1. Define vapor

chemicals or solids or liquids that are now in the gas phase are referred to as vapors.

2. Define pressure and give the equation.

   \[ P = \frac{F}{A} \]

   force per unit area

3. Define force.

   \[ F = ma \]

   \( m = \) mass \hspace{1cm} \( a = \) acceleration

4. Define pascal, bar, psi, mmHg, atmosphere, and torr

   \( 1 \text{ atm} = 760 \text{ mmHg} = 1.01325 \times 10^5 \text{ Pa} = 101.325 \text{ kPa} = 1.01325 \text{ bar} \)

5. Define standard atmospheric pressure

   The typical pressure at sea level sufficient to make if 760 mm Hg.

6. Torricelli used mercury in his barometer because it has a very high density, which makes it possible to make a more compact barometer than one based on a less dense fluid. Calculate the density of mercury, \( d_{\text{Hg}} \), using the observation that the column of mercury is 760 mm high when the atmospheric pressure is 1.01 * 10^5 Pa. Assume the tube containing the mercury is a cylinder with a constant cross-sectional area.

   \[ \rho_{\text{Hg}} = \frac{F}{A} = \frac{m_{\text{Hg}}g}{A} \]

   \[ V = \pi r^2 h \]

   \[ \rho_{\text{Hg}} = \frac{\pi r^2 h g}{A} \]

   \[ \frac{\pi}{4} \rho_{\text{Hg}} \frac{r^2 h}{A} = \frac{1.35 \times 10^{-4} \text{ kg/m}^3}{(0.760 \text{ mm})(760 \text{ mm})} \]

7. Gallium melts just above room temperature and is liquid over a very wide temperature range (30–2204 °C), which means it would be a suitable fluid for a high-temperature barometer. Given its density, \( d_{\text{Ga}} = 6.0 \) g/cm³, what would be the height of the column if gallium is used as the barometer fluid and the external pressure is 9.5 * 10^4 Pa?

   \[ h = \frac{\rho_{\text{atm}}}{d_{\text{Ga}}} \]

   \[ h = \frac{9.5 \times 10^4 \text{ Pa}}{(6.0 \text{ g/cm}^3)(\frac{1 \text{ kg}}{1000 \text{ g}})(\frac{1 \text{ m}}{1000 \text{ cm}})} \]

   units are in m and kg


   \[ \frac{p_1 V_1}{T_1} = \frac{p_2 V_2}{T_2} \]

   \[ \frac{p_1 V_1}{T_1} = \frac{p_2 V_2}{T_2} \]

   any gas law can be derived from \( \frac{pV}{nRT} \)
9. Suppose we have a gas confined to a cylinder with a movable piston that is sealed so there are no leaks. How will each of the following changes affect (i) the pressure of the gas, (ii) the number of moles of gas in the cylinder, (iii) the average distance between molecules: (a) Heating the gas while maintaining a constant pressure; (b) Reducing the volume while maintaining a constant temperature; (c) Injecting additional gas while keeping the temperature and volume constant.

- a) i) same ii) same iii) farther
- b) i) inc ii) same iii) closer
- c) i) inc ii) inc iii) decrease

10. What is the gas constant?

\[
R = \frac{0.08206 \text{ atm} \cdot \text{L}}{\text{mol} \cdot \text{K}}
\]

11. Calcium carbonate, CaCO\(_3\), the principal compound in limestone, decomposes upon heating to CaO(s) and CO\(_2\)(g). A sample of CaCO\(_3\) is decomposed, and the carbon dioxide is collected in a 250-mL flask. After decomposition is complete, the gas has a pressure of 1.3 atm at a temperature of 31 °C. How many moles of CO\(_2\) gas were generated?

\[
P_1 = 1.3 \text{ atm} \quad V = 0.250 \text{ L} \\
31^\circ C = 304 \text{ K}
\]

\[
\frac{n}{R} = \frac{PV}{RT} = 0.013 \text{ mol CO}_2
\]

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

\[
\frac{P_1}{T_1} = \frac{P_2}{T_2} \\
1.5 \text{ atm} = \frac{x}{298 \text{ K}}
\]

\[
x = 723 \text{ K}
\]

