Chapter 5 Homework Answers 6, 8, 16, 24, 28, 30, 34, 39, 48, 54

6. Refer to Section 5.1.

$$P_1 = 977 \text{ mm Hg} \times \frac{1 \text{ atm}}{760.0 \text{ mm Hg}} = 1.29 \text{ atm}$$

$T_1 = 25°C = 25 + 273 = 298$ K

$P_2 = 1.50$ atm

$T_2 = ?$

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

$$\frac{1.29 \text{ atm}}{298 \text{ K}} = \frac{1.50 \text{ atm}}{T_2}$$

$T_2 = 347$ K $= 74°C$

8. Refer to Section 5.3 and Example 5.2.

$T_1 = 32°C = 22 + 273 = 305$ K

a. $V_2 = 0.75V_1$

$$\frac{T_2}{T_1} = \frac{V_2}{V_1} \Rightarrow T_2 = T_1\frac{V_2}{V_1} = 305 \text{ K} \times \frac{0.75V_1}{V_1} = 229 \text{ K} = -44^0\text{C}$$

b. $V_2 = 0.25V_1$

$$\frac{T_2}{T_1} = \frac{V_2}{V_1} \Rightarrow T_2 = T_1\frac{V_2}{V_1} = 305 \text{ K} \times \frac{0.25V_1}{V_1} = 76 \text{ K} = -197^0\text{C}$$

16. Refer to Section 5.3 and Example 5.3.

Calculate the moles of O_2 from the mass, then use the ideal gas law to calculate pressure.

$T = 37°C + 273 = 310\ K$

$$n = \frac{0.25\ mg\ O_2}{1\ mL} \times \frac{1000\ mL}{1\ L} \times \frac{1\ g}{1000\ mg} \times \frac{1\ mol.\ O_2}{32.00\ g\ O_2} = 7.8 \times 10^{-3}\ mol.\ O_2 \text{ per liter of blood}$$

$$P = \frac{nRT}{V} = \frac{(7.8 \times 10^{-3}\ mol.)(0.0821\ L \cdot atm/mol. \cdot K)(310\ K)}{(1.00\ L)} = 0.20\ atm$$

24. Refer to Section 5.3 and Example 5.4.

$$P_1 = 755\ mm\ Hg \times \frac{1\ atm}{760.0\ mm\ Hg} = 0.993\ atm$$

$T_1 = 0°C = 0 + 273 = 273\ K$

$$P_2 = 210\ mm\ Hg \times \frac{1\ atm}{760.0\ mm\ Hg} = 0.276\ atm$$

$T_2 = 0°C = 0 + 273 = 273\ K$

$$d = \frac{(MM)P}{RT}$$

$$d_1 = \frac{(29.0\ g/mol.)(0.993\ atm)}{(0.0821\ L \cdot atm/mol. \cdot K)(273\ K)} = 1.28\ g/L$$

$$d_2 = \frac{(29.0 g/mol.)(0.276\ atm)}{(0.0821\ L \cdot atm/mol. \cdot K)(273\ K)} = 0.357\ g/L$$

$\frac{1.28\ g/L}{0.357\ g/L} = 3.59$ At sea level, the density of air is 3.59 times more dense than the air at the top of Mount Everest.

28. *Refer to Section 5.3.*

The molar mass of exhaled air is simply the sum of the molar masses of the individual components multiplied by their abundances. Calculate the moles of air in 1 L, then the mass of 1 L to get density.

a. $(28.02 \text{ g/mol. } N_2)(0.745) + (32.00 \text{ g/mol. } O_2)(0.157) + (44.01 \text{ g/mol. } CO_2)(0.036)$
$+ (18.02 \text{ g/mol. } H_2O)(0.062) = 28.6 \text{ g/mol.}$

b. $T = 37°C + 273 = 310 \text{ K}$

$$757 \text{ mm Hg} \times \frac{1 \text{ atm}}{760.0 \text{ mm Hg}} = 0.996 \text{ atm}$$

$$n = \frac{PV}{RT} = \frac{(0.996 \text{ atm})(1.00 \text{ L})}{(0.0821 \text{ L} \cdot \text{atm/mol.} \cdot \text{K})(310 \text{ K})} = 0.0391 \text{ mol.}$$

