

Chapter 10. Gases

Media Resources

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Figure 10.3 A Mercury Manometer
Figure 10.7 Boyle's Law
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Figure 10.18 The Effect of Molecular Mass on Molecular Speed at 25 °C.
Figure 10.22 The Effect of Pressure on the Behavior of Several Real Gases
Figure 10.23 The Effect of Temperature and Pressure on the Behavior of Nitrogen Gas

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Gas Laws
Density of Gases

Partial Pressures
Motions of a Gas
Gas Phase: Boltzmann Distribution
Gas Diffusion and Effusion
Nonideal Gas Behavior

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P - V Relationships
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Kinetic Energy of a Gas

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Diffusion of Bromine Vapor

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Other Resources

Further Readings:

Gases and Their Behavior
 Carbon Dioxide Flooding: A Classroom Case Study
 Derived from Surgical Practice
 Gay-Lussac: Chemist Extraordinary
 Gay-Lussac after 200 Years
 The Chemistry behind the Air Bag

Chemistry of Air Bags

Cinema, Flirts, Snakes, and Gases
 Toy Flying Saucers and Molecular Speeds

Live Demonstrations:

Boiling at Reduced Pressure
 Boyle's Law
 Boyle's Law and the Monster Marshmallow
 Robert Boyle: The Founder of Modern Chemistry
 Boyle's Law and the Mass of a Textbook
 Effect of Pressure on the Size of a Balloon
 Thermal Expansion of Gases
 Charles' Law of Gases: A Simple Experimental
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Chapter 10. Gases

Common Student Misconceptions

- Students need to be told to *always* use temperature in Kelvin in gas problems.
- Due to several systems of units, students often use ideal gas constants with units inconsistent with values.
- Students often confuse the standard conditions for gas behavior (STP) with the standard conditions in thermodynamics.
- Ideal gas behavior should be discussed as just that, *ideal*; students should be reminded that real gases do not behave ideally, especially at high pressures and/or low temperatures.
- Students expect a change in the gas particle distribution upon temperature changes at constant V .
- Students commonly confuse effusion and diffusion.

Teaching Tips

- Students should always use units in gas-law problems to keep track of required conversions. Encourage them to use dimensional analysis to detect conversion errors.
- Students may be confused by seeing densities given in units of g/mL; this is the case for liquids and solids. The densities of gases are usually expressed in g/L.
- It is helpful to remind students that the sum of the mole fractions of a mixture must equal 1, i.e., $\sum X_i = 1$.

Lecture Outline

10.1 Characteristics of Gases^{1,2}

- All substances have three phases: solid, liquid and gas.
- Substances that are liquids or solids under ordinary conditions may also exist as gases.
 - These are often referred to as **vapors**.
- Many of the properties of gases differ from those of solids and liquids:
 - Gases are highly compressible and occupy the full volume of their containers.
 - When a gas is subjected to pressure, its volume decreases.
 - Gases always form homogeneous mixtures with other gases.
- Gases only occupy a small fraction of the volume of their containers.
 - As a result, each molecule of gas behaves largely as though other molecules were absent.

FORWARD REFERENCES

- Thermodynamics of phase changes will be discussed in Chapter 19.
- Such important gaseous reactions as the Haber process or equilibria involving nitrogen oxides will be covered in Chapter 15.

10.2 Pressure

- **Pressure** is the force acting on an object per unit area:

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

¹ “Gases and Their Behavior” from Further Readings

Atmospheric Pressure and the Barometer^{3,4,5,6,7}

- The SI unit of force is the *newton* (N).
 - $1 \text{ N} = 1 \text{ kg}\cdot\text{m}/\text{s}^2$
- The SI unit of pressure is the **pascal** (Pa).
 - $1 \text{ Pa} = 1 \text{ N}/\text{m}^2$
 - A related unit is the **bar**, which is equal to 10^5 Pa .
 - Another pressure unit is pounds per square inch (psi, lbs/in^2).
- Gravity exerts a force on the Earth's atmosphere.
 - A column of air 1 m^2 in cross section extending to the upper atmosphere exerts a force of 10^5 N .
 - Thus, the pressure of a 1 m^2 column of air extending to the upper atmosphere is 100 kPa .
 - Atmospheric pressure at sea level is about 100 kPa or 1 bar or 14.7 psi .
 - The actual atmospheric pressure at a specific location depends on the altitude and weather conditions.
- Atmospheric pressure is measured with a *barometer*.
 - If a tube is completely filled with mercury and then inverted into a container of mercury open to the atmosphere, the mercury will rise 760 mm up the tube.
 - **Standard atmospheric pressure** is the pressure required to support 760 mm of Hg in a column.
 - Important non-SI units used to express gas pressure include:
 - **atmospheres** (atm)
 - *millimeter of mercury* (mm Hg) or **torr**
 - $1 \text{ atm} = 760 \text{ mm Hg} = 760 \text{ torr} = 1.01325 \times 10^5 \text{ Pa} = 101.325 \text{ kPa} = 1.01325 \text{ bar}$.
- The pressure of enclosed gases is measured with a *manometer*.

FORWARD REFERENCES

- Osmotic pressure (in atm) will be calculated in Chapter 13 (section 13.5).
- K_p 's and thermodynamic equilibrium constants in Chapter 15 will use pressure (in atm).
- Pressure and Le Châtelier's principle will be discussed in Chapter 15 (section 15.7).

10.3 The Gas Laws

- The equations that express the relationships among T (temperature), P (pressure), V (volume), and n (number of moles of gas) are known as *gas laws*.

The Pressure-Volume Relationship: Boyle's Law^{8,9,10,11,12,13}

- Weather balloons are used as a practical application of the relationship between pressure and volume of a gas.
 - As the weather balloon ascends, the volume increases.
 - As the weather balloon gets further from Earth's surface, the atmospheric pressure decreases.

² "Carbon Dioxide Flooding: A Classroom Case Study Derived from Surgical Practice" from Further Readings

³ Figure 10.1 from Transparency Pack

⁴ Figure 10.2 from Transparency Pack

⁵ "Boiling at Reduced Pressure" from Live Demonstrations

⁶ "Manometer" Activity from Instructor's Resource CD/DVD

⁷ Figure 10.3 from Transparency Pack

⁸ "Boyle's Law" from Live Demonstrations

⁹ "Boyle's Law and the Monster Marshmallow" from Live Demonstrations

¹⁰ "Robert Boyle: The Founder of Modern Chemistry" from Live Demonstrations

¹¹ Figure 10.7 from Transparency Pack

¹² "Boyle's Law and the Mass of a Textbook" from Live Demonstrations

¹³ " P - V Relationships" Animation from Instructor's Resource CD/DVD

- **Boyle's law:** The volume of a fixed quantity of gas, at constant temperature, is inversely proportional to its pressure.
- Mathematically:

$$V = \text{constant} \times \frac{1}{P} \text{ or } PV = \text{constant}$$

- A plot of V versus P is a hyperbola.
- A plot of V versus $1/P$ must be a straight line passing through the origin.
- The working of the lungs illustrates Boyle's law.
 - As we breathe in, the diaphragm moves down, and the ribs expand; therefore, the volume of the lungs increases.
 - According to Boyle's law, when the volume of the lungs increases, the pressure decreases; therefore, the pressure inside the lungs is less than atmospheric pressure.
 - Atmospheric pressure then forces air into the lungs until the pressure once again equals atmospheric pressure.
 - As we breathe out, the diaphragm moves up and the ribs contract. Therefore, the volume of the lungs decreases.
 - By Boyle's law, the pressure increases and air is forced out.

