

Unit 9 includes the following OpenStax sections

- 9.1 Gas Pressure
- 9.2 Relating Pressure, Volume, Amount, and Temperature: The Ideal Gas Law
- 9.3 Stoichiometry of Gaseous Substances, Mixtures, and Reactions
- 9.4 Effusion and Diffusion of Gases
- 9.5 The Kinetic-Molecular Theory

Figure 9.1



The hot air inside these balloons is less dense than the surrounding cool air. This results in a buoyant force that causes the balloons to rise when their guy lines are untied. (credit: modification of work by Anthony Quintano)

Learning Objectives

- Gas Pressure
 - Define the property of pressure
 - Define and convert among the units of pressure measurements
 - Describe the operation of common tools for measuring gas pressure
 - Calculate pressure from manometer data

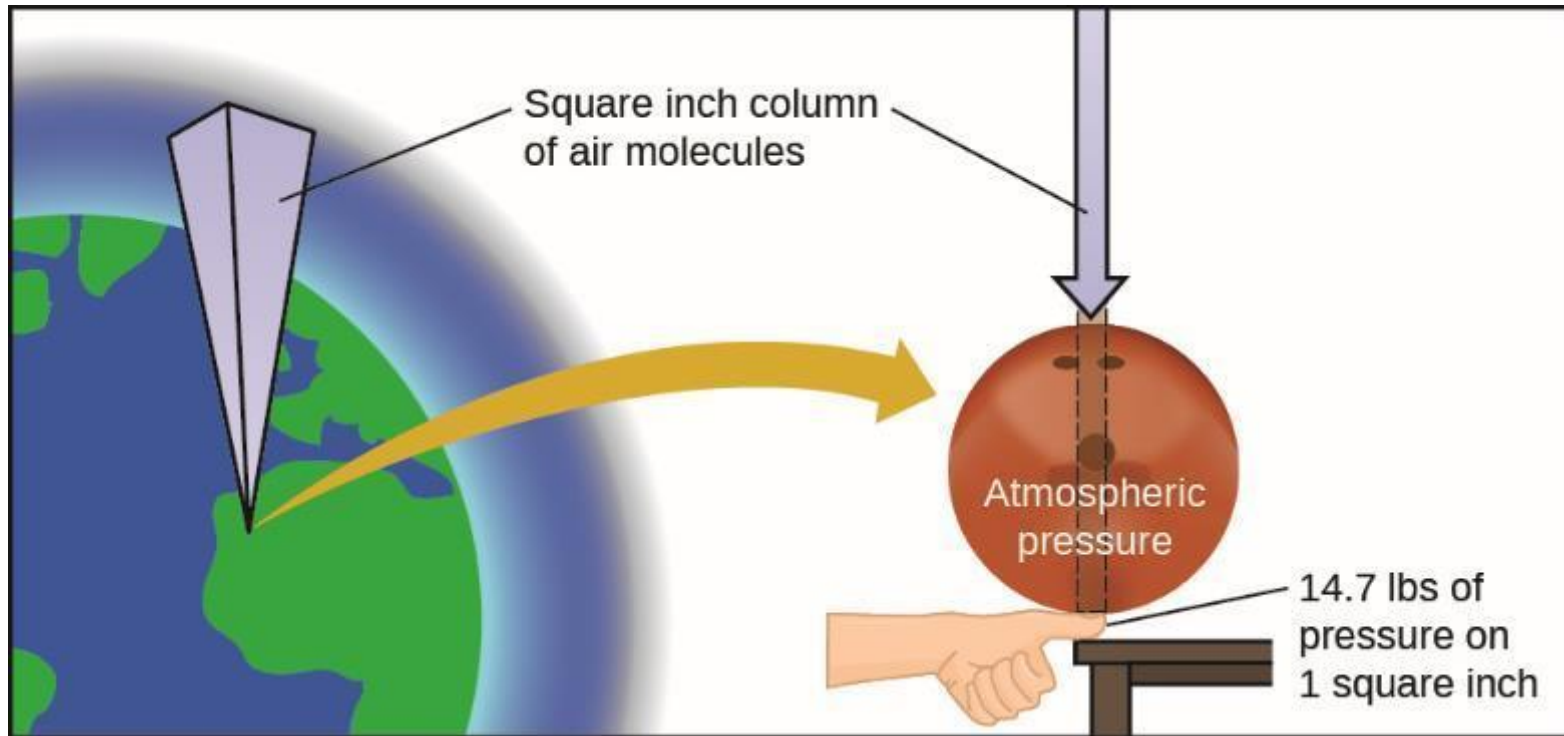
Air—Our Most Familiar Gas

- Air has been known to exist since ancient times.
 - The Greeks considered air one of the four fundamental elements of nature along with earth, water, and fire.
- 18th century—Lavoisier, Cavendish, and Priestley:
 - Determined that air is actually a mixture of gases.

Air and Contemporary Issues

- Air in the news:
 - Ozone depletion in the stratosphere.
 - Carbon dioxide and global warming.

Figure 9.2



The atmosphere above us exerts a large pressure on objects at the surface of the earth, roughly equal to the weight of a bowling ball pressing on an area the size of a human thumbnail.

Gas Pressure

- **Pressure** is defined as the force exerted on a given area.
- Gas pressure is a result of the force exerted by gas molecules colliding with the surface of objects.
- English system units for pressure: **pounds per square inch (psi)**
- A **mercury barometer** is a device used to measure the atmospheric pressure.

Figure 9.3



(a)



(b)

Although (a) an elephant's weight is large, creating a very large force on the ground, (b) the figure skater exerts a much higher pressure on the ice due to the small surface area of her skates. (credit a: modification of work by Guido da Rozze; credit b: modification of work by Ryosuke Yagi)

Units of Pressure

- Due to the way gas pressure is measured, **millimeters of mercury (mm Hg)** is a commonly used unit of pressure.
- Probably the most common unit of gas pressure is the **atmosphere (atm)**.
 - 1 atm is the “normal atmospheric pressure”
 - 1 atm = 760 mm Hg

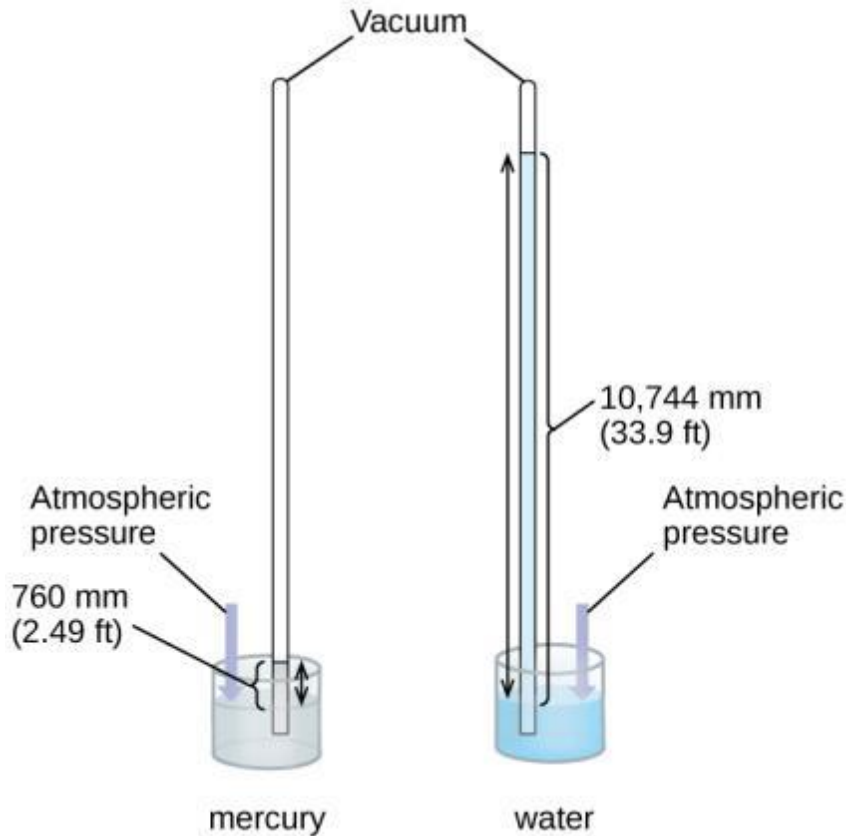
Units of Pressure

- SI unit for pressure is the **Pascal (Pa)**.
 - 1 Pa is defined as the pressure exerted by a 0.1 mm high film of water on the surface beneath it.
 - **Kilopascal (kPa)**
- A related unit of pressure is the **bar**.
 - 1 bar = 105 Pa
- Pressure conversions:
 - 1.013 bar = 1 atm = 760 mm Hg = 760 torr
 - = 14.7 psi = 101.3 kPa

Measuring Atmospheric Pressure

- The barometer measures pressure in terms of the height of a column of liquid mercury.
- The atmosphere exerts a force on the liquid mercury causing it to rise.
- Commonly, the pressure of the atmosphere at sea level pushes the column of mercury ~760 mm high.

Figure 9.4



In a barometer, the height, h , of the column of liquid is used as a measurement of the air pressure. Using very dense liquid mercury (left) permits the construction of reasonably sized barometers, whereas using water (right) would require a barometer more than 30 feet tall.

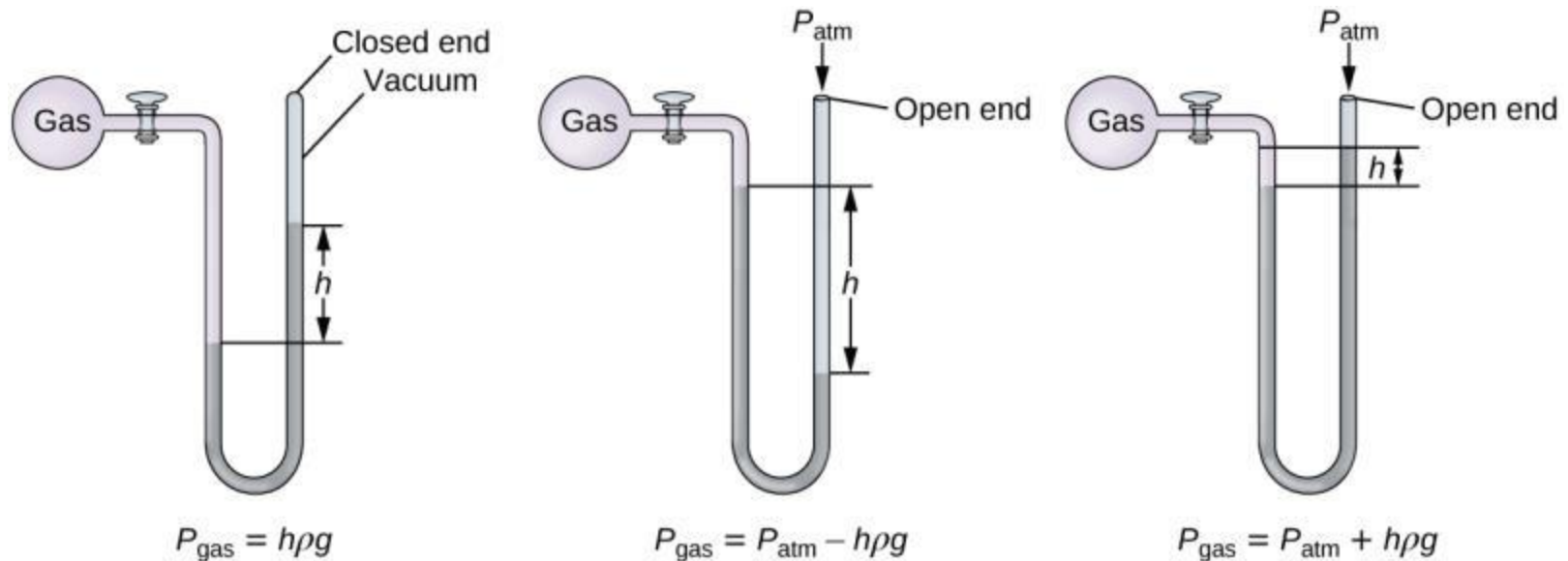
Atmospheric Pressure and Altitude

- Pressure varies with the height of the column of air above.
- We don't notice atmospheric pressure but do notice changes in atmospheric pressure

