Example Exercise 11.1 Gas Pressure Conversion

Meteorologists state that a "falling" barometer indicates an approaching storm. Given a barometric pressure of 27.5 in. Hg, express the pressure in each of the following units of pressure:

(a) atm (b) mm Hg
(c) psi (d) kPa

Solution

For each conversion, we apply a unit conversion factor related to units of standard pressure.

(a) To express the pressure in atmospheres, we derive a unit factor related to the equivalent relationship 29.9 in. Hg = 1 atm.

$$27.5\ \cancel{\text{in. Hg}} \times \frac{1\ \text{atm}}{29.9\ \cancel{\text{in. Hg}}} = 0.920\ \text{atm}$$

(b) To convert to millimeters of mercury, we derive a unit factor related to the equivalent relationship 29.9 in. Hg = 760 mm Hg.

$$27.5\ \cancel{\text{in. Hg}} \times \frac{760\ \text{mm Hg}}{29.9\ \cancel{\text{in. Hg}}} = 699\ \text{mm Hg}$$

(c) To calculate the pressure in pounds per square inch, we derive a unit factor related to the equivalent relationship 29.9 in. Hg = 14.7 psi.

$$27.5\ \cancel{\text{in. Hg}} \times \frac{14.7\ \text{psi}}{29.9\ \cancel{\text{in. Hg}}} = 13.5\ \text{psi}$$

(d) To find the pressure in kilopascals, we derive a unit factor related to the equivalent relationship 29.9 in. Hg = 101 kPa.

$$27.5\ \cancel{\text{in. Hg}} \times \frac{101\ \text{kPa}}{29.9\ \cancel{\text{in. Hg}}} = 92.9\ \text{kPa}$$

Example Exercise 11.1 Gas Pressure Conversion

Continued

Since 1 atm and 760 mm Hg are exact values, the answers have been rounded to three significant digits, which is consistent with the given value, 27.5 in. Hg.

Practice Exercise

Given that the air pressure inside an automobile tire is 34.0 psi, express the pressure in each of the following units:

(a)	atm	(b)	cm Hg
(c)	torr	(d)	kPa

Answers:

(a)	2.31 atm	(b)	176 cm Hg
(c)	1760 torr	(d)	234 kPa

Concept Exercise

Which of the following expresses the greatest gas pressure: 1 atm, 1 in. Hg, 1 torr, 1 cm Hg, 1 mm Hg, or 1 psi?

Answer: See Appendix G.

Example Exercise 11.2 Gas Pressure Changes

State whether the pressure of a gas in a sealed container increases or decreases with each of the following changes:

(a) The volume changes from 250 mL to 500 mL.

(b) The temperature changes from 20 °C to –80 °C.

(c) The moles of gas change from 1.00 mol to 1.50 mol.

Solution

For each of the changes we must consider whether the number of molecular collisions increases or decreases.

(a) The volume increases, and so the number of collisions decreases; thus the *pressure decreases*.

(b) The temperature decreases, and so the number of collisions decreases; thus, the *pressure decreases*.

(c) When the moles of gas increase, the number of molecules increases. With more molecules, there are more collisions, and the *pressure increases*.

Practice Exercise

Indicate whether gas pressure increases or decreases with each of the following changes in a sealed container:

(a) increasing the temperature

(b) increasing the volume

(c) increasing the number of gas molecules

Answers: (a) pressure increases; (b) pressure decreases; (c) pressure increases

Concept Exercise

Which of the following decreases gas pressure: increasing volume, increasing temperature, or increasing the number of molecules?

Answer: See Appendix G.

Example Exercise 11.3 Boyle's Law

A 1.50 L sample of methane gas exerts a pressure of 1650 mm Hg. Calculate the final pressure if the volume changes to 7.00 L. Assume the temperature remains constant.