The Temperature-Volume Relationship: Charles's Law^{14,15,16}

- We know that hot-air balloons expand when they are heated.
- **Charles's law:** The volume of a fixed quantity of gas at constant pressure is directly proportional to its absolute temperature.
- Mathematically:

$$V = \text{constant} \times T \text{ or } \frac{V}{T} = \text{constant}$$

- Note that the value of the constant depends on the pressure and number of moles of gas.
- A plot of V versus T is a straight line.
- When T is measured in °C, the intercept on the temperature axis is -273.15 °C.
- We define *absolute zero*, $0 \text{ K} = -273.15$ °C.

The Quantity-Volume Relationship: Avogadro's Law^{17,18}

- Gay-Lussac's *law of combining volumes*: At a given temperature and pressure the volumes of gases that react with one another are ratios of small whole numbers.
- **Avogadro's hypothesis:** Equal volumes of gases at the same temperature and pressure contain the same number of molecules.
- **Avogadro's law:** The volume of gas at a given temperature and pressure is directly proportional to the number of moles of gas.
 - Mathematically:

$$V = \text{constant} \times n$$

- We can show that 22.4 L of any gas at 0 °C and 1 atmosphere contains 6.02×10^{23} gas molecules.

FORWARD REFERENCES

- Vapor pressure vs. temperature will be discussed in Chapter 13 (section 13.5).

¹⁴ "Effect of Pressure on the Size of a Balloon" from Live Demonstrations

¹⁵ "Charles' Law of Gases" from Live Demonstrations

¹⁶ "Charles' Law: The Relationship Between Volume and Temperature of a Gas" from Live Demonstrations

¹⁷ "Thermal Expansion of Gases" from Live Demonstrations

¹⁸ "Gas Laws" Activity from Instructor's Resource CD/DVD

- Increasing entropy of gases with temperature as well as entropy of gases vs. other states of matter will be discussed in Chapter 19 (section 19.3).

10.4 The Ideal-Gas Equation^{19,20,21,22}

- Summarizing the gas laws:
 - Boyle: $V \propto 1/P$ (constant n, T)
 - Charles: $V \propto T$ (constant n, P)
 - Avogadro: $V \propto n$ (constant P, T)
 - Combined: $V \propto nT/P$
- Ideal-gas equation** or **ideal-gas law**: $PV = nRT$
 - An **ideal gas** is a hypothetical gas whose $P, V,$ and T behavior is completely described by the ideal-gas equation.
 - $R = \text{gas constant} = 0.08206 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K}$
 - Other numerical values of R in various units are given in Table 10.2.
- Define **STP (standard temperature and pressure)** = $0^\circ\text{C}, 273.15 \text{ K}, 1 \text{ atm}$.
 - The molar volume of 1 mol of an ideal gas at STP is 22.41 L.

Relating the Ideal-Gas Equation and the Gas Laws

- If $PV = nRT$ and n and T are constant, then PV is constant and we have Boyle's law.
 - Other laws can be generated similarly.
- In general, if we have a gas under two sets of conditions, then

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

- We often have a situation in which $P, V,$ and T all change for a fixed number of moles of gas.
 - For this set of circumstances,

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

- Which gives the *combined gas law*.

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

FORWARD REFERENCES

- The ideal gas constant will be used in Chapter 14 in the Arrhenius equation (section 14.5).
- The ideal gas constant will be used in Chapter 15 in conversions between K_c and K_p (section 15.2) and to relate Gibbs free energy with the equilibrium constant in Chapter 19 (section 19.7) as well as with the cell potentials in Chapter 20 (sections 20.5 and 20.6).

10.5 Further Applications of the Ideal-Gas Equation

Gas Densities and Molar Mass²³

- Density has units of mass over volume.