Measuring the Pressure of a Confined Gas—The Manometer openstax™

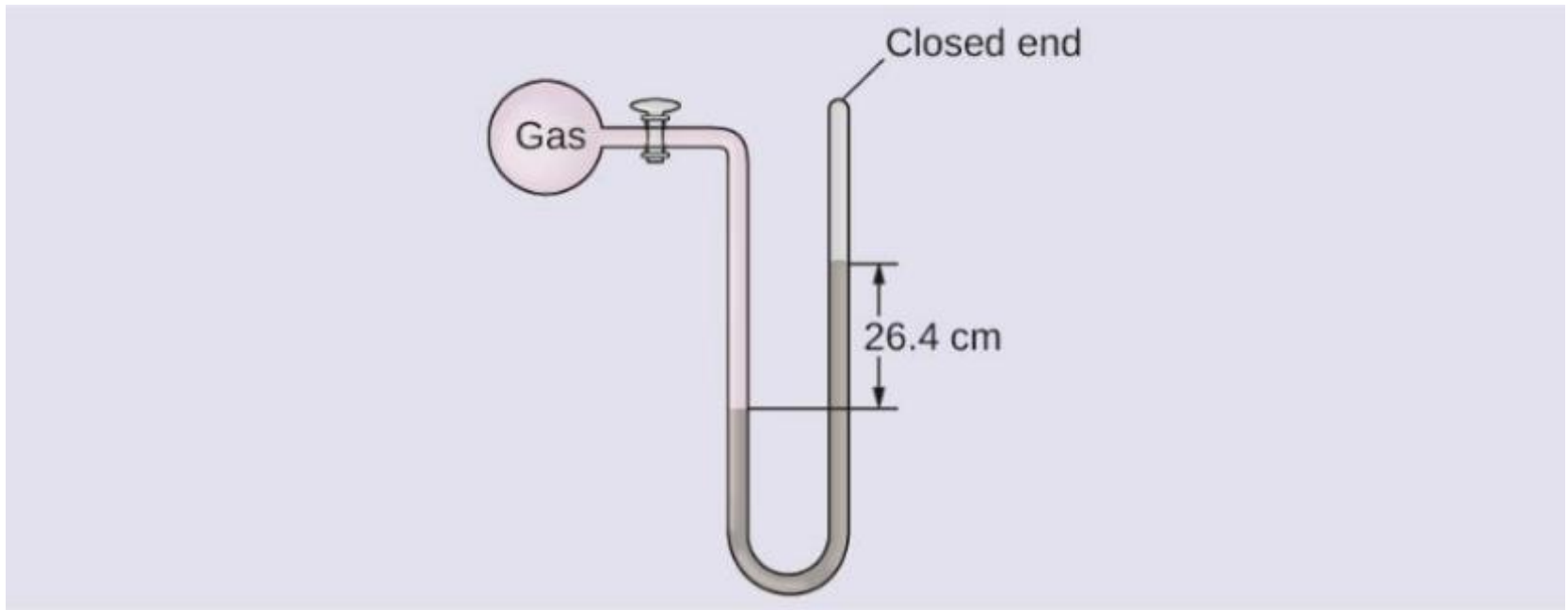
- The force of the gas in the flask is measured relative to the force of the atmosphere (atmospheric pressure).
- The distance and direction that the Hg travels is related to the pressure of the gas.

Figure 9.5

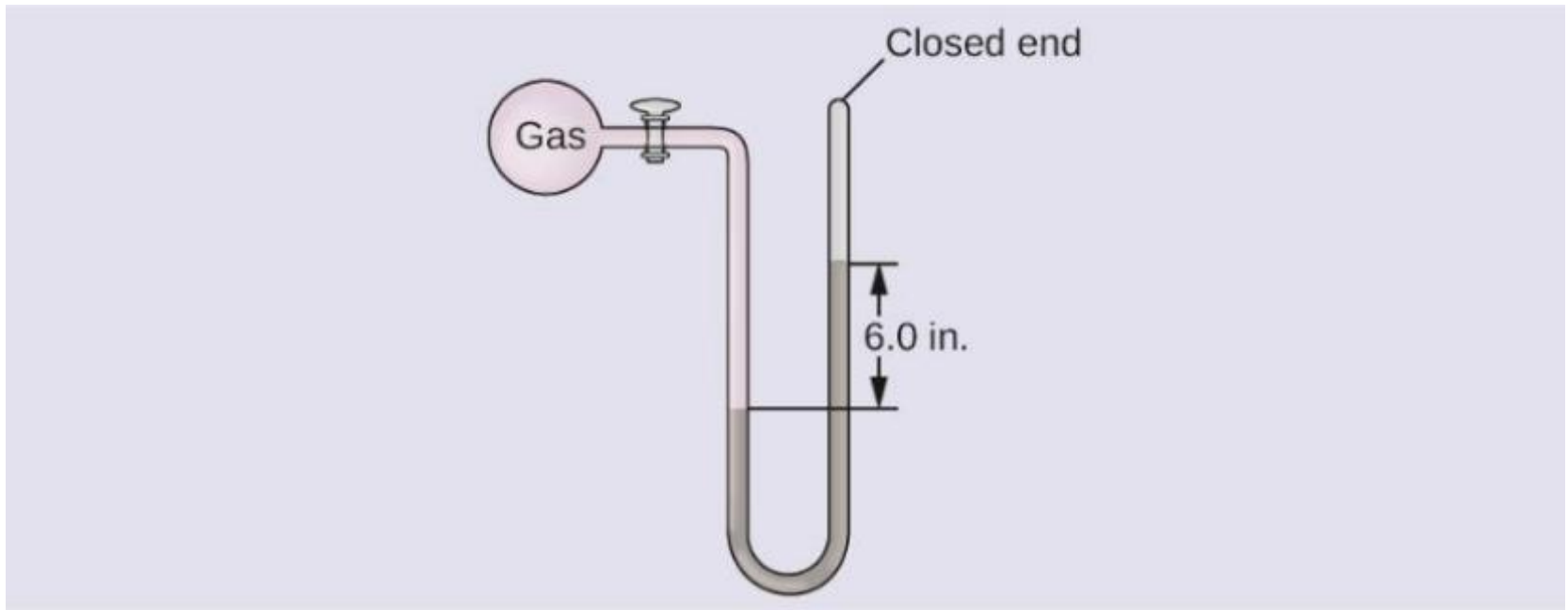


A manometer can be used to measure the pressure of a gas. The (difference in) height between the liquid levels (h) is a measure of the pressure. Mercury is usually used because of its large density.

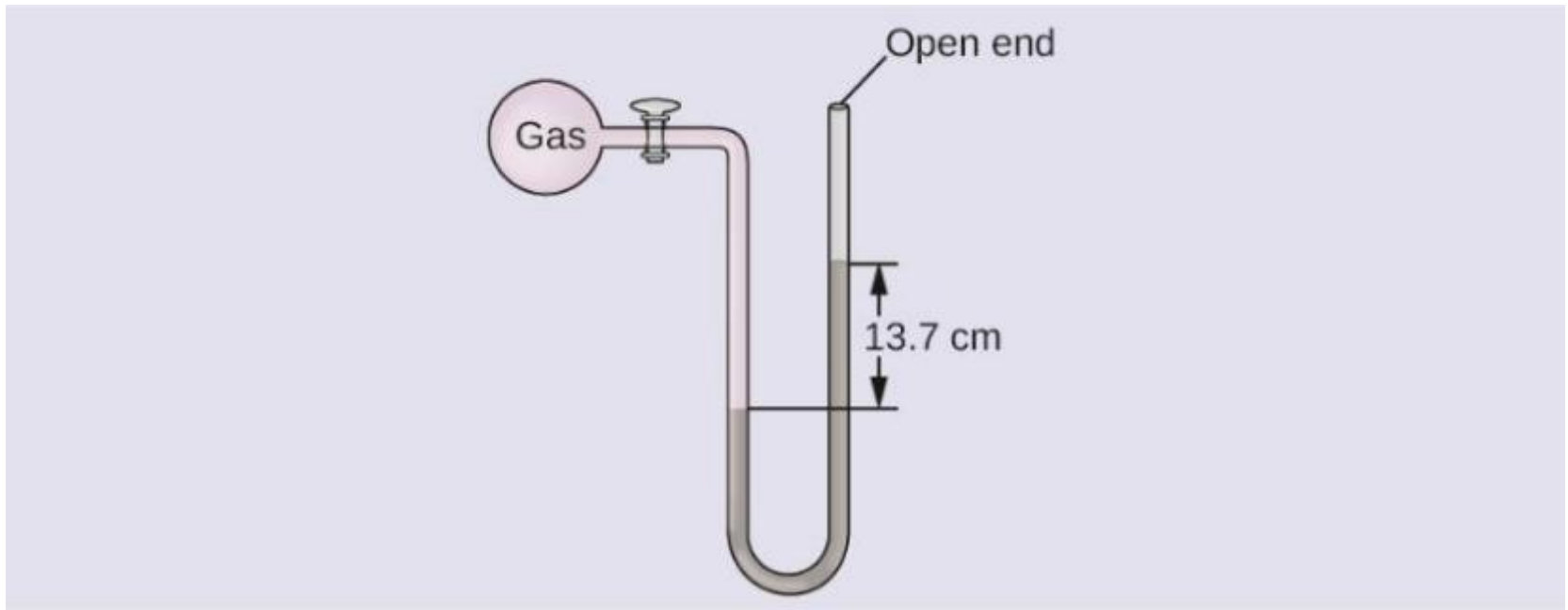
Example 9.3



Example 9.3



Example 9.4



Example 9.4

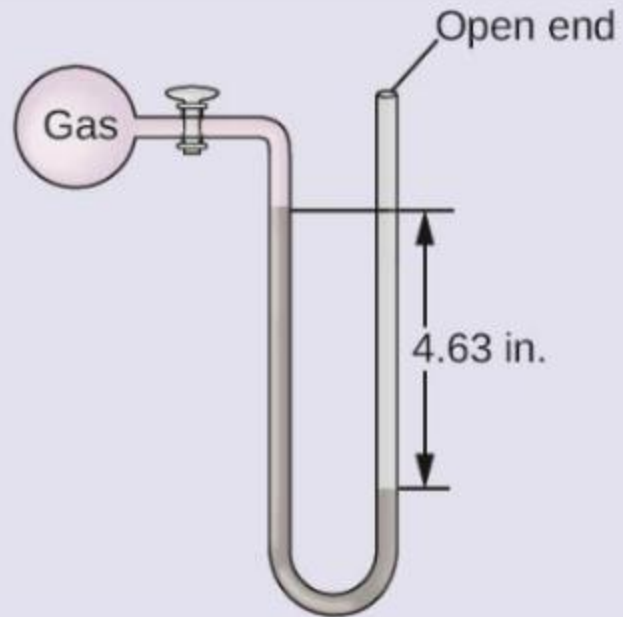


Figure 9.6



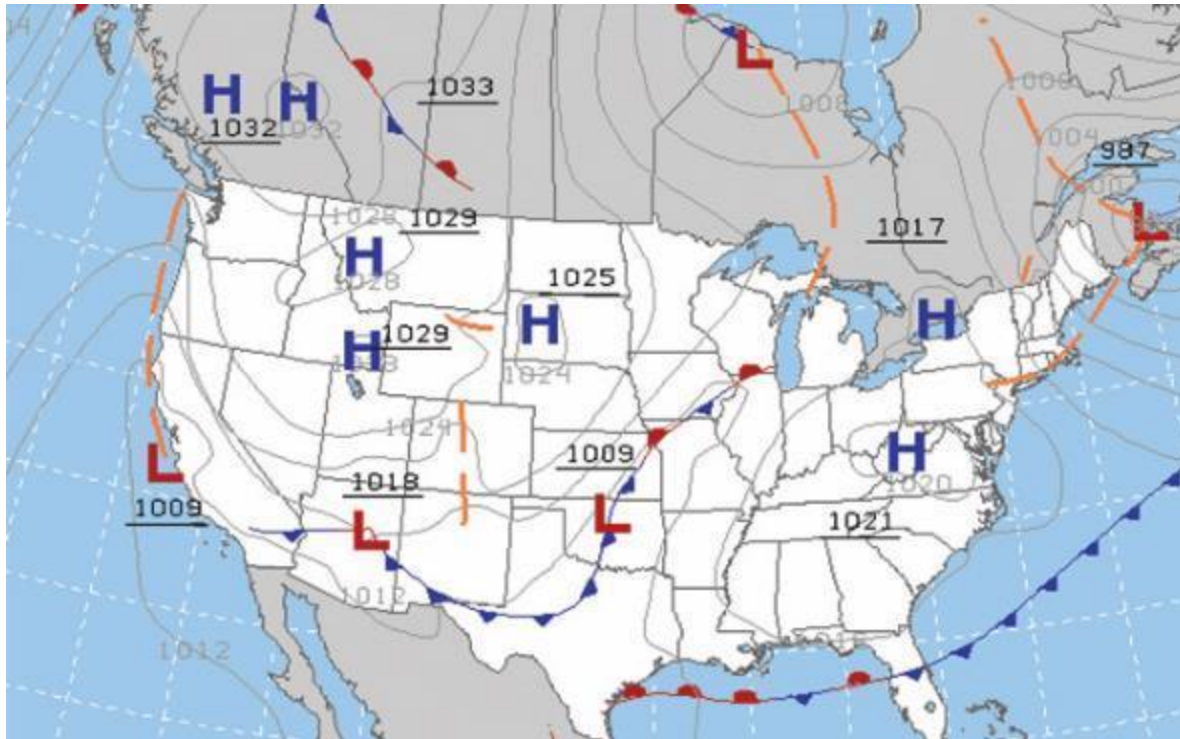
(a)



(b)

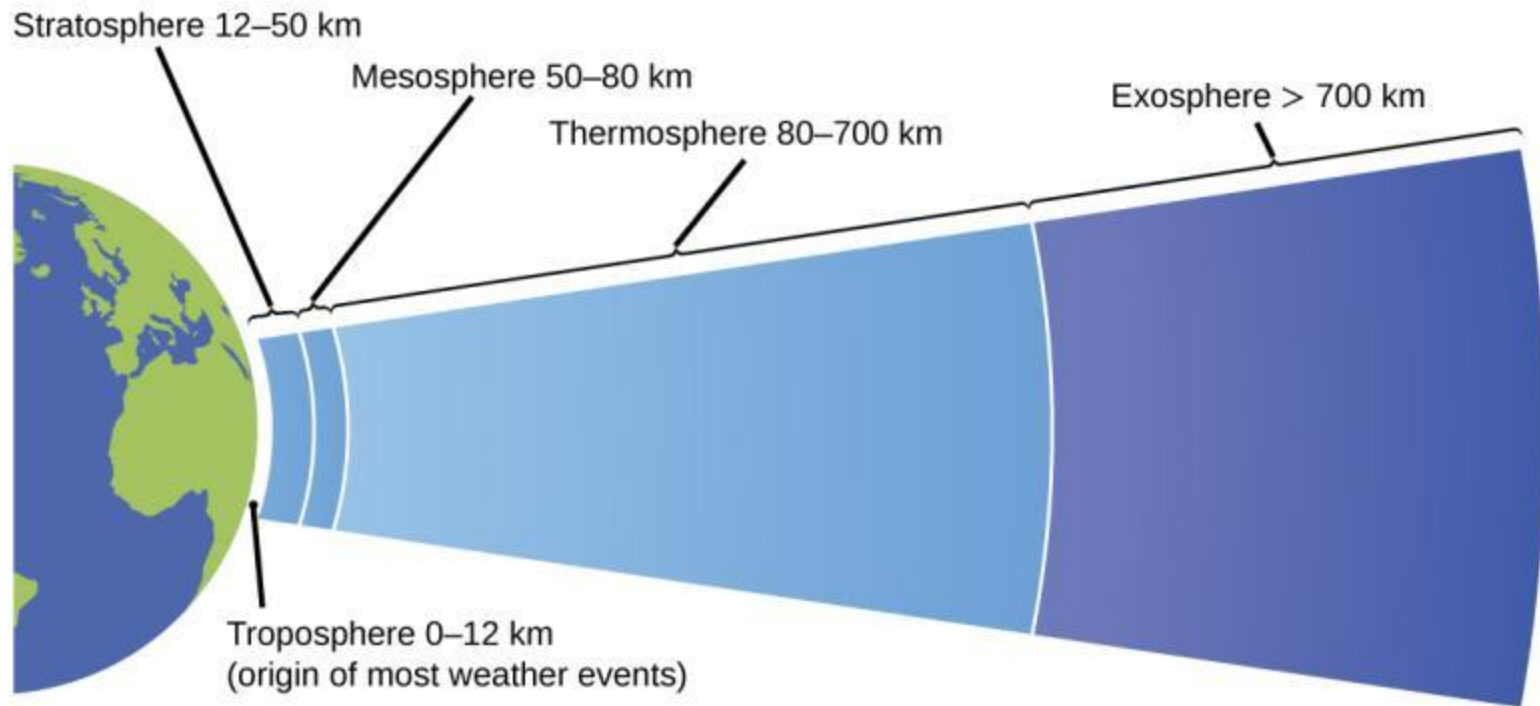
(a) A medical technician prepares to measure a patient's blood pressure with a sphygmomanometer. (b) A typical sphygmomanometer uses a valved rubber bulb to inflate the cuff and a diaphragm gauge to measure pressure. (credit a: modification of work by Master Sgt. Jeffrey Allen)

Figure 9.7



Meteorologists use weather maps to describe and predict weather. Regions of high (H) and low (L) pressure have large effects on weather conditions. The gray lines represent locations of constant pressure known as isobars. (credit: modification of work by National Oceanic and Atmospheric Administration)

Figure 9.8



Earth's atmosphere has five layers: the troposphere, the stratosphere, the mesosphere, the thermosphere, and the exosphere.

Learning Objectives

- 9.2 Relating Pressure, Volume, Amount, Temperature: The Ideal Gas Law
 - Identify the mathematical relationships between the various properties of gases
 - Use the ideal gas law, and related gas laws, to compute the values of various gas properties under specified conditions

Volume of a Gas



- At constant temperature and pressure, a gas expands uniformly to fill the container in which it is contained.

Relating Pressure, Volume, Amount, Temperature: The Ideal Gas Law



- There are many different types of gases.
 - Elements: Ar, He, H₂, N₂, O₂
 - Compounds: CO₂, CO, H₂O, NH₃
- *All gases* have the same dependence on those four properties.
 - Volume
 - Amount
 - Temperature
 - Pressure

Figure 9.9



(a)



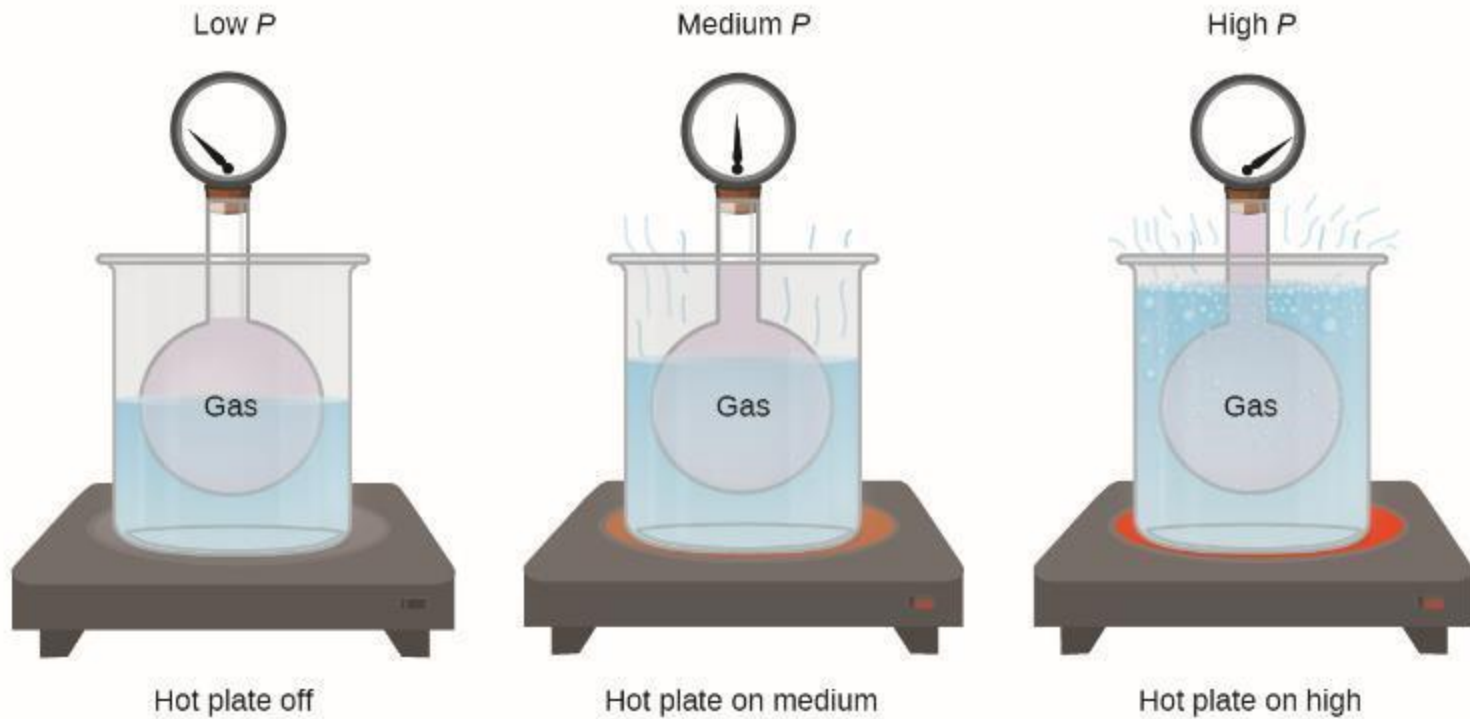
(b)



(c)

In 1783, the first (a) hydrogen-filled balloon flight, (b) manned hot air balloon flight, and (c) manned hydrogen-filled balloon flight occurred. When the hydrogen-filled balloon depicted in (a) landed, the frightened villagers of Gonesse reportedly destroyed it with pitchforks and knives. The launch of the latter was reportedly viewed by 400,000 people in Paris.

Figure 9.10



The effect of temperature on gas pressure: When the hot plate is off, the pressure of the gas in the sphere is relatively low. As the gas is heated, the pressure of the gas in the sphere increases.

Pressure and Temperature

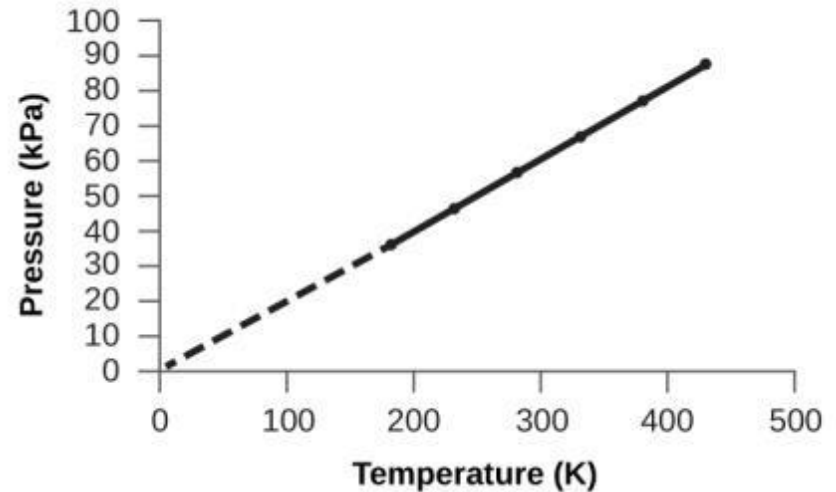
- The pressure of a gas is directly proportional to its temperature in K.
- At constant V and n

$$P = kT$$

- This is Amonton's Law or Gay-Lussac's Law.

Figure 9.11

Temperature (°C)	Temperature (K)	Pressure (kPa)
-100	173	36.0
-50	223	46.4
0	273	56.7
50	323	67.1
100	373	77.5
150	423	88.0



For a constant volume and amount of air, the pressure and temperature are directly proportional, provided the temperature is in kelvin. (Measurements cannot be made at lower temperatures because of the condensation of the gas.) When this line is extrapolated to lower pressures, it reaches a pressure of 0 at $-273\text{ }^{\circ}\text{C}$, which is 0 on the kelvin scale and the lowest possible temperature, called absolute zero.

Volume and Temperature

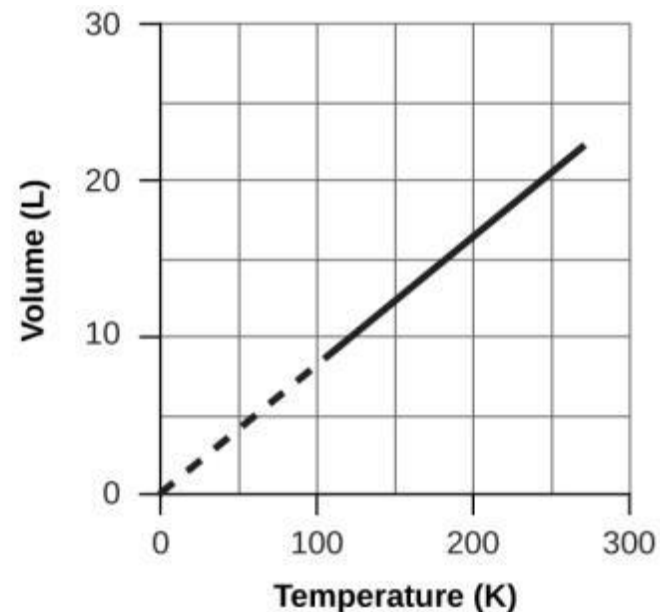
- The volume of a gas is directly proportional to its temperature in K.
- At constant P and n

$$V = kT$$

- This is Charles's Law.

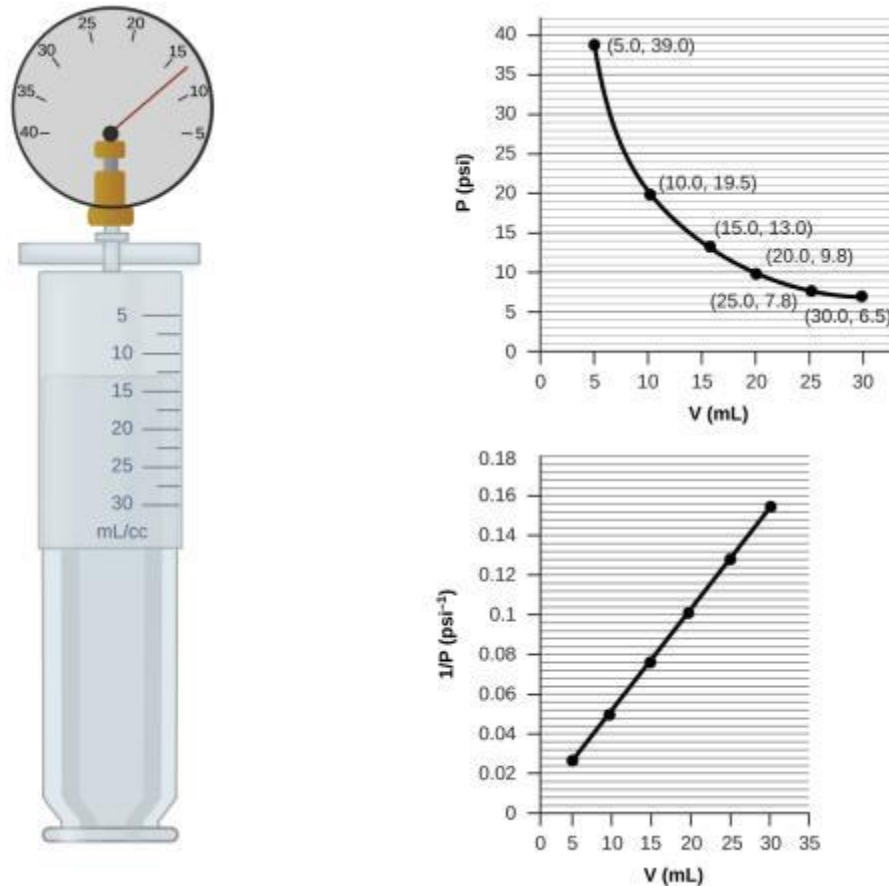
Figure 9.12

Temperature (°C)	Temperature (K)	Volume (L)
-3	270	22
-23	250	21
-53	220	18
-162	111	9



The volume and temperature are linearly related for 1 mole of methane gas at a constant pressure of 1 atm. If the temperature is in kelvin, volume and temperature are directly proportional. The line stops at 111 K because methane liquefies at this temperature; when extrapolated, it intersects the graph's origin, representing a temperature of absolute zero.