Conceptual Solution

We can find the final pressure (P_2) by applying Boyle's law and using the relationship

$$P_1 \times V_{\text{factor}} = P_2$$

The volume increases from 1.50 L to 7.00 L. Thus, the pressure decreases. The V_{factor} must be less than 1. Hence,

$$1650 \text{ mm Hg} \times \frac{1.50\ \cancel{L}}{7.00\ \cancel{L}} = 354 \text{ mm Hg}$$

We can visually summarize the Boyle's law solution as follows:

Algebraic Solution

Alternatively, we can solve this problem using the equation

$$P_1 V_1 = P_2 V_2$$

Solving for P_2 gives

$$\frac{P_1 V_1}{V_2} = P_2$$

Example Exercise 11.3 Boyle's Law

Continued

Substituting for each variable and simplifying, we obtain

$$\frac{1650 \text{ mm Hg} \times 1.50 \cancel{\text{L}}}{7.00 \cancel{\text{L}}} = 354 \text{ mm Hg}$$

Practice Exercise

A sample of ethane gas has a volume of 125 mL at 20 °C and 725 torr. What is the volume of the gas at 20 °C when the pressure decreases to 475 torr?

Answer: 191 mL

Concept Exercise

When air in a steel cylinder is compressed from 10 L to 5 L, and temperature remains constant, the gas pressure inside the cylinder (increases/decreases).

Answer: See Appendix G.

Example Exercise 11.4 Charles's Law

A 275 L helium balloon is heated from 20 °C to 40 °C. Calculate the final volume assuming the pressure remains constant.

Conceptual Solution

We first convert the Celsius temperatures to Kelvin by adding 273 units.

$$20\ °\text{C} + 273 = 293\ \text{K}$$
$$40\ °\text{C} + 273 = 313\ \text{K}$$

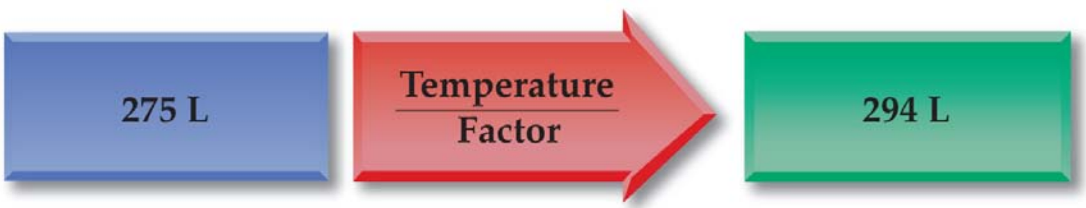
We can find the final volume, (V_2), by applying Charles's law and using the relationship

$$V_1 \times T_{\text{factor}} = V_2$$

The temperature increases from 293 K to 313 K. It follows that the volume increases and the *T* factor must be greater than 1.

$$275\ \text{L} \times \frac{313\ \cancel{\text{K}}}{293\ \cancel{\text{K}}} = 294\ \text{L}$$

We can visually summarize the Charles's law solution as follows:



Algebraic Solution

Alternatively, we can solve this problem using the equation

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

Example Exercise 11.4 Charles's Law

Continued

Solving for V_2 gives

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

Substituting for each variable and simplifying, we obtain

$$\frac{275\ \text{L} \times 313\ \cancel{\text{K}}}{293\ \cancel{\text{K}}} = 294\ \text{L}$$

Practice Exercise

A krypton balloon has a volume of 555 mL at 21 °C. If the balloon is cooled and the volume decreases to 475 mL, what is the final temperature? Assume that the pressure remains constant.

Answer: –21 °C

Concept Exercise

When air in an elastic balloon cools from 25 °C to 20 °C, the volume of the balloon (increases/decreases).

Answer: See Appendix G.

Example Exercise 11.5 Gay-Lussac's Law

A steel container filled with nitrous oxide at 15.0 atm is cooled from 2 °C to –40 °C. Calculate the final pressure assuming the volume remains constant.

Conceptual Solution

We must first convert the Celsius temperatures to Kelvin by adding 273.

$$25\ °\text{C} + 273 = 298\ \text{K}$$
$$-40\ °\text{C} + 273 = 233\ \text{K}$$

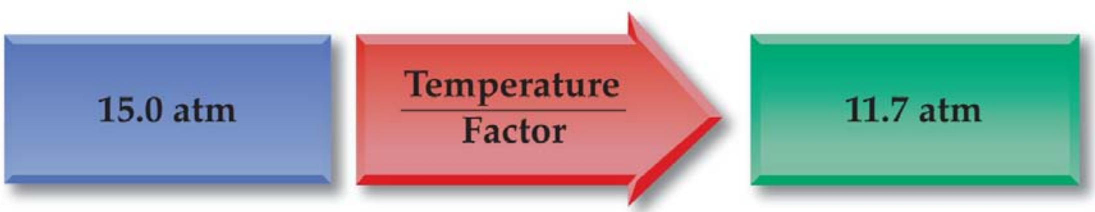
We can find the final pressure (P_2) by applying Gay-Lussac's law and using the relationship