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

\[
\frac{V_1P_1}{T_1} = \frac{V_2P_2}{T_2} \\
\frac{(0.04 \text{ L})(1 \text{ atm})}{295 \text{ K}} = \frac{x(0.45 \text{ atm})}{252 \text{ K}}
\]

14. Give the equations for gas densities and molar masses related to the ideal gas laws.

\[
\begin{align*}
\rho &= \frac{nM}{V} = \frac{P \cdot M}{RT} \\
M &= \text{molar mass} \\
d &= \text{density}
\end{align*}
\]
15. What is the density of carbon tetrachloride vapor at 714 torr and 125 °C?

\[ \rho = \frac{(0.939 \text{ atm}) \left( \frac{153.8 \text{ g/mol}}{0.08206 \text{ L atm/mol K}} \right)}{(398 \text{ K})} = 1.425 \text{ g/L} \]

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

\[ M = \frac{\rho V T}{\rho} \left( \frac{3.041 \text{ J}}{\text{mol K}} \right) \left( \frac{0.08206 \text{ L atm/mol K}}{\text{mol atm}} \right) 364 \text{ K} \]

\[ = 79.7 \text{ g/mol} \]

17. Automobile air bags are inflated by nitrogen gas generated by the rapid decomposition of sodium azide, NaN₃:

\[ \text{NaN₃(s)} \rightarrow 2 \text{ Na(s)} + 3 \text{ N₂(g)} \]

If an air bag has a volume of 36 L and is to be filled with nitrogen gas at 1.15 atm and 26 °C, how many grams of NaN₃ must be decomposed?

\[ n = \frac{pV}{RT} \]

\[ n = 1.15 \text{ atm} \times 36 \text{ L} \]

\[ = 1.12 \text{ mol NaN₃} \]

18. Define partial pressure

The pressure exerted by a particular component of a mixture of gases

19. Dalton’s Law of partial pressure

Same as above

\[ P = \sum \frac{n_i (RT)}{V} \text{ etc.} \]

20. A mixture of 6.00 g of O₂(g) and 9.00 g of CH₄(g) is placed in a 15.0-L vessel at 0 °C. What is the partial pressure of each gas, and what is the total pressure in the vessel?

\[ \text{O}_2: \frac{6.00 \text{ g}}{32} = 0.188 \text{ mol} \]

\[ P = \frac{nRT}{V} = \frac{0.188 \text{ mol}}{15.0 \text{ L}} = 0.0128 \text{ atm} \]

\[ \text{CH}_4: \frac{9.00 \text{ g}}{16.04} = 0.560 \text{ mol} \]

\[ 0.841 \text{ atm} \]
21. Give an equation for partial pressures and mole fractions from page 417.

\[ P_i = \left( \frac{n_i}{n_t} \right) P_t = X_i P_t \]

22. A study of the effects of certain gases on plant growth requires a synthetic atmosphere composed of 1.5 mol % CO₂, 18.0 mol % O₂, and 80.5 mol % Ar. (a) Calculate the partial pressure of O₂ in the mixture if the total pressure of the atmosphere is to be 745 torr. (b) If this atmosphere is to be held in a 121-L space at 295 K, how many moles of O₂ are needed?

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

\[ n_{O_2} = P_{O_2} \left( \frac{N}{2} \right) \]

\[ = 8.80 \text{ mol} \]

23. Define kinetic molecular theory.

The behavior of gases defined by 5 postulates

24. Summarize the 5 statements about the kinetic molecular theory.

1) Gases consist of large numbers of molecules that are in continuous, random motion,

2) the combined volume of all molecules of the gas is negligible relative to the total volume in which the gas is contained

3) Attractive and repulsive forces between gas molecules are negligible

4) Energy can be transferred between molecules during collisions but, as long as temperature remains constant, the average kinetic energy of the molecules does not change with time

5) The average kinetic energy of the molecules is proportional to the absolute temperature. At any given temp, the molecules of all gases have the same average kinetic energy

25. Define root-mean-square (RMS) speed. Give the equation and draw and label the diagram provided on page 419.

\[ \text{RMS speed} = \sqrt{\frac{3RT}{M}} \]