¹⁹ "Gay-Lussac: Chemist Extraordinary" from Further Readings

²⁰ "Gay-Lussac after 200 Years" from Further Readings

²¹ "Collapsing Can" from Live Demonstrations

²² Figure 10.11 from Transparency Pack

²³ "Density of Gases" Activity from Instructor's Resource CD/DVD

- Rearranging the ideal-gas equation with M as molar mass we get

$$\frac{n}{V} = \frac{P}{RT}$$

$$\frac{nM}{V} = \frac{PM}{RT}$$

$$\therefore d = \frac{nM}{V} = \frac{PM}{RT}$$

- The molar mass of a gas can be determined as follows:

$$M = \frac{dRT}{P}$$

Volumes of Gases in Chemical Reactions^{24,25,26}

- The ideal-gas equation relates P , V , and T to number of moles of gas.
- The n can then be used in stoichiometric calculations.

FORWARD REFERENCES

- Solubility of gases vs. temperature (Henry's law) will be covered in Chapter 13 (section 13.3).

10.6 Gas Mixtures and Partial Pressures²⁷

- Since gas molecules are so far apart, we can assume they behave independently.
- Dalton observed:
 - The total pressure of a mixture of gases equals the sum of the pressures that each would exert if present alone.
 - Partial pressure** is the pressure exerted by a particular component of a gas mixture.
- Dalton's law of partial pressures:** In a gas mixture the total pressure is given by the sum of partial pressures of each component:

$$P_t = P_1 + P_2 + P_3 + \dots$$

- Each gas obeys the ideal gas equation.
 - Thus,

$$P_t = (n_1 + n_2 + n_3 + \dots) \frac{RT}{V} = n_t \frac{RT}{V}$$

Partial Pressures and Mole Fractions

- Let n_1 be the number of moles of gas 1 exerting a partial pressure P_1 , then

$$P_1 = X_1 P_t$$
 - Where X_1 is the **mole fraction** (n_1/n_t).
 - Note that a mole fraction is a dimensionless number.

Collecting Gases over Water²⁸

- It is common to synthesize gases and collect them by displacing a volume of water.

²⁴ "The Chemistry Behind the Air Bag" from Further Readings

²⁵ "The Chemistry of Air Bags" from Further Readings

²⁶ "Air Bags" Animation from Instructor's Resource CD/DVD

²⁷ "Partial Pressures" Activity from Instructor's Resource CD/DVD

²⁸ "Determining the Molecular Weight of a Gas" from Live Demonstrations

- To calculate the amount of gas produced, we need to correct for the partial pressure of the water:

$$P_{\text{total}} = P_{\text{gas}} + P_{\text{water}}$$

- The vapor pressure of water varies with temperature.
 - Values can be found in Appendix B.

FORWARD REFERENCES

- Vapor pressure, volatility, and temperature relationships will be introduced in Chapter 11 (section 11.5) and further applied to Raoult's Law in Chapter 13 (section 13.5).
- Air – a mixture of gases – will be discussed in Chapter 18 (section 18.1) and 22 (section 22.7).

10.7 Kinetic-Molecular Theory of Gases^{29,30,31,32,33,34}

- The **kinetic molecular theory of gases** was developed to *explain* gas behavior.
 - It is a theory of moving molecules.
- Summary:
 - Gases consist of a large number of molecules in constant random motion.
 - The combined volume of all the molecules is negligible compared with the volume of the container.
 - Intermolecular forces (forces between gas molecules) are negligible.
 - Energy can be transferred between molecules during collisions, but the average kinetic energy is constant at constant temperature.
 - Energy can be transferred between molecules during collisions but, as long as the temperature remains constant, the *average* kinetic energy of the molecules does not change with time (the collisions are perfectly elastic).
 - The average kinetic energy of the gas molecules is proportional to the absolute temperature.
- Kinetic molecular theory gives us an *understanding* of pressure and temperature on the molecular level.
 - The pressure of a gas results from the collisions with the walls of the container.
 - The magnitude of the pressure is determined by how often and how hard the molecules strike.
- The absolute temperature of a gas is a measure of the average kinetic energy.
 - Some molecules will have less kinetic energy or more kinetic energy than the average (distribution).
 - There is a spread of individual energies of gas molecules in any sample of gas.
 - As the temperature increases, the average kinetic energy of the gas molecules increases.