$$0.0391 \text{ mol.} \times \frac{28.6 \text{ g air}_{\text{Exhaled}}}{1 \text{ mol. air}_{\text{Exhaled}}} = 1.12 \text{ g air}_{\text{Exhaled}}$$

$$d_{\text{air}_{\text{Exhaled}}} = \frac{\text{g}}{\text{L}} \times \frac{1.12 \text{ g air}_{\text{Exhaled}}}{1.00 \text{ L}} = 1.12 \text{ g/L}$$

$$0.0391 \text{ mol.} \times \frac{29.0 \text{ g air}_{\text{ordinary}}}{1 \text{ mol. air}_{\text{ordinary}}} = 1.13 \text{ g air}_{\text{ordinary}}$$

$$d_{\text{air}_{\text{ordinary}}} = \frac{\text{g}}{\text{L}} \times \frac{1.13 \text{ g air}_{\text{ordinary}}}{1.00 \text{ L}} = 1.13 \text{ g/L}$$

Ordinary air is slightly more dense than exhaled air.

30. *Refer to Section 5.3.*

Use the ideal gas equation to calculate the number of moles in the sample, then the molar mass of the compound. Then calculate the atomic mass of X and you have the identity.

$T = 20°C + 273 = 293 \text{ K}$

$$329.5 \text{ cm}^3 \times \frac{1 \text{ mL}}{1 \text{ cm}^3} \frac{1 \text{ L}}{1000 \text{ mL}} = 0.3295 \text{ L}$$

$$n = \frac{PV}{RT} = \frac{(1.00 \text{ atm})(0.3295 \text{ L})}{(0.0821 \text{ L} \cdot \text{atm/mol.} \cdot \text{K})(293 \text{ K})} = 0.0137 \text{ mol.}$$

$$\text{molar mass} = \frac{2.00\text{ g}}{0.0137\text{ mol.}} = 146\text{ g/mol.}$$

146 g/mol. = 1(S) + 6(X)
146 g/mol. = (32.07 g/mol.) + 6(X)
$6X$ = 114 g/mol.
X = 19 g/mol.
X = F
Thus the compound is SF_6, sulfur hexafluoride.

34. *Refer to Section 5.4 and Example 5.5.*

a. This equation can be balanced by inspection.

$NH_4NO_3(s) \rightarrow N_2O(g) + 2H_2O(l)$

b. T = 250°C + 273 = 523 K

$$V = \frac{nRT}{P} = \frac{(0.0625\text{ mol.})(0.0821\text{ L}\cdot\text{atm/mol.}\cdot\text{K})(298\text{ K})}{1.0\text{ atm}} = 2.7\text{ L}$$

39. $n_{\text{tot}} = \frac{(0.986)(1.00)}{(0.0821)(315)} = 0.0381\text{ mol}$

$n_{\text{dry}} = \frac{(1.00)(1.04)}{(0.0821)(363)} = 0.0349\text{ mol}$

$n_{H_2O} = 0.0381\text{ mol} - 0.0349\text{ mol} = 0.0032\text{ mol}$

$P_{H_2O} = \frac{(0.0032\text{ mol})(0.0821\text{ L}\cdot\text{atm/mol}\cdot\text{K})(315\text{ K})}{1.00\text{ L}} = 0.083\text{ atm} = 63\text{ mm Hg}$

48. *Refer to Section 5.6 and Example 5.11.*

Calculate the ratio of the effusion rates of the two gases.

$$\frac{\text{rate of effusion of } H_2}{\text{rate of effusion of } N_2} = \left(\frac{MM_{N_2}}{MM_{H_2}}\right)^{\frac{1}{2}} = \left(\frac{28.02}{2.016}\right)^{\frac{1}{2}} = 3.73$$

Thus the hydrogen balloon will deflate 3.73 times faster than the nitrogen balloon.

54. *Refer to Section 5.7.*

Deviations from ideal behavior tend to be largest at high pressures and low temperatures.

a. If pressure is reduced from 20 atm to 1 atm, CH_4 should behave **more** ideally.

b. If temperature is reduced from 50°C to –50°C, CH_4 should behave **less** ideally.