Figure 9.13



When a gas occupies a smaller volume, it exerts a higher pressure; when it occupies a larger volume, it exerts a lower pressure (assuming the amount of gas and the temperature do not change). Since P and V are inversely proportional, a graph of $\frac{1}{P}$ vs. V is linear.

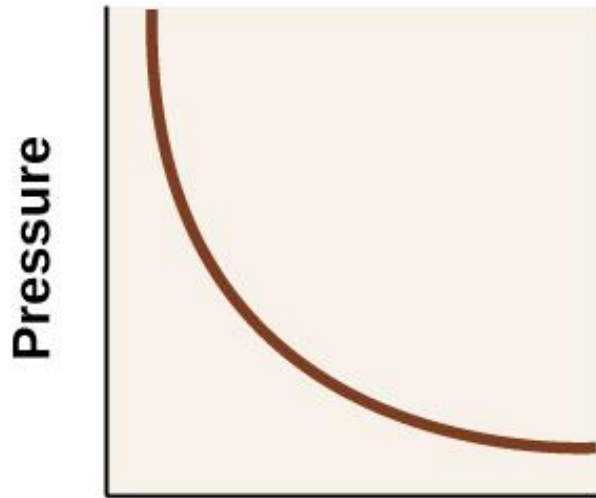
Volume and Pressure

- The volume of a gas is inversely proportional to its pressure.
- At constant T and n

$$V = \frac{k}{P}$$

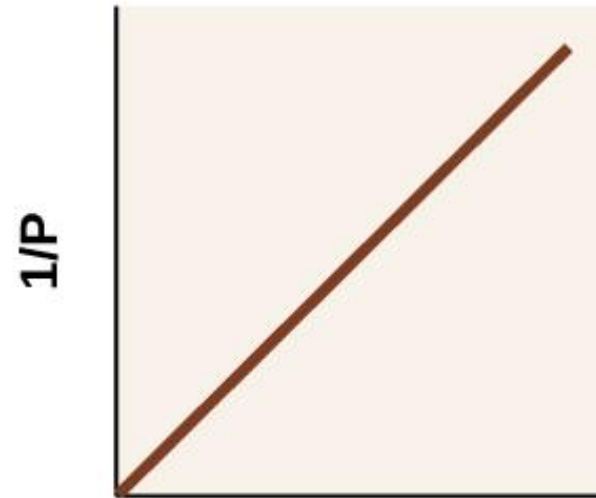
- This is Boyle's Law.

Figure 9.14



Volume

(a)

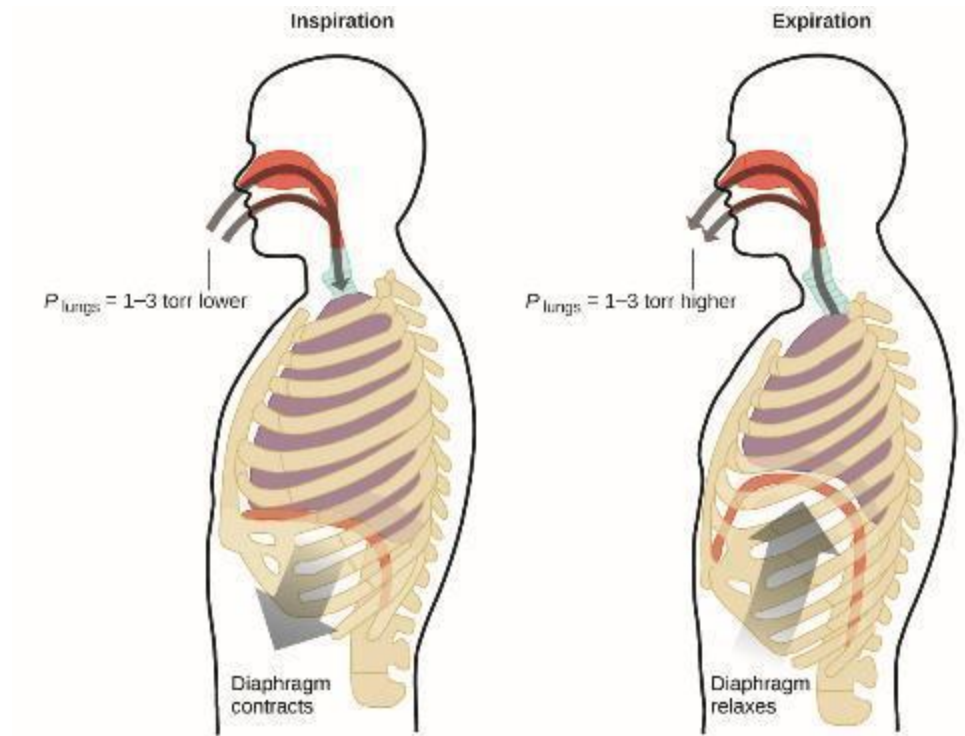


Volume

(b)

The relationship between pressure and volume is inversely proportional. (a) The graph of P vs. V is a parabola, whereas (b) the graph of $\frac{1}{P}$ vs. V is linear.

Figure 9.15



Breathing occurs because expanding and contracting lung volume creates small pressure differences between your lungs and your surroundings, causing air to be drawn into and forced out of your lungs.

Volume and Amount (Moles)

- The volume of a gas is directly proportional to its amount (moles of gas).
- At constant T and P

$$V = kn$$

- This is Avogadro's Law.

The Ideal Gas Law



$$P = kT \quad V = kT \quad V = \frac{k}{P} \quad V = kn$$

- These four equations can be combined into a single law consisting of all four properties.
- **The Ideal Gas Law:** $PV = nRT$
- The four constants (k) were combined into a new constant, R , called the **ideal gas constant**.

The Ideal Gas Law

$$PV = nRT$$

- P is the pressure in atm.
- V is the volume in L.
- n is the moles of gas.
- T is the temperature in K.
- R is the ideal gas constant.

$$R = 0.08206 \frac{\text{L} \times \text{atm}}{\text{mol} \times \text{K}}$$

The Ideal Gas Law



$$PV = nRT$$

- The ideal gas law accurately describes the properties of an ideal gas.
- What is an ideal gas? What assumptions are being made?

Figure 9.16



Scuba divers use compressed air to breathe while underwater. (credit: modification of work by Mark Goodchild)

Figure 9.17



Scuba divers, whether at the Great Barrier Reef or in the Caribbean, must be aware of buoyancy, pressure equalization, and the amount of time they spend underwater, to avoid the risks associated with pressurized gases in the body. (credit: Kyle Taylor)

Standard Temperature and Pressure

- Sometimes it is important to define standard conditions.
- **Standard temperature and pressure (STP)**
 - $0\text{ }^{\circ}\text{C} = 273\text{ K}$
 - 1 atm

Note: The IUPAC recommended value of standard pressure was changed to 1 bar in 1982, but many texts continue to use the 1 atm value.

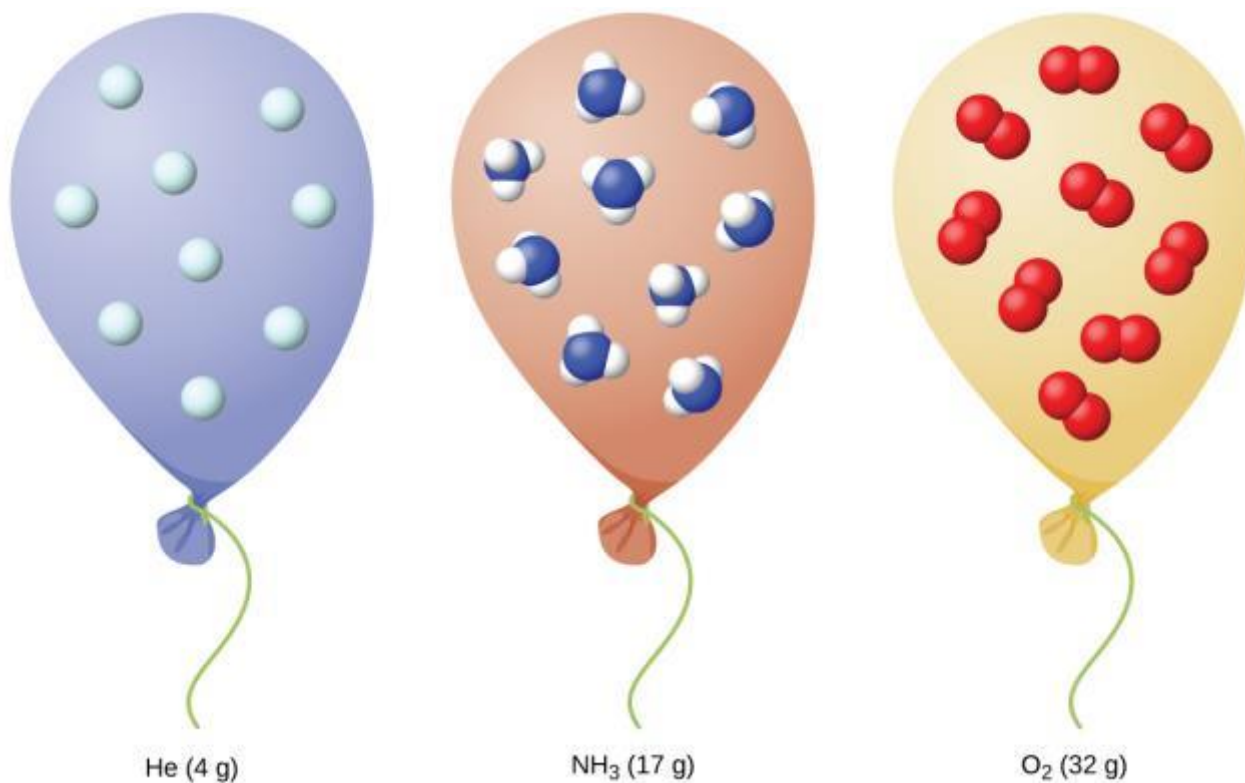
- At STP, the volume of one mole of any gas can be calculated.

Modified Ideal Gas Law

$$PV = nRT$$

- Modified version of the ideal gas law:
 - Moles kept constant
 - But other properties change

Figure 9.18



Regardless of its chemical identity, one mole of gas behaving ideally occupies a volume of ~22.4 L at STP.

Learning Objectives

- 9.3 Stoichiometry of Gaseous Substances, Mixtures, and Reactions
 - Use the ideal gas law to compute gas densities and molar masses
 - Perform stoichiometric calculations involving gaseous substances
 - State Dalton's law of partial pressures and use it in calculations involving gaseous mixtures

Calculating the Density of a Gas

- Recall, the density of a substance

$$d = \frac{\text{mass}}{V}$$

- We can rearrange the following equation

$$PV = \frac{\text{grams}}{M} RT$$

- Substitute into the density equation to obtain

$$d = M \frac{P}{RT}$$

Density of Gases

$$d = M \frac{P}{RT}$$

- Recall that density is an intensive property.
 - Does not depend on the amount of substance
- But, the density of a gas does depend on
 - Pressure
 - Temperature
 - Molar mass

Calculating the Molar Mass of a Gas

- Recall, the molar mass (M) of a substance.

$$M = \frac{\text{grams}}{\text{mole}} = \frac{m}{n}$$

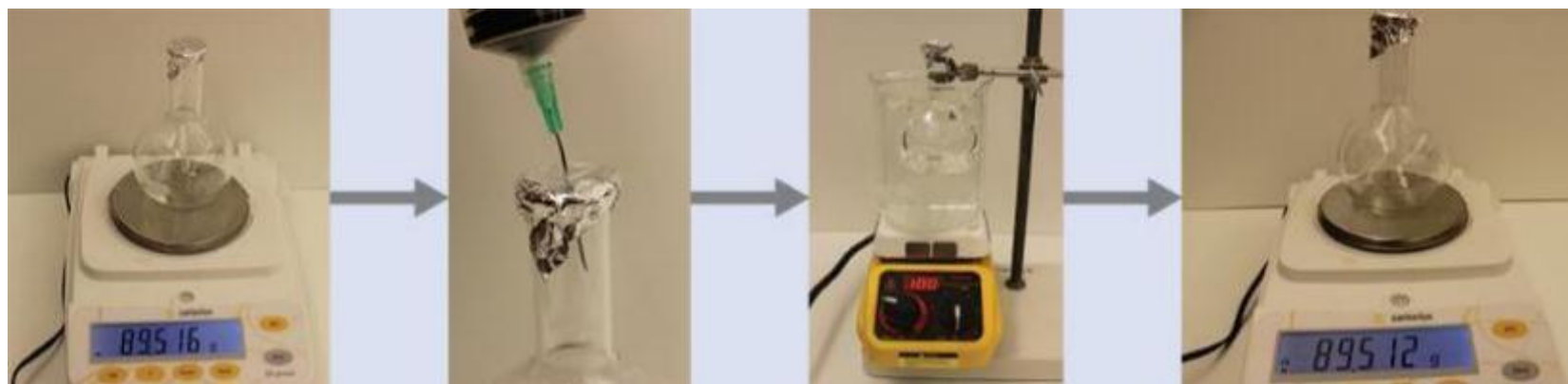
- This expression can be rearranged.

$$n = \frac{m}{M}$$

- And then substituted into the Ideal Gas Law.

$$PV = \frac{m}{M} RT$$

Figure 9.19



When the volatile liquid in the flask is heated past its boiling point, it becomes gas and drives air out of the flask. At $t_{l \rightarrow g}$, the flask is filled with volatile liquid gas at the same pressure as the atmosphere. If the flask is then cooled to room temperature, the gas condenses and the mass of the gas that filled the flask, and is now liquid, can be measured. (credit: modification of work by Mark Ott)

Dalton's Law

- The ideal gas law also applies to mixtures of gases if the gases don't react with each other.
- **Dalton's law of partial pressure:** The total pressure of a mixture of ideal gases is equal to the sum of the partial pressures of the component gases.

$$P_T = P_A + P_B + P_C \dots$$

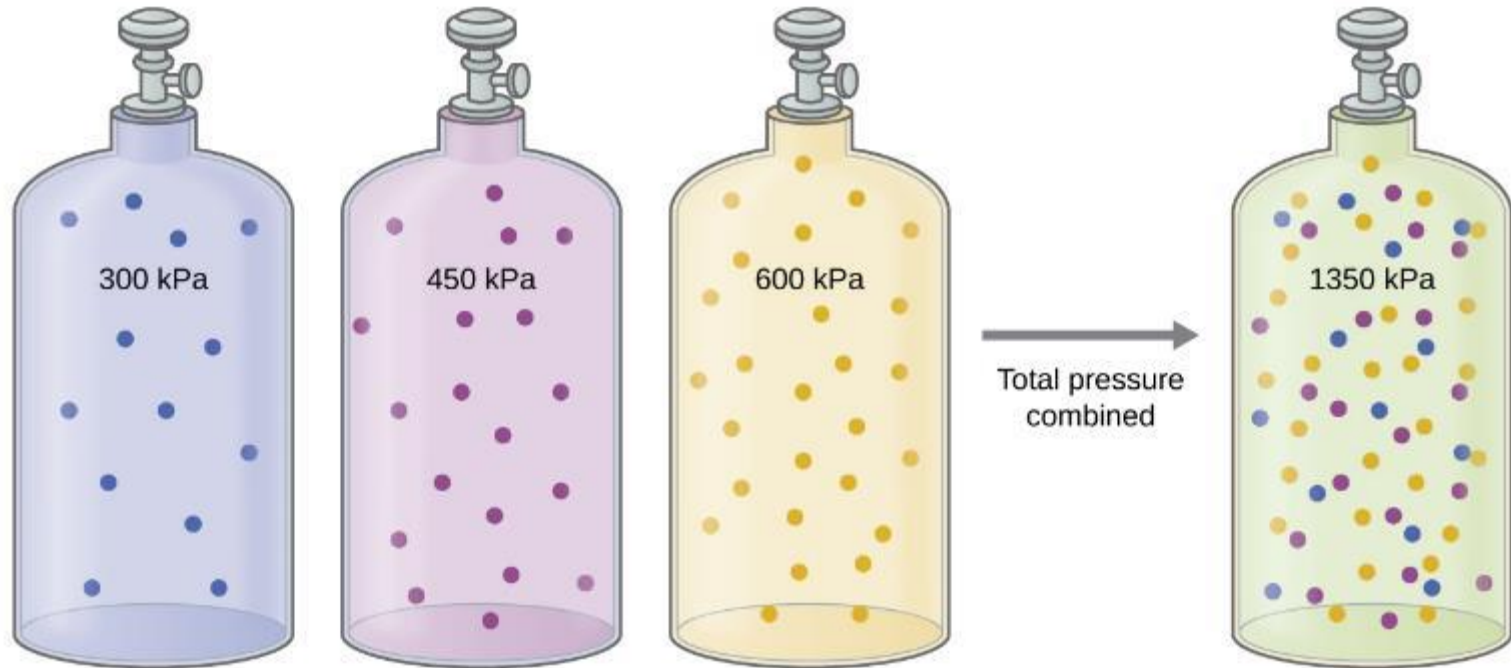
Dalton's Law of Partial Pressure

$$P_T = P_A + P_B + P_C \dots$$

- For a mixture of gases A, B, C, ...
- P_T is the total pressure of the gas mixture
- P_A is the partial pressure of gas A
- P_B is the partial pressure of gas B
- P_C is the partial pressure of gas C

- The pressure exerted by each individual gas in a mixture is called its **partial pressure**.
 - The partial pressure is equal to the pressure that the gas would exert if it were by itself.

Figure 9.20



If equal-volume cylinders containing gasses at pressures of 300 kPa, 450 kPa, and 600 kPa are all combined in the same-size cylinder, the total pressure of the gas mixture is 1350 kPa.

Partial Pressure and Mole Fraction

- We can express the partial pressure of a gas in a mixture in terms of its mole fraction.
- Consider a mixture of gases A and B.

$$P_A = X_A P_T \qquad P_B = X_B P_T$$

- The mole fraction of gas A (X_A) is the number of moles of A divided by the total number of moles of gas in the mixture.

Collecting a Gas Over Water

- A common way to collect and quantify a gas produced in a chemical reaction is to capture the gas in an inverted bottle that had been filled with water.
- The volume of water displaced can then be related back to the amount of gas produced.
- As the gas travels through the water, it picks up water vapor, $\text{H}_2\text{O} (g)$.
- The collected gas mixture has a P_T equal to the atmospheric pressure.

Figure 9.21



When a reaction produces a gas that is collected above water, the trapped gas is a mixture of the gas produced by the reaction and water vapor. If the collection flask is appropriately positioned to equalize the water levels both within and outside the flask, the pressure of the trapped gas mixture will equal the atmospheric pressure outside the flask (see the earlier discussion of manometers).

Wet Gases



- We refer to the gas collected over water as a wet gas mixture.
- Say, for example, that the gas being collected was H_2 .

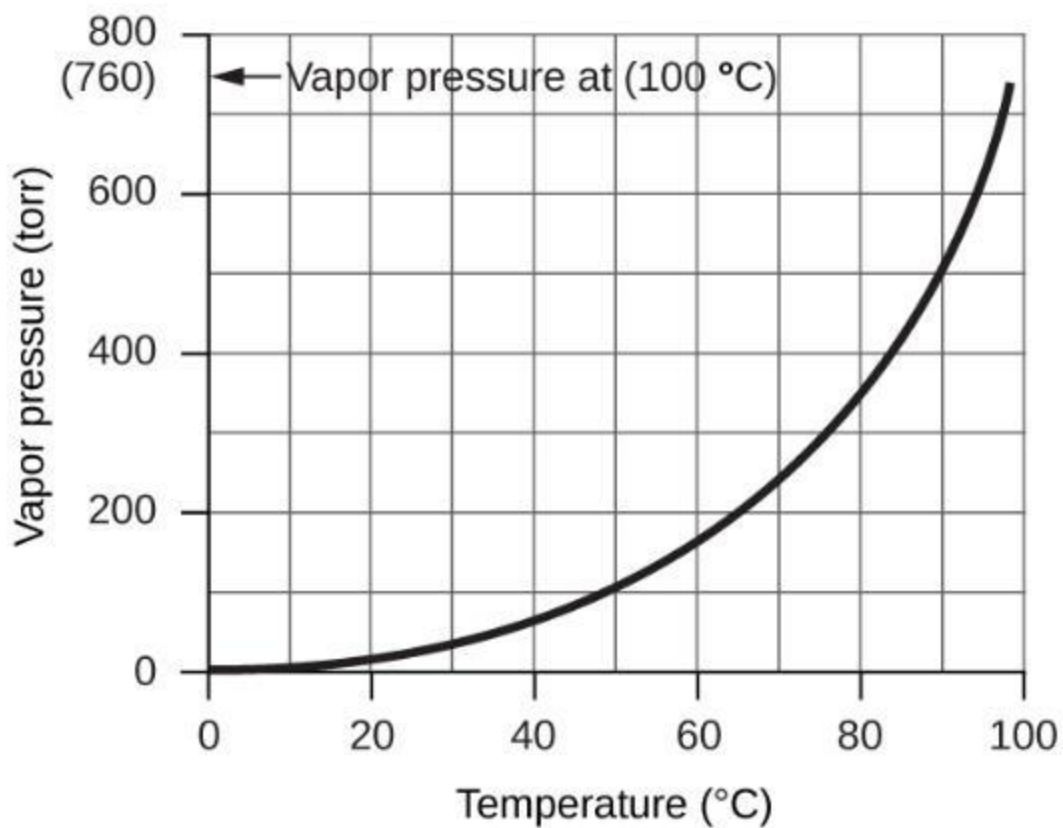
$$P_T = P_{H_2O} + P_{H_2}$$

- P_T is the atmospheric pressure.
- P_{H_2} is the partial pressure of the H_2 (g) collected.
- P_{H_2O} is the **vapor pressure** of water.

Vapor Pressure

- **Vapor pressure of water:** The pressure exerted by water vapor in equilibrium with liquid water in a closed container.
 - Intensive property—does not depend on the amount of water
 - Temperature dependent

Figure 9.22



This graph shows the vapor pressure of water at sea level as a function of temperature.

Stoichiometry in Gaseous Reactions



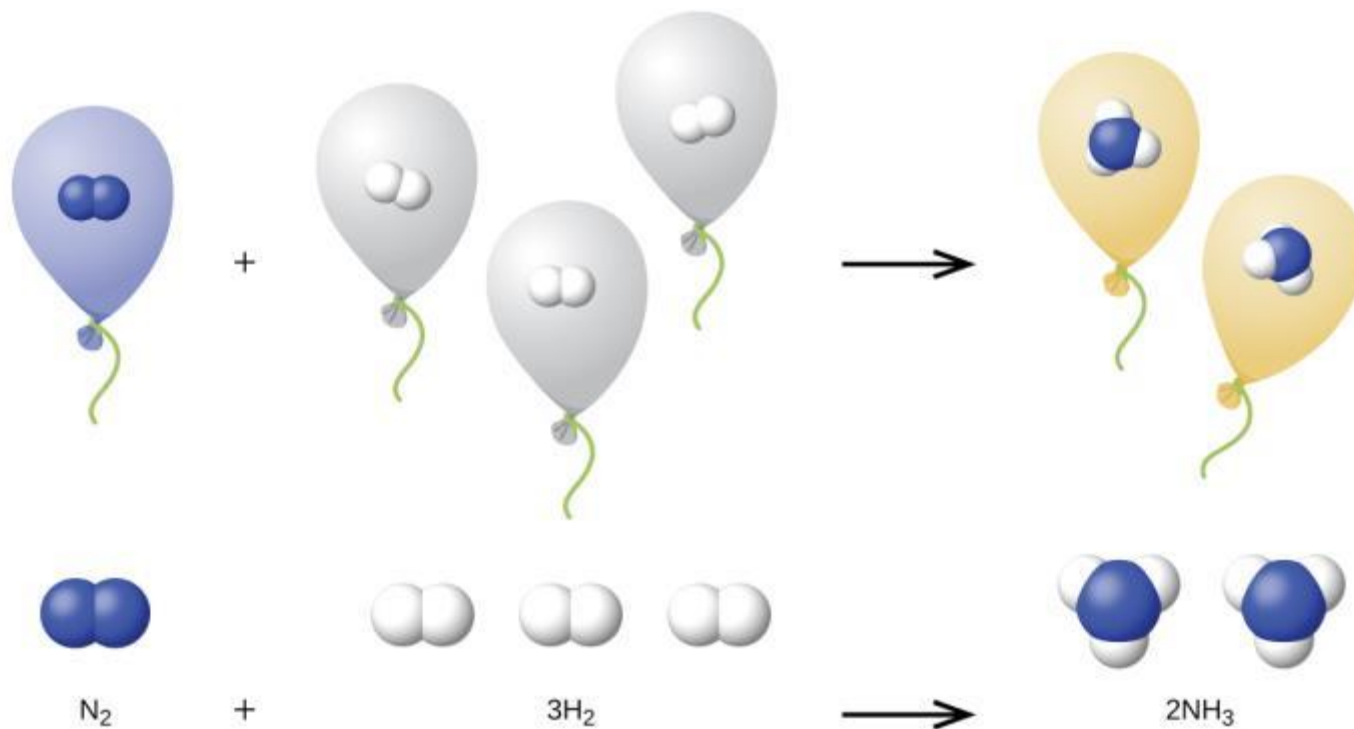
- Gases may appear as reactants or products in a chemical reaction.
- The ideal gas law and a balanced chemical equation can be used to relate the P , V , T , n , and grams of a gas that takes part in a chemical reaction.
- **Stoichiometric factors** can be used to relate the moles of one substance to the moles of any other substance in a *balanced* chemical equation.

Stoichiometry with Gas Volumes



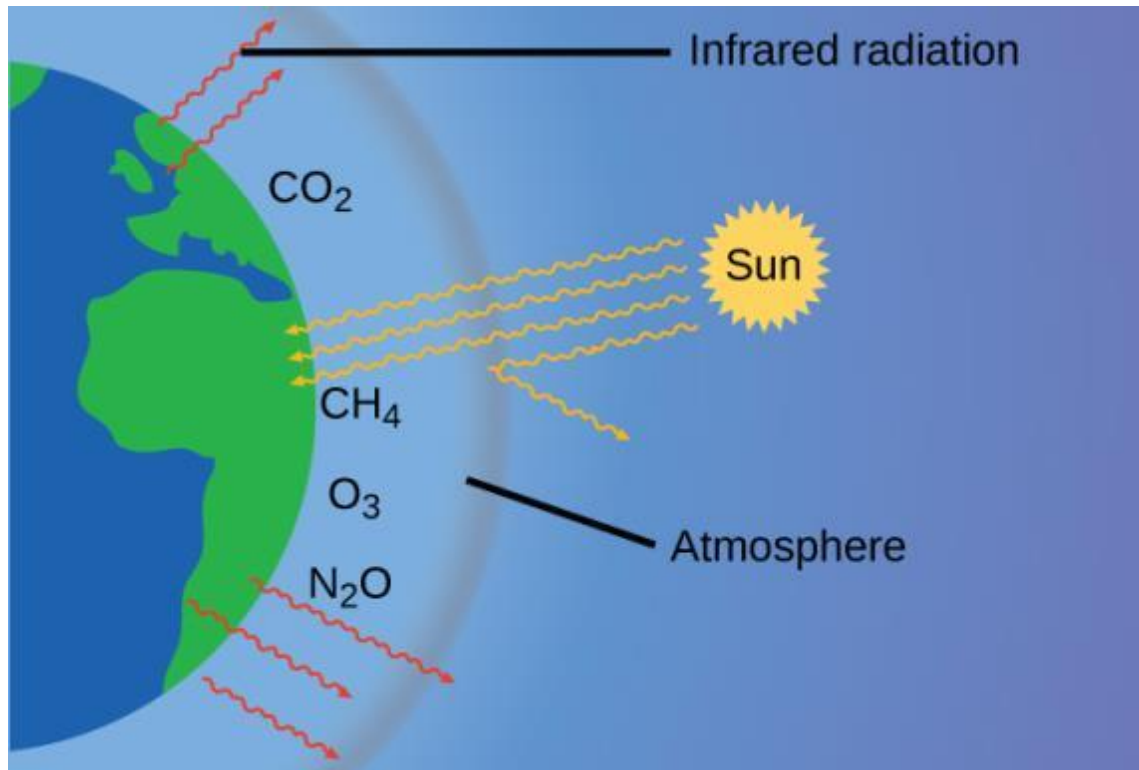
- The **volume ratio** of any two gases in a reaction at constant temperature and pressure is the same as the stoichiometric ratio (mole ratio) in the balanced chemical equation.
- Recall that $V = kn$ (constant T, P). The volume of a gas is directly proportional to the moles of gas.

Figure 9.23



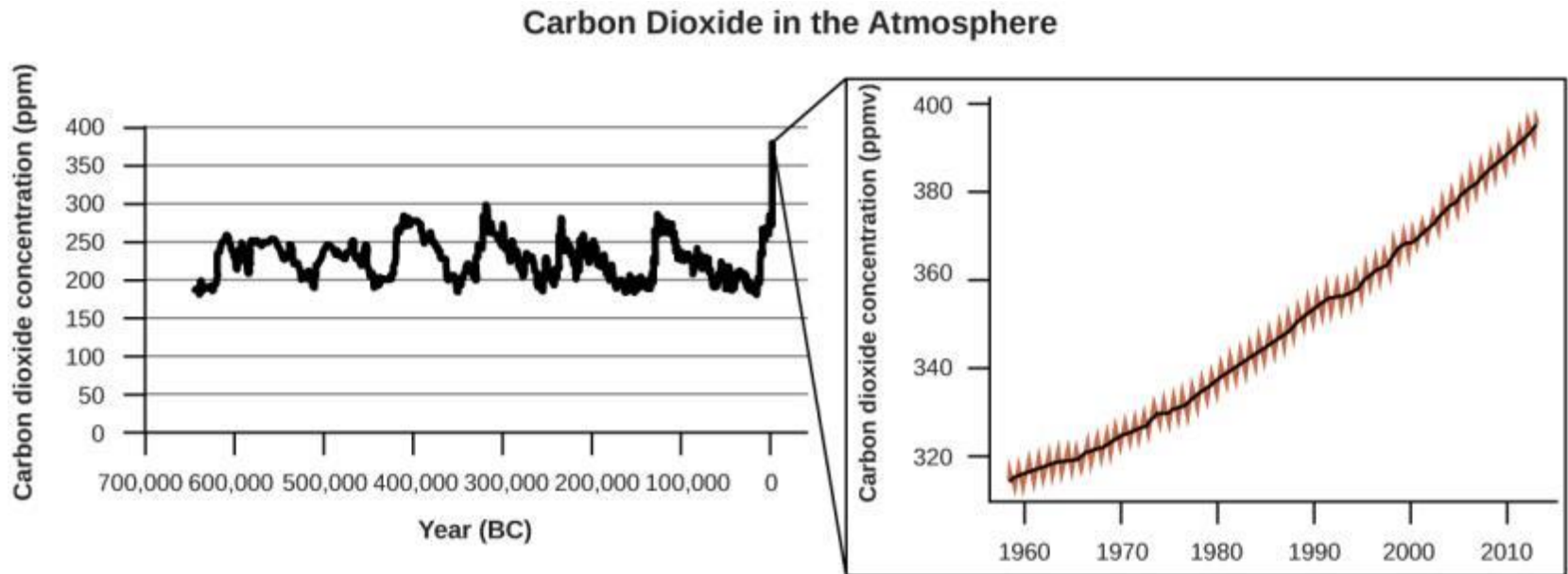
One volume of N_2 combines with three volumes of H_2 to form two volumes of NH_3 .

Figure 9.24



Greenhouse gases trap enough of the sun's energy to make the planet habitable—this is known as the greenhouse effect. Human activities are increasing greenhouse gas levels, warming the planet and causing more extreme weather events.

Figure 9.25



CO₂ levels over the past 700,000 years were typically from 200–300 ppm, with a steep, unprecedented increase over the past 50 years.

Figure 9.26

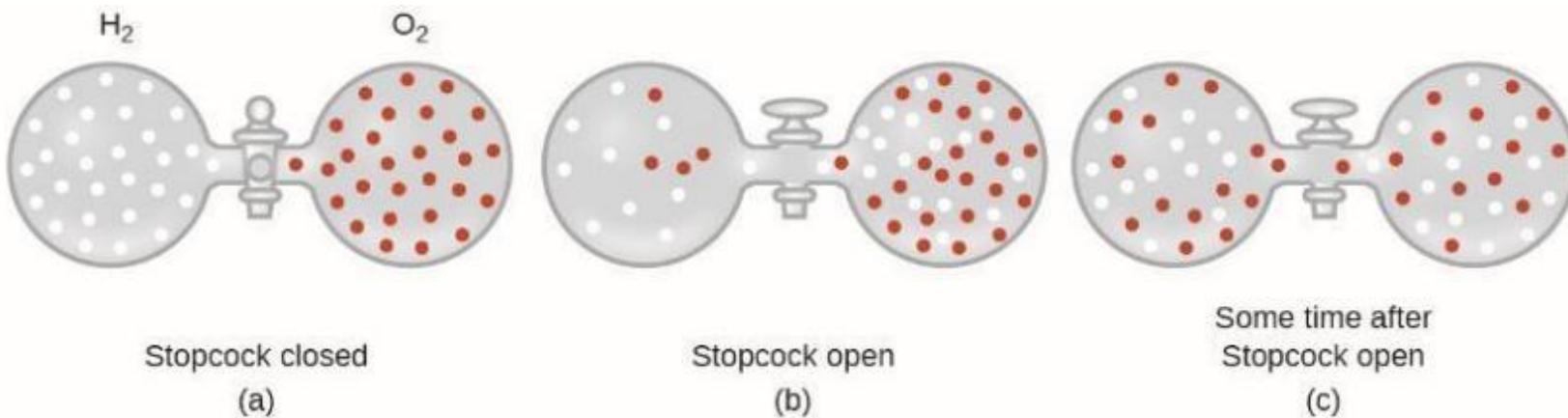


Susan Solomon's research focuses on climate change and has been instrumental in determining the cause of the ozone hole over Antarctica. (credit: National Oceanic and Atmospheric Administration)

Learning Objectives

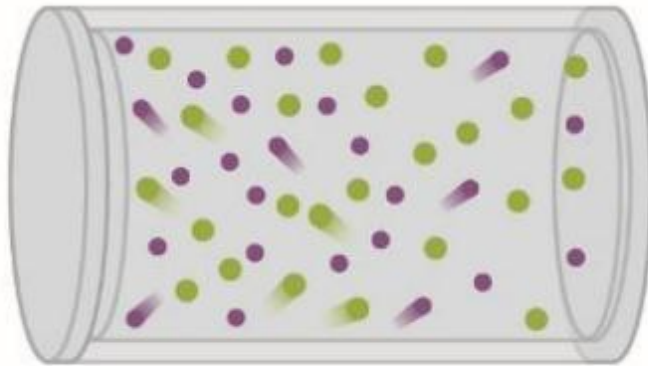
- 9.4 Effusion and Diffusion of Gases
 - Define and explain effusion and diffusion
 - State Graham's law and use it to compute relevant gas properties

Figure 9.27

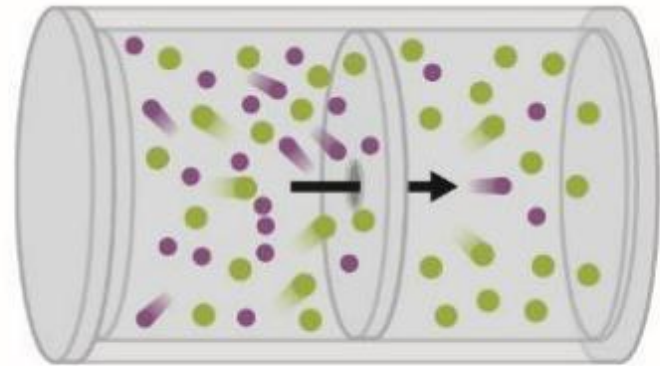


(a) Two gases, H₂ and O₂, are initially separated. (b) When the stopcock is opened, they mix together. The lighter gas, H₂, passes through the opening faster than O₂, so just after the stopcock is opened, more H₂ molecules move to the O₂ side than O₂ molecules move to the H₂ side. (c) After a short time, both the slower-moving O₂ molecules and the faster-moving H₂ molecules have distributed themselves evenly on both sides of the vessel.

Figure 9.28



Diffusion



Effusion

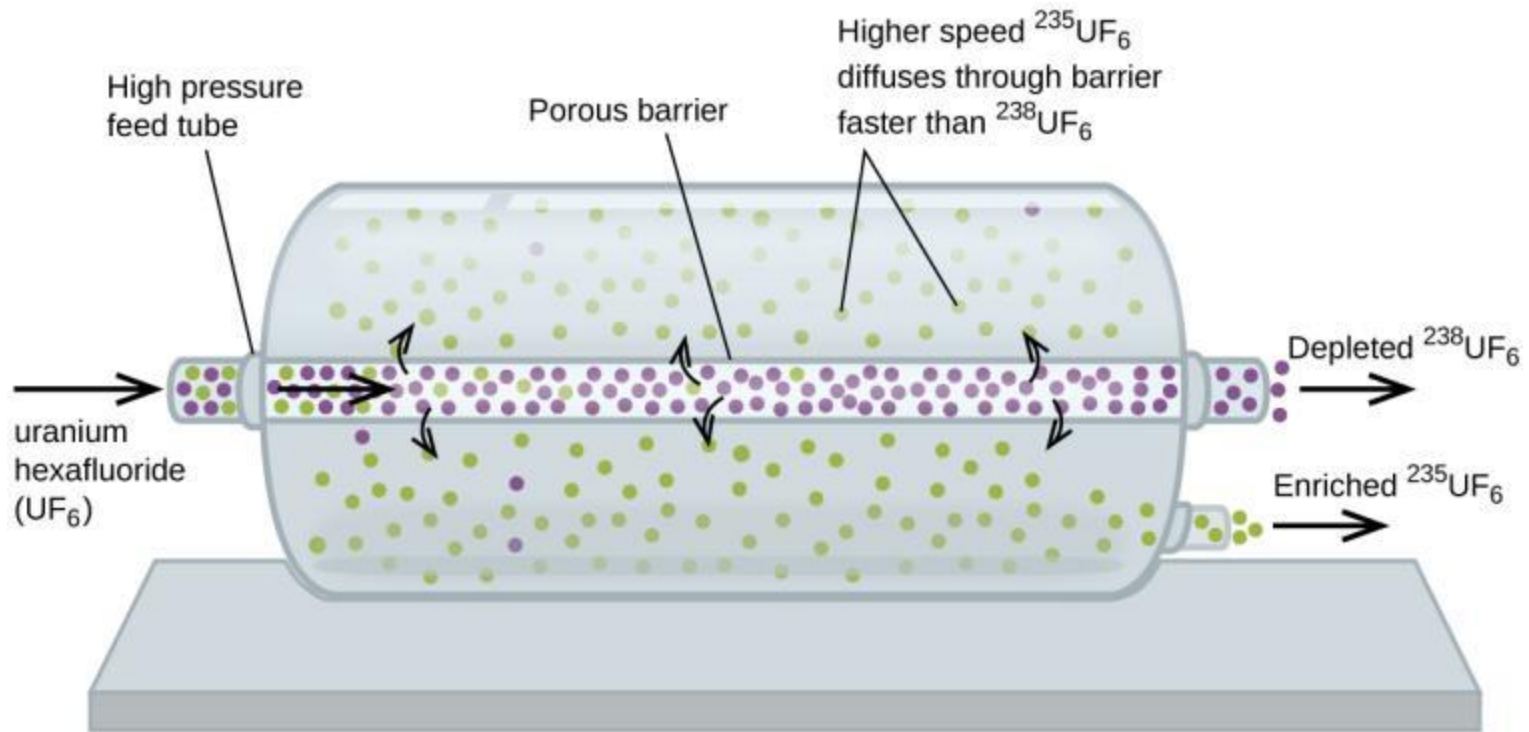
Diffusion occurs when gas molecules disperse throughout a container. Effusion occurs when a gas passes through an opening that is smaller than the mean free path of the particles, that is, the average distance traveled between collisions. Effectively, this means that only one particle passes through at a time.

Figure 9.29



The left photograph shows two balloons inflated with different gases, helium (orange) and argon (blue). The right-side photograph shows the balloons approximately 12 hours after being filled, at which time the helium balloon has become noticeably more deflated than the argon balloon, due to the greater effusion rate of the lighter helium gas. (credit: modification of work by Paul Flowers)

Figure 9.30

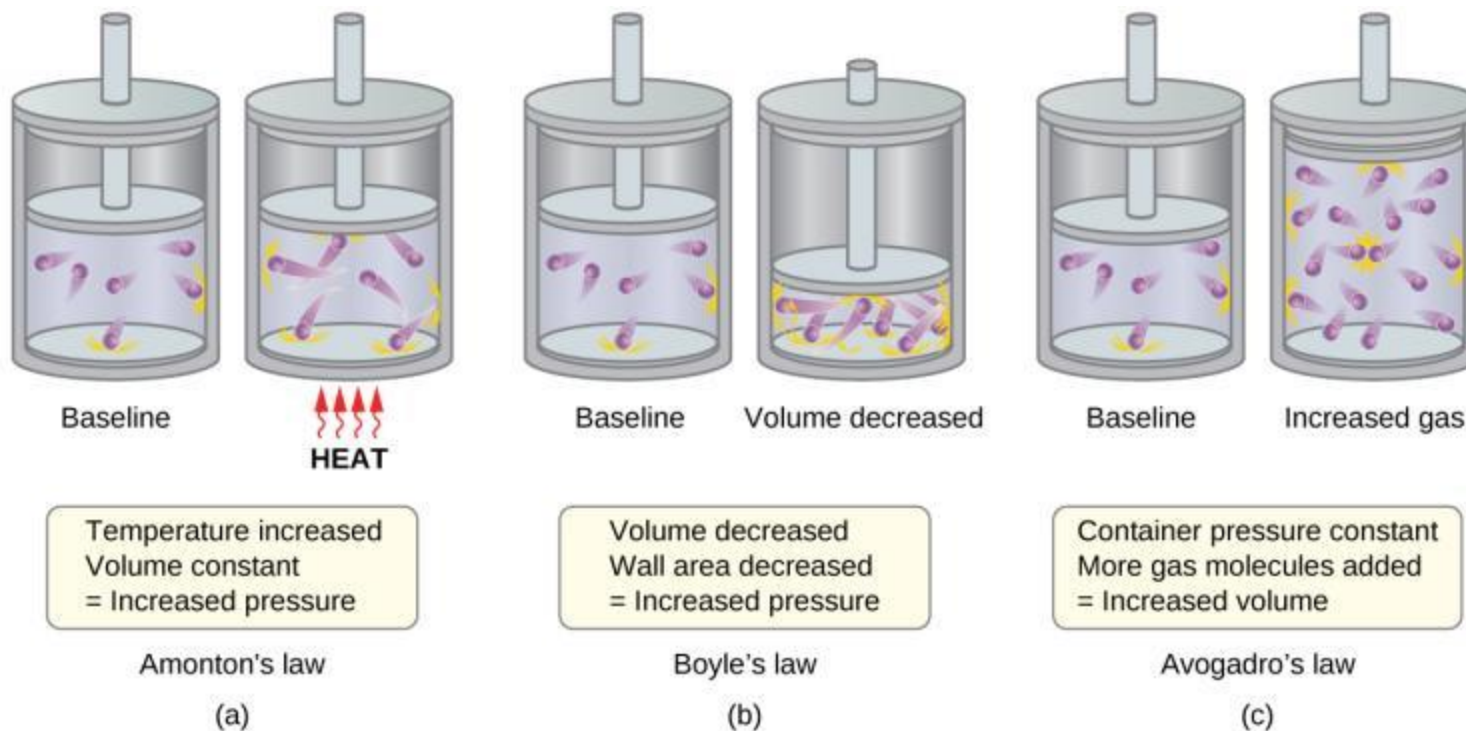


In a diffuser, gaseous UF_6 is pumped through a porous barrier, which partially separates $^{235}\text{UF}_6$ from $^{238}\text{UF}_6$. The UF_6 must pass through many large diffuser units to achieve sufficient enrichment in ^{235}U .

Learning Objectives

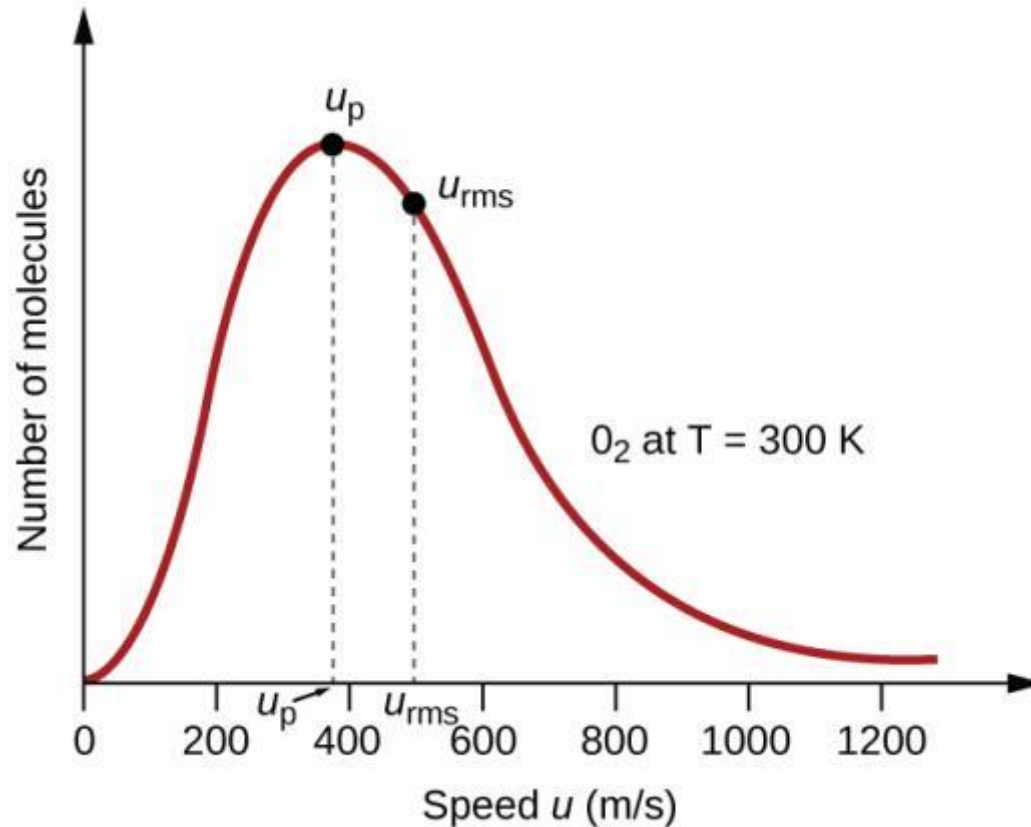
- 9.5 The Kinetic-Molecular Theory
 - State the postulates of the kinetic-molecular theory
 - Use this theory's postulates to explain the gas laws

Figure 9.31



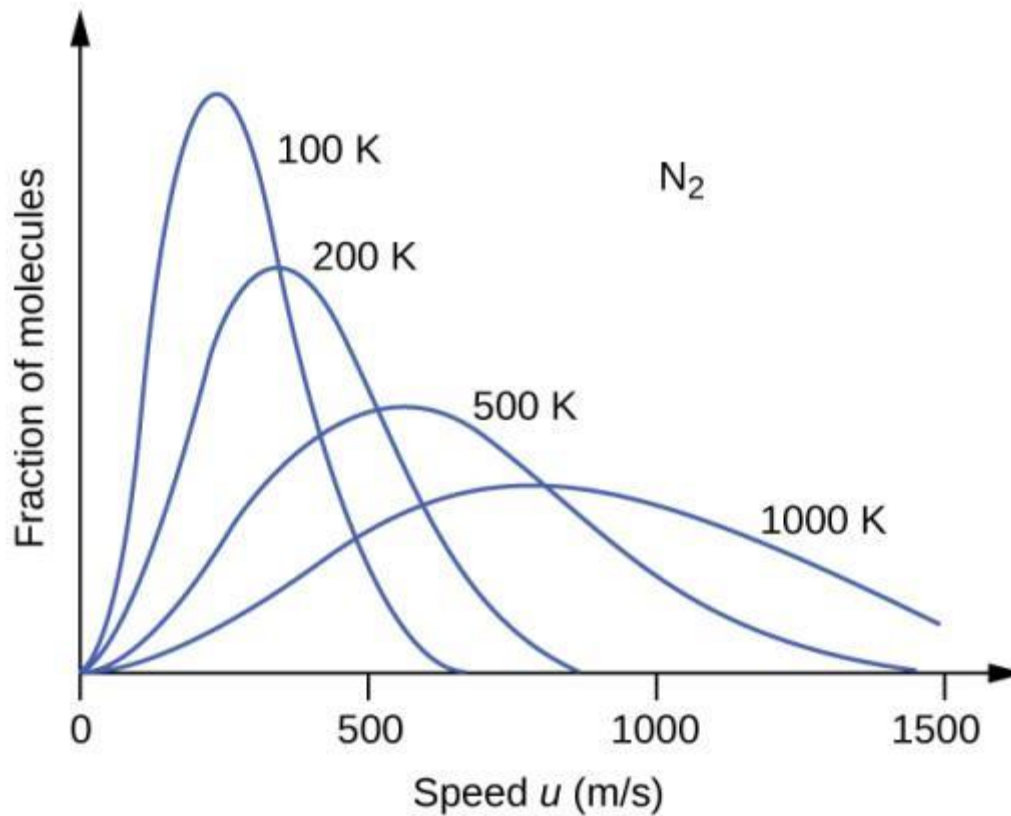
(a) When gas temperature increases, gas pressure increases due to increased force and frequency of molecular collisions. (b) When volume decreases, gas pressure increases due to reduced frequency of molecular collisions. (c) When the amount of gas increases at a constant pressure, volume increases to yield a constant number of collisions per unit wall area.

Figure 9.32



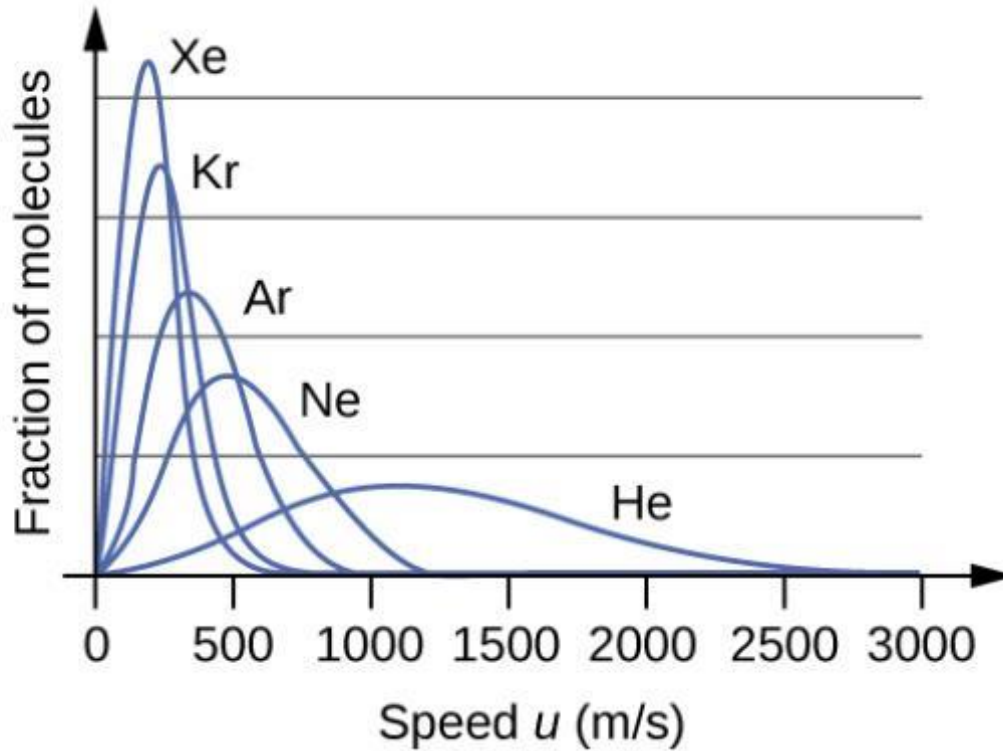
The molecular speed distribution for oxygen gas at 300 K is shown here. Very few molecules move at either very low or very high speeds. The number of molecules with intermediate speeds increases rapidly up to a maximum, which is the most probable speed, then drops off rapidly. Note that the most probable speed, v_p , is a little less than 400 m/s, while the root mean square speed, u_{rms} , is closer to 500 m/s.