[Diagram showing RMS speed with labeled parts]
26. Describe the applications of the kinetic molecular theory.
   1) An increase in volume at constant temp causes pressure to decrease
   2) A temperature increase at constant volume causes pressure to increase

27. A sample of O₂ gas initially at STP is compressed to a smaller volume at constant temperature. What effect does this change have on (a) the average kinetic energy of the molecules, (b) their average speed, (c) the number of collisions they make with the container walls per unit time, (d) the number of collisions they make with a unit area of container wall per unit time, (e) the pressure?
   A) unchanged  c) increases  E) increases
   B) unchanged  d) decreases

28. What is the definition for most probable speed equation and rms equation.
   \[ \bar{v}_{\text{rms}} = \sqrt{\frac{3RT}{M}} \]
   \[ \bar{v}_{\text{max}} = \sqrt{\frac{2RT}{M}} \]

29. Calculate the rms speed of the molecules in a sample of N₂ gas at 25 °C.
   \[ T = 298 \text{ K} \quad M = 28 \quad R = 8.314 \text{ J/mol·K} \]
   \[ \bar{v}_{\text{rms}} = \sqrt{\frac{3RT}{M}} = 5.15 \times 10^{3} \text{ m/s} \]

   \[ \frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}} \]
   \[ r = \text{rate of effusion} \]

31. An unknown gas composed of homonuclear diatomic molecules effuses at a rate that is 0.355 times the rate at which O₂ gas effuses at the same temperature. Calculate the molar mass of the unknown and identify it.
   \[ \frac{r_x}{r_{O_2}} = \sqrt{\frac{M_{O_2}}{M_x}} \quad r_x = 0.355(r_{O_2}) \]
   \[ M_x = 2.56 \text{ g/mol} \]

32. Define diffusion.
   the spread of one substance throughout a space or throughout a second substance
   effusion - the escape of gas molecules through a tiny hole

33. Define mean free path and give its equation.
   the average distance traveled by a molecule between collisions
34. When do gases deviate from ideal behavior?

very high pressure, very low temperature

35. How do real gases behave compared to ideal gases?

have finite volume and attract one another

36. Define the van der Waals equation.

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

\[ a = \text{attraction} \]
\[ b = \text{volume of molecule} \]

37. If 10.00 mol of an ideal gas were confined to 22.41 L at 0.0 °C, it would exert a pressure of 10.00 atm. Use the van der Waals equation and Table 10.3 to estimate the pressure exerted by 1.000 mol of Cl₂(g) in 22.41 L at 0.0 °C.

\[ \frac{aL^2}{mol} = 6.04 \times 10^4 \]
\[ \frac{bL}{mol} = 0.05 \]

\[ P = 273 \text{ k} \]

<table>
<thead>
<tr>
<th>Table 10.3 Van der Waals Constants for Gas Molecules</th>
</tr>
</thead>
<tbody>
<tr>
<td>Substance</td>
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<tr>
<td>He</td>
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<td>Ne</td>
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<tr>
<td>Ar</td>
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<td>CH₄</td>
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<tr>
<td>CO₂</td>
</tr>
<tr>
<td>CCl₄</td>
</tr>
</tbody>
</table>

38. Cyanogen, a highly toxic gas, is 46.2% C and 53.8% N by mass. At 25 °C and 751 torr, 1.05 g of cyanogen occupies 0.500 L. (a) What is the molecular formula of cyanogen? Predict (b) its molecular structure and (c) its polarity.

(a) Molecular formula:

\[ \text{C} \quad \frac{4.49 \text{ g}}{12.01 \text{ g/mol}} = 3.58 \text{ mol} \]

\[ \text{N} \quad \frac{5.58 \text{ g}}{14.01 \text{ g/mol}} = 0.398 \text{ mol} \]

\[ \text{CN} \]

\[ M = \frac{RT}{P} = \frac{1(0.052 \text{ atm})(0.042 \text{ amagat})(207 \text{ K})}{0.98 \text{ atm}} = 52.09 \text{ amagat} \]

(b) Molecular structure:

\[ \text{C}_2\text{N}_2 : \text{N} = \text{C} - \text{N} - \text{C} = \text{C}_2\text{N}_2 \]

(c) Polarity:

Nonpolar
39. How does a gas compare with a liquid for each of the following properties: (a) density, (b) compressibility, (c) ability to mix with other substances of the same phase to form homogeneous mixtures, (d) ability to conform to the shape of its container?

