Chemistry, The Central Science, 11th edition
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Chapter 10
Gases

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Announcements

- Office hours
  - Mon, Wed, 11:30-12:30 am
  - Sun,Tue,Thu 11:30-12:30 pm
- Office Location
  - D1-L0
- Reading
  - Chapter 10, Sections (10.2),(10.3), (10.40, (8.5, (10.6)and (10.8)
- Suggested Problems
  15,19,23,31,33,37,41,45,49,53,57,59,61,65,69 and 81

Gases

Characteristics of Gases

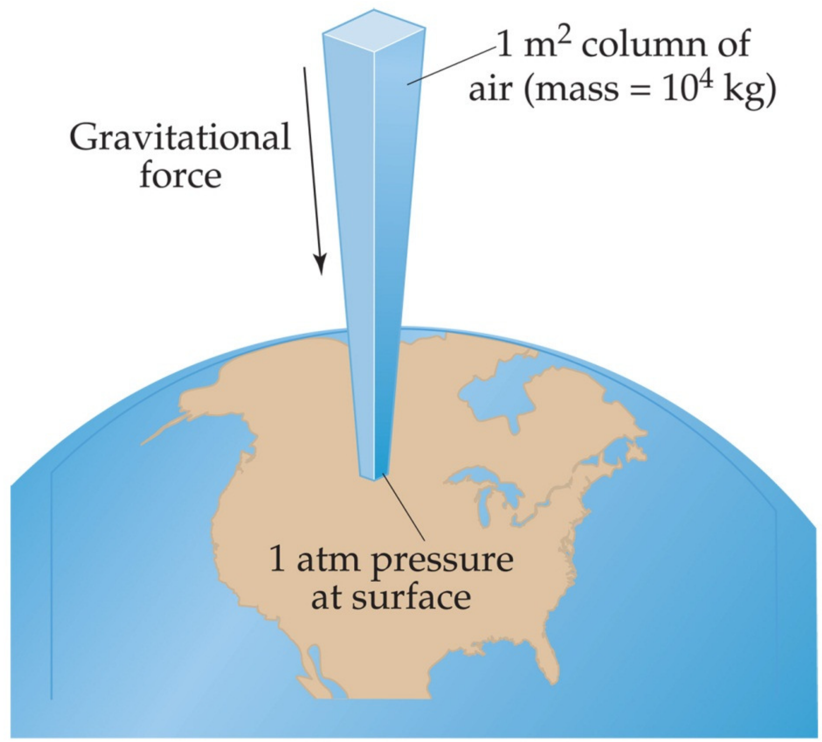
- