued

Solve The number of moles of N_2 is determined using the ideal-gas equation:

$$n = \frac{PV}{RT} = \frac{(1.15 \text{ atm})(36 \text{ L})}{(0.08206 \text{ L-atm/mol-K})(299 \text{ K})} = 1.69 \text{ mol } N_2$$

We use the coefficients in the balanced equation to calculate the number of moles of NaN_3:

$$(1.69 \text{ mol } N_2)\left(\frac{2 \text{ mol } NaN_3}{3 \text{ mol } N_2}\right) = 1.12 \text{ mol } NaN_3$$

Finally, using the molar mass of NaN_3, we convert moles of NaN_3 to grams:

$$(1.12 \text{ mol } NaN_3)\left(\frac{65.0 \text{ g } NaN_3}{1 \text{ mol } NaN_3}\right) = 73 \text{ g } NaN_3$$

Check The units cancel properly at each step in the calculation, leaving us with the correct units in the answer, g NaN_3.

Practice Exercise

In the first step in the industrial process for making nitric acid, ammonia reacts with oxygen in the presence of a suitable catalyst to form nitric oxide and water vapor:

$$4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g)$$

How many liters of $NH_3(g)$ at 850 °C and 5.00 atm are required to react with 1.00 mol of $O_2(g)$ in this reaction?

Answer: 14.8 L

Sample Exercise 10.10 Applying Dalton's Law of Partial Pressures

A mixture of 6.00 g $O_2(g)$ and 9.00 g $CH_4(g)$ is placed in a 15.0-L vessel at 0 °C. What is the partial pressure of each gas, and what is the total pressure in the vessel?

Solution

Analyze We need to calculate the pressure for two gases in the same volume and at the same temperature.

Plan Because each gas behaves independently, we can use the ideal gas equation to calculate the pressure each would exert if the other were not present. The total pressure is the sum of these two partial pressures.

Solve We first convert the mass of each gas to moles:

$$n_{O_2} = (6.00 \text{ g } O_2)\left(\frac{1 \text{ mol } O_2}{32.0 \text{ g } O_2}\right) = 0.188 \text{ mol } O_2$$

$$n_{CH_4} = (9.00 \text{ g } CH_4)\left(\frac{1 \text{ mol } CH_4}{16.0 \text{ g } CH_4}\right) = 0.563 \text{ mol } CH_4$$

We use the ideal-gas equation to calculate the partial Pressure of each gas:

$$P_{O_2} = \frac{n_{O_2}RT}{V} = \frac{(0.188 \text{ mol})(0.08206 \text{ L-atm/mol-K})(273\text{K})}{15.0\text{L}} = 0.281 \text{ atm}$$

$$P_{CH_4} = \frac{n_{CH_2}RT}{V} = \frac{(0.563 \text{ mol})(0.08206 \text{ L-atm/mol-K})(273 \text{ K})}{15.0 \text{ L}} = 0.841 \text{ atm}$$

According to Dalton's law of partial pressures (Equation 10.12), the total pressure in the vessel is the sum of the partial pressures:

$$P_t = P_1 + P_2 + P_3 + \cdots \qquad [10.12]$$

$$Pt = P_{O_2} + P_{CH_4} = 0.281 \text{ atm} + 0.841 \text{ atm} = 1.122 \text{ atm}$$

Sample Exercise 10.10 Applying Dalton's Law of Partial Pressures

Continued

Check A pressure of roughly 1 atm seems right for a mixture of about 0.2 mol O_2 and a bit more than 0.5 mol CH_4, together in a 15-L volume, because 1 mol of an ideal gas at 1 atm pressure and 0 °C occupies about 22 L.

Practice Exercise

What is the total pressure exerted by a mixture of 2.00 g of $H_2(g)$ and 8.00 g of $N_2(g)$ at 273 K in a 10.0-L vessel?

Answer: 2.86 atm

Sample Exercise 10.11 Relating Mole Fractions and Partial Pressures

A study of the effects of certain gases on plant growth requires a synthetic atmosphere composed of 1.5 mol percent CO_2, 18.0 mol percent O_2, and 80.5 mol percent Ar. **(a)** Calculate the partial pressure of O_2 in the mixture if the total pressure of the atmosphere is to be 745 torr. **(b)** If this atmosphere is to be held in a 121-L space at 295 K, how many moles of O_2 are needed?

Solution

Analyze For (a) we need to calculate the partial pressure of O_2 given its mole percent and the total pressure of the mixture. For (b) we need to calculate the number of moles of O_2 in the mixture given its volume (121 L), temperature (745 torr), and partial pressure from part (a).

