Assignment 6 Solutions

Chapter 6, #6.4, 6.12, 6.32, 6.36, 6.43, 6.60, 6.70, 6.80, 6.88, 6.90, 6.100, 6.104, 6.108.

6.4. **Collect and Organize**

When the temperature of the balloon Figure P6.3 increases, does the balloon (a) shrink while keeping the molecules randomly distributed in the balloon, (b) expand, or (c) shrink and condense the molecules of gas into liquid?

Analyze

Charles's law states that volume and temperature of a gas are directly proportional.

Solve

If we increase the temperature, Charles's law states that the volume of the balloon increases. This is represented by diagram b.

Think About It

If the temperature is increased too much, the balloon will burst.

6.12. **Collect and Organize**

A mixture of gases is a homogeneous system of one or more gases. We are asked to choose the drawing in Figure P6.12 that best represents this definition.

Analyze

In order for the mixture to be homogeneous the two gases, helium and neon, for example, must be unreactive with each other and must be uniformly and randomly dispersed throughout the container.

Solve

Drawing b shows He and Ne randomly dispersed as a mixture of gases. Drawing a shows one component in the gas phase and the other in a condensed phase (liquid or solid). Drawing c shows the two components as gases, but they are not uniformly dispersed in the container. Drawing d shows the product of a reaction (HeNe) in which the atoms have bonded to each other. This is not a mixture—it is a sample of pure HeNe gas.

Think About It

Gases form uniform mixtures because the gas particles are in constant, random motion leading to uniform distribution of the atoms or molecules of each gas throughout the sample.

6.32. **Collect and Organize**

The pressure due to gravity for a block of gold is given by

$$P = \frac{F}{A} = \frac{ma}{A}$$

where m = mass of the gold cube (given as 38.6 g), a = acceleration due to gravity, and A = area over which the force is distributed. The block measures 2.00 cm × 1.00 cm × 1.00 cm, and we are asked to calculate the pressure exerted by a square face and a rectangular face.

Analyze

The area over which the force is distributed is different depending on which face the block is lying on. The area of the square face is 1.00 cm × 1.00 cm = 1.00 cm^2 whereas the area of the rectangular face is 2.00 cm × 1.00 cm = 2.00 cm^2. We also must watch our units here so as to get pressure in Pa ($kg/m \cdot s^2$) for comparison.

Solve

(a) On the square face:

$$P = \frac{\left(38.6\ \text{g} \times \dfrac{1\ \text{kg}}{1000\ \text{g}} \times 9.8\ \text{m/s}^2\right)}{1.00\ \text{cm}^2 \times \left(\dfrac{1\ \text{m}}{100\ \text{cm}}\right)^2} = 3.8 \times 10^3\ \text{Pa}$$

(b) On the rectangular face:

$$P = \frac{\left(38.6\ \text{g} \times \dfrac{1\ \text{kg}}{1000\ \text{g}}\ \text{kg} \times 9.8\ \text{m/s}^2\right)}{2.00\ \text{cm}^2 \times \left(\dfrac{1\ \text{m}}{100\ \text{cm}}\right)^2} = 1.9 \times 10^3\ \text{Pa}$$

Think About It

Because the area doubled in going from the square face to rectangular face, the pressure decreased by a factor of two.

6.36. **Collect and Organize**

We are to express the difference of the lowest measured pressure between Hurricanes Katrina and Wilma of 2 kPa in millimeters of mercury, atmospheres, and millibars.

Analyze

The conversion factors needed are

$$1\ \text{atm} = 101.325\ \text{kPa} = 101{,}325\ \text{Pa}$$
$$1\ \text{atm} = 760\ \text{mmHg}$$
$$10\ \text{mbar} = 1\ \text{kPa}$$

Solve

(a) $2\ \text{kPa} \times \dfrac{1\ \text{atm}}{101.325\ \text{kPa}} \times \dfrac{760\ \text{mmHg}}{1\ \text{atm}} = 15\ \text{mmHg}$

(b) $2\ \text{kPa} \times \dfrac{1\ \text{atm}}{101.325\ \text{kPa}} = 0.020\ \text{atm}$

(c) $2\ \text{kPa} \times \dfrac{10\ \text{mbar}}{1\ \text{kPa}} = 20\ \text{mbar}$

Think About It

When thinking about differences in pressure, it is often easier to express that difference in terms of smaller units such as millibars or millimeters of mercury as seen in this problem.

6.43. **Collect and Organize**

Using Boyle's law, find the pressure of the ammonia gas when 1.00 mol at 1.00 atm in 78.0 mL is compressed to 39.0 mL (or one-half of its original volume).