Distributions of Molecular Speed³⁵

- As kinetic energy increases, the velocity of the gas molecules increases.
 - Root-mean-square (rms) speed**, u_{rms} , is the speed of a gas molecule having average kinetic energy.
- Average kinetic energy, ε , is related to rms speed:

$$\varepsilon = \frac{1}{2} mu^2$$
 - where m = mass of the molecule.

²⁹ “Cinema, Flirts, Snakes, and Gases” from Further Readings

³⁰ “Kinetic Energy of a Gas” Animation from Instructor’s Resource CD/DVD

³¹ “Motions of a Gas” Activity from Instructor’s Resource CD/DVD

³² Figure 10.17 from Transparency Pack

³³ “Gas Phase: Boltzmann Distribution” Activity from Instructor’s Resource CD/DVD

³⁴ “Toy Flying Saucers and Molecular Speeds” from Further Readings

³⁵ “Nitrogen” 3-D Model from Instructor’s Resource CD/DVD

Application of Kinetic-Molecular Theory to the Gas-Laws

- We can understand empirical observations of gas properties within the framework of the kinetic-molecular theory.
- *Effect of an increase in volume (at constant temperature):*
 - As volume increases at constant temperature, the average kinetic of the gas remains constant.
 - Therefore, u is constant.
 - However, volume increases, so the gas molecules have to travel further to hit the walls of the container.
 - Therefore, pressure decreases.
- *Effect of an increase in temperature (at constant volume):*
 - If temperature increases at constant volume, the average kinetic energy of the gas molecules increases.
 - There are more collisions with the container walls.
 - Therefore, u increases.
 - The change in momentum in each collision increases (molecules strike harder).
 - Therefore, pressure increases.

FORWARD REFERENCES

- The collision model in Chapter 14 (section 14.5) will be based on the kinetic-molecular theory.

10.8 Molecular Effusion and Diffusion^{36,37}

- The average kinetic energy of a gas is related to its mass:

$$\varepsilon = \frac{1}{2} mu^2$$
- Consider two gases at the same temperature: the lighter gas has a higher rms speed than the heavier gas.
 - Mathematically:

$$u_{rms} = \sqrt{\frac{3RT}{M}}$$

- The lower the molar mass, M , the higher the rms speed for that gas at a constant temperature.
- Two consequences of the dependence of molecular speeds on mass are:
 - **Effusion** is the escape of gas molecules through a tiny hole into an evacuated space.
 - **Diffusion** is the spread of one substance throughout a space or throughout a second substance.

Graham's Law of Effusion^{38,39}

- The rate of effusion can be quantified.
- Consider two gases with molar masses M_1 and M_2 , with effusion rates, r_1 and r_2 , respectively:
 - The relative rate of effusion is given by **Graham's law**:

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$

- Only those molecules that hit the small hole will escape through it.
- Therefore, the higher the rms speed, the more likely that a gas molecule will hit the hole.

³⁶ Figure 10.18 from Transparency Pack

³⁷ "Diffusion of Gases" from Live Demonstrations

³⁸ "Gas Diffusion and Effusion" Activity from Instructor's Resource CD/DVD

³⁹ "Overhead Projection of Graham's Law of Gaseous Diffusion" from Live Demonstrations

- We can show that:

$$\frac{r_1}{r_2} = \frac{u_1}{u_2} = \sqrt{\frac{3RT/M_1}{3RT/M_2}} = \sqrt{\frac{M_2}{M_1}}$$

Diffusion and Mean Free Path^{40,41}

- Diffusion is faster for light gas molecules.
- Diffusion is significantly slower than the rms speed.
 - Diffusion is slowed by collisions of gas molecules with one another.
 - Consider someone opening a perfume bottle: It takes awhile to detect the odor, but the average speed of the molecules at 25 °C is about 515 m/s (1150 mi/hr).
- The average distance traveled by a gas molecule between collisions is called the **mean free path**.
- At sea level, the mean free path for air molecules is about 6×10^{-6} cm.