Figure 9.33



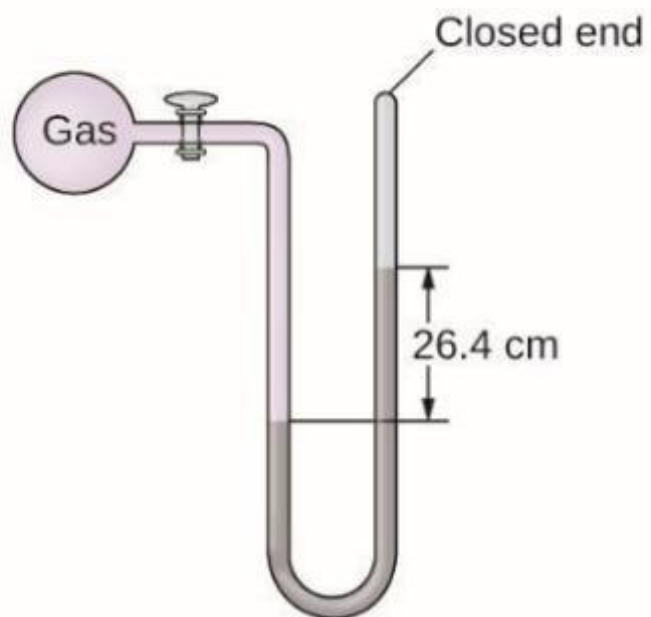
The molecular speed distribution for nitrogen gas (N_2) shifts to the right and flattens as the temperature increases; it shifts to the left and heightens as the temperature decreases.

Figure 9.34

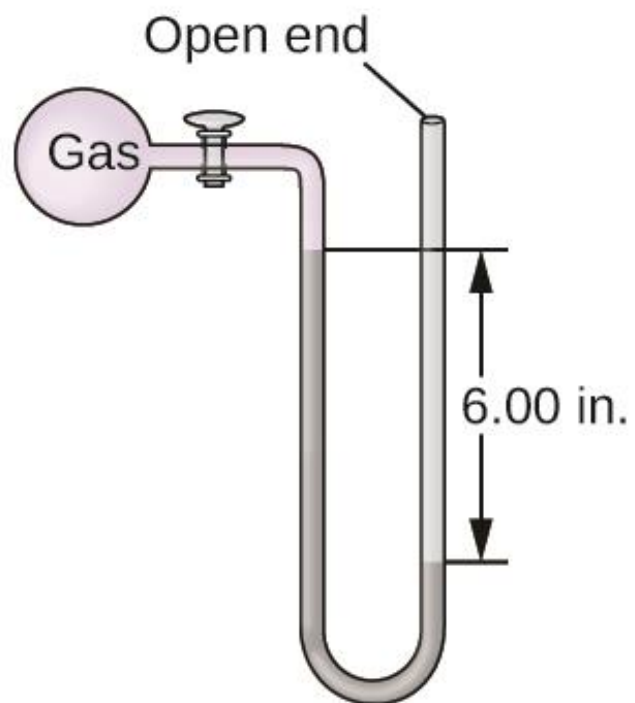


Molecular speed is directly related to molecular mass. At a given temperature, lighter molecules move faster on average than heavier molecules.

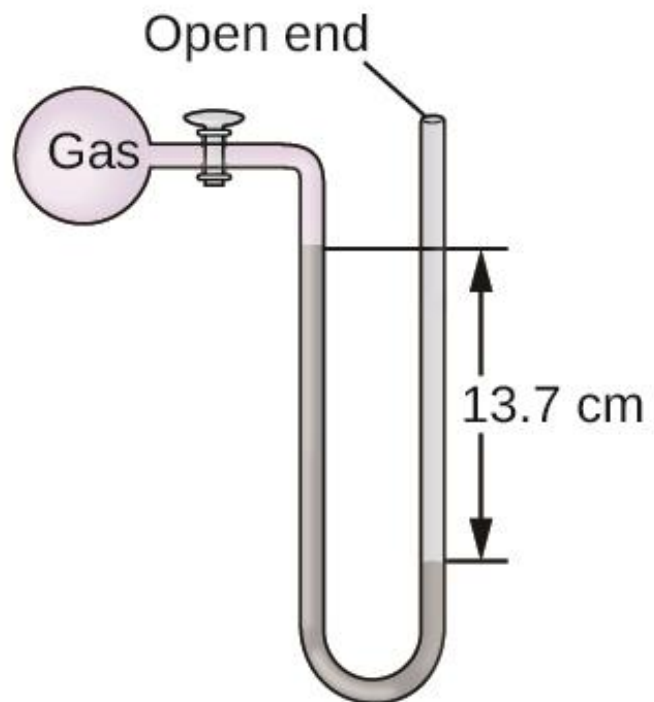
Exercise 13



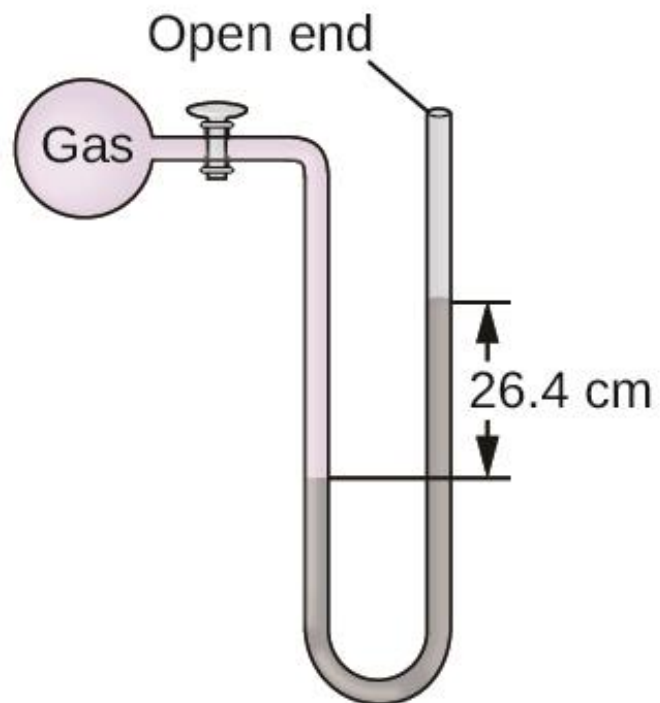
Exercise 14



Exercise 15



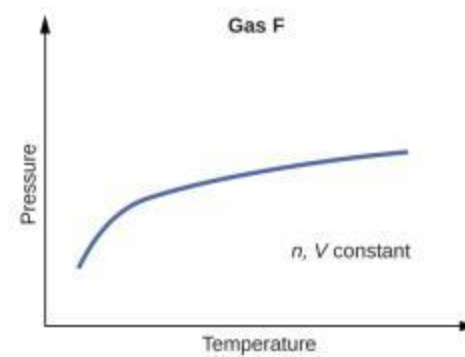
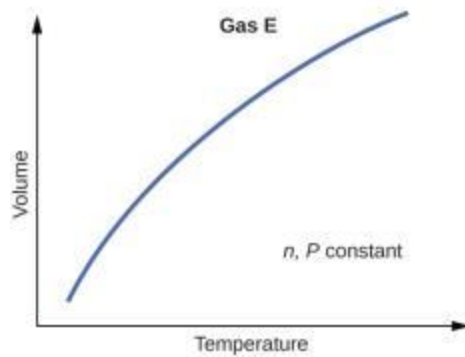
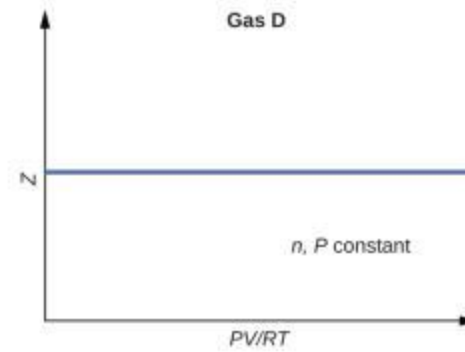
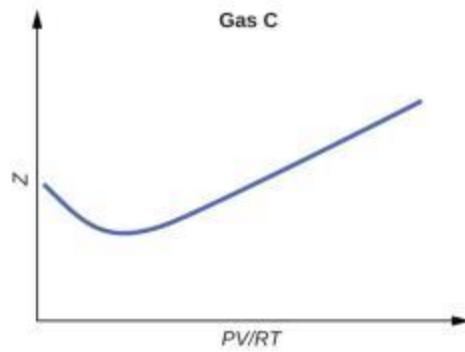
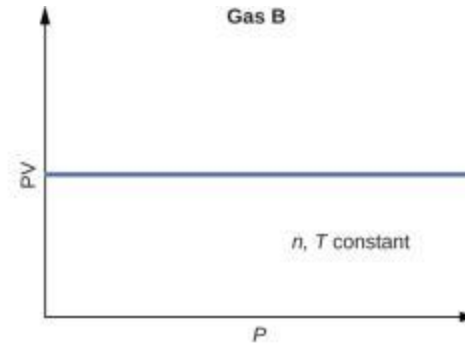
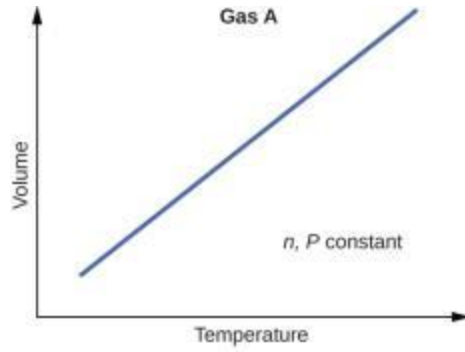
Exercise 16



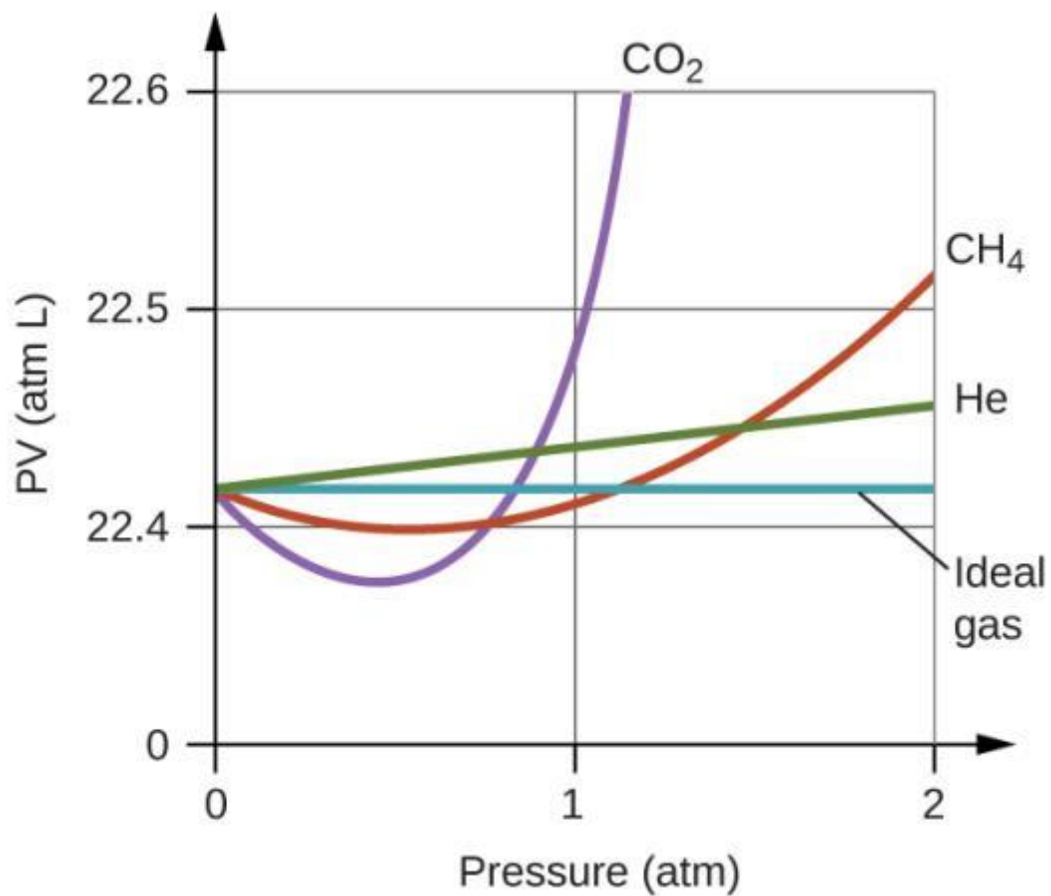
Exercise 31



Exercise 99



Exercise 100



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