A) gas is lower density
B) gas is more compressible
C) gases always mix, liquids don’t always mix
D) gases take entire shape liquids only take volume of the liquid container

40. Suppose that a woman weighing 130 lb and wearing high-heeled shoes momentarily places all her weight on the heel of one foot. If the area of the heel is 0.50 in.², calculate the pressure exerted on the underlying surface in (a) kilopascals, (b) atmospheres, and (c) pounds per square inch.

\[ P = \frac{130 \text{ lb}}{0.50 \text{ in.}^2} = 260 \text{ lb/in}^2 \]

\[ A = 1.8 \times 10^{-3} \text{ km}^2 \]

\[ P = \frac{260 \text{ lb/in}^2}{1.8 \times 10^{-3} \text{ km}^2} \]

\[ 141.2 \text{ psi} = 1 \text{ atm} \]

41. (a) How high in meters must a column of water be to exert a pressure equal to that of a 760-mm column of mercury? The density of water is 1.0 g/mL, whereas that of mercury is 13.6 g/mL. (b) What is the pressure, in atmospheres, on the body of a diver if he or she is 39 ft below the surface of the water when atmospheric pressure at the surface is 0.97 atm?

\[ d_{w} = \frac{P_{o}}{\rho g} \]

\[ (1 \text{ g/cm}^3) \left( \frac{1 \text{ kg}}{1000 \text{ g}} \right) \left( \frac{1000 \text{ cm}}{1 \text{ m}} \right)^3 = \frac{760 \text{ torr}}{9.8 \text{ m/s}} \]

\[ 760 \text{ torr} \]

\[ 14.2 \text{ psi} \]

42. The typical atmospheric pressure on top of Mt. Everest (29,028 ft) is about 265 torr. Convert this pressure to (a) atm, (b) mm Hg, (c) pascals, (d) bars, (e) psi.

\[ 265 \text{ torr} \]

\[ 265 \text{ torr} \]

\[ 1.01325 \times 10^5 \text{ Pa} \]

\[ 14.2 \text{ psi} \]

43. In the United States, barometric pressures are generally reported in inches of mercury (in. Hg). On a beautiful summer day in Chicago, the barometric pressure is 30.45 in. Hg. (a) Convert this pressure to torr. (b) Convert this pressure to atm.

\[ 30.45 \text{ in. Hg} \]

\[ 416 \text{ cm Hg} \]

\[ \frac{1000 \text{ mm Hg}}{760 \text{ torr}} = \frac{26.4 \text{ in. Hg}}{760 \text{ torr}} \]

\[ 1 \text{ atm} = \frac{416 \text{ cm Hg}}{760 \text{ torr}} \]
44. (a) Amonton’s law expresses the relationship between pressure and temperature. Use Charles’s law and Boyle’s law to derive the proportionality relationship between $P$ and $T$. (b) If a car tire is filled to a pressure of 32.0 lbs/in$^2$ (psi) measured at 75 °F, what will be the tire pressure if the tires heat up to 120 °F during driving?

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

\[ \begin{align*}
\text{As } T \uparrow \quad P \uparrow & \quad \text{If } T = 75 \degree F = 297 \text{ K} \\
\text{at } 120 \degree F = 322 \text{ K} & \quad \frac{297 \text{ K}}{297 \text{ K}} = \frac{2.0 \text{ psi}}{x} \\
\text{x} &= \frac{322 \text{ K}(2.0 \text{ psi})}{297 \text{ K}} \\
\end{align*} \]

45. (a) What conditions are represented by the abbreviation STP? (b) What is the molar volume of an ideal gas at STP? (c) Room temperature is often assumed to be 25 °C. Calculate the molar volume of an ideal gas at 25 °C and 1 atm pressure. (d) If you measure pressure in bars instead of atmospheres, calculate the corresponding value of $R$ in L-bar/mol-K.