FORWARD REFERENCES

- Similar molar mass related issues (e.g., passing of particles of solute through semipermeable membranes) for solutions will be discussed in Chapter 13 (section 13.5).

10.9 Real Gases: Deviations from Ideal Behavior^{42,43}

- From the ideal gas equation:

$$\frac{PV}{RT} = n$$

- For 1 mol of an ideal gas, $PV/RT = 1$ for all pressures.
 - In a real gas, PV/RT varies from 1 significantly.
 - The higher the pressure, the more the deviation from ideal behavior.
- For 1 mol of an ideal gas, $PV/RT = 1$ for all temperatures.
 - As temperature increases, the gases behave more ideally.
- The assumptions in the kinetic-molecular theory show where ideal gas behavior breaks down:
 - The molecules of a gas *have* finite volume.
 - Molecules of a gas *do* attract each other.
- As the pressure on a gas increases, the molecules are forced closer together.
 - As the molecules get closer together, the free space in which the molecules can move gets smaller.
 - The smaller the container, the more of the total space the gas molecules occupy.
 - Therefore, the higher the pressure, the less the gas resembles an ideal gas.
 - As the gas molecules get closer together, the intermolecular distances decrease.
 - The smaller the distance between gas molecules, the more likely that attractive forces will develop between the molecules.
 - Therefore, the less the gas resembles an ideal gas.
- As temperature increases, the gas molecules move faster and further apart.
 - Also, higher temperatures mean more energy available to break intermolecular forces.
 - As temperature increases, the negative departure from ideal-gas behavior disappears.

⁴⁰ “Diffusion of Bromine Vapor” Movie from Instructor’s Resource CD/DVD

⁴¹ “Relative Velocity of Sound Propagation: Musical Molecular Weights” from Live Demonstrations

⁴² Figure 10.22 from Transparency Pack

⁴³ Figure 10.23 from Transparency Pack

The van der Waals Equation⁴⁴

- We add two terms to the ideal gas equation to correct for
 - The volume of molecules: $(V - nb)$
 - For molecular attractions:

$$\left(\frac{n^2 a}{V^2} \right)$$

- The correction terms generate the **van der Waals equation**:

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

- where a and b are empirical constants (van der Waals constants) that differ for each gas.
- van der Waals constants for some common gases can be found in Table 10.3.
- To understand the effect of intermolecular forces on pressure, consider a molecule that is about to strike the wall of the container.
 - The striking molecule is attracted by neighboring molecules.
 - Therefore, the impact on the wall is lessened.

FORWARD REFERENCES

- The name of van der Waals will come up again in Chapter 11 for van der Waals forces.

⁴⁴ “Nonideal Gas Behavior” Activity from Instructor’s Resource CD/DVD

Further Readings:

1. Joseph S. Schmuckler, "Gases and Their Behavior," *J. Chem. Educ.*, Vol. 57, **1980**, 885. A collection of gas law references from past editions of the *Journal of Chemical Education*
2. Robert C. Kerber, "Carbon Dioxide Flooding: A Classroom Case Study Derived from Surgical Practice," *J. Chem. Educ.*, Vol. 80, **2003**, 1437–1438.
3. R. P. Graham, "Gay-Lussac: Chemist Extraordinary," *J. Chem. Educ.*, Vol. 58, **1981**, 789.
4. Harold Goldwhite, "Gay-Lussac after 200 Years," *J. Chem. Educ.*, Vol. 55, **1978**, 366–368.
5. Andreas Madlung, "The Chemistry Behind the Air Bag," *J. Chem. Educ.*, Vol. 73, **1996**, 347–348.
6. William L. Bell, "Chemistry of Air Bags," *J. Chem. Educ.*, Vol. 67, **1990**, 61.
7. Dacio R. Hartwig and Romeu C. Rocha Filho, "Cinema, Flirts, Snakes, and Gases," *J. Chem. Educ.*, Vol. 59, **1982**, 295. The kinetic theory of gases is explored with an analogy in this short reference.
8. Reggie L. Hudson, "Toy Flying Saucers and Molecular Speeds," *J. Chem. Educ.*, Vol. 59, **1982**, 1025–1026. An analogy for molecular speed distributions features a common toy: the Frisbee®.

