CHAPTER 10

Gases

Section 10.1 *Properties of Gases*

2. The following are observed properties of gases:
   (a) Gases have a variable volume.
   (b) Gases expand infinitely.
   (c) Gases compress uniformly.
   (d) Gases have low densities.

Section 10.2 *Atmospheric Pressure*

4.

	Units	Standard Pressure
(a)	inches of mercury	29.9 in. Hg
(b)	pounds per square inch	14.7 psi

6. (a) $9.50\ \cancel{\text{atm}} \times \dfrac{29.9\ \text{in. Hg}}{1\ \cancel{\text{atm}}} = 284\ \text{in. Hg}$

   (b) $9.50\ \cancel{\text{atm}} \times \dfrac{760\ \text{mm Hg}}{1\ \cancel{\text{atm}}} = 7220\ \text{mm Hg}$

8. (a) $104\ \cancel{\text{kPa}} \times \dfrac{14.7\ \text{psi}}{101\ \cancel{\text{kPa}}} = 15.1\ \text{psi}$

   (b) $104\ \cancel{\text{kPa}} \times \dfrac{76\ \text{cm Hg}}{101\ \cancel{\text{kPa}}} = 78.3\ \text{cm Hg}$

Section 10.3 *Variables Affecting Gas Pressure*

10.

	Change	Observation	Explanation
(a)	volume decreases	pressure increases	molecules are closer together and collide more frequently
(b)	temperature decreases	pressure decreases	molecules are moving slower and collide with lower frequency and less energy
(c)	moles of gas decrease	pressure decreases	fewer molecules have fewer collisions

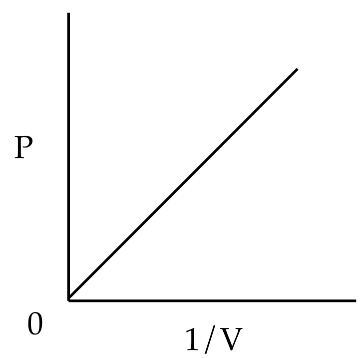
12.

	Change	Observation
(a)	volume decreases	pressure increases
(b)	temperature decreases	pressure decreases
(c)	moles of gas increase	pressure increases

14. Increasing the temperature of a gas causes gas molecules to move faster, and they collide with the container more frequently and with more energy. As a result, the pressure of the gas increases.

Section 10.4 *Boyle's Law: Pressure–Volume Relationships*

16. Pressure vs. Reciprocal of Volume



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

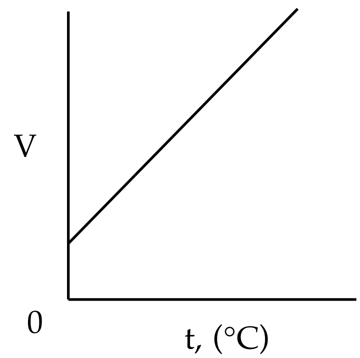
$$15.0 \text{ psi} \times \frac{555 \cancel{\text{mL}}}{275 \cancel{\text{mL}}} = 30.3 \text{ psi}$$

20. $V_1 \times P_{\text{factor}} = V_2$

$$125 \text{ mL} \times \frac{705 \cancel{\text{mm Hg}}}{385 \cancel{\text{mm Hg}}} = 229 \text{ mL}$$

Section 10.5 *Charles's Law: Volume–Temperature Relationships*

22. <u>Volume vs. Celsius Temperature</u>



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

$100\ °C + 273 = 373\ K$
$20\ °C + 273 = 293\ K$

$$2.50\ L \times \frac{293\ \cancel{K}}{373\ \cancel{K}} = 1.96\ L$$

26. $T_1 \times V_{\text{factor}} = T_2$

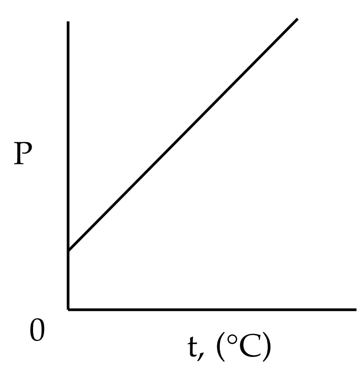
$50\ °C + 273 = 323\ K$

$$323\ K \times \frac{3.75\ \cancel{L}}{5.00\ \cancel{L}} = 242\ K$$

$242\ K - 273 = -31\ °C$

Section 10.6 *Gay–Lussac's Law: Pressure–Temperature Relationships*

28. <u>Pressure vs. Celsius Temperature</u>



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

$100\ °C + 273 = 373\ K$
$10\ °C + 273 = 283\ K$

$$455\ \text{cm Hg} \times \frac{283\ \cancel{K}}{373\ \cancel{K}} = 345\ \text{cm Hg}$$

32. $T_1 \times P_{factor} = T_2$

$0\ °C + 273 = 273\ K$

$$273\ K \times \frac{225\ \cancel{\text{mm Hg}}}{375\ \cancel{\text{mm Hg}}} = 164\ K$$

$164\ K - 273 = -109\ °C$

Section 10.7 *Combined Gas Law*

34.

	P	V	T
initial	772 mm Hg	100.0 mL	21 °C + 273 = 294 K
final	760 mm Hg	V_2	273 K

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

$$100.0\ \text{mL} \times \frac{772\ \cancel{\text{mm Hg}}}{760\ \cancel{\text{mm Hg}}} \times \frac{273\ \cancel{K}}{294\ \cancel{K}} = 94.3\ \text{mL}$$

36.

	P	V	T
initial	650 mm Hg	25.0 mL	–25 °C + 273 = 248 K
final	350 mm Hg	V_2	25 °C + 273 = 298 K

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

$$25.0\ \text{mL} \times \frac{650\ \cancel{\text{mm Hg}}}{350\ \cancel{\text{mm Hg}}} \times \frac{298\ \cancel{K}}{248\ \cancel{K}} = 55.8\ \text{mL}$$

38.

	P	V	T
initial	760 mm Hg	1250 mL	273 K
final	P_2	255 mL	300 °C + 273 = 573 K

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

$$760 \text{ mm Hg} \times \frac{1250 \cancel{\text{mL}}}{255 \cancel{\text{mL}}} \times \frac{573 \cancel{\text{K}}}{273 \cancel{\text{K}}} = 7820 \text{ mm Hg}$$

40.

	P	V	T
initial	760 mm Hg	50.0 mL	273 K
final	350 mm Hg	350.0 mL	T_2

$$T_1 \times P_{factor} \times V_{factor} = T_2$$

$$273 \text{ K} \times \frac{350 \cancel{\text{mm Hg}}}{760 \cancel{\text{mm Hg}}} \times \frac{350.0 \cancel{\text{mL}}}{50.0 \cancel{\text{mL}}} = 880 \text{ K}$$

$$880 \text{ K} - 273 = 607 \text{ °C}$$

42.

	P	V	T
initial	75.0 cm Hg	500.0 mL	–185 °C + 273 = 88 K
final	55.0 cm Hg	225.0 mL	T_2

$$T_1 \times P_{factor} \times V_{factor} = T_2$$

$$88 \text{ K} \times \frac{55.0 \cancel{\text{cm Hg}}}{75.0 \cancel{\text{cm Hg}}} \times \frac{225.0 \cancel{\text{mL}}}{500.0 \cancel{\text{mL}}} = 29 \text{ K}$$

$$29 \text{ K} - 273 = -244 \text{ °C}$$

Section 10.8 *The Vapor Pressure Concept*

44. The vapor pressure of water is much greater than the vapor pressure of mercury at the same temperature. Thus, the vapor pressure of water can be measured by placing a drop of liquid water on a column of mercury in a barometer.

46.

	Temperature	Vapor Pressure
(a)	75 °C	289.1 mm Hg = 0.3804 atm
(b)	100 °C	760.0 mm Hg = 1.000 atm

Section 10.9 *Dalton's Law of Partial Pressures*

48. $P_{\text{nitrogen}} = 587 \text{ mm Hg}$
$P_{\text{oxygen}} = 158 \text{ mm Hg}$
$P_{\text{argon}} = 7 \text{ mm Hg}$

$P_{\text{nitrogen}} + P_{\text{oxygen}} + P_{\text{argon}} = P_{\text{atmosphere}}$
$587 \text{ mm Hg} + 158 \text{ mm Hg} + 7 \text{ mm Hg} = 752 \text{ mm Hg}$

50. $P_{\text{total}} = 5.00 \cancel{\text{atm}} \times \dfrac{760 \text{ in. Hg}}{1 \cancel{\text{atm}}} = 3800 \text{ mm Hg}$
$P_{\text{nitrogen}} = 1850 \text{ mm Hg}$
$P_{\text{hydrogen}} = 1150 \text{ mm Hg}$

$P_{\text{nitrogen}} + P_{\text{hydrogen}} + P_{\text{ammonia}} = P_{\text{total}}$
$P_{\text{ammonia}} = P_{\text{total}} - P_{\text{nitrogen}} - P_{\text{hydrogen}}$
$P_{\text{ammonia}} = 3800 \text{ mm Hg} - 1850 \text{ mm Hg} - 1150 \text{ mm Hg} = 800 \text{ mm Hg}$

52. The term "wet" gas refers to a gas collected over water. A "wet" gas contains water vapor in the gaseous sample; a "dry" gas does not contain water vapor. The pressure exerted by a "dry" gas is equal to the pressure of the "wet" gas minus the vapor pressure of water at the given temperature.