$$P_1 \times T_{factor} = P_2$$

The volume of the container remains constant, but the temperature decreases from 298 K to 233 K. Therefore, the pressure decreases. The T_{factor} must be less than 1. Hence,

$$15.0\ \text{atm} \times \frac{233\ \cancel{\text{K}}}{298\ \cancel{\text{K}}} = 11.7\ \text{atm}$$

We can visually summarize the Gay-Lussac's law solution as follows:



Example Exercise 11.5 Gay-Lussac's Law

Continued

Algebraic Solution

Alternatively, we can solve this problem using the equation

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

Solving for P_2 gives

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

Substituting for each variable and simplifying, we have

$$\frac{15.0 \text{ atm} \times 233 \cancel{\text{K}}}{298 \cancel{\text{K}}} = 11.7 \text{ atm}$$

Practice Exercise

A copper container has a volume of 555 mL and is filled with air at 25 °C. The container is immersed in dry ice, and the pressure of the gas drops from 761 torr to 495 torr. What is the final temperature of the air in the copper container?

Answer: 194 K (–79 °C)

Concept Exercise

When air in a rigid steel tank cools from 25 °C to 20 °C, the pressure inside the tank (increases/decreases).

Answer: See Appendix G.

Example Exercise 11.6 Combined Gas Law

A nitrogen gas sample occupies 50.5 mL at –80 °C and 1250 torr. What is the volume at STP?

Conceptual Solution

Although the final conditions are not given, we know that STP conditions are 273 K and 760 torr. We can summarize the information as follows:

CONDITIONS	P	V	T
initial	1250 torr	50.5 mL	$-80\ °C + 273 = 193\ K$
final	760 torr	V_2	273 K

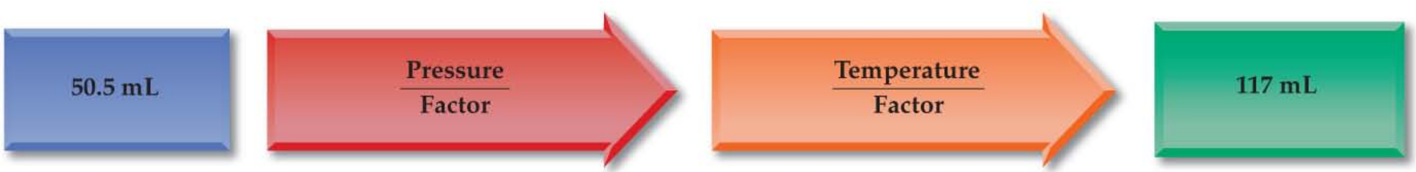
We can calculate the final volume by applying a P_{factor} and a T_{factor} to the initial volume.

$$V_1 \times P_{factor} \times T_{factor} = V_2$$

The pressure decreases, and so the volume increases; thus, the P_{factor} is greater than 1. The temperature increases, and so the volume increases; thus, the T_{factor} is also greater than 1.

$$50.5\ \text{mL} \times \frac{1250\ \cancel{\text{torr}}}{760\ \cancel{\text{torr}}} \times \frac{273\ \cancel{\text{K}}}{193\ \cancel{\text{K}}} = 117\ \text{mL}$$

We can visually summarize the combined gas law solution as follows:



Example Exercise 11.6 Combined Gas Law

Continued

Algebraic Solution

Alternatively, we can solve this problem using the equation

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Rearranging variables and solving for V_2,

$$\frac{P_1 V_1 T_2}{T_1 P_2} = V_2$$

Substituting for each variable and simplifying, we obtain

$$\frac{1250\ \cancel{\text{torr}} \times 50.5\ \text{mL} \times 273\ \cancel{\text{K}}}{193\ \cancel{\text{K}} \times 760\ \cancel{\text{torr}}} = 117\ \text{mL}$$

Practice Exercise

An oxygen gas sample occupies 50.0 mL at 27 °C and 765 mm Hg. What is the final temperature if the gas is cooled to a volume of 35.5 mL and a pressure of 455 mm Hg?

