Chapter 9: Gases

9.1 Gas Pressure

Question 50-1.

Why are sharp knives more effective than dull knives? (Hint: Think about the definition of pressure.)

Solution

The cutting edge of a knife that has been sharpened has a smaller surface area than a dull knife. Since pressure is force per unit area, a sharp knife will exert a higher pressure with the same amount of force and cut through material more effectively.

Question 50-2.

Why do some small bridges have weight limits that depend on how many wheels or axles the crossing vehicle has?

Solution

The more axles/wheels a vehicle has, the more surface area it has that is in contact with the road. For a given weight of vehicle (force applied to the road), a higher area corresponds to a lower pressure, making vehicles with more axles/wheels less likely to damage the road surface.

Question 50-3.

Why should you roll or belly crawl rather than walk across a thinly frozen pond?

Solution

Lying down distributes your weight over a larger surface area, exerting less pressure on the ice compared to standing up. If you exert less pressure, you are less likely to break through thin ice.

Question 50-4.

A typical barometric pressure in Redding, California, is about 750 mm Hg. Calculate this pressure in atm and kPa.

Solution

$$750\ \cancel{\text{mm Hg}} \times \frac{1\ \text{atm}}{760\ \cancel{\text{mm Hg}}} = 0.987\ \text{atm}$$

$$0.987\ \cancel{\text{atm}} \times \frac{101.325\ \text{kPa}}{1\ \cancel{\text{atm}}} = 100\ \text{kPa}$$

Question 50-5.

A typical barometric pressure in Denver, Colorado, is 615 mm Hg. What is this pressure in atmospheres and kilopascals?

Solution

Convert 615 mm Hg to atmospheres using 760 mm Hg = 1 atm. Use 1 atm = 101.325 kPa in the second part.

$$615 \text{ mm Hg} \times \frac{1\text{atm}}{760\text{mmHg}} = 0.809 \text{ atm}$$

$$0.809 \text{ atm} \times \frac{101.325\text{kPa}}{1\text{atm}} = 82.0 \text{ kPa}$$

Question 50-6.

A typical barometric pressure in Kansas City is 740 torr. What is this pressure in atmospheres, in millimeters of mercury, and in kilopascals?

Solution

Convert 740 torr to the units required: 760 torr = 1 atm = 760 mmHg = 101.3 kPa.

$$P(\text{atm}) = 740 \cancel{\text{torr}} \times \frac{1 \text{ atm}}{760 \cancel{\text{torr}}} = 0.974 \text{ atm}$$

$$P(\text{mmHg}) = 740 \text{ torr} \times \frac{760 \text{ mmHg}}{760 \text{ torr}} = 740 \text{ mmHg}$$

$$P(\text{kPa}) = 740 \cancel{\text{torr}} \times \frac{101.3 \text{ kPa}}{760 \cancel{\text{torr}}} = 98.6 \text{ kPa}$$

Question 50-7.

Canadian tire pressure gauges are marked in units of kilopascals. What reading on such a gauge corresponds to 32 psi?

Solution

$$32.0 \cancel{\text{lb}} \cancel{\text{in}^{-2}} \times \frac{1 \cancel{\text{atm}}}{14.7 \cancel{\text{lb}} \cancel{\text{in}^{-2}}} \times \frac{101.325 \text{ kPa}}{1 \cancel{\text{atm}}} = 2.2 \times 10^2 \text{ kPa}$$

Question 50-8.

During the Viking landings on Mars, the atmospheric pressure was determined to be on the average about 6.50 millibars (1 bar = 0.987 atm). What is that pressure in torr and kPa?

Solution

$$6.50 \cancel{\text{mbar}} \times \frac{1 \cancel{\text{bar}}}{1000 \cancel{\text{mbar}}} \times \frac{0.987 \cancel{\text{atm}}}{1 \cancel{\text{bar}}} \times \frac{760 \text{ torr}}{1 \cancel{\text{atm}}} = 4.88 \text{ torr}$$

$$6.50 \cancel{\text{mbar}} \times \frac{1 \cancel{\text{bar}}}{1000 \cancel{\text{mbar}}} \times \frac{0.987 \cancel{\text{atm}}}{1 \cancel{\text{bar}}} \times \frac{101.325 \text{ kPa}}{1 \cancel{\text{atm}}} = 0.650 \text{ kPa}$$

Question 50-9.

The pressure of the atmosphere on the surface of the planet Venus is about 88.8 atm. Compare that pressure in psi to the normal pressure on earth at sea level in psi.

Solution

Identify: 14.7 psi = 1 atm

$$88.8 \cancel{\text{atm}} \times \frac{14.7 \text{ psi}}{1 \cancel{\text{atm}}} = 1.30 \times 10^3 \text{ psi}$$

Question 50-10.

A medical laboratory catalog describes the pressure in a cylinder of a gas as 14.82 MPa. What is the pressure of this gas in atmospheres and torr?

Solution

Convert 14.82 MPa into atm.

$$P(\text{atm}) = 14.82\ \cancel{\text{MPa}} \times \frac{1000\ \cancel{\text{kPa}}}{1\ \cancel{\text{MPa}}} \times \frac{1\ \text{atm}}{101.3\ \cancel{\text{kPa}}} = 146.3\ \text{atm}$$

$$P(\text{torr}) = 146.3\ \cancel{\text{atm}} \times \frac{760\ \text{torr}}{1\ \cancel{\text{atm}}} = 1.112 \times 10^5\ \text{torr}$$

Question 50-11.

Consider this scenario and answer the following questions: On a mid-August day in the northeastern United States, the following information appeared in thelocal newspaper: atmospheric pressure at sea level 29.97 in., 1013.9 mbar.

(a) What was the pressure in kPa?

(b) The pressure near the seacoast in the northeastern United States is usually reported near 30.0 in. Hg. During a hurricane, the pressure may fall to near 28.0 in. Hg. Calculate the drop in pressure in torr.

Solution

(a) $29.97\ \cancel{\text{in. Hg}} \times \frac{101.325\ \text{kPa}}{29.92\ \cancel{\text{in. Hg}}} = 101.5\ \text{kPa}$; (b) $28.0\ \cancel{\text{in. Hg}} \times \frac{760\ \text{torr}}{29.92\ \cancel{\text{in. Hg}}} = 711\ \text{torr}$; 762 – 711 = 51 torr drop

Question 50-12.

Why is it necessary to use a nonvolatile liquid in a barometer or manometer?

Solution

A volatile liquid evaporates easily, which would produce some gas pressure in enclosed volumes, thus affecting the measurements.

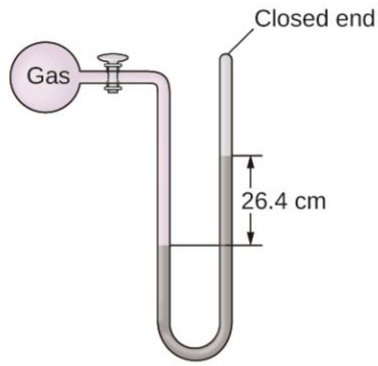
Question 50-13.

The pressure of a sample of gas is measured at sea level with a closed-end manometer. The liquid in the manometer is mercury. Determine the pressure of the gas in:

(a) torr

(b) Pa

(c) bar



Solution

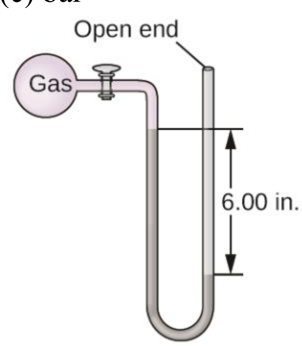
(a) $26.4\ \cancel{\text{cm}} \times \dfrac{10\ \cancel{\text{mm}}}{1\ \cancel{\text{cm}}} \times \dfrac{1\ \text{torr}}{1\ \cancel{\text{mm}}} = 264\ \text{torr}$; (b) $264\ \text{torr} \times \dfrac{101{,}325\ \text{Pa}}{760\ \text{torr}} = 35{,}200\ \text{Pa}$; (c)

$$264\ \cancel{\text{torr}} \times \frac{1.01325\ \text{bar}}{760\ \cancel{\text{torr}}} = 0.352\ \text{bar}$$

Question 50-14.

The pressure of a sample of gas is measured with anopen-end manometer, partially shown to the right. The liquid in the manometer is mercury. Assuming atmospheric pressure is 29.92 in. Hg, determine the pressure of the gas in:

(a) torr
(b) Pa
(c) bar



Solution

Since the mercury level is lower in the open arm of the manometer tube, atmospheric pressure is greater than the trapped gas pressure. The pressure of the gas is therefore calculated by subtracting the hydrostatic pressure corresponding to a 6 in. column of mercury from atmospheric pressure. (a) In in. Hg, this is: 29.92 in. Hg – 6.00 in. Hg = 23.92 in. Hg,

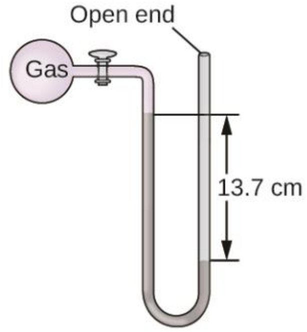
$23.92\ \text{in. Hg} \times \dfrac{760\ \text{torr}}{29.92\ \text{in. Hg}} = 608\ \text{torr}$; (b) $608\ \text{torr} \times \dfrac{101325\ \text{Pa}}{760\ \text{torr}} = 81100\ \text{Pa}$; (c)

$$608\ \cancel{\text{torr}} \times \frac{1.01325\ \text{bar}}{760\ \cancel{\text{torr}}} = 0.811\ \text{bar}$$

Question 50-15.

The pressure of a sample of gas is measured at sea level with an open-end mercury manometer. Assuming atmospheric pressure is 760.0 mm Hg, determine the pressure of the gas in:

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



Solution

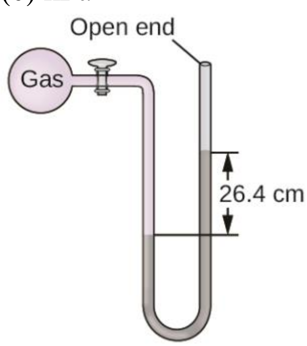
The pressure of the gas equals the hydrostatic pressure due to the pressure of the atmosphere at sea level minus a column of mercury of height 13.7 cm. The pressure on the left is due to the gas and the pressure on the right is due to the atmospheric pressure minus 13.7 cm Hg).(a) In mm Hg, this is: 760 mm Hg – 137 mmHg = 623 mm Hg; (b)

$$623 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}} = 0.820 \text{ atm}; \text{(c)}\ 0.820 \text{ atm} \times \frac{101.325 \text{ kPa}}{1 \text{ atm}} = 83.1 \text{ kPa}$$

Question 50-16.

The pressure of a sample of gas is measured at sea level with an open-end mercury manometer. Assuming atmospheric pressure is 760 mm Hg, determine the pressure of the gas in:

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



Solution

(a) In mm Hg, this is: 760 mm Hg + 264 mm Hg = 1024 mm Hg; (b)

$$1024 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}} = 1.34 \text{ atm}; \text{(c)}\ 1024 \text{ mm Hg} \times \frac{101.325 \text{ kPa}}{760 \text{ mm Hg}} = 136 \text{ kPa}$$

Question 50-17.

How would the use of a volatile liquid affect the measurement of a gas using open-ended manometers vs. closed-end manometers?

Solution

With a closed-end manometer, no change would be observed, since the vaporized liquid would contribute equal, opposing pressures in both arms of the manometer tube. However, with an open-ended manometer, a higher pressure reading of the gas would be obtained than expected, since $P_{gas} = P_{atm} + P_{vol\ liquid}$.

9.2 Relating Pressure, Volume, Amount, and Temperature: The Ideal Gas Law

Question 51-1.

Sometimes leaving a bicycle in the sun on a hot day will cause a blowout. Why?

Solution

The temperature of the air in the tire increases, which causes the pressure and/or volume to increase. The volume cannot increase substantially, and high enough pressure can cause the tire to burst.

Question 51-2.

Explain how the volume of the bubbles exhausted by a scuba diver (Figure 9.16) change as they rise to the surface, assuming that they remain intact.

Solution

As the bubbles rise, the pressure decreases, so their volume increases as suggested by Boyle's law.

Question 51-3.

One way to state Boyle's law is "All other things being equal, the pressure of a gas is inversely proportional to its volume."

(a) What is the meaning of the term "inversely proportional?"

(b) What are the "other things" that must be equal?

Solution

(a) The pressure increases as the volume decreases and vice-versa. Mathematically, this can be expressed $P \times V =$ a constant, or $P =$ constant $\times \dfrac{1}{V}$. (b) amount of gas and temperature; the temperature and the amount (number of moles) of gas must not change

Question 51-4.