Analyze

The equation for Boyle's law for the compression (or expansion) of a gas is

$$P_1V_1 = P_2V_2$$

The number of moles of ammonia gas does not change so we do not need it in the calculation.

Solve

$$1.00\ \text{atm} \times 78.0\ \text{mL} = P_2 \times 39.0\ \text{mL}$$
$$P_2 = 2.00\ \text{atm}$$

Think About It

Because the volume was halved, the pressure increased by a factor of two.

6.60. **Collect and Organize**

For various changes in temperature and volume, we are to predict how the pressure of a gas sample changes.

Analyze

We must keep in mind the direct proportionality between V and T as shown by Charles's law and the inverse proportionality between V and P as shown by Boyle's law.

Solve

(a) When the absolute temperature is halved, the pressure is halved. When the volume is doubled, the pressure is halved. Combining these gives a decrease in the pressure to 1/4 the original pressure.
(b) When the absolute temperature is doubled, the pressure is doubled. When the volume is doubled, the pressure is halved. Combining these gives no change in the pressure.
(c) Combining these effects is best looked at mathematically:

$$P_2 = \frac{P_1V_1 \times 1.75T_1}{T_1 \times 1.50V_1} = 1.17P_1 \text{ or an increase of } 17\%$$

Think About It

Notice the way to simplify this problem is to consider each change separately, then "add" the effects.

6.70. **Collect and Organize**

Using the ideal gas law, we are to calculate the temperature at which 1.00 mole of gas in a 1.00 L vessel gives 1.00 atm of pressure.

Analyze

Rearranging the ideal gas law to solve for temperature,

$$PV = nRT$$

$$T = \frac{PV}{nR}$$

Solve

$$T = \frac{1.00 \text{ atm} \times 1.00 \text{ L}}{1.00 \text{ mol} \times 0.08206 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K}} = 12.2 \text{ K}$$

Think About It

This temperature is quite cold, –261˚C.

6.80. **Collect and Organize**

We can compare the volume of CO_2 produced in the combustion of propane to that of methane. The balanced combustion equations are as follows:

$$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$$
$$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(g)$$

Analyze

First calculate the moles of CO_2 produced in each reaction by converting 1.00 kg (1000 g) of each fuel into CO_2. For this, we need the molar masses of methane (16.04 g/mol) and propane (44.10 g/mol) and the molar ratio of the fuel to CO_2 from the balanced chemical equations. Then, the volume of CO_2 can be calculated from the moles of CO_2 using the ideal gas equation where P = 1.00 atm and T = 20˚C (292 K).

Solve

For methane:

$$n_{CO_2} = 1000 \text{ g } CH_4 \times \frac{1 \text{ mol } CH_4}{16.04 \text{ g}} \times \frac{1 \text{ mol } CO_2}{1 \text{ mol } CH_4} = 62.3 \text{ mol } CO_2$$

$$V_{CO_2} = \frac{62.3 \text{ mol} \times 0.08206 \text{ L} \cdot \text{atm/mol} \cdot \text{K} \times 293 \text{ K}}{1.00 \text{ atm}} = 1500 \text{ L } CO_2$$

For propane:

$$n_{CO_2} = 1000 \text{ g } C_3H_8 \times \frac{1 \text{ mol } C_3H_8}{44.10 \text{ g}} \times \frac{3 \text{ mol } CO_2}{1 \text{ mol } C_3H_8} = 68.0 \text{ mol } CO_2$$

$$V_{CO_2} = \frac{68.0 \text{ mol} \times 0.08206 \text{ L} \cdot \text{atm/mol} \cdot \text{K} \times 293 \text{ K}}{1.00 \text{ atm}} = 1630 \text{ L } CO_2$$

Think About It

Burning 1 kg of propane releases more CO_2 than burning 1 kg of methane.

6.88. **Collect and Organize**

In order for a balloon to float in the air, the density of the balloon must be less than the density of the surrounding air. The balloons for this problem are large (20 L) and have a mass of 10.0 g when filled.

Analyze

For each gas, we need to calculate the density of the gas in the balloon using

$$d = \frac{m}{V}$$

where m is the mass of the filled balloon (10.0 g) and V is the volume of the balloon (20.0 L). This gives 0.5 g/L as the density of each balloon. Since each balloon's density is less than the density of air (0.00117 g/mL), the balloons will all float.

6.90. **Collect and Organize**

Given the mass of an unknown gas (0.193 g) in a known volume (100.0 mL) measured at a particular temperature (17˚C) and pressure (760 mmHg), we are to determine the identity of the gas through calculation of the gas's molar mass.