54. $\text{P}_{\text{total}} = 755 \text{ mm Hg}$
$\text{P}_{\text{water vapor}} = 31.8 \text{ mm Hg at } 30\ °\text{C}$

$\text{P}_{\text{water vapor}} + \text{P}_{\text{xenon}} = \text{P}_{\text{total}}$
$\text{P}_{\text{xenon}} = \text{P}_{\text{total}} - \text{P}_{\text{water vapor}}$
$\text{P}_{\text{xenon}} = 755 \text{ mm Hg} - 31.8 \text{ mm Hg} = 723 \text{ mm Hg}$

Section 10.10 *Ideal Gas Behavior*

56. An ideal gas strictly obeys the gas laws under all conditions. A real gas deviates from ideal gas behavior; this deviation is especially severe under conditions of *low temperature* and *high pressure.*

58. At a temperature of –273 °C (absolute zero), a gas possesses zero kinetic energy.

60.

	Description	Gas
(a)	highest kinetic energy	all gases are equal
(b)	lowest kinetic energy	all gases are equal
(c)	highest velocity	H_2 molecules (lightest molecules)
(d)	lowest velocity	O_2 molecules (heaviest molecules)

Note: Since all three gases are at the same temperature, each has the same kinetic energy.

62. An ideal gas at 0 K (absolute zero) theoretically occupies zero volume.

Section 10.11 *Ideal Gas Law*

64. $V = \frac{n\,R\,T}{P}$

$n = 1.25 \text{ mol } O_2$

$T = 25\text{ °C} + 273 = 298\text{ K}$

$P = 1200 \cancel{\text{mm Hg}} \times \frac{1 \text{ atm}}{760 \cancel{\text{mm Hg}}} = 1.58 \text{ atm}$

$$V = \frac{1.25 \cancel{\text{mol}} \times 298 \cancel{\text{K}}}{1.58 \cancel{\text{atm}}} \times \frac{0.0821 \cancel{\text{atm}} \cdot \text{L}}{1 \cancel{\text{mol}} \cdot \cancel{\text{K}}} = 19.4 \text{ L}$$

66. $T = \frac{P\,V}{n\,R}$

$P = 725 \cancel{\text{mm Hg}} \times \frac{1 \text{ atm}}{760 \cancel{\text{mm Hg}}} = 0.954 \text{ atm}$

$V = 2.15 \text{ L}$

$n = 0.100 \text{ mol}$

$$T = \frac{0.954 \cancel{\text{atm}} \times 2.15 \cancel{\text{L}}}{0.100 \cancel{\text{mol}}} \times \frac{1 \cancel{\text{mol}} \cdot \text{K}}{0.0821 \cancel{\text{atm}} \cdot \cancel{\text{L}}} = 250 \text{ K}$$

$$250 \text{ K} - 273 = -23 \text{ °C}$$

General Exercises

68. $165\ \cancel{\text{cm}^2} \times \left(\dfrac{1\ \cancel{\text{in.}}}{2.54\ \cancel{\text{cm}}}\right)^2 \times \dfrac{14.7\ \text{lb}}{1\ \cancel{\text{in.}^2}} = 376\ \text{lb}$

70. $76.0\ \cancel{\text{cm Hg}} \times \dfrac{19.2\ \text{cm H}_2\text{O}}{1\ \cancel{\text{cm Hg}}} = 1460\ \text{cm alcohol}$

72. $P_{\text{nitrogen}} = 51.0\ \cancel{\text{psi}} \times \dfrac{1\ \text{atm}}{14.7\ \cancel{\text{psi}}} = 3.47\ \text{atm}$

$P_{\text{oxygen}} = 13.5\ \cancel{\text{psi}} \times \dfrac{1\ \text{atm}}{14.7\ \cancel{\text{psi}}} = 0.918\ \text{atm}$

$P_{\text{argon}} = 0.5\ \cancel{\text{psi}} \times \dfrac{1\ \text{atm}}{14.7\ \cancel{\text{psi}}} = 0.03\ \text{atm}$

$P_{\text{total}} = 3.47\ \text{atm} + 0.918\ \text{atm} + 0.03\ \text{atm} = 4.42\ \text{atm}$