An alternate way to state Avogadro's law is "All other things being equal, the number of molecules in a gas is directly proportional to the volume of the gas."

(a) What is the meaning of the term "directly proportional?"

(b) What are the "other things" that must be equal?

Solution

(a) The number of particles in the gas increases as the volume increases. This relationship may be written as n = constant × V. It is a direct relationship. (b) The temperature and pressure must be kept constant.

Question 51-5.

How would the graph in Figure 9.12 change if the number of moles of gas in the sample used to determine the curve were doubled?

Solution

The slope of the line would increase by a factor of 2.

Question 51-6.

How would the graph in Figure 9.13 change if the number of moles of gas in the sample used to determine the curve were doubled?

Solution

The curve would be farther to the right and higher up, but the same basic shape.

Question 51-7.

In addition to the data found in Figure 9.13, what other information do we need to find the mass of the sample of air used to determine the graph?

Solution

The temperature and the (average) molar mass of air.

Question 51-8.

Determine the volume of 1 mol of CH_4 gas at 150 K and 1 atm, using Figure 9.12.

Solution

The figure shows the change in volume for 1 mol of CH_4 gas as a function of temperature. The graph shows that the volume is about 12.5 L.

Question 51-9.

Determine the pressure of the gas in the syringe shown in Figure 9.13 when its volume is 12.5 mL, using:

(a) the appropriate graph

(b) Boyle's law

Solution

(a) The graphical approach requires interpolating a value of pressure, and so the linear plot is preferred. The volume of 12.5 mL corresponds to an approximate value of 0.065 psi^{-1} for $\frac{1}{P}$, and so the pressure is approximately $(0.065\ \text{psi}^{-1})^{-1}$ = 15 psi. (b) Any of the pressure-volume data pairs provided in the graph may be used to compute the requested pressure via Boyle's law. For example, using the (15.0 mL, 13.0 psi) data point yields $P_1V_1 = P_2V_2$

$$P_2 = \frac{13.0\ \text{psi} \times 15.0\ \cancel{\text{mL}}}{12.5\ \cancel{\text{mL}}} = 15.6\ \text{psi}$$

Question 51-10.

A spray can is used until it is empty except for the propellant gas, which has a pressure of 1344 torr at 23 °C. If the can is thrown into a fire (T = 475 °C), what will be the pressure in the hot can?

Solution

The first thing to recognize about this problem is that the volume and moles of gas remain constant. Thus, we can use the combined gas law equation in the form:

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

$$P_2 = \frac{P_1T_2}{T_1} = 1344 \text{ torr} \times \frac{475 + 273.15}{23 + 273.15} = 3.40 \times 10^3 \text{ torr}$$

Question 51-11.

What is the temperature of an 11.2-L sample of carbon monoxide, CO, at 744 torr if it occupies 13.3 L at 55 °C and 744 torr?

Solution

This scenario involves a constant amount of gas (unspecified) undergoing a temperature induced change in volume at a constant pressure (744 torr). Charles law may therefore be used.

V_1 = 13.3 L

T_1 = 55 °C = 328 K

V_2= 11.2 L

T_2 = ?

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

$$T_2 = \frac{V_2T_1}{V_1} = \frac{11.2 \text{ L} \times 328 \text{ K}}{13.3 \text{ L}} = 276 \text{ K}$$

Question 51-12.

A 2.50-L volume of hydrogen measured at –196 °C is warmed to 100 °C. Calculate the volume of the gas at the higher temperature, assuming no change in pressure.

Solution

Apply Charles's law to compute the volume of gas at the higher temperature:

V_1 = 2.50 L

T_1 = –196 °C = 77.15 K

V_2= ?

T_2 = 100 °C = 373.15 K

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

$$V_2 = \frac{V_1T_2}{T_1} = \frac{2.50\ \text{L} \times 373.15\ \cancel{\text{K}}}{77.15\ \cancel{\text{K}}} = 12.1\ \text{L}$$

Question 51-13.

A balloon inflated with three breaths of air has a volume of 1.7 L. At the same temperature and pressure, what is the volume of the balloon if five more same-sized breaths are added to the balloon?

Solution

Use Avogadro's law to compute the requested volume. The gas volume added with each breath is 0.567 L, $\left(\frac{1.7\ \text{L}}{3\ \text{breaths}} = \frac{0.567\ \text{L}}{1\ \text{breath}}\right)$, keeping a nonsignificant "guard digit" to avoid a rounding error. If a total of eight breaths are in the balloon, the total volume of gas is:

$$\left(8\ \cancel{\text{breaths}} \times \frac{0.567\ \text{L}}{1\ \cancel{\text{breath}}}\right) = 4.5\ \text{L}.$$

Question 51-14.

A weather balloon contains 8.80 moles of helium at a pressure of 0.992 atm and a temperature of 25° C at ground level. What is the volume of the balloon under these conditions?

Solution

$PV = nRT$

$$V = \frac{nRT}{P} = \frac{8.80\ \cancel{\text{mol}} \times 0.08206\ \text{L}\ \cancel{\text{atm}}\ \cancel{\text{mol}^{-1}}\ \cancel{\text{K}^{-1}} \times 298.15\ \cancel{\text{K}}}{0.992\ \cancel{\text{atm}}} = 217\ \text{L}$$

Question 51-15.

The volume of an automobile air bag was 66.8 L when inflated at 25 °C with 77.8 g of nitrogen gas. What was the pressure in the bag in kPa?

Solution

$$P = \frac{nRT}{V} = \frac{\left(\frac{77.8\ \text{g}}{28.0135\ \text{g mol}^{-1}}\right)\left(0.08206\ \text{L atm mol}^{-1}\ \text{K}^{-1}\right)\left(298.1\ \text{K}\right)}{66.8\ \text{L}} = 1.02\ \text{atm}$$

$$P = 1.02\ \cancel{\text{atm}} \times \frac{101.325\ \text{kPa}}{1\ \cancel{\text{atm}}} = 103\ \text{kPa}$$

Question 51-16.

How many moles of gaseous boron trifluoride, BF_3, are contained in a 4.3410-L bulb at 788.0 K if the pressure is 1.220 atm? How many grams of BF_3?

Solution

$$n = \frac{PV}{RT} = \frac{1.220\ \cancel{\text{atm}}\ (4.3410\ \text{L})}{(0.08206\ \text{L}\ \cancel{\text{atm}}\ \text{mol}^{-1}\ \cancel{\text{K}^{-1}})(788.0\ \cancel{\text{K}})} = 0.08190\ \text{mol} = 8.190 \times 10^{-2}\ \text{mol}$$

$$n \times \text{molar mass} = 8.190 \times 10^{-2}\ \cancel{\text{mol}} \times 67.8052\ \text{g}\ \cancel{\text{mol}^{-1}} = 5.553\ \text{g}$$

Question 51-17.

Iodine, I_2, is a solid at room temperature but sublimes (converts from a solid into a gas) when warmed. What is the temperature in a 73.3-mL bulb that contains 0.292 g of I_2 vapor at a pressure of 0.462 atm?

Solution

The molar mass of I_2 is 2 × 126.90447 = 253.8089 g/mol;

$$T = \frac{PV}{nR} = \frac{0.462\ \text{atm}(0.0733\ \text{L})}{\left(\frac{0.292\ \text{g}}{253.8\ \text{g mol}^{-1}}\right)(0.08206\ \text{L atm mol}^{-1}\ \text{K}^{-1}} = 359\ \text{K}$$

Question 51-18.

How many grams of gas are present in each of the following cases?

(a) 0.100 L of CO_2 at 307 torr and 26 °C

(b) 8.75 L of C_2H_4, at 378.3 kPa and 483 K

(c) 221 mL of Ar at 0.23 torr and –54 °C

Solution

In each of these problems, we are given a volume, pressure, and temperature. We can obtain moles from this information using the molar mass, $m = nM$, where M is the molar mass:

$$P, V, T \xrightarrow{n=PV/RT} n, \xrightarrow{m=n\ \text{(molar mass)}} \text{grams}$$

or we can combine these equations to obtain:

$$\text{mass} = m = \frac{PV}{RT} \times M$$

(a)

$$307\ \cancel{\text{torr}} \times \frac{1\ \text{atm}}{760\ \cancel{\text{torr}}} = 0.4039\ \text{atm} \qquad 26\ °\text{C} = 299.1\ \text{K}$$

$$\text{Mass} = m = \frac{0.4039\ \cancel{\text{atm}}\ (0.100\ \cancel{\text{L}})}{0.08206\ \cancel{\text{L}}\ \cancel{\text{atm}}\ \text{mol}^{-1}\ \cancel{\text{K}^{-1}}\ (299.1\ \cancel{\text{K}})} \times 44.01\ \text{g mol}^{-1} = 7.24 \times 10^{-2}\ \text{g};$$

(b)

$$\text{Mass} = m = \frac{378.3\ \cancel{\text{kPa}}\ (8.75\ \cancel{\text{L}})}{8.314\ \cancel{\text{L}}\ \cancel{\text{kPa}}\ \text{mol}^{-1}\ \cancel{\text{K}^{-1}}\ (483\ \cancel{\text{K}})} \times 28.05376\ \text{g mol}^{-1} = 23.1\ \text{g};$$

(c)

$$221\ \cancel{mL} \times \frac{1\ L}{1000\ \cancel{mL}} = 0.221\ L \qquad -54\ °C + 273.15 = 219.15\ K$$

$$0.23\ \cancel{torr} \times \frac{1\ atm}{760\ \cancel{torr}} = 3.03 \times 10^{-4}\ atm$$

$$\text{Mass} = m = \frac{3.03 \times 10^{-4}\ \cancel{atm}\ (0.221\ \cancel{L})}{0.08206\ \cancel{L}\ \cancel{atm}\ mol^{-1}\ \cancel{K^{-1}}\ (219.15\ \cancel{K})} \times 39.948\ g\ mol^{-1} = 1.5 \times 10^{-4}\ g$$

Question 51-19.

A high-altitude balloon is filled with 1.41×10^4 L of hydrogen at a temperature of 21 °C and a pressure of 745 torr. What is the volume of the balloon at a height of 20 km, where the temperature is –48 °C and the pressure is 63.1 torr?

Solution

Write the two sets of conditions in tabular form:

P_1: 745 torr
P_2: 63.1 torr
T_1: 21 °C (294.15 K)
T_2: –48 °C (225.15 K)
V_1: 1.41×10^4 L
V_2: ?

Writing the data given in this way makes it easy to find the unknown. We can use the ideal gas law with the moles of gas (n) held constant:

$PV = nRT$

$$\frac{P_1V_1}{T_1} = nR = \frac{P_2V_2}{T_2}$$

$$V_2 = \frac{P_1V_1T_2}{T_1P_2} = \frac{75\ \text{torr}\ (1.41 \times 10^2\ \text{L})(225.15\ \text{K})}{(294.15\ \text{K})(63.1\ \text{torr})} = 1.27 \times 10^5\ \text{L}$$

Note that since R is not required to work the problem, the pressure can be in any pressure unit. The temperature must be in kelvin.

Question 51-20.

A cylinder of medical oxygen has a volume of 35.4 L, and contains O_2 at a pressure of 151 atm and a temperature of 25 °C. What volume of O_2 does this correspond to at normal body conditions, that is, 1 atm and 37 °C?

Solution

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$V_2 = \frac{P_1V_1T_2}{T_1P_2}$$

$$= \frac{(151\ \text{atm})(35.4\ \text{L})(310\ \text{K})}{(298\ \text{K})(1\ \text{atm})} = 5561\ \text{L}$$

Question 51-21.

A large scuba tank with a volume of 18 L is rated for a pressure of 220 bar. The tank is filled at 20 °C and contains enough air to supply 1860 L of air to a diver at a pressure of 2.37 atm (a depth of 45 feet).Was the tank filled to capacity at 20 °C?

Solution

$$\frac{2.37\ \cancel{\text{atm}}}{1} \times \frac{1.01325\ \text{bar}}{1\ \cancel{\text{atm}}} = 2.40\ \text{bar}$$

$P_1V_1 = P_2V_2$

$$P_1 = \frac{P_2V_2}{V_1}$$

$$P_1 = \frac{(2.40\ \text{bar})(1860\ \cancel{\text{L}})}{(18\ \cancel{\text{L}})} = 248\ \text{bar}$$

The pressure when the tank was filled is 248 bar, which exceeds the rated pressure of 220 bar, so the tank was actually overfilled.

Question 51-22.

A 20.0-L cylinder containing 11.34 kg of butane, C_4H_{10}, was opened to the atmosphere. Calculate the mass of the gas remaining in the cylinder if it were opened and the gas escaped until the pressure in the cylinder was equal to the atmospheric pressure, 0.983 atm, and a temperature of 27 °C.