\[ \begin{align*}
\text{A) } & \quad 0 \degree C \text{ and atm} \\
\text{B) } & \quad 22.4 L \\
\text{C) } & \quad \frac{22.4 L}{273 \text{ K}} = 0.08205 \frac{\text{L-atm}}{\text{mol-K}} \\
\text{D) } & \quad \frac{1 \text{ atm}}{1.01325 \text{ bar}} \\
\end{align*} \]

46. Suppose you are given two 1-L flasks and told that one contains a gas of molar mass 30, the other a gas of molar mass 60, both at the same temperature. The pressure in flask A is $X$ atm, and the mass of gas in the flask is 1.2 g. The pressure in flask B is $0.5X$ atm, and the mass of gas in that flask is 1.2 g. Which flask contains the gas of molar mass 30, and which contains the gas of molar mass 60?

\[ \begin{align*}
\text{Flask A: } & \quad \frac{1.2 g}{1 \text{ L}} = \frac{x(30)}{R(1 \text{ K})} = 0.00328 \text{ atm} \\
\text{Flask B: } & \quad \frac{1.2 g}{1 \text{ L}} = \frac{(0.5X \text{ atm})(60)}{R(1 \text{ K})} = 0.00328 \text{ atm} \\
\end{align*} \]

47. The Goodyear blimps, which frequently fly over sporting events, hold approximately 175,000 ft$^3$ of helium. If the gas is at 23 °C and 1.0 atm, what mass of helium is in a blimp?

\[ \begin{align*}
\frac{(1 \text{ atm})(49,000 \text{ ft}^3)}{273 \text{ K}} & = n(0.08205 \frac{\text{L-atm}}{\text{mol-K}})(300 \text{ K}) \\
\frac{49,000 \text{ ft}^3}{273 \text{ K}} & = \frac{204,000 \text{ mol He}}{1 \text{ L}} \\
\frac{49,000 \text{ ft}^3}{273 \text{ K}} & = 81,000 \text{ ft}^3 \text{ He} \\
\end{align*} \]

48. (a) Calculate the number of molecules in a deep breath of air whose volume is 2.25 L at body temperature, 37 °C, and a pressure of 735 torr. (b) The adult blue whale has a lung capacity of 5.0 * 10$^3$ L. Calculate the mass of air (assume an average molar mass of 28.98 g/mol) contained in an adult blue whale’s lungs at 0.0 °C and 1.00 atm, assuming the air behaves ideally.

\[ \begin{align*}
\text{(a) } & \quad (0.96 \text{ atm})(2.25 \text{ L}) = n(0.08205 \frac{\text{L-it}m}{\text{mol-K}})(273 \text{ K}) \\
\frac{n}{0.08205 \frac{\text{L-atm}}{\text{mol-K}}} & = 0.0906 \text{ mol} \\
\text{(b) } & \quad (1 \text{ atm})(5000 \text{ L}) = n(0.08205 \frac{\text{L-atm}}{\text{mol-K}})(273 \text{ K}) \\
\frac{223 \text{ mol}}{0.08205 \frac{\text{L-atm}}{\text{mol-K}}} & = 6470 \text{ g} \\
\end{align*} \]

49. A scuba diver’s tank contains 0.29 kg of O$_2$ compressed into a volume of 2.3 L. (a) Calculate the gas pressure inside the tank at 9 °C. (b) What volume would this oxygen occupy at 26 °C and 0.95 atm?

\[ \begin{align*}
\frac{290 \text{ g}}{1 \text{ mol}} & = 9.09 \text{ mol} \\
\rho & = \frac{290 \text{ g mol}^{-1}}{2.3 \text{ L}} \\
V & = \frac{9.09 \text{ mol}(282 \text{ K})}{0.95 \text{ atm}} \\
\end{align*} \]
50. A 35.1 g sample of solid CO2 (dry ice) is added to a container at a temperature of 100 K with a volume of 4.0 L. If the container is evacuated (all of the gas removed), sealed and then allowed to warm to room temperature $T = 298$ K so that all of the solid CO2 is converted to a gas, what is the pressure inside the container?