Plan We calculate the partial pressures using Equation 10.16, and then use P_{O_2}, V, and T in the ideal-gas equation to calculate the number of moles of O_2.

$$P_1 = \left(\frac{n_1}{n_t}\right)P_t = X_1 P_t \qquad [10.16]$$

Solve

(**a**) The mole percent is the mole fraction times 100. Therefore, the mole fraction of O_2 is 0.180. Equation 10.16 gives

$$P_{O_2} = (0.180)(745 \text{ torr}) = 134 \text{ torr}$$

Sample Exercise 10.11 Relating Mole Fractions and Partial Pressures

Continued

(**b**) Tabulating the given variables and converting To appropriate units, we have

$$P_{O_2} = (134\ \text{torr})\left(\frac{1\ \text{atm}}{760\ \text{torr}}\right) = 0.176\ \text{atm}$$

$$V = 121\ \text{L}$$

$$n_{O_2} = ?$$

$$R = 0.08206\ \frac{\text{L-atm}}{\text{mol-K}}$$

$$T = 295\ \text{K}$$

Solving the ideal-gas equation for n_{O_2}, we have

$$n_{O_2} = P_{O_2}\left(\frac{V}{RT}\right)$$

$$= (0.176\ \text{atm})\frac{121\ \text{L}}{(0.08206\ \text{L-atm/mol-K})(295\ \text{K})} = 0.879\ \text{mol}$$

Check The units check out, and the answer seems to be the right order of magnitude.

Practice Exercise

From data gathered by *Voyager 1*, scientists have estimated the composition of the atmosphere of Titan, Saturn's largest moon. The pressure on the surface of Titan is 1220 torr. The atmosphere consists of 82 mol percent N_2, 12 mol percent Ar, and 6.0 mol percent CH_4. Calculate the partial pressure of each gas.

Answer: 1.0×10^3 torr N_2, 1.5×10^2 torr Ar, and 73 torr CH_4

Sample Exercise 10.12 Calculating the Amount of Gas Collected over Water

When a sample of $KClO_3$ is partially decomposed in the setup shown in Figure 10.15, the volume of gas collected is 0.250 L at 26 °C and 765 torr total pressure. **(a)** How many moles of O_2 are collected? **(b)** How many grams of $KClO_3$ were decomposed?

Solution

(a) Analyze We need to calculate the number of moles of O_2 gas in a container that also contains water vapor.

Plan We are given values for V and T. To use the ideal-gas equation to calculate the unknown, n_{O_2}, we must know the partial pressure of O_2 in the system. We can calculate this partial pressure from the total pressure (765 torr) and the vapor pressure of water.

Solve The partial pressure of the O_2 gas is the difference between the total pressure and the pressure of the water vapor at 26 °C, 25 torr (Appendix B):

$$P_{O_2} = 765 \text{ torr} - 25 \text{ torr} = 740 \text{ torr}$$

We use the ideal-gas equation to calculate the number of moles of O_2:

$$n_{O_2} = \frac{P_{O_2}V}{RT} = \frac{(740 \text{ torr})(1 \text{ atm}/760 \text{ torr})(0.250 \text{ L})}{(0.08206 \text{ L-atm/mol-K})(299 \text{ K})} = 9.92 \times 10^{-3} \text{ mol } O_2$$

Sample Exercise 10.12 Calculating the Amount of Gas Collected over Water

Continued

(b) Analyze We need to calculate the number of moles of reactant $KClO_3$ decomposed.

Plan We can use the number of moles of O_2 formed and the balanced chemical equation to determine the number of moles of $KClO_3$ decomposed, which we can then convert to grams of $KClO_3$.

Solve From Equation 10.17, we have 2 mol $KCl_3 \simeq 3 \text{ mol } O_2$.

$$2\,KClO_3(s) \longrightarrow 2\,KCl(s) + 3\,O_2(g) \qquad [10.17]$$

The molar mass of $KClO_3$ is 122.6 g/mol. Thus, we can convert the number of moles of O_2 from part (a) to moles of $KClO_3$ and then to grams of $KClO_3$:

$$(9.92 \times 10^{-3} \text{ mol } O_2)\left(\frac{2 \text{ mol } KClO_3}{3 \text{ mol } O_2}\right)\left(\frac{122.6 \text{ g } KClO_3}{1 \text{ mol } KClO_3}\right) = 0.811 \text{ g } KClO_3$$

Check The units cancel appropriately in the calculations. The numbers of moles of O_2 and $KClO_3$ seem reasonable, given the small volume of gas collected.