Solution

Calculate the amount of butane in 20.0 L at 0.983 atm and 27 °C. The original amount in the container does not matter.

$$n = \frac{PV}{RT} = \frac{0.983\ \cancel{\text{atm}} \times 20.0\ \cancel{\text{L}}}{0.08206\ \cancel{\text{L}}\ \cancel{\text{atm}}\ \text{mol}^{-1}\ \cancel{\text{K}^{-1}}\ (300.1\ \cancel{\text{K}})} = 0.798\ \text{mol}$$

Mass of butane = 0.798 mol × 58.1234 g/mol = 46.4 g

Question 51-23.

While resting, the average 70-kg human man consumes 14 L of pure O_2 per hour at 25 °C and 100 kPa. How many moles of O_2 are consumed by a 70-kg man while resting for 1.0 h?

Solution

Use the ideal gas law: $PV = nRT$

0.99 atm × 14 L = n × 0.08226 atm L/mol K × 298 K

$$n = \frac{(0.99\ \text{atm} \times 14\ \text{L})}{(0.08226\ \text{atm L/mol K} \times 298\ \text{K})} = 0.57\ \text{mol}$$

Question 51-24.

For a given amount of gas showing ideal behavior, draw labeled graphs of:

(a) the variation of P with V

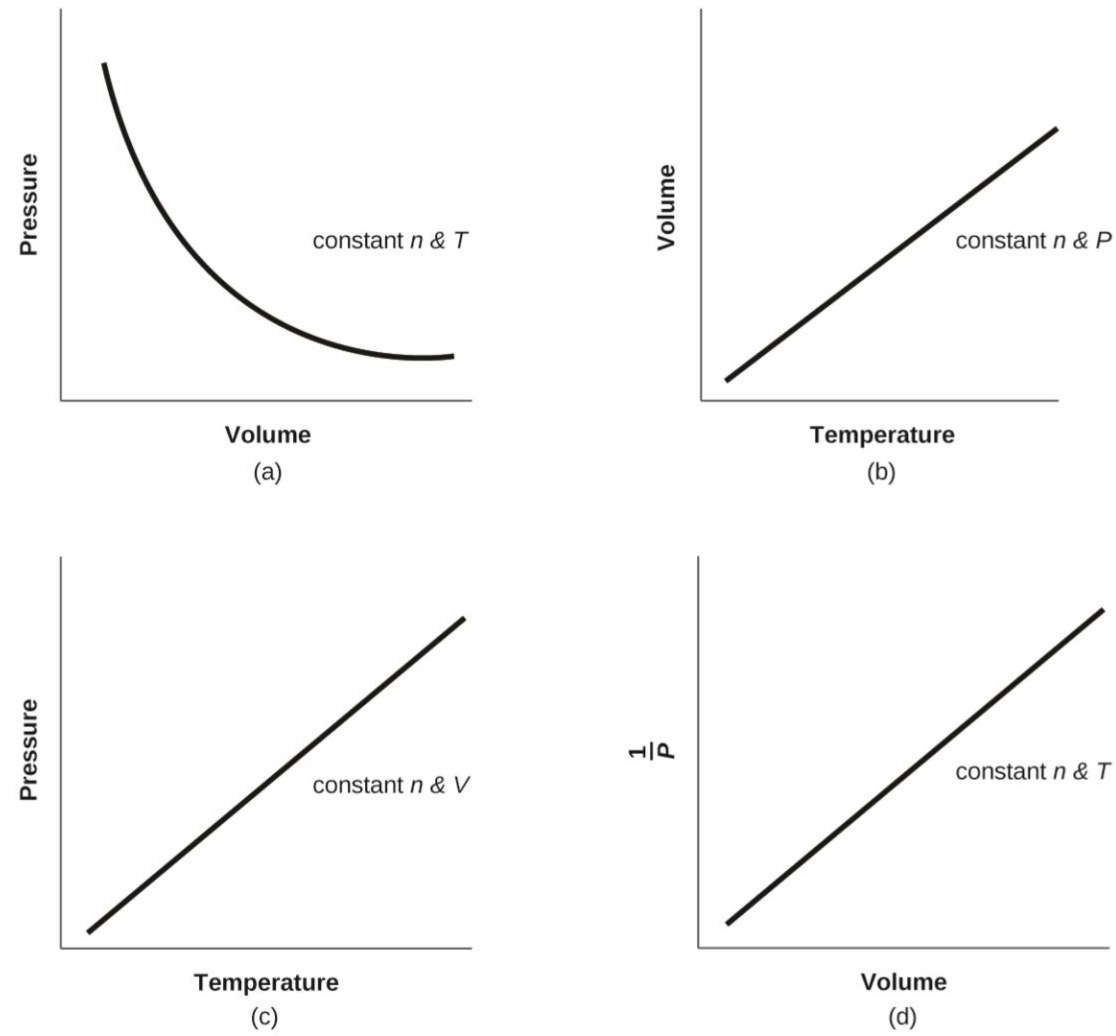
(b) the variation of V with T

(c) the variation of P with T

(d) the variation of $\frac{1}{P}$ with V

Solution

For a gas exhibiting ideal behavior:



Question 51-25.

A liter of methane gas, CH_4, at STP contains more atoms of hydrogen than does a liter of pure hydrogen gas, H_2, at STP. Using Avogadro's law as a starting point, explain why.

Solution

One liter of CH_4 and one liter of H_2 contain the same number of *molecules* at STP (Avogadro's law), but each CH_4 molecule contains four H atoms while each H_2 molecule contains two H atoms.

Question 51-26.

The effect of chlorofluorocarbons (such as CCl_2F_2) on the depletion of the ozone layer is well known. The use of substitutes, such as CH_3CH_2F(*g*), for the chlorofluorocarbons, has largely corrected the problem. Calculate the volume occupied by 10.0 g of each of these compounds at STP:

(a) $CCl_2F_2(g)$

(b) $CH_3CH_2F(g)$

Solution

(a) Determine the molar mass of CCl_2F_2 then calculate the moles of $CCl_2F_2(g)$ present. Use the ideal gas law $PV = nRT$ to calculate the volume of $CCl_2F_2(g)$:

$$10.0\text{ g } CCl_2F_2 \times \frac{1\text{ mol } CCl_2F_2}{120.91\text{ g } CCl_2F_2} = 0.0827\text{ mol } CCl_2F_2$$

$PV = nRT$, where n = # mol CCl_2F_2

$$1\text{ atm} \times V = 0.0827\text{ mol} \times \frac{0.0821\text{ L atm}}{\text{mol K}} \times 273\text{ K} = 1.85\text{ L } CCl_2F_2;$$

$$\text{(b) } 10.0\text{ g } CH_3CH_2F \times \frac{1\text{ mol } CH_3CH_2F}{48.07\text{ g } CH_3CH_2F} = 0.208\text{ mol } CH_3CH_2F$$

$PV = nRT$, with n = # mol CH_3CH_2F

1 atm $\times V$ = 0.208 mol × 0.0821 L atm/mol K × 273 K = 4.66 L CH_3CH_2F

Question 51-27.

As 1 g of the radioactive element radium decays over 1 year, it produces 1.16×10^{18} alpha particles (helium nuclei). Each alpha particle becomes an atom of helium gas. What is the pressure in pascal of the helium gas produced if it occupies a volume of 125 mL at a temperature of 25 °C?

Solution

After 1 year, there are:

$$\frac{1.16 \times 10^{18}}{6.022 \times 10^{28}\text{ mol}^{-1}} = 1.926 \times 10^{-6}\text{ mol}$$

From the ideal gas law, $PV = nRT$:

$$P = \frac{nRT}{V} = \frac{1.926 \times 10^{-6}\ \cancel{\text{mol}} \times 8.314\ \cancel{\text{L}}\text{ kPa}\ \cancel{\text{mol}^{-1}}\ \cancel{\text{K}^{-1}} \times 298\ \cancel{\text{K}}}{0.125\ \cancel{\text{L}}} = 3.82 \times 10^{-2}\text{ kPa}$$

Question 51-28.

A balloon with a volume of 100.21 L at 21 °C and 0.981 atm is released and just barely clears the top of Mount Crumpet in British Columbia. If the final volume of the balloon is 144.53 L at a temperature of 5.24 °C, what is the pressure experienced by the balloon as it clears Mount Crumpet?

Solution

Identify the variables in the problem and determine that the combined gas law $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ is the necessary equation to use to solve the problem. Then solve for P_2:

$$\frac{0.981\text{ atm} \times 100.21\text{ L}}{294\text{ K}} = \frac{P_2 \times 144.53\text{ L}}{278.24\text{ K}}$$

$$P_2 = 0.644\text{ atm}$$

Question 51-29.

If the temperature of a fixed amount of a gas is doubled at constant volume, what happens to the pressure?

Solution

The pressure is doubled.

Question 51-30.

If the volume of a fixed amount of a gas is tripled at constant temperature, what happens to the pressure?

Solution

The pressure decreases by a factor of 3.

9.3 Stoichiometry of Gaseous Substances, Mixtures, and Reactions

Question 52-1.

What is the density of laughing gas, dinitrogen monoxide, N_2O, at a temperature of 325 K and a pressure of 113.0 kPa?

Solution

$$\rho = \frac{PM}{RT} = \frac{113.0\ \cancel{\text{kPa}}\ (2 \times 14.00674 + 15.9994)\ \text{g}\ \cancel{\text{mol}^{-1}}}{8.314\ \text{L}\ \cancel{\text{kPa}}\ \cancel{\text{mol}^{-1}}\ \cancel{\text{K}^{-1}} \times 325\ \cancel{\text{K}}} = 1.84\ \text{g L}^{-1}$$

Question 52-2.

Calculate the density of Freon 12, CF_2Cl_2, at 30.0 °C and 0.954 atm.

Solution

$$\rho = \frac{PM}{RT} = \frac{0.954\ \cancel{\text{atm}}\ \left[12.011 + 2(18.9954) + 2(35.453)\right]\ \text{g}\ \cancel{\text{mol}^{-1}}}{0.08206\ \text{L}\ \cancel{\text{atm}}\ \cancel{\text{mol}^{-1}}\ \cancel{\text{K}^{-1}} \times 303.15\ \cancel{\text{K}}} = 4.64\ \text{g L}^{-1}$$

Question 52-3.

Which is denser at the same temperature and pressure, dry air or air saturated with water vapor? Explain.

Solution

Dry air is denser; it has a larger molecular mass. Since density is mass per unit volume, at a given temperature and pressure, *n* is a constant. Thus, a given volume will have the same number of moles of gas. Therefore, if the air becomes moist while the pressure, temperature, and volume remain the same, then some of the O_2 and N_2 molecules will be replaced by H_2O molecules to keep the total number of moles of gas the same. Because the molar mass of H_2O is less than that of O_2 and N_2, the air becomes less dense. Thus, air saturated with water vapor is less dense. Therefore, dry air is denser.

Question 52-4.

A cylinder of $O_2(g)$ used in breathing by patients with emphysema has a volume of 3.00 L at a pressure of 10.0 atm. If the temperature of the cylinder is 28.0 °C, what mass of oxygen is in the cylinder?

Solution

$$\text{mass } O_2 = \frac{(31.9988 \text{ g } \cancel{\text{mol}^{-1}})(10.0 \cancel{\text{atm}})(3.00 \cancel{\text{L}})}{(0.08206 \cancel{\text{L}} \cancel{\text{atm}} \cancel{\text{mol}^{-1}} \cancel{\text{K}^{-1}})(301.15 \cancel{\text{K}})} = 38.8 \text{ g}$$

Question 52-5.

What is the molar mass of a gas if 0.0494 g of the gas occupies a volume of 0.100 L at a temperature 26 °C and a pressure of 307 torr?

Solution

From the ideal gas law, $PV = nRT$, set $n = \frac{\text{mass}}{\text{molar mass}}$ and solve for the molar mass.

$$\text{molar mass} = \frac{(0.0494 \text{ g})(0.08206 \cancel{\text{L}} \cancel{\text{atm}} \text{ mol}^{-1} \cancel{\text{K}^{-1}})(299.15 \cancel{\text{K}})}{\left(\frac{307 \cancel{\text{torr}}}{760 \cancel{\text{torr}} \cancel{\text{atm}^{-1}}}\right)(0.1 \cancel{\text{L}})} = 30.0 \text{ g mol}^{-1}$$

Question 52-6.

What is the molar mass of a gas if 0.281 g of the gas occupies a volume of 125 mL at a temperature 126 °C and a pressure of 777 torr?