74. $P_{\text{hydrogen}} + P_{\text{water vapor}} = P_{\text{total}}$

$P_{\text{hydrogen}} = P_{\text{total}} - P_{\text{water vapor}}$

$P_{\text{hydrogen}} = 758\ \text{mm Hg} - 17.5\ \text{mm Hg} = 741\ \text{mm Hg}$

	P	V	T
initial	741 mm Hg	95.0 mL	20 °C + 273 = 293 K
final	760 mm Hg	V_2	273 K

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

$$95.0\ \text{mL} \times \frac{741\ \cancel{\text{mm Hg}}}{760\ \cancel{\text{mm Hg}}} \times \frac{273\ \cancel{\text{K}}}{293\ \cancel{\text{K}}} = 86.3\ \text{mL}$$

76. $P = \frac{n\,R\,T}{V}$

$$n = 3.38 \times 10^{22}\ \cancel{\text{molecules}} \times \frac{1\ \text{mol}}{6.02 \times 10^{23}\ \cancel{\text{molecules}}} = 0.0561\ \text{mol}$$

$$T = 100\ °\text{C} + 273 = 373\ \text{K} \qquad P = 255\ \cancel{\text{torr}} \times \frac{1\ \text{atm}}{760\ \cancel{\text{torr}}} = 0.336\ \text{atm}$$

$$V = \frac{0.0561\ \cancel{\text{mol}} \times 373\ \cancel{\text{K}}}{0.336\ \cancel{\text{atm}}} \times \frac{0.0821\ \cancel{\text{atm}} \cdot \text{L}}{1\ \cancel{\text{mol}} \cdot \cancel{\text{K}}} = 5.11\ \text{L}$$

78. Propane, C_3H_8, molecules have a lower molecular mass than butane, C_4H_{10}; therefore, C_3H_8 molecules have the fastest velocity.

80. $n = \frac{P\,V}{R\,T}$

$$P = 710\ \cancel{\text{mm Hg}} \times \frac{1\ \text{atm}}{760\ \cancel{\text{mm Hg}}} = 0.934\ \text{atm}$$

$$V = 855\ \cancel{\text{mL}} \times \frac{1\ \text{L}}{1000\ \cancel{\text{mL}}} = 0.855\ \text{L}$$

$$T = 155\ °\text{C} + 273 = 428\ \text{K}$$

$$n = \frac{0.934\ \cancel{\text{atm}} \times 0.855\ \cancel{\text{L}}}{428\ \cancel{\text{K}}} \times \frac{1\ \text{mol} \cdot \cancel{\text{K}}}{0.0821\ \cancel{\text{atm}} \cdot \cancel{\text{L}}} = 0.0227\ \text{mol}$$

$$0.0227\ \cancel{\text{mol F}_2} \times \frac{38.00\ \text{g F}_2}{1\ \cancel{\text{mol F}_2}} = 0.863\ \text{g F}_2$$

Challenge Exercises

82. $MM = \frac{g\,R\,T}{P\,V}$

$g = 5.40\text{ g}$
$V = 1.00\text{ L}$
$P = 1\text{ atm}$
$T = 0\text{ °C} + 273 = 273\text{ K}$

$$MM = \frac{5.40\text{ g} \times 273\ \cancel{\text{K}}}{1\ \cancel{\text{atm}} \times 1.00\ \cancel{\text{L}}} \times \frac{0.0821\ \cancel{\text{atm}} \cdot \cancel{\text{L}}}{1\text{ mol} \cdot \cancel{\text{K}}} = 121\text{ g/mol}$$

The molar mass of $\mathbf{CF_2Cl_2}$ is closest to the calculated value of 121 g/mol.

84. $MM = \frac{g\,R\,T}{P\,V}$

$g = 2.85\text{ g}$

$V = 750\ \cancel{\text{mL}} \times \frac{1\text{ L}}{1000\ \cancel{\text{mL}}} = 0.750\text{ L}$

$P = 760\text{ mm Hg} = 1\text{ atm}$

$T = 100\text{ °C} + 273 = 373\text{ K}$

$$MM = \frac{2.85\text{ g} \times 373\ \cancel{\text{K}}}{1\ \cancel{\text{atm}} \times 0.750\ \cancel{\text{L}}} \times \frac{0.0821\ \cancel{\text{atm}} \cdot \cancel{\text{L}}}{1\text{ mol} \cdot \cancel{\text{K}}} = 116\text{ g/mol}$$

Online Exercises

86. There is 2.0 g of oxygen gas, and 98.0 g of helium gas, in 100 g of a standard "heliox" mixture.