Sample Exercise 10.12 Calculating the Amount of Gas Collected over Water

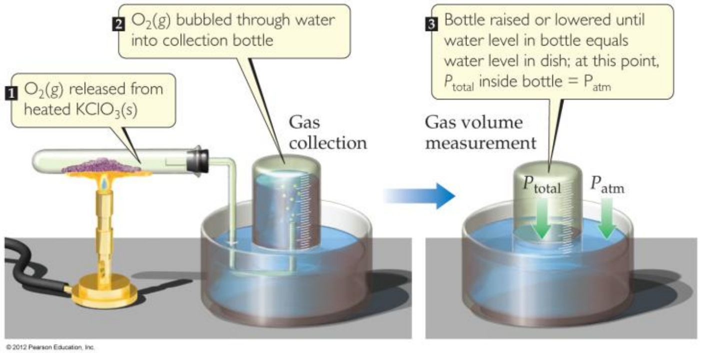
Continued

Practice Exercise

Ammonium nitrite, NH_4NO_2, decomposes on heating to form N_2 gas:

$$NH_4NO_2(s) \rightarrow N_2(g) + 2\ H_2O(l)$$

When a sample of NH_4NO_2 is decomposed in the apparatus of Figure 10.15, 511 mL of N_2 gas is collected over water at 26 °C and 745 torr total pressure. How many grams of NH_4NO_2 were decomposed?



Answer: 1.26 g

Sample Exercise 10.13 Applying the Kinetic-Molecular Theory

A sample of O_2 gas initially at STP is compressed to a smaller volume at constant temperature. What effect does this change have on **(a)** the average kinetic energy of the molecules, **(b)** their average speed, **(c)** the number of collisions they make with the container walls per unit time, **(d)** the number of collisions they make with a unit area of container wall per unit time?

Solution

Analyze We need to apply the concepts of the kinetic-molecular theory of gases to a gas compressed at constant temperature.

Plan We will determine how each of the quantities in (a)–(d) is affected by the change in volume at constant temperature.

Solve **(a)** Because the average kinetic energy of the O_2 molecules is determined only by temperature, this energy is unchanged by the compression. **(b)** Because the average kinetic energy of the molecules does not change, their average speed remains constant. **(c)** The number of collisions with the walls per unit time increases because the molecules are moving in a smaller volume but with the same average speed as before. Under these conditions they must encounter a wall more frequently. **(d)** The number of collisions with a unit area of wall per unit time increases because the total number of collisions with the walls per unit time increases and the area of the walls decreases.

Check In a conceptual exercise of this kind, there is no numerical answer to check. All we can check in such cases is our reasoning in the course of solving the problem.

Sample Exercise 10.13 Applying the Kinetic-Molecular Theory

Continued

Practice Exercise

How is the rms speed of N_2 molecules in a gas sample changed by **(a)** an increase in temperature, **(b)** an increase in volume, **(c)** mixing with a sample of Ar at the same temperature?

Answers: **(a)** increases, **(b)** no effect, **(c)** no effect

Sample Exercise 10.14 Calculating a Root-Mean-Square Speed

Calculate the rms speed of the molecules in a sample of N_2 gas at 25 °C.

Solution

Analyze We are given the identity of a gas and the temperature, the two quantities we need to calculate the rms speed.

Plan We calculate the rms speed using Equation 10.22.

$$u_{rms} = \sqrt{\frac{3RT}{\mathcal{M}}} \qquad [10.22]$$

Solve We must convert each quantity in our equation to SI units. We will also use R in units of J/mol-K (Table 10.2) to make the units cancel correctly.

TABLE 10.2 • Numerical Values of the Gas Constant R in Various Units	
Units	**Numerical Value**
L-atm/mol-K	0.08206
J/mol-K*	8.314
cal/mol-K	1.987
m^3-Pa/mol-K*	8.314
L-torr/mol-K	62.36

*SI unit

$$T = 25 + 273 = 298 \text{ K}$$

$$\mathcal{M} = 28.0 \text{ g/mol} = 28.0 \times 10^{-3} \text{ kg/mol}$$

$$R = 8.314 \text{ J/mol-K} = 8.314 \text{ kg-m}^2/\text{s}^2\text{-mol-K} \qquad (\text{Since } 1 \text{ J} = 1 \text{ kg-m}^2/\text{s}^2)$$

$$u_{rms} = \sqrt{\frac{3(8.314 \text{ kg-m}^2/\text{s}^2\text{-mol-K})(298 \text{ K})}{28.0 \times 10^{-3} \text{ kg/mol}}} = 5.15 \times 10^2 \text{ m/s}$$

Comment This corresponds to a speed of 1150 mi/hr. Because the average molecular weight of air molecules is slightly greater than that of N_2, the rms speed of air molecules is a little lower than that for N_2.

Sample Exercise 10.14 Calculating a Root-Mean-Square Speed

Continued

Practice Exercise

What is the rms speed of an atom in a sample of He gas at 25 °C?

Answer: 1.36×10^3 m/s

Sample Exercise 10.15 Applying Graham's Law

An unknown gas composed of homonuclear diatomic molecules effuses at a rate that is 0.355 times the rate at which O_2 gas effuses at the same temperature. Calculate the molar mass of the unknown and identify it.