Solution

From the ideal gas law, $PV = nRT$, set $n = \frac{\text{mass}}{\text{molar mass}}$ and solve the molar mass.

$$\text{molar mass} = \frac{(0.281 \text{ g})(0.08206 \cancel{\text{L}} \cancel{\text{atm}} \text{ mol}^{-1} \cancel{\text{K}^{-1}})(399.15 \cancel{\text{K}})}{\left(\frac{777 \cancel{\text{torr}}}{760 \cancel{\text{torr}} \cancel{\text{atm}^{-1}}}\right)(0.125 \cancel{\text{L}})} = 72.0 \text{ g mol}^{-1}$$

Question 52-7.

How could you show experimentally that the molecular formula of propene is C_3H_6, not CH_2?

Solution

A person could measure the mass, volume, and pressure of the gas and calculate the molar mass from the ideal gas equation. C_3H_6 will have three times the molar mass expected for CH_2.

Question 52-8.

The density of a certain gaseous fluoride of phosphorus is 3.93 g/L at STP. Calculate the molar mass of this fluoride and determine its molecular formula.

Solution

$$\text{M} = \frac{mRT}{PV} \qquad D = \frac{m}{V} \qquad \text{M} = \frac{DRT}{P}$$

$$M = \frac{3.93\ \text{g L}^{-1}\ (0.08206\ \text{L}\ \cancel{\text{atm}}\ \text{mol}^{-1}\ \cancel{\text{K}^{-1}})(273.15\ \cancel{\text{K}})}{1.00\ \cancel{\text{atm}}} = 88.1\ \text{g mol}^{-1}$$

$M_{phosphorous}$ = 30.97376 g/mol
$M_{fluorine}$ = 18.998403 g/mol
molecular formula: phosphorous: 30.97376

$$\begin{array}{lr} \text{fluorine:} & \underline{3(18.998403)} \\ & 87.968969 \end{array}$$

The molecular formula is PF_3.
To find this answer you can either use trial and error, or you can realize that since phosphorus is in group 5, it can fill its valence shell by forming three bonds. Fluorine, being in group 7, needs to form only one bond to fill its shell. Thus it makes sense to start with PF_3 as a probable formula.

Question 52-9.

Consider this question: What is the molecular formula of a compound that contains 39% C, 45% N, and 16% H if 0.157 g of the compound occupies 125 mL with a pressure of 99.5 kPa at 22 °C?

(a) Outline the steps necessary to answer the question.
(b) Answer the question.

Solution
(a) Determine the empirical formula of the compound from the percent composition. Determine the molar mass from the mass of the sample and its volume, temperature, and pressure. Then determine the molecular formula from the molar mass and the empirical formula.
(b) Assume that a 100.0-g sample is present and the percentages translate directly to grams of the sample.

$$\text{C: } 39\ \text{g} \times \frac{1\ \text{mol C}}{12.011\ \text{g mol}^{-1}} = 3.247\ \text{mol C}$$

$$\text{H: } 16\ \text{g} \times \frac{1\ \text{mol H}}{1.00794\ \text{g mol}^{-1}} = 15.874\ \text{mol H}$$

$$\text{N: } 45\ \text{g} \times \frac{1\ \text{mol N}}{14.00674\ \text{g mol}^{-1}} = 3.213\ \text{mol N}$$

Divide each quantity by the smallest value. After rounding, the values are 1C, 5H, and 1N. The empirical formula is CH_5N. The empirical mass is CH_5N = 31. The likely arrangement of hydrogen atoms is CH_3NH_2.
Now the molar mass from the ideal gas law:

$$\begin{aligned}\text{molar mass} &= \frac{\text{mass} \times RT}{PV} \\ &= \frac{(0.157\ \text{g})(8.314\ \cancel{\text{L}}\ \cancel{\text{kPa}}\ \text{mol}^{-1}\ \cancel{\text{K}^{-1}})(295.15\ \cancel{\text{K}})}{99.5\ \cancel{\text{kPa}} \times 0.125\ \cancel{\text{L}}} \\ &= 31.0\ \text{g mol}^{-1}\end{aligned}$$

The empirical formula is the molar formula, CH_3NH_2.

Question 52-10.

A 36.0–L cylinder of a gas used for calibration of blood gas analyzers in medical laboratories contains 350 g CO_2, 805 g O_2, and 4,880 g N_2. At 25° C, what is the pressure in the cylinder in atmospheres, in torr, and in kilopascals?

Solution

Calculate the moles of each gas present and from that, calculate the pressure from the ideal gas law. Assume 25 °C. The calibration gas contains:

$$\frac{350\ \cancel{g}\ CO_2}{44.0098\ \cancel{g}\ mol^{-1}\ CO_2} = 7.953\ mol\ CO_2$$

$$\frac{805\ \cancel{g}\ O_2}{31.9988\ \cancel{g}\ mol^{-1}\ O_2} = 25.157\ mol\ O_2$$

$$\frac{4880\ \cancel{g}\ N_2}{28.01348\ \cancel{g}\ mol^{-1}\ N_2} = 174.202\ mol\ N_2$$

Total moles = 7.953 + 25.157+ 174.202 = 207.312 mol

$$P = \frac{nRT}{V} = \frac{207.312\ \cancel{mol} \times 0.08206\ \cancel{L}\ atm\ \cancel{mol^{-1}\ K^{-1}} \times 298.15\ \cancel{K}}{36.0\ \cancel{L}} = 141\ atm$$

P in torr = 107,000 torr

P in kPa = 14,300 kPa

Question 52-11.

A cylinder of a gas mixture used for calibration of blood gas analyzers in medical laboratories contains 5.0% CO_2, 12.0% O_2, and the remainder N_2 at a total pressure of 146 atm. What is the partial pressure of each component of this gas? (The percentages given indicate the percent of the total pressure that is due to each component.)

Solution

The calibration gas contains:

5.0% CO_2

12.0% O_2

100 – (5 + 12) = 83% N_2

Since these are percentages of the total pressure, the partial pressures can be calculated as follows:

CO_2: 5% of 146 atm = 0.05 ×146 = 7.3 atm = 7 atm

O_2: 12% of 146 atm = 0.12 × 146 = 17.5 atm = 18 atm

N_2: 83% of 146 atm = 0.83 × 146 = 121 atm = 1.2×10^2 atm

Question 52-12.

A sample of gas isolated from unrefined petroleum contains 90.0% CH_4, 8.9% C_2H_6, and 1.1% C_3H_8 at a total pressure of 307.2 kPa. What is the partial pressure of each component of this gas? (The percentages given indicate the percent of the total pressure that is due to each component.)

Solution

Since these are percentages of the total pressure, the partial pressure can be calculated as follows:

CH_4: 90% of 307.2 kPa = 0.900 × 307.2 = 276 kPa

C_2H_6: 8.9% of 307.2 kPa = 0.089 × 307.2 = 27 kPa

C_3H_8: 1.1% of 307.2 kPa = 0.011 × 307.2 = 3.4 kPa

Question 52-13.

A mixture of 0.200 g of H_2, 1.00 g of N_2, and 0.820 g of Ar is stored in a closed container at STP. Find the volume of the container, assuming that the gases exhibit ideal behavior.

Solution

First find the number of moles of gases present.

H_2: $\dfrac{0.200\ \cancel{g}}{2.016\ \cancel{g}\ mol^{-1}} = 0.0992\ mol$

N_2: $\dfrac{1.00\ \cancel{g}}{28.013\ \cancel{g}\ mol^{-1}} = 0.0357\ mol$

Ar: $\dfrac{0.820\ \cancel{g}}{39.948\ \cancel{g}\ mol^{-1}} = 0.0205\ mol$

The total amount of substance is 0.1555 mol.

$$V = \frac{nRT}{P} = \frac{0.155\ \cancel{mol} \times 0.08206\ L\ \cancel{atm}\ \cancel{mol^{-1}}\ \cancel{K^{-1}} \times 273.15\ \cancel{K}}{1.00\ \cancel{atm}} = 3\ .474\ L$$

Question 52-14.

Most mixtures of hydrogen gas with oxygen gas are explosive. However, a mixture that contains less than 3.0 % O_2 is not. If enough O_2 is added to a cylinder of H_2 at 33.2 atm to bring the total pressure to 34.5 atm, is the mixture explosive?

Solution

The oxygen increases the pressure within the tank to (34.5 atm – 33.2 atm =) 1.3 atm. The percentage O_2 on a mole basis is $\dfrac{1.3}{34.5} \times 100\% = 3.77\%$. The mixture is explosive. However, the percentage is given as a weight percent. Converting to a mass basis increases the percentage of oxygen even more, so the mixture is still explosive.

Question 52-15.

A commercial mercury vapor analyzer can detect, in air, concentrations of gaseous Hg atoms (which are poisonous) as low as 2×10^{-6} mg/L of air. At this concentration, what is the partial pressure of gaseous mercury if the atmospheric pressure is 733 torr at 26 °C?

Solution

The total pressure is the pressure of the normal atmospheric components plus that due to mercury. Use the ideal gas equation to calculate the total number of moles of gas required to produce the stated conditions. The pressure of mercury is the mole fraction of mercury times the total pressure:

$$\text{mol (Hg)} = 2 \times 10^{-6}\ \cancel{\text{mg}}\ \text{L}^{-1} \times \frac{1\ \cancel{\text{g}}}{1000\ \cancel{\text{mg}}} \times \frac{1\ \text{mol}}{200.59\ \cancel{\text{g}}} = 9.97 \times 10^{-12}\ \text{mol L}^{-1}$$

$$n_T = \frac{PV}{RT} = \frac{\left(\frac{733}{760\ \cancel{\text{atm}}}\right)(1.0\ \cancel{\text{L}})}{(0.08206\ \cancel{\text{L}}\ \cancel{\text{atm}}\ \text{mol}^{-1}\ \cancel{\text{K}^{-1}})(294.15\ \cancel{\text{K}})} = 0.0400\ \text{mol}$$

$$P(\text{Hg}) = \frac{9.97 \times 10^{-12}\ \cancel{\text{mol}}}{0.0400\ \cancel{\text{mol}}} \times \frac{733\ \cancel{\text{torr}}}{760\ \cancel{\text{torr}}\ \text{atm}^{-1}} = 2.4 \times 10^{-10}\ \text{atm}$$

$$= 2.4 \times 10^{-10}\ \text{atm} \times 760\ \text{torr atm}^{-1}$$

$$= 2 \times 10^{-7}\ \text{torr}$$

Question 52-16.

A sample of carbon monoxide was collected over water at a total pressure of 756 torr and a temperature of 18 °C. What is the pressure of the carbon monoxide? (See Table 9.2 for the vapor pressure of water.)

Solution

The vapor pressure of water at 18 °C is 15.5 torr. Subtract the vapor pressure of water from the total pressure to find the pressure of the carbon monoxide:

$P_T = P_{gas} + P_{water}$

Rearrangement gives:

$P_T - P_{water} = P_{gas}$

756 torr – 15.5 torr = 740 torr

Question 52-17.

In an experiment in a general chemistry laboratory, a student collected a sample of a gas over water. The volume of the gas was 265 mL at a pressure of 753 torr and a temperature of 27 °C. The mass of the gas was 0.472 g. What was the molar mass of the gas?

Solution

First, find the water vapor pressure at 27 °C and subtract it from the measured pressure. From a rearranged form of the ideal gas equation, find the molar mass of the gas:

$$\text{molar mass} = \frac{0.472\ \text{g} \times 0.08206\ \cancel{\text{L}}\ \cancel{\text{atm}}\ \text{mol}^{-1}\ \cancel{\text{K}^{-1}} \times 303.15\ \cancel{\text{K}}}{\frac{726\ \cancel{\text{torr}}}{760\ \cancel{\text{torr}}}\ \cancel{\text{atm}^{-1}} \times 0.265\ \cancel{\text{L}}} = 46.4\ \text{g mol}^{-1}$$

Question 52-18.

Joseph Priestley first prepared pure oxygen by heating mercuric oxide, HgO:

$$2\text{HgO}(s) \longrightarrow 2\text{Hg}(l) + \text{O}_2(g)$$

(a) Outline the steps necessary to answer the following question: What volume of O_2 at 23 °C and 0.975 atm is produced by the decomposition of 5.36 g of HgO?

(b) Answer the question.