$$\frac{1}{4} \times \frac{\text{mol CO}_2}{\text{g CO}_2} \times \frac{1}{3510} \times \frac{1}{44.0} \times \frac{1}{\text{mol CO}_2} = \frac{0.797 \text{ mol CO}_2}{4 \text{ L}}$$

$$P = \frac{(0.797 \text{ mol})(R)(298 \text{ K})}{4 \text{ L}}$$

51. Chlorine is widely used to purify municipal water supplies and to treat swimming pool waters. Suppose that the volume of a particular sample of Cl₂ gas is 8.70 L at 895 torr and 24 °C. (a) How many grams of Cl₂ are in the sample? (b) What volume will the Cl₂ occupy at STP? (c) At what temperature will the volume be 15.00 L if the pressure is 8.76 x 10² torr? (d) At what pressure will the volume equal 5.00 L if the temperature is 58 °C?

$$\text{a) } \frac{8.70 \text{ L}(1.17 \text{ atm})}{291 \text{ K}} = n$$

$$n = \frac{4.18 \text{ mol}}{1.01 \text{ mol}} \times 29.5 \text{ g Cl}_2$$

$$\text{b) } V = \frac{(0.441 \text{ mol})(273 \text{ K})}{(1 \text{ atm})(8.70 \text{ L})}$$

52. In an experiment reported in the scientific literature, a male cockroach was weighed while running at different speeds on a miniature treadmill while their oxygen consumption was measured. In 1 hr the average cockroach running at 0.08 km/hr consumed 0.8 mL of O₂ at 1 atm pressure and 24 °C per gram of insect mass. (a) How many moles of O₂ would be consumed in 1 hr by a 5.2-g cockroach moving at this speed? (b) This same cockroach is caught by a child and placed in a 1-qt fruit jar with a tight lid. Assuming the same level of continuous activity as in the research, will the cockroach consume more than 20% of the available O₂ in a 48-hr period? (Air is 21 mol % O₂.)

$$\text{a) } 5.2 \text{ g} = 11.1 \text{ mL O}_2 / \text{ hr}$$

$$\frac{11.1 \text{ mL}}{48 \text{ min}} = 0.23 \text{ mol O}_2 / \text{ hr}$$

$$\frac{0.23 \text{ mol}}{1 \text{ hr}} \times 48 \text{ min} = 199.48 \text{ mL}$$

53. Which gas is most dense at 1.00 atm and 298 K: CO₂, N₂O, or Cl₂? Explain.

The one with the highest molar mass assuming all other properties of the system is the same.
54. Which of the following statements best explains why a closed balloon filled with helium gas rises in air? (a) Helium is a monatomic gas, whereas nearly all the molecules that make up air, such as nitrogen and oxygen, are diatomic. (b) The average speed of helium atoms is greater than the average speed of air molecules, and the greater speed of collisions with the balloon walls propels the balloon upward. (c) Because the helium atoms are of lower mass than the average air molecule, the helium gas is less dense than air. The mass of the balloon is thus less than the mass of the air displaced by its volume. (d) Because helium has a lower molar mass than the average air molecule, the helium atoms are in faster motion. This means that the temperature of the helium is greater than the air temperature. Hot gases tend to rise.

\[ C \text{ is best answer.} \]

55. (a) Calculate the density of NO₂ gas at 0.970 atm and 35 °C. (b) Calculate the molar mass of a gas if 2.50 g occupies 0.875 L at 685 torr and 35 °C.

\[ \text{(a)} \quad \rho = \frac{PM}{RT} = \left(\frac{0.970 \ atm}{0.875 \ L}\right) \left(\frac{44.0 \ g}{mol}\right) \left(\frac{0.08205 \ L \ atm}{mol \ K}\right) \left(308 \ K\right) \]

\[ \text{(b)} \quad \frac{2.50 \ g}{0.875 \ L} = \left(\frac{0.901 \ atm}{0.08205 \ L \ atm}\right) \left(308 \ K\right) \]

56. Magnesium can be used as a “getter” in evacuated enclosures to react with the last traces of oxygen. (The magnesium is usually heated by passing an electric current through a wire or ribbon of the metal.) If an enclosure of 0.452 L has a partial pressure of O₂ of 3.5 * 10⁻⁶ torr at 27 °C, what mass of magnesium will react according to the following equation? 2 Mg(s) + O₂(g) → 2 MgO(s)

\[ \left(4.40 \times 10^{-5} \ atm\right) \times \left(0.452 \ L\right) = n \left(0.08205 \ L \ atm \ K\right) \left(300 \ K\right) \]

\[ n = 8.4 \times 10^{-10} \ mol \ O_2 \]

1.62 x 10⁻¹⁰ mol Mg is required.