Live Demonstrations:

1. Lee R. Summerlin, Christie L. Borgford and Julie B. Ealy, "Boiling at Reduced Pressure," *Chemical Demonstrations, A Sourcebook for Teachers, Volume 2* (Washington: American Chemical Society, **1988**), pp 24-25. The volume of a gas-filled balloon is changed by immersion in an ice bath or a warm water bath in this demonstration of Charles's Law.
2. Rick Broniec, "Boyle's Law and the Monster Marshmallow," *J. Chem. Educ.*, Vol. 59, **1982**, 974. A quick demonstration of Boyle's law.
3. Kathryn R. Williams, "Robert Boyle: The Founder of Modern Chemistry," *J. Chem. Educ.*, Vol. 86, **2010**, 148-149.
4. Bassam Z. Shakhshiri, "Effect of Pressure on the Size of a Balloon", *Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 2* (Madison: The University of Wisconsin Press, **1985**), pp. 12–13.
5. Bassam Z. Shakhshiri, "Boyle's Law and the Mass of a Textbook," *Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 2* (Madison: The University of Wisconsin Press, **1985**), pp. 20–23.
6. Bassam Z. Shakhshiri, "Boyle's Law," *Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 2* (Madison: The University of Wisconsin Press, **1985**), pp. 14–19. The relationship between gas pressure and volume at constant temperature is explored.
7. Bassam Z. Shakhshiri, "Thermal Expansion of Gases," *Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 2* (Madison: The University of Wisconsin Press, **1985**), pp. 24–27.

8. John T. Petty, "Charles' Law of Gases: A Simple Experimental Demonstration," *J. Chem. Educ.*, Vol. 72, **1995**, 257. A short demonstration of Charles's Law.
9. Lee R. Summerlin, Christie L. Borgford, and Julie B. Ealy, "Charles' Law: The Relationship Between Volume and Temperature of a Gas," *Chemical Demonstrations, A Sourcebook for Teachers, Volume 2* (Washington: American Chemical Society, **1988**), p. 23. The volume of a gas-filled balloon is changed by immersion in an ice bath or a warm water bath in this demonstration of Charles's Law.
10. Bassam Z. Shakhashiri, "Collapsing Can", *Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 2* (Madison: The University of Wisconsin Press, **1985**), pp. 6–8.
11. Lee R. Summerlin and James L. Ealy, Jr., "Determining the Molecular Weight of a Gas," *Chemical Demonstrations, A Sourcebook for Teachers, Volume 1, 2nd edition* (Washington: American Chemical Society, **1988**), pp. 19–20. The molar mass of butane is determined from its mass and the volume of water it displaces.
12. Lee R. Summerlin and James L. Ealy, Jr., "Diffusion of Gases," *Chemical Demonstrations, A Sourcebook for Teachers, Volume 1, 2nd edition* (Washington: American Chemical Society, **1988**), pp. 14–15. Graham's Law is checked by timing color changes in pH paper caused by HCl(g) or NH₃(g).
13. Dianne N. Epp. "Overhead Projection of Graham's Law of Gaseous Diffusion," *J. Chem. Educ.*, Vol. 67, **1990**, 1061. HCl(g) or NH₃(g) are used in this demonstration of Graham's Law.
14. Bassam Z. Shakhashiri, "Relative Velocity of Sound Propagation: Musical Molecular Weights," *Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 2* (Madison: The University of Wisconsin Press, **1985**), pp. 88–89. The relationship between the pitch of a pipe organ and the molar mass of gas passed through it is explored.