Solution

(a) Determine the moles of HgO that decompose; using the chemical equation, determine the moles of O_2 produced by decomposition of this amount of HgO; and determine the volume of O_2 from the moles of O_2, temperature, and pressure.
(b)

$$5.36\ \cancel{\text{g HgO}} \times \frac{1\text{ mol HgO}}{(200.59 + 15.9994)\ \cancel{\text{g HgO}}} = 0.0247\text{ mol HgO}$$

$$0.0247\ \cancel{\text{mol HgO}} \times \frac{1\text{ mol O}_2}{2\ \cancel{\text{mol HgO}}} = 0.01235\text{ mol O}_2$$

$PV = nRT$
$P = 0.975$ atm
$T = (23.0 + 273.15)$ K

$$V = \frac{nRT}{P} = \frac{0.01235\ \cancel{\text{mol}}\ (0.08206\text{ L}\ \cancel{\text{atm}}\ \cancel{\text{mol}^{-1}}\ \cancel{\text{K}^{-1}})(296.15\ \cancel{\text{K}})}{0.975\ \cancel{\text{atm}}} = 0.308\text{ L}$$

Question 52-19.

Cavendish prepared hydrogen in 1766 by the novel method of passing steam through a red-hot gun barrel:

$$4H_2O(g) + 3Fe(s) \longrightarrow Fe_3O_4(s) + 4H_2(g)$$

(a) Outline the steps necessary to answer the following question: What volume of H_2 at a pressure of 745 torr and a temperature of 20°C can be prepared from the reaction of 15.O g of H_2O?
(b) Answer the question.

Solution
(a) Determine the moles of water in 15.0 g. Using the chemical equation, determine the moles of H_2 produced from this amount of water; then determine the volume of H_2 from the number of moles of H_2, its temperature, and its pressure.
(b)

$$15.0\ \cancel{\text{g H}_2\text{O}} \times \frac{1\ \cancel{\text{mol H}_2\text{O}}}{18.015\ \cancel{\text{g H}_2\text{O}}} \times \frac{4\text{ mol H}_2}{4\ \cancel{\text{mol H}_2\text{O}}} = 0.8326\text{ mol H}_2$$

$$V = \frac{nRT}{P} = \frac{(0.8326\ \cancel{\text{mol}})\ (0.08206\text{ L}\ \cancel{\text{atm}}\ \cancel{\text{mol}^{-1}}\ \cancel{\text{K}^{-1}})(296.15\ \cancel{\text{K}})}{\dfrac{745\ \cancel{\text{torr}}}{760\ \cancel{\text{torr}}\ \cancel{\text{atm}}}} = 20.4\text{ L}$$

Question 52-20.

The chlorofluorocarbon CCl_2F_2 can be recycled into a different compound by reaction with hydrogen to produce $CH_2F_2(g)$, a compound useful in chemical manufacturing:

$$CCl_2F_2(g) + 4H_2(g) \longrightarrow CH_2F_2(g) + 2HCl(g).$$

(a) Outline the steps necessary to answer the following question: What volume of hydrogen at 225 atm and 35.5 °C would be required to react with 1 ton (1.000×10^3 kg) of CCl_2F_2?
(b) Answer the question.

Solution

(a) Determine the molar mass of CCl_2F_2. From the balanced equation, calculate the moles of H_2 needed for the complete reaction. From the ideal gas law, convert moles of H_2 into volume.
(b) Molar mass of CCl_2F_2 = 12.011 + 2 × 18.9984 + 2 × 35.4527 = 120.913 g/mol

$$\text{mol } H_2 = 1.000 \times 10^6 \text{ g} \times \frac{1 \text{ mol } CCL_2F_2}{120.913 \text{ g}} \times \frac{4 \text{ mol } H_2}{1 \text{ mol } CCl_2F_2} = 3.308 \times 10^4 \text{ mol}$$

$$V = \frac{nRT}{P} = \frac{(3.308 \times 10^4 \cancel{\text{mol}})(0.08206 \text{ L} \cancel{\text{atm}} \cancel{\text{mol}^{-1}} \cancel{\text{K}^{-1}})(308.65 \cancel{\text{K}})}{225 \cancel{\text{atm}}} = 3.72 \times 10^3 \text{ L}$$

Question 52-21.

Automobile air bags are inflated with nitrogen gas, which is formed by the decomposition of solid sodium azide (NaN_3). The other product is sodium metal. Calculate the volume of nitrogen gas at 27 °C and 756 torr formed by the decomposition of 125 g of sodium azide.

Solution

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

First, find the moles of N_2 produced. Molar mass = NaN_3 = 65.01 g/mol.

$$\text{mol } N_2 = 125 \cancel{\text{g NaN}_3} \times \frac{1 \cancel{\text{mol NaN}_3}}{65.01 \cancel{\text{g NaN}_3}} \times \frac{3 \text{ mol } N_2}{2 \cancel{\text{mol NaN}_3}} = 2.884 \text{ mol}$$

$$V = \frac{nRT}{P} = \frac{(2.884 \cancel{\text{mol}})(0.08206 \text{ L} \cancel{\text{atm}} \cancel{\text{mol}^{-1}} \cancel{\text{K}^{-1}})(300.15 \cancel{\text{K}})}{\dfrac{756 \cancel{\text{torr}}}{760 \cancel{\text{torr}} \text{ atm}^{-1}}} = 71.4 \text{ L}$$

Question 52-22.

Lime, CaO, is produced by heating calcium carbonate, $CaCO_3$; carbon dioxide is the other product.

(a) Outline the steps necessary to answer the following question: What volume of carbon dioxide at 875 K and 0.966 atm is produced by the decomposition of 1 ton (1.000 × 10^3 kg) of calcium carbonate?
(b) Answer the question.

Solution

(a) Balance the equation. Determine the grams of CO_2 produced and the number of moles. From the ideal gas law, determine the volume of gas.

(b) $CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$

$$\text{mass } CO_2 = 1.00 \times 10^6 \text{ g} \times \frac{1 \text{ mol } CaCO_2}{100.087 \text{ g}} \times \frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2} \times \frac{1 \text{ mol } CO_2}{1 \text{ mol } CaCO_2} = 4.397 \times 10^5 \text{ g}$$

$$\text{mol } CO_2 = \frac{4.397 \times 10^5 \text{ g}}{44.01 \text{ g mol}^{-1}} = 9991 \text{ mol}$$

$$V = \frac{nRT}{P} = \frac{(9991 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(875 \text{ K})}{0.966 \text{ atm}} = 7.43 \times 10^5 \text{ L}$$

Question 52-23.

Before small batteries were available, carbide lamps were used for bicycle lights. Acetylene gas, C_2H_2, and solid calcium hydroxide were formed by the reaction of calcium carbide, CaC_2, with water. The ignition of the acetylene gas provided the light. Currently, the same lamps are used by some cavers, and calcium carbide is used to produce acetylene for carbide cannons.

(a) Outline the steps necessary to answer the following question: What volume of C_2H_2 at 1.005 atm and 12.2 °C is formed by the reaction of 15.48 g of CaC_2 with water?

(b) Answer the question.

Solution

(a) First, write a balanced equation for the reaction. Then determine the moles of CaC_2 in 15.48 g. From the equation, determine the moles of C_2H_2 produced, and, from the ideal gas law, find the volume of the gas.

(b)

$$CaC_2 + 2H_2O \longrightarrow C_2H_2 + Ca(OH)_2$$

Molar mass of CaC_2 = (40.078 + 2 × 12.011) = 64.100 g

$$\text{mol } CaC_2 = \frac{15.48\text{ g}}{64.100\text{ g mol}^{-1}} = 0.2415\text{ mol}$$

From the balanced equation, 1 mol CaC_2 produces 1 mol C_2H_2. Consequently, 0.2415 mol of C_2H_2 is present.

$$V = \frac{nRT}{P} = \frac{(0.2415\text{ mol})(0.08206\text{ L atm mol}^{-1}\text{ K}^{-1})(285.35\text{ K})}{1.005\text{ atm}} = 5.627\text{ L}$$

Question 52-24.

Calculate the volume of oxygen required to burn 12.00 L of ethane gas, C_2H_6, to produce carbon dioxide and water, if the volumes of C_2H_6 and O_2 are measured under the same conditions of temperature and pressure.

Solution

$$2C_2H_6(g) + 7O_2(g) \longrightarrow 4CO_2(g) + 6H_2O(g)$$

From the balanced equation, we see that 2 mol of C_2H_6 requires 7 mol of O_2 to burn completely. Gay-Lussac's law states that gases react in simple proportions by volume. As the number of liters is proportional to the number of moles,

$$\frac{12.00\text{ L}}{2\text{ mol } C_2H_6} = \frac{V(O_2)}{7\text{ mol } O_2}$$

$$V(O_2) = \frac{12.00\text{ L} \times 7}{2} = 42.00\text{ L}$$

Question 52-25.

What volume of O_2 at STP is required to oxidize 8.0 L of NO at STP to NO_2? What volume of NO_2 is produced at STP?

Solution

Write the balanced equation for the reaction:

$$2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$$

Here, 2 mol of NO requires 1 mol of O_2 and 2 mol of NO_2 is produced. Gay-Lussac's law states that gases react in simple proportions by volume. Therefore, 8.0 L of NO requires 4.0 L of O_2 for reaction to produce 8.0 L of NO_2.

Question 52-26.

Consider the following questions:

(a) What is the total volume of the $CO_2(g)$ and $H_2O(g)$ at 600 °C and 0.888 atm produced by the combustion of 1.00 L of $C_2H_6(g)$ measured at STP?

(b) What is the partial pressure of H_2O in the product gases?

Solution

(a) The scheme to solve this problem is:

$$\text{volume } C_2H_6(g) \xrightarrow[\text{equation}]{\text{ideal gas}} \text{mol } C_2H_6(g) \xrightarrow[\text{stoichiometry}]{\text{reaction}} \text{mol } CO_2 + H_2O$$

$$\xrightarrow[\text{equation}]{\text{ideal gas}} \text{volume } CO_2 + H_2O$$

$$C_2H_6(g) + 3\frac{1}{2}O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(g)$$

$$1.\ n(C_2H_6) = \frac{PV}{RT} = \frac{1.00\ \cancel{\text{atm}} \times 1.00\ \cancel{\text{L}}}{0.08206\ \cancel{\text{L}}\ \cancel{\text{atm}}\ \text{mol}^{-1}\ \cancel{\text{K}^{-1}}(273.15\ \cancel{\text{K}})} = 0.0446\ \text{mol}$$

$$2.\ 0.0446\ \cancel{\text{mol } C_2H_6} \times \frac{5\ \text{mol products}}{1\ \cancel{\text{mol } C_2H_6}} = 0.223\ \text{mol products}$$

$$3.\ V = nRT = \frac{(0.223\ \cancel{\text{mol}})(0.08206\ \text{L}\ \cancel{\text{atm}}\ \cancel{\text{mol}^{-1}}\ \cancel{\text{K}^{-1}})(873.15\ \cancel{\text{K}})}{0.888\ \cancel{\text{atm}}} = 18.0\ \text{L}$$

(b) First, calculate the mol H_2O produced:

$$0.0446\ \text{mol } C_2H_6 \times \frac{3\ \text{mol products}}{1\ \text{mol } C_2H_6} = 0.1338\ \text{mol}$$

Second, calculate the pressure of H_2O:

$$P = \frac{nRT}{V} = \frac{(0.1338\ \cancel{\text{mol}})(0.08206\ \cancel{\text{L}}\ \text{atm}\ \cancel{\text{mol}^{-1}}\ \cancel{\text{K}^{-1}})(873.15\ \cancel{\text{K}})}{18.0\ \cancel{\text{L}}} = 0.533\ \text{atm}$$

Question 52-27.

Methanol, CH_3OH, is produced industrially by the following reaction:

$$CO(g) + 2H_2(g) \xrightarrow{\text{copper catalyst } 300\ ^\circ\text{C},\ 300\ \text{atm}} CH_3OH(g)$$

Assuming that the gases behave as ideal gases, find the ratio of the total volume of the reactants to the final volume.

Solution

In this reaction, 3 mol of gas forms 1 mol of gas.

$$PV = nRT$$

$$\frac{V_1}{n_1} = \frac{RT}{P} = \frac{V_2}{n_2}$$

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

This tells us that the value, at constant P and T, is directly proportional to the moles of gas. The volume ratio of reactants to final volume, V_1/V_2, will be equal to the mole ratio of reactants to products:

$$\frac{n_1}{n_2} = \frac{3}{1}$$

The ratio is 3:1.

Question 52-28.

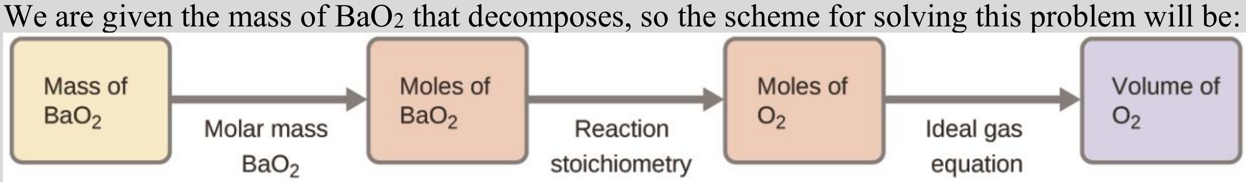
What volume of oxygen at 423.0 K and a pressure of 127.4 kPa is produced by the decomposition of 129.7 g of BaO_2 to BaO and O_2?