57. The metabolic oxidation of glucose, C₆H₁₂O₆, in our bodies produces CO₂, which is expelled from our lungs as a gas: C₆H₁₂O₆(aq) + 6 O₂(g) → 6 CO₂(g) + 6 H₂O(l) (a) Calculate the volume of dry CO₂ produced at body temperature 37 °C and 0.970 atm when 24.5 g of glucose is consumed in this reaction. (b) Calculate the volume of oxygen you would need, at 1.00 atm and 298 K, to completely oxidize 50.0 g of glucose.

\[ \frac{24.5 \ g}{180 \ g} \times \frac{1 \ mol \ C₆H₁₂O₆}{1 \ mol \ CO₂} = 0.817 \ mol \ CO₂ \]

\[ (0.970 \ atm) \times (0.817 \ mol) \times (310 \ K) \]

\[ V = 21.4 \ L \]

58. A mixture containing 0.765 mol He(g), 0.330 mol Ne(g), and 0.110 mol Ar(g) is confined in a 10.00-L vessel at 25 °C. (a) Calculate the partial pressure of each of the gases in the mixture. (b) Calculate the total pressure of the mixture.

\[ 4 \text{c} \quad P_\text{He} = \frac{0.765 \ mol \times (268 \ atm)}{10.00 \ L} = 1.87 \ atm \]

\[ \text{(a) add together} \]

\[ P_\text{Ne} = 0.807 \ atm \]

\[ P_\text{Ar} = 0.263 \ atm \]

\[ \text{(b) total} \]

\[ 2.95 \ atm \]
59. The atmospheric concentration of CO₂ gas is presently 390 ppm (parts per million, by volume; that is, 390 L of every 100,000 L of the atmosphere are CO₂). What is the mole fraction of CO₂ in the atmosphere?

\[
\frac{390 \text{ L}}{100,000 \text{ L}} = 0.0039
\]

60. A piece of dry ice (solid carbon dioxide) with a mass of 5.50 g is placed in a 10.0-L vessel that already contains air at 705 torr and 24 °C. After the carbon dioxide has totally sublimed, what is the partial pressure of the resultant CO₂ gas, and the total pressure in the container at 24 °C?

\[
(0.928 \text{ atm})(10.0 \text{ L}) = n R (298 \text{ K})
\]

\[
n = 0.381 \text{ mol}
\]

\[
\frac{5.50 \text{ g}}{44.0} = 0.125 \text{ mol}
\]

\[
x (10.0 \text{ L}) = 0.500 \text{ mol} R (298 \text{ K})
\]

\[
x = 1.23 \text{ atm}
\]

61. A rigid vessel containing a 3:1 mol ratio of carbon dioxide and water vapor is held at 200 °C where it has a total pressure of 2.00 atm. If the vessel is cooled to 10 °C so that all of the water vapor condenses, what is the pressure of carbon dioxide? Neglect the volume of the liquid water that forms on cooling.

\[
\frac{P_1}{T_1} = \frac{P_2}{T_2}
\]

\[
\frac{2.00 \text{ atm}}{273 \text{ K}} = \frac{x}{283 \text{ K}}
\]

\[
x = 0.897 \text{ atm}
\]

62. At an underwater depth of 250 ft, the pressure is 8.38 atm. What should the mole percent of oxygen be in the diving gas for the partial pressure of oxygen in the mixture to be 0.21 atm, the same as in air at 1 atm?

\[
\frac{0.21 \text{ atm}}{8.38 \text{ atm}} = 0.025
\]

63. Determine whether each of the following changes will increase, decrease, or not affect the rate with which gas molecules collide with the walls of their container: (a) increasing the volume of the container, (b) increasing the temperature, (c) increasing the molar mass of the gas.