Solution

Analyze We are given the rate of effusion of an unknown gas relative to that of O_2 and asked to find the molar mass and identity of the unknown. Thus, we need to connect relative rates of effusion to relative molar masses.

Plan We use Equation 10.24, to determine the molar mass of the unknown gas.

$$\frac{r_1}{r_2} = \sqrt{\frac{\mathcal{M}_2}{\mathcal{M}_1}} \qquad [10.24]$$

If we let r_x and M_x represent the rate of effusion and molar mass of the gas, we can write

$$\frac{r_x}{r_{O_2}} = \sqrt{\frac{\mathcal{M}_{O_2}}{\mathcal{M}_x}}$$

Solve From the information given,

$$r_x = 0.355 \times r_{O_2}$$

Thus,

$$\frac{r_x}{r_{O_2}} = 0.355 = \sqrt{\frac{32.0\ \text{g/mol}}{\mathcal{M}_x}}$$

$$\frac{32.0\ \text{g/mol}}{\mathcal{M}_x} = (0.355)^2 = 0.126$$

$$\mathcal{M}_x = \frac{32.0\ \text{g/mol}}{0.126} = 254\ \text{g/mol}$$

Sample Exercise 10.15 Applying Graham's Law

Continued

Because we are told that the unknown gas is composed of homonuclear diatomic molecules, it must be an element. The molar mass must represent twice the atomic weight of the atoms in the unknown gas. We conclude that the unknown gas is I_2.

Practice Exercise

Calculate the ratio of the effusion rates of N_2 gas and O_2 gas.

Answer: $r_{N_2}/r_{O_2} = 1.07$

Sample Exercise 10.16 Using the van der Waals Equation

If 1.000 mol of an ideal gas were confined to 22.41 L at 0.0 °C, it would exert a pressure of 1.000 atm. Use the van der Waals equation and Table 10.3 to estimate the pressure exerted by 1.000 mol of $Cl_2(g)$ in 22.41 L at 0.0 °C.

TABLE 10.3 • Van der Waals Constants for Gas Molecules

Substance	a(L^2-atm/mol^2)	b (L/mol)
He	0.0341	0.02370
Ne	0.211	0.0171
Ar	1.34	0.0322
Kr	2.32	0.0398
Xe	4.19	0.0510
H_2	0.244	0.0266
N_2	1.39	0.0391
O_2	1.36	0.0318
Cl_2	6.49	0.0562
H_2O	5.46	0.0305
CH_4	2.25	0.0428
CO_2	3.59	0.0427
CCl_4	20.4	0.1383

Solution

Analyze We need to determine a pressure. Because we will use the van der Waals equation, we must identify the appropriate values for the constants that appear there.

Plan Solving Equation 10.27 for *P*, we have

$$P = \frac{nRT}{V - nb} - \frac{n^2a}{V^2}$$

$$\left(P + \frac{n^2a}{V^2}\right)(V - nb) = nRT \qquad [10.27]$$

Solve Substituting n = 1.000 mol, R = 0.08206 L-atm/mol-K, T = 273.2 K, V = 22.41 L, a = 6.49 L^2-atm/mol^2, and b = 0.0562 L/mol:

$$P = \frac{(1.000\ \text{mol})(0.08206\ \text{L-atm/mol-K})(273.2\ \text{K})}{22.41\ \text{L} - (1.000\ \text{mol})(0.0562\ \text{L/mol})} - \frac{(1.000\ \text{mol})^2(6.49\ \text{L}^2\text{-atm/mol}^2)}{(22.14\ \text{L})^2}$$

$$= 1.003\ \text{atm} - 0.013\ \text{atm} = 0.990\ \text{atm}$$

Sample Exercise 10.16 Using the van der Waals Equation

Continued

Check We expect a pressure not far from 1.000 atm, which would be the value for an ideal gas, so our answer seems very reasonable.

Comment Notice that the term 1.003 atm is the pressure corrected for molecular volume. This value is higher than the ideal value, 1.000 atm, because the volume in which the molecules are free to move is smaller than the container volume, 22.41 L. Thus, the molecules collide more frequently with the container walls. The term 0.013 atm corrects for intermolecular forces. The intermolecular attractions between molecules reduce the pressure to 0.990 atm.We conclude, therefore, that the intermolecular attractions are the main cause of the slight deviation of $Cl_2(g)$ from ideal behavior under the stated experimental conditions.

Practice Exercise

A sample of 1.000 mol of $CO_2(g)$ is confined to a 3.000-L container at 0.000 °C. Calculate the pressure of the gas using **(a)** the ideal-gas equation and **(b)** the van der Waals equation.

Answers: **(a)** 7.47 atm, **(b)** 7.18 atm