Solution

First, we must write a balanced equation to establish the stoichiometry of the reaction:

$$2BaO_2 \longrightarrow 2BaO + O_2$$

We are given the mass of BaO_2 that decomposes, so the scheme for solving this problem will be:



Mass (BaO_2) = 137.33 + 2(15.9994) = 169.33 g/mol

$$n(O_2) = 129.7 \text{ g } BaO_2 \times \frac{1 \text{ mol } BaO_2}{169.33 \text{ g } BaO_2} \times \frac{1 \text{ mol } O_2}{2 \text{ mol } BaO_2} = 0.3830 \text{ mol } O_2$$

$$V(O_2) = \frac{nRT}{P} = \frac{0.3830 \text{ mol}(8.314 \text{ L kPa mol}^{-1} \text{ K}^{-1})(423.0 \text{ K})}{127.4 \text{ kPa}} = 10.57 \text{ L } O_2$$

Question 52-29.

A 2.50-L sample of a colorless gas at STP decomposed to give 2.50 L of N_2 and 1.25 L of O_2 at STP. What is the colorless gas?

Solution

Because gas volumes are directly related to the number of moles present, the production of 2.50 L of N_2 indicates that two nitrogen atoms occur in the compound. The formation of 1.25 L of O_2 gives one-half the original volume, indicating that only one oxygen atom occurs in the original compound. The compound is N_2O.

Question 52-30.

Ethanol, C_2H_5OH, is produced industrially from ethylene, C_2H_4, by the following sequence of reactions:

$$3C_2H_4 + 2H_2SO_4 \longrightarrow C_2H_5HSO_4 + (C_2H_5)_2SO_4$$

$$C_2H_5HSO_4 + (C_2H_5)_2SO_4 + 3H_2O \longrightarrow 3C_2H_5OH + 2H_2SO_4$$

What volume of ethylene at STP is required to produce 1.000 metric ton (1000 kg) of ethanol if the overall yield of ethanol is 90.1%?

Solution

At 90.1% conversion, a 1.000×10^6 g final yield would require a $\left(\frac{1.000 \times 10^6}{0.901}\right) = 1.1099 \times 10^6$ g theoretical yield.

$3C_2H_4$ produces $3C_2H_5OH$, giving a 1:1 ratio:

$$\text{mol}\ (C_2H_4) = 1.1099 \times 10^6\ \cancel{\text{g } C_2H_5OH} \times \frac{1\ \cancel{\text{mol } C_2H_5OH}}{46.069\ \cancel{\text{g } C_2H_5OH}} \times \frac{1\ \text{mol } C_2H_4}{1\ \cancel{\text{mol } C_2H_5OH}}$$

$$= 2.409 \times 10^4\ \text{mol}$$

$V(C_2H_4) = 22.4$ L/mol$\times$ 2.409×10^4 mol $= 5.40 \times 10^5$ L

Question 52-31.

One molecule of hemoglobin will combine with four molecules of oxygen. If 1.0 g of hemoglobin combines with 1.53 mL of oxygen at body temperature (37 °C) and a pressure of 743 torr, what is the molar mass of hemoglobin?

Solution

The oxygen reacts on a 1:4 mol basis. First, find the number of moles of O_2 from the ideal gas law; then, divide by 4.

$$n = \frac{PV}{RT} = \frac{\left(\frac{743\ \cancel{\text{torr}}}{763\ \cancel{\text{torr}}\ \cancel{\text{atm}^{-1}}}\right)(1.53\ \cancel{\text{mL}})\left(\frac{1\ \cancel{\text{L}}}{1000\ \cancel{\text{mL}}}\right)}{(0.08206\ \cancel{\text{L}}\ \cancel{\text{atm}^{-1}}\ \text{mol}^{-1}\ \cancel{\text{K}^{-1}})\ (310.15\ \cancel{\text{K}})} = 5.854 \times 10^{-5}\ \text{mol}$$

$$\frac{1}{4}(5.854 \times 10^{-5}\ \text{mol}) = 1.464 \times 10^{-5}\ \text{mol hemoglobin}$$

Divide 1.0 g by the corresponding number of moles to find the molar mass.

$$\frac{1.0\ \text{g}}{1.464 \times 10^{-5}\ \text{mol}} = 6.8 \times 10^4\ \text{g mol}^{-1}$$

Question 52-32.

A sample of a compound of xenon and fluorine was confined in a bulb with a pressure of 18 torr. Hydrogen was added to the bulb until the pressure was 72 torr. Passage of an electric spark through the mixture produced Xe and HF. After the HF was removed by reaction with solid KOH, the final pressure of xenon and unreacted hydrogen in the bulb was 36 torr. What is the empirical formula of the xenon fluoride in the original sample? (Note: Xenon fluorides contain only one xenon atom per molecule.)

Solution

The reaction is:

$$XeF_x + \frac{x}{2}H_2 \longrightarrow Xe + xHF$$

Immediately after the H_2 is added (before the reaction):

$$P_{\text{Total}} = P_{\text{XeF}_2} + P_{\text{H}_2}$$

$$\begin{aligned} P_{\text{H}_2} &= P_{\text{Total}} - P_{\text{XeF}_2} \\ &= 72 \text{ torr} - 18 \text{ torr} \\ &= 54 \text{ torr} \end{aligned}$$

After the reaction:

$P_{\text{Xe}} = 18$ torr (1 mol $\text{XeF}_x \longrightarrow$ 1 mol Xe)

And the partial pressure of unreacted H_2 is:

$$\begin{aligned} P_{\text{H}_2} &= P_{\text{Total}} - P_{\text{Xe}} \\ &= 36 \text{ torr} - 18 \text{ torr} \\ &= 18 \text{ torr} \end{aligned}$$

The pressure of H_2 that reacts is:

54 torr – 18 torr = 36 torr

The number of moles of gas is proportional to the partial pressures. The reaction used 18 torr of XeF_x and 36 torr of H_2 so:

$$\frac{\text{mol H}_2}{\text{mol XeF}_x} = \frac{x/2}{1} = \frac{x}{2} = \frac{36 \text{ torr}}{18 \text{ torr}} \longrightarrow x = \frac{72 \text{ torr}}{18 \text{ torr}} = 4$$

The empirical formula for the compound is XeF_4.

Question 52-33.

One method of analyzing amino acids is the van Slyke method. The characteristic aminogroups ($-NH_2$) in protein material are allowed to react with nitrous acid, HNO_2, to form N_2 gas. From the volume of the gas, the amount of amino acid can be determined. A 0.0604-g sample of a biological sample containing glycine, $CH_2(NH_2)COOH$, was analyzed by the van Slyke method and yielded 3.70 mL of N_2 collected over water at a pressure of 735 torr and 29 °C. What was the percentage of glycine in the sample?

$$\text{CH}_2(\text{NH}_2)\text{CO}_2\text{H} + \text{HNO}_2 \longrightarrow \text{CH}_2(\text{OH})\text{CO}_2\text{H} + \text{H}_2\text{O} + \text{N}_2$$

Solution

$$\text{volume N}_2 \text{ moist} \xrightarrow[\text{partial pressure}]{\text{(1) law of}} \xrightarrow[\text{gas equation}]{\text{(2) ideal}} \text{mol N}_2 \xrightarrow[\text{stoichiometry}]{\text{(3) reaction}} \text{mol glycine} \xrightarrow{(4)} \text{wt \% glycine in sample}$$

The N_2 was collected over water, so the partial pressure of water at 29 °C, which is 30.0 torr, must be taken into account:

$$P_{\text{T}} = P_{\text{N}_2} + P_{\text{H}_2\text{O}}$$

$$P_{\text{N}_2} = P_{\text{T}} - P_{\text{H}_2\text{O}} = (735 - 30.0) \text{ torr} = 705 \text{ torr}$$

$$2.\ n(\text{N}_2) = \frac{PV}{RT} = \frac{\left(\frac{705}{760}\right) \cancel{\text{atm}}\ (0.00370\ \cancel{\text{L}})}{(0.08206\ \cancel{\text{L}}\ \cancel{\text{atm}}\ \text{mol}^{-1}\ \cancel{\text{K}^{-1}})\ (302.15\ \cancel{\text{K}})} = 1.38 \times 10^{-4} \text{ mol}$$

$$3.\ 1.38 \times 10^{-4}\ \cancel{\text{mol N}_2} \times \frac{1\ \text{mol glycine}}{1\ \cancel{\text{mol N}_2}} = 1.38 \times 10^{-4}\ \text{glycine}$$

$$4.\ M_{glycine} = \begin{array}{r} 2(12.011) \\ 5(1.00794) \\ 2(15.9994) \\ \underline{(14.0067)} \\ 75.067\ \text{g mol}^{-1} \end{array}$$

$$\text{mass glycine} = 1.38 \times 10^{-4}\ \cancel{\text{mol}} \times \frac{75.067\ \text{g}}{\cancel{\text{mol}}} = 1.04 \times 10^{-2}\ \text{g}$$

$$\text{wt \% glycine} = \frac{\text{mass glycine}}{\text{mass sample}} \times 100\% = \frac{1.04 \times 10^{-2}\ \cancel{\text{g}}}{0.0604\ \cancel{\text{g}}} \times 100\% = 17.2\%$$

9.4 Effusion and Diffusion of Gases

Question 53-1.

A balloon filled with helium gas takes 6 hours to deflate to 50% of its original volume. How long will it take for an identical balloon filled with the same volume of hydrogen gas (instead of helium) to decrease its volume by 50%?

Solution

Use the rate of effusion equation:

$$\frac{6\ \text{hours}}{t} = \sqrt{\frac{4}{2}}$$

$$t = \frac{6\ \text{hours}}{1.4} = 4.2\ \text{hours}$$

Question 53-2.

Explain why the numbers of molecules are not identical in the left- and right-hand bulbs shown in the center illustration of Figure 9.27.

Solution

When the stopcock is opened, the faster-moving H_2 will effuse through the small opening faster than the slower-moving O_2. Thus, in a unit of time, more hydrogen will effuse through the hole into the oxygen compartment faster than oxygen will move into the hydrogen compartment.

Question 53-3.

Starting with the definition of rate of effusion and Graham's finding relating rate and molar mass, show how to derive the Graham's law equation, relating the relative rates of effusion for two gases to their molecular masses.

Solution

Effusion can be defined as the process by which a gas escapes through a pinhole into a vacuum. Graham's law states that with a mixture of two gases A and B:

$\left(\dfrac{\text{rate A}}{\text{rate B}}\right) = \left(\dfrac{\text{molar mass of B}}{\text{molar mass of A}}\right)^{1/2}$. Both A and B are in the same container at the same temperature and, therefore, will have the same kinetic energy:

$KE_A = KE_B$

$$KE = \frac{1}{2}mv^2$$

Therefore, $\frac{1}{2}m_A v_A^2 = \frac{1}{2}m_B v_B^2$

$$\frac{v_A^2}{v_B^2} = \frac{m_B}{m_A}$$

$$\left(\frac{v_A^2}{v_B^2}\right)^{1/2} = \left(\frac{m_B}{m_A}\right)^{1/2}$$

$$\frac{v_A}{v_B} = \left(\frac{m_B}{m_A}\right)^{1/2}$$

Question 53-4.

Heavy water, D_2O (molar mass = 20.03 g mol^{-1}), can be separated from ordinary water, H_2O (molar mass = 18.01), as a result of the difference in the relative rates of diffusion of the molecules in the gas phase. Calculate the relative rates of diffusion of H_2O and D_2O.

Solution

In this problem, we use Graham's law for rates of diffusion of gases:

$$\frac{R_1}{R_2} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

$$\frac{R_{H_2O}}{R_{D_2O}} = \frac{\sqrt{M_{D_2O}}}{\sqrt{M_{H_2O}}} = \frac{\sqrt{20.03}}{\sqrt{18.01}} = 1.055$$

Therefore, H_2O diffuses 1.055-times faster than does D_2O.

Question 53-5.

Which of the following gases diffuse more slowly than oxygen? F_2, Ne, N_2O, C_2H_2, NO, Cl_2, H_2S

Solution

Gases with molecular masses greater than that of oxygen (31.9988 g/mol) will diffuse more slowly than O_2. These gases are F_2 (37.9968 g/mol), N_2O (44.0128 g/mol), Cl_2 (70.906 g/mol), and H_2S (34.082 g/mol).