A) decrease  B) increase  C) decrease

64. Which assumptions are common to both kinetic-molecular theory and the ideal-gas equation?

i) The volume of gas molecules is negligible relative to the container.

ii) Attractive and repulsive forces among molecules are negligible.

65. WF₆ is one of the heaviest known gases. How much slower is the root-mean-square speed of WF₆ than He at 300 K?

\[
u_{\text{rms}} = \sqrt{\frac{3kT}{M}} = \sqrt{\frac{3(8.31 \text{ J/mol} \cdot \text{K})(300 \text{ K})}{297.837 \text{ g/mol}}}
\]

\[
u_{\text{rms}} \approx \sqrt{\frac{3(8.32)(300 \text{ K})}{4}}
\]

WF₆ moves 8 times slower.
66. The temperature of a 5.00-L container of N₂ gas is increased from 20 °C to 250 °C. If the volume is held constant, predict qualitatively how this change affects the following: (a) the average kinetic energy of the molecules; (b) the root-meansquare speed of the molecules; (c) the strength of the impact of an average molecule with the container walls; (d) the total number of collisions of molecules with walls per second.

A) Increases
B) Increases
C) Increases
D) Increases

67. (a) Place the following gases in order of increasing average molecular speed at 25 °C: Ne, HBr, SO₂, NF₃, CO. (b) Calculate the rms speed of NF₃ molecules at 25 °C. (c) Calculate the most probable speed of an ozone molecule in the stratosphere, where the temperature is 270 K.

A) HBr < NF₃ < SO₂ < CO < N₂
B) \( v_{\text{rms}} = \sqrt{\frac{3RT}{M}} = 341.2 \text{ m/s} \)
C) \( v_{\text{mp}} = \sqrt{\frac{2RT}{2}} = 306 \text{ m/s} \)

68. Explain the difference between effusion and diffusion.

- Effusion: Flow of a gas through a small hole
- Diffusion: Migration of a gas throughout a substance

69. Hydrogen has two naturally occurring isotopes, \(^1\text{H}\) and \(^2\text{H}\). Chlorine also has two naturally occurring isotopes, \(^35\text{Cl}\) and \(^37\text{Cl}\). Thus, hydrogen chloride gas consists of four distinct types of molecules: \(^1\text{H}^35\text{Cl}\), \(^1\text{H}^37\text{Cl}\), \(^2\text{H}^35\text{Cl}\), and \(^2\text{H}^37\text{Cl}\). Place these four molecules in order of increasing rate of effusion.

\[ ^2\text{H}^37\text{Cl} < ^1\text{H}^37\text{Cl} < ^2\text{H}^35\text{Cl} < ^1\text{H}^35\text{Cl} \]

70. Arsenic(III) sulfide sublimes readily, even below its melting point of 320 °C. The molecules of the vapor phase are found to effuse through a tiny hole at 0.28 times the rate of effusion of Ar atoms under the same conditions of temperature and pressure. What is the molecular formula of arsenic(III) sulfide in the gas phase?
71. (a) List two experimental conditions under which gases deviate from ideal behavior. (b) List two reasons why the gases deviate from ideal behavior.

A) high pressure and low temp

B) real volume of gas molecules and attractive intermolecular forces between molecules cause gases to behave non-ideally

72. Based on their respective van der Waals constants (Table 10.3), is Ar or CO$_2$ expected to behave more nearly like an ideal gas at high pressures? Explain.

Ar because both a and b are smaller than CO$_2$

73. Table 10.3 shows that the van der Waals $b$ parameter has units of L/mol. This implies that we can calculate the size of atoms or molecules from $b$. Using the value of $b$ for Xe, calculate the radius of a Xe atom and compare it to the value found in Figure 7.7, that is, 1.40 Å. Recall that the volume of a sphere is $\frac{4}{3}\pi r^3$.

\[
b = 0.0510 \text{ L/mol} \implies \frac{1 \text{ mol}}{1.02 \times 10^{23} \text{ particles}} = \sqrt[3]{\frac{4}{3}\pi r^3}
\]