Answer: 127 K (–146 °C)

Concept Exercise

An oxygen gas sample occupies 50.0 mL at 27 °C and 765 mm Hg. What is the final temperature if the gas is cooled to a volume of 35.5 mL and a pressure of 455 mm Hg?

Answer: See Appendix G.

Example Exercise 11.7 Dalton's Law of Partial Pressures

An atmospheric sample contains nitrogen, oxygen, argon, and traces of other gases. If the partial pressure of nitrogen is 587 mm Hg, oxygen is 158 mm Hg, and argon is 7 mm Hg, what is the observed pressure as read on the barometer?

Solution

The sum of the individual partial pressures equals the total atmospheric pressure; therefore,

$$P_{\text{nitrogen}} + P_{\text{oxygen}} + P_{\text{argon}} = P_{\text{total}}$$

Substituting the values for the partial gas pressures, we have

$$587 \text{ mm Hg} + 158 \text{ mm Hg} + 7 \text{ mm Hg} = 752 \text{ mm Hg}$$

Thus, the atmospheric pressure as read on the barometer is 752 mm Hg.

Practice Exercise

The regulator on a steel scuba tank containing compressed air indicates that the pressure is 2250 psi. If the partial pressure of nitrogen is 1755 psi and that of argon is 22 psi, what is the partial pressure of oxygen in the tank?

Answer: 473 psi

Concept Exercise

A rigid steel cylinder contains N_2, O_2, and NO at a total pressure of 2.50 atm. What is the partial pressure of NO gas if N_2 and O_2 are each 1.00 atm?

Answer: See Appendix G.

Example Exercise 11.8 Ideal Gas Behavior

Suppose we have two 5.00 L samples of gas at 25 °C. One sample is ammonia, NH_3, and the other nitrogen dioxide, NO_2. Which gas has the greater kinetic energy? Which gas has the faster molecules?

Solution

Since the temperature of each gas is 25 °C, we know that the kinetic energy is the same for NH_3 and NO_2. At the same temperature, we know that lighter molecules move faster than heavier molecules. The molecular mass of NH_3 is 17 amu and NO_2 is 46 amu. Since NH_3 is lighter than NO_2, the ammonia molecules have a higher velocity than the nitrogen dioxide molecules.

Practice Exercise

Which of the following statements is *not* true according to the kinetic theory of gases?

(a) Molecules occupy a negligible volume.
(b) Molecules move in straight-line paths.
(c) Molecules are attracted to each other.
(d) Molecules undergo elastic collisions.
(e) Molecules of different gases at the same temperature have the same average kinetic energy.

Answer: All these statements are true except (c). Molecules of an ideal gas are not attracted to each other and behave as independent particles.

Concept Exercise

Which of the following is an example of an ideal gas: H_2, O_2, or He?

Answer: See Appendix G.

Example Exercise 11.9 Ideal Gas Law

How many moles of hydrogen gas occupy a volume of 0.500 L at STP?

Solution

We begin by rearranging the ideal gas equation, *PV* = *nRT*, and solving for *n*.

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

The temperature at STP is 273 K, and the pressure is 1 atm. Then we substitute for each variable in the ideal gas equation.

$$n = \frac{1\ \cancel{\text{atm}} \times 0.500\ \cancel{\text{L}}}{\dfrac{0.0821\ \cancel{\text{L}} \cdot \cancel{\text{atm}}}{1\ \text{mol} \cdot \cancel{\text{K}}} \times 273\ \cancel{\text{K}}} = 0.0223\ \text{mol}$$

Alternatively, we can display the value for the ideal gas constant separately to more clearly show the cancellation of units.

$$n = \frac{1\ \cancel{\text{atm}} \times 0.500\ \cancel{\text{L}}}{273\ \cancel{\text{K}}} \times \frac{1\ \text{mol} \cdot \cancel{\text{K}}}{0.0821\ \cancel{\text{L}} \cdot \cancel{\text{atm}}} = 0.0223\ \text{mol}$$

Practice Exercise

What is the temperature of 0.250 mol of chlorine gas at 655 torr if the volume is 3.50 L?

Answer: 147 K (–126 °C)

Concept Exercise

Solve for the proportionality constant (*R*) in the ideal gas law.

Answer: See Appendix G.