Question 53-6.

During the discussion of gaseous diffusion for enriching uranium, it was claimed that $^{235}UF_6$ diffuses 0.4% faster than $^{238}UF_6$. Show the calculation that supports this value. The molar mass of $^{235}UF_6$ = 235.043930 + 6 × 18.998403 = 349.034348 g/mol, and the molar mass of $^{238}UF_6$ = 238.050788 + 6 × 18.998403 = 352.041206 g/mol.

Solution

The diffusion rate for ^{235}U is proportional to $\dfrac{1}{\sqrt{349.034348}}$ = 0.053526139, whereas the diffusion rate for ^{238}U is proportional to $\dfrac{1}{\sqrt{352.041206}}$ = 0.05329706.

$\dfrac{\text{diffusion rate for } ^{235}\text{U}}{\text{diffusion rate for } ^{238}\text{U}} = \dfrac{0.053526139}{0.05329706} = 1.0043$. Therefore, the rate for ^{235}U is 0.43% faster.

Question 53-7.

Calculate the relative rate of diffusion of $^1H_2$ (molar mass 2.0 g/mol) compared with that of $^2H_2$ (molar mass 4.0 g/mol) and the relative rate of diffusion of O_2 (molar mass 32 g/mol) compared to that of O_3 (molar mass 48 g/mol).

Solution

$$\frac{R_{\mathrm{H_2}}}{R_{\mathrm{D_2}}} = \frac{\sqrt{M_{\mathrm{D_2}}}}{\sqrt{M_{\mathrm{H_2}}}} = \frac{\sqrt{4.0}}{\sqrt{2.0}} = \frac{2.000}{1.414} = 1.4$$

$$\frac{R_{\mathrm{O_2}}}{R_{\mathrm{O_3}}} = \frac{\sqrt{M_{\mathrm{O_3}}}}{\sqrt{M_{\mathrm{O_2}}}} = \frac{\sqrt{48}}{\sqrt{32}} = 1.2$$

Question 53-8.

A gas of unknown identity diffuses at a rate of 83.3 mL/s in a diffusion apparatus in which carbon dioxide diffuses at the rate of 102 mL/s. Calculate the molecular mass of the unknown gas.

Solution

According to Graham's law of effusion, the rate of diffusion is inversely proportional to the square root of its molar mass or molecular mass. Thus,

$$\frac{R_1}{R_2} = \frac{\sqrt{\text{molar mass}_1}}{\sqrt{\text{molar mass}_1}}$$

$$\frac{83.3\text{ mL/s}}{102\text{ mL/s}} = \frac{\sqrt{44}}{\sqrt{\text{molar mass}_1}}$$

It is perhaps easiest to square both sides in this case and then solve for molar mass$_1$. Thus,

$$\text{molar mass}_1 = 44.0\text{ g/mol} \times \frac{(102\text{ mL s}^{-1})^2}{(83.3\text{ mL s}^{-1})^2} = 44.0 \times \frac{102^2\text{ g}}{83.3^2\text{ g mol}} = 66.0\text{ g mol}^{-1}$$

Question 53-9.

When two cotton plugs, one moistened with ammonia and the other with hydrochloric acid, are simultaneously inserted into opposite ends of a glass tube that is 87.0 cm long, a white ring of NH_4Cl forms where gaseous NH_3 and gaseous HCl first come into contact. (Hint: Calculate the rates of diffusion for both NH_3 and HCl, and find out how much faster NH_3 diffuses than HCl.)
$NH_3(g) + HCl(g) \longrightarrow NH_4Cl(s)$ At approximately what distance from the ammonia-moistened plug does this occur?

Solution

Rate of diffusion for NH_3 is proportional to $\dfrac{1}{17.04^{1/2}} = 0.242250792$

Rate of diffusion for HCl is proportional to $\dfrac{1}{36.46^{1/2}} = 0.165611949$, $\left(\dfrac{0.242250792}{0.165611949}\right) = 1.4627$.

Set up an algebraic expression, letting x represent the distance travelled by the HCl: $x + 1.4627x = 87$, $x = 35.3$, so the distance travelled by the NH_3 is $(1.4627)x = 51.7$ cm.

9.5 The Kinetic Molecular Theory

Question 54-1.

Using the postulates of the kinetic-molecular theory, explain why a gas uniformly fills a container of any shape.

Solution

From the first postulate, the gas molecules travel in straight lines and change direction only when they collide. With a large sample of gas, many collisions occur with the walls of the container as well as between gas molecules. These events are completely random. Consequently, the direction of motion of the gases becomes completely random.

Question 54-2.

Can the speed of a given molecule in a gas double at constant temperature? Explain your answer.

Solution

Yes. At any given instant, there are a range of values of molecular speeds in a sample of gas. Any single molecule can speed up or slow down as it collides with other molecules. The average speed of all the molecules is constant at constant temperature.

Question 54-3.

Describe what happens to the average kinetic energy of ideal gas molecules when the conditions are changed as follows:

(a) The pressure of the gas is increased by reducing the volume at constant temperature.
(b) The pressure of the gas is increased by increasing the temperature at constant volume.
(c) The average speed of the molecules is increased by a factor of 2.

Solution

(a) The average kinetic energy remains the same at constant temperature. (b) A temperature increase corresponds to an increase in kinetic energy. (c) As the average speed is doubled, a corresponding increase must occur in temperature and, therefore, also in kinetic energy.

Question 54-4.

The distribution of molecular speeds in a sample of helium is shown in Figure 9.34. If the sample is cooled, will the distribution of speeds look more like that of H_2 or of H_2O? Explain your answer.

Solution

H_2O. Cooling slows the speeds of the He atoms, causing them to behave as though they were heavier.

Question 54-5.

What is the ratio of the average kinetic energy of a SO_2 molecule to that of an O_2 molecule in a mixture of two gases? What is the ratio of the root mean square speeds, u_{rms}, of the two gases?

Solution

According to the kinetic-molecular theory, the kinetic energies of the two gases are equal as long as they remain at the same temperature, as in this case. Because their kinetic energies are equal, we can write:

$$\frac{u_{\text{rms}}(\text{SO}_2)}{u_{\text{rms}}(\text{O}_2)} = \frac{\sqrt{M_{\text{O}_2}}}{\sqrt{M_{\text{SO}_2}}}$$

$$\frac{u_{\text{rms}}(\text{SO}_2)}{u_{\text{rms}}(\text{O}_2)} = \frac{\sqrt{M_{\text{O}_2}}}{\sqrt{M_{\text{SO}_2}}} = \frac{\sqrt{31.9988}}{\sqrt{64.0648}} = \frac{5.6567}{8.0040} = 0.70673$$

Thus, the average kinetic energies are equal (1:1 ratio), and the ratio is $\frac{u_{\text{SO}_2}}{u_{\text{O}_2}} = 0.70673$.

Question 54-6.

A 1-L sample of CO initially at STP is heated to 546 K, and its volume is increased to 2 L.

(a) What effect do these changes have on the pressure exerted by the gas?
(b) What is the effect on the average kinetic energy of the molecules?
(c) What is the effect on the root mean square speed of the molecules?

Solution

Both the temperature and the volume are doubled for this gas (n constant), so P remains constant. (a) The pressure of the gas remains constant. (b) The average kinetic energy doubles; it is proportional to temperature. (c) The root mean square speed increases to $\sqrt{2}$ times its initial value; u_{rms} is proportional to $\sqrt{\text{KE}_{\text{avg}}}$.

Question 54-7.

The root mean square speed of H_2 molecules at 25 °C is about 1.6 km/s. What is the root mean square speed of a N_2 molecule at 25 °C?

Solution

As the temperature and, thus, the kinetic energy remain the same,

$$\frac{u_{rms}(H_2)}{u_{rms}(N_2)} = \frac{\sqrt{M_{N_2}}}{\sqrt{M_{H_2}}} = \frac{\sqrt{2.01588}}{\sqrt{28.01348}} = \frac{1.4198169}{5.2927762} = 0.2682556$$

$u_{rms}(N_2)$ = 1.6 km/s × 0.2683 = 0.43 km/s

Question 54-8.

Answer the following questions:

(a) Is the pressure of the gas in the hot-air balloon shown at the opening of this chapter greater than, less than, or equal to that of the atmosphere outside the balloon?

(b) Is the density of the gas in the hot-air balloon shown at the opening of this chapter greater than, less than, or equal to that of the atmosphere outside the balloon?

(c) At a pressure of 1 atm and a temperature of 20 °C, dry air has a density of 1.2256 g/L. What is the (average) molar mass of dry air?

(d) The average temperature of the gas in a hot-air balloon is 1.30×10^2 °F. Calculate its density, assuming the molar mass equals that of dry air.

(e) The lifting capacity of a hot-air balloon is equal to the difference in the mass of the cool air displaced by the balloon and the mass of the gas in the balloon. What is the difference in the mass of 1.00 L of the cool air in part (c) and the hot air in part (d)?

(f) An average balloon has a diameter of 60 feet and a volume of 1.1×10^5 ft^3. What is the lifting power of such a balloon? If the weight of the balloon and its rigging is 500 pounds, what is its capacity for carrying passengers and cargo?

(g) A balloon carries 40.0 gallons of liquid propane (density 0.5005 g/L). What volume of CO_2 and H_2O gas is produced by the combustion of this propane at STP?

(h) A balloon flight can last about 90 minutes. If all of the fuel is burned during this time, what is the approximate rate of heat loss (in kJ/min) from the hot air in the bag during the flight?

Solution

(a) equal, because the balloon is free to expand until the pressures are equalized; (b) less than the density outside; (c) assume three-place accuracy throughout unless greater accuracy is stated:

$$\text{molar mass} = \frac{DRT}{P} = 1.2256 \text{ g } \cancel{\text{L}^{-1}} \times \frac{0.08206 \cancel{\text{ L}} \cancel{\text{ atm}} \text{ mol}^{-1} \cancel{\text{K}^{-1}} \times 293.15 \cancel{\text{ K}}}{1.00 \cancel{\text{ atm}}};$$

$$= 29.48 \text{ g mol}^{-1}$$

(d) convert the temperature to °C; then use the ideal gas law:

$$°\text{C} = \frac{5}{9}(\text{F} - 32) = \frac{5}{9}(130 - 32) = 54.44 \text{ °C} = 327.6 \text{ K}$$

$$D = \frac{\text{M}P}{RT} = 29.48 \text{ g } \cancel{\text{mol}^{-1}} \times \frac{1.00 \cancel{\text{ atm}}}{0.08206 \text{ L} \cancel{\text{ atm}} \cancel{\text{ mol}^{-1}} \cancel{\text{ K}^{-1}} \times 327.6 \cancel{\text{ K}}} = 1.0966 \text{ g L}^{-1};$$

(e) 1.2256 g/L – 1.09966 g/L = 0.129 g/L; (f) calculate the volume in liters, multiply the volume by the density difference to find the lifting capacity of the balloon, subtract the weight of the balloon after converting to pounds:

$$1.1 \times 105 \text{ ft}^3 \times \left(\frac{12\text{in}}{91\text{ft}}\right)^3 \times \left(\frac{2.54\text{cm}}{\text{in}}\right)^3 \times \frac{1 \text{ L}}{1000 \text{ cm}^3} = 3.11 \times 10^6 \text{ L}$$

3.11 × 106 L × 0.129 g/L = 4.01×10^5 g

$$\frac{4.01\ 10^5\,\text{g}}{453.59\ \text{g lb}^{-1}} = 884\ \text{lb};\ 884\ \text{lb} - 500\ \text{lb} = 384\ \text{lb}$$

net lifting capacity = 384 lb; (g) First, find the mass of propane contained in 40.0 gal. Then calculate the moles of $CO_2(g)$ and $H_2O(g)$ produced from the balanced equation.

$$40.0\ \cancel{\text{gal}} \times \frac{4(0.9463\ \text{L})}{1\ \cancel{\text{gal}}} = 151.4\ \text{L}$$

$151.4\ \text{L} \times 0.5005\ \text{g L}^{-1} = 75.8\text{g}$

Molar mass of propane = 3(12.011) + 8(1.00794) = 36.033 + 8.064 = 44.097 g mol^{-1}

$$\frac{75.8\ \cancel{\text{g}}}{44.097\ \cancel{\text{g}}\ \text{mol}^{-1}} = 1.72\ \text{mol}$$

The reaction is $C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(g)$

For each 1.72 mol propane, there are 3 × 1.72 mol = 5.15 mol of CO_2 and 4 × 1.72 mol = 6.88 mol H_2O. The total volume at STP = 22.4 L × 12.04 = 270 L; (h) The total heat released is determined from the heat of combustion of the propane. Using the equation in part (g),

$$\begin{aligned}\Delta H^\circ_{\text{combustion}} &= 3\Delta H^\circ_{CO_2(g)} + 4\Delta H^\circ_{H_2O(g)} - \Delta H^\circ_{\text{propane}} \\ &= 3(-393.51) + 4(-241.82) - (-103.85) \\ &= -1180.52 - 967.28 + 103.85 = -2043.96\ \text{kJ mol}^{-1}\end{aligned}$$

Since there is 1.72 mol propane, 1.72 × 2043.96 kJ mol^{-1} = 3.52 × 10^3 kJ is used for heating.

This heat is used over 90 minutes, so $\dfrac{3.52 \times 10^3\ \text{kJ}}{90\ \text{min}} = 39.1\ \text{kJ min}^{-1}$ is released.

Question 54-9.

Show that the ratio of the rate of diffusion of Gas 1 to the rate of diffusion of Gas 2, $\dfrac{R_1}{R_2}$, is the same at 0 °C and 100 °C.

Solution

The rate at which a gas will diffuse, *R,* is proportional to u_{rms}, the root mean square speed of its molecules. The square of this value, in turn, is proportional to the average kinetic energy. The average kinetic energy is:

$\overline{\text{KE}}_{\text{avg}} = kT$.

For two different gases, 1 and 2, the constant of proportionality can be represented as k_1 and k_2, respectively. Thus,

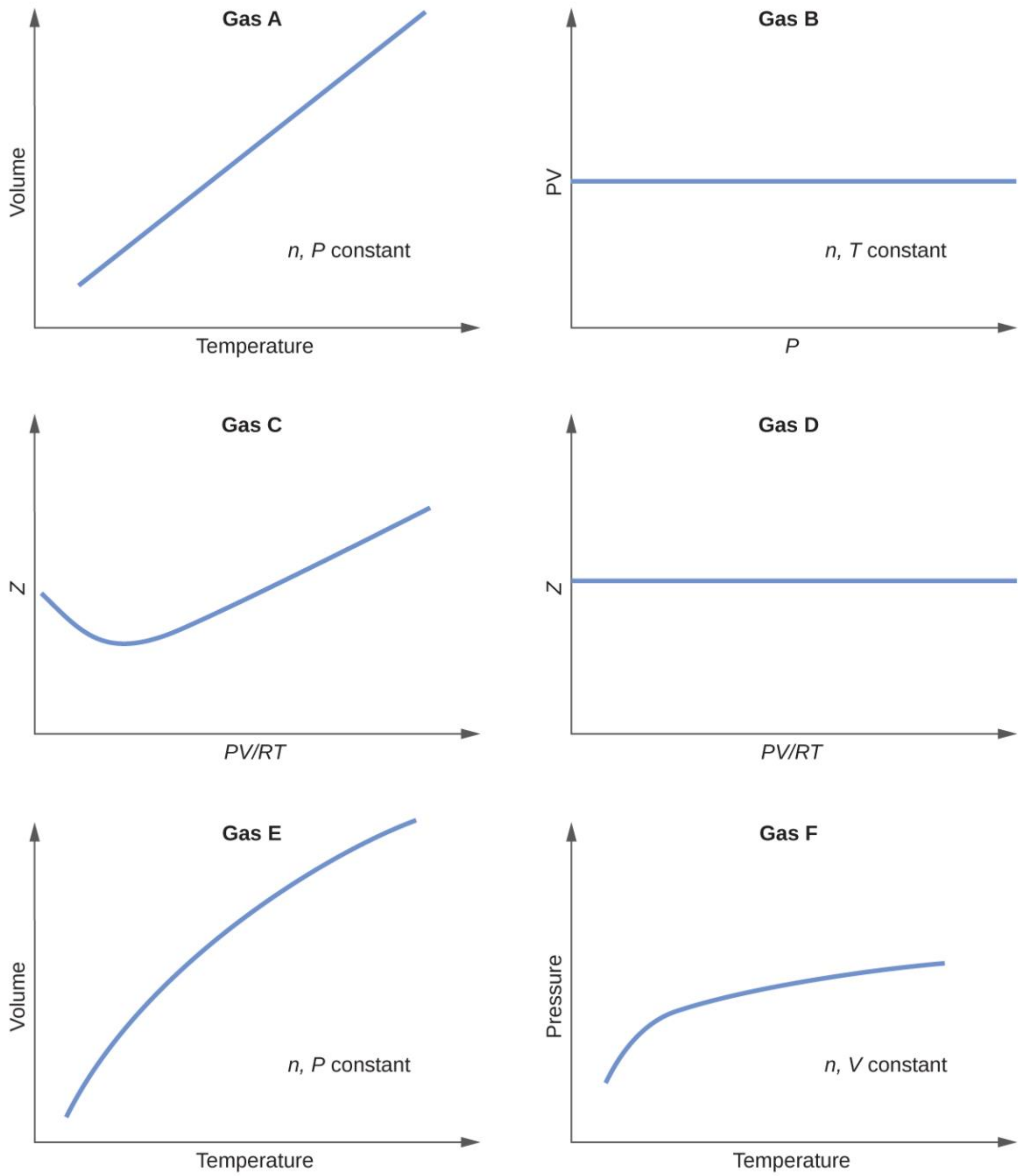
$$\frac{R_1}{R_2} = \frac{k_1\sqrt{T}}{k_2\sqrt{T}}.$$

As a result of this relationship, no matter at which temperature diffusion occurs, the temperature term will cancel out of the equation and the ratio of rates will be the same.

9.6 Non Ideal Gas Behavior

Question 55-1.

Graphs showing the behavior of several different gases follow. Which of these gases exhibit behavior significantly different from that expected for ideal gases?



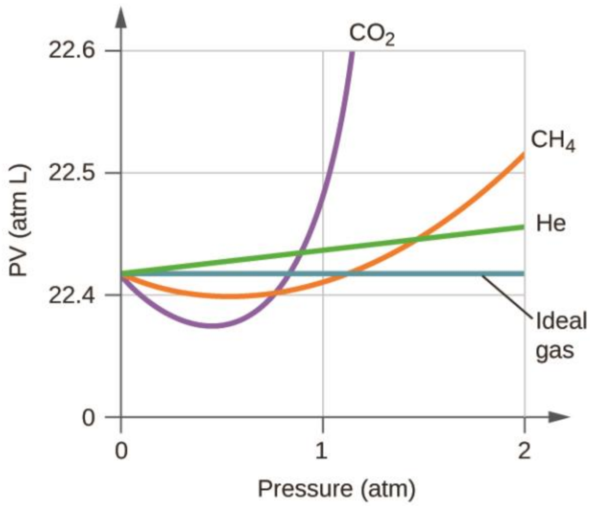
Solution

Gas A: volume increases linearly as temperature increases with moles and pressure held constant, as expected by the ideal gas law $V = (\boldsymbol{nR}/\boldsymbol{P})T$; Gas B: PV stays constant as pressure increases with moles and temperature held constant, as expected by the ideal gas law $PV = \boldsymbol{nRT}$;

Gas C: compressibility factor (Z) varies as *PV*/*RT* increases, as expected of a real gas; Gas D: compressibility factor (Z) stays constant as *PV*/*RT* increases with moles and pressure held constant, as expected of an ideal gas; Gas E: as temperature increases, volume increases, but not linearly with moles and pressure held constant, as would *not* be expected by the ideal gas law $V = (\boldsymbol{nR}/\boldsymbol{P})T$, as seen in Gas A; Gas F: as temperature increases, pressure increases with moles and volume held constant, but not linearly, as would *not* be expected by the ideal gas law $P = (\boldsymbol{nR}/V)T$, as seen in Gas A; Gases C, E, and F exhibit behavior significantly different from that expected for an ideal gas.

Question 55-2.

Explain why the plot of *PV* for CO_2 differs from that of an ideal gas.



Solution
All real gases have finite volume, which shows itself as larger than ideal volume (and consequently *PV* value) at high pressures. The gases that are capable of van der Waals forces (CO_2, CH_4) also show a smaller than ideal volume at some pressures due to the attraction between molecules.

Question 55-3.

Under which of the following sets of conditions does a real gas behave most like an ideal gas, and for which conditions is a real gas expected to deviate from ideal behavior? Explain.

(a) high pressure, small volume

(b) high temperature, low pressure

(c) low temperature, high pressure

Solution
The gas behavior most like an ideal gas will occur under the conditions that minimize the chances of significant interactions between the gaseous atoms/molecules, namely, low pressures (fewer atoms/molecules per unit volume) and high temperatures (greater kinetic energies of atoms/molecules make them less susceptible to attractive forces). The conditions described in (b), high temperature and low pressure, are therefore most likely to yield ideal gas behavior.

Question 55-4.

Describe the factors responsible for the deviation of the behavior of real gases from that of an ideal gas.

Solution

Anything that causes the gas particles to move close together, or have attractions, will cause deviations. Lowering the temperature will enhance this deviant behavior. In addition, the larger the molecules, the larger the volume occupied by the gas, which, in turn, causes deviations.

Question 55-5.

For which of the following gases should the correction for the molecular volume be largest: CO, CO_2, H_2, He, NH_3, SF_6?

Solution

We would expect the molecule with the largest volume to need the largest correction. SF_6 would need the largest correction.

Question 55-6.

A 0.245-L flask contains 0.467 mol CO_2 at 159 °C. Calculate the pressure:

(a) using the ideal gas law

(b) using the van der Waals equation

(c) Explain the reason for the difference.

(d) Identify which correction (that for P or V) is dominant and why.

Solution

(a) $PV = nRT$

$P \times 0.245\text{ L} = 0.467\text{ mol} \times 0.0821\text{ atm L/mol K} \times 432\text{ K}$

$P = \dfrac{(0.467\ \cancel{\text{mol}} \times 0.0821\text{ atm}\ \cancel{\text{L}}/\cancel{\text{mol}}\ \cancel{\text{K}} \times 432\ \cancel{\text{K}})}{0.245\ \cancel{\text{L}}} = 67.6\text{ atm}$; (b) For CO_2, a = 3.59 L^2 atm mol^2 and b = 0.427 L/mol. The rearranged equation for P is:

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

$$= \frac{0.467\ \cancel{\text{mol}} \times 0.0821\text{ atm}\ \cancel{\text{L}}/\cancel{\text{mol}^{-1}}\ \cancel{\text{K}^{-1}} \times 432\ \cancel{\text{K}}}{0.245\ \cancel{\text{L}} - 0.467\ \cancel{\text{mol}} \times 0.0427\ \cancel{\text{L}}/\cancel{\text{mol}^{-1}}};$$

$$- \frac{(0.467\ \cancel{\text{mol}})^2 \times 3.59\ \cancel{\text{L}^2}\text{ atm}\ \cancel{\text{mol}^{-2}}}{(0.245\ \cancel{\text{L}})^2}$$

$= 60.5\text{ atm}$

(c) Differences are due to the intermolecular attractions between gas molecules and the volume of the gas molecules themselves. (d) The pressure correction factor due to intermolecular attractions is dominant, which is usually the case at relatively low pressures; the volume correction would result in a higher pressure than expected, but a lower pressure than the value obtained with $PV = nRT$.

Question 55-7.

Answer the following questions:

(a) If XX behaved as an ideal gas, what would its graph of Z vs. P look like?

(b) For most of this chapter, we performed calculations treating gases as ideal. Was this justified?

(c) What is the effect of the volume of gas molecules on Z? Under what conditions is this effect small? When is it large? Explain using an appropriate diagram.

(d) What is the effect of intermolecular attractions on the value of Z? Under what conditions is this effect small? When is it large? Explain using an appropriate diagram.

(e) In general, under what temperature conditions would you expect Z to have the largest deviations from the Z for an ideal gas?

Solution

Answer: (a) A straight horizontal line at 1.0 (see Figure 9.35) for the line representing an ideal gas. (b) When real gases are at low pressures and high temperatures, they behave close enough to ideal gases that they are approximated as such; however, in some cases, we see that at a high pressure and temperature, the ideal gas approximation breaks down and is significantly different from the pressure calculated by the van der Waals equation. (c) The greater the compressibility, the more the volume matters. At low pressures, the correction factor for intermolecular attractions is more significant, and the effect of the volume of the gas molecules on Z would be a small lowering compressibility. At higher pressures, the effect of the volume of the gas molecules themselves on Z would increase compressibility (see Figure 9.35). (d) Once again, at low pressures, the effect of intermolecular attractions on Z would be more important than the correction factor for the volume of the gas molecules themselves, though perhaps still small. At higher pressures and low temperatures, the effect of intermolecular attractions would be larger. See Figure 9.35. (e) Low temperatures