Analyze

Use the gas density equation to calculate

$$d = \frac{m}{V} = \frac{\} \ P}{RT}$$

$$\} = \frac{mRT}{VP}$$

Be sure to use units of volume in liters, pressure in atmospheres, and temperature in kelvins.

Solve

$$\} = \frac{0.193 \text{ g} \times 0.08206 \text{ L} \cdot \text{atm/mol} \cdot \text{K} \times 290 \text{ K}}{0.1000 \text{ L} \times 1.00 \text{ atm}} = 45.9 \text{ g/mol}$$

The molar masses of NO, NO_2, and N_2O_5 are 30.01, 46.01, and 108.02 g/mol, respectively. The gas in the flask is NO_2.

6.100. **Collect and Organize**

The mole fractions of the gases in Problem 6.98 are $\chi(N_2) = 0.11$, $\chi(H_2) = 0.44$, and $\chi(CH_4) = 0.44$. In this problem, we are to use this information to calculate the partial pressure of each gas and the total pressure of the gas mixture when the volume is 1.00 L at 0˚C (273 K).

Analyze

The partial pressure of each gas is related to the total pressure: $P_x = \chi_x P_{\text{total}}$. We can calculate P_{total} from the ideal gas law where V = 1.00 L, n = total moles of gas in the mixture, and T = 273 K.

Solve

$$P_{\text{total}} = \frac{2.25 \text{ mol} \times 0.08206 \text{ L}\cdot\text{atm/mol}\cdot\text{K} \times 273 \text{ K}}{1.00 \text{ L}} = 50.4 \text{ atm}$$
$$P_{N_2} = 0.11 \times 50.4 \text{ atm} = 5.5 \text{ atm}$$
$$P_{H_2} = 0.44 \times 50.4 \text{ atm} = 22 \text{ atm}$$
$$P_{CH_4} = 0.44 \times 50.4 \text{ atm} = 22 \text{ atm}$$

Think About It

Notice that the sum of the partial pressures approximately equals P_{total}. Any differences are due to rounding.

6.104. **Collect and Organize**

We are given the balanced chemical equations for three reactions in which reactants and products are all gases. Based on the moles of gas consumed and the moles of gas produced we are to determine how the pressure changes for each reaction.

Analyze

The greater the moles of gas in a sealed container, the higher the pressure. If the reaction produces more moles of gas than it consumes, then the pressure after the reaction is complete must be greater than the pressure before the reaction took place. On the other hand, is a reaction produces fewer moles of gas than it consumes, the pressure must be lower. If there is no change in the moles of gas between reactants and products, the pressure does not change.

Solve

(a) Two moles of gas are consumed, and two moles of gas are produced; $\Delta n = n_{\text{products}} - n_{\text{reactants}} = 0$ so the pressure does not change throughout the reaction.
(b) Nine moles of gas are consumed, and ten moles of gas are produced; $\Delta n = 1$ and the pressure is greater at the end of the reaction.
(c) Three moles of gas are consumed, and two moles of gas are produced; $\Delta n = -1$ and the pressure is lower at the end of the reaction.

Think About It

Mathematically, if $\Delta n = n_{\text{products}} - n_{\text{reactants}}$ is positive, the pressure increases; if $\Delta n = 0$, the pressure stays the same; if Δn is negative, the pressure decreases.

6.108. **Collect and Organize**

The reaction of NO with O_3 to produce NO_2 and O_2 is described by the balanced equation

$$NO(g) + O_3(g) \rightarrow NO_2(g) + O_2(g)$$

The reactant gases are confined in a 10.0 L vessel in the amounts of 0.280 mol of NO and 0.280 mol of O_3 at 350 K. From this information we are to calculate the partial pressure of each of the products NO_2 and O_2 and the total pressure at the end of the reaction.

Analyze

Because the stoichiometry of the reaction shows that for every mole of NO and O_3 consumed one mole of NO and O_2 are produced, the moles of NO_2 and O_2 will be 0.280 mol each. Because the reaction yields equimolar amounts of NO_2 and O_2, the partial pressure of each will be the same. We can calculate that partial pressure using $PV = nRT$.

Solve

Partial pressure of both NO_2 and O_2 at the end of the reaction:

$$P_x = \frac{0.280 \text{ mol} \times 0.08206 \text{ L} \cdot \text{atm/mol} \cdot \text{K} \times 350 \text{ K}}{10.0 \text{ L}} = 0.804 \text{ atm}$$

Total pressure

$$P_{\text{total}} = P_{NO_2} + P_{O_2} = 0.804 + 0.804 = 1.608 \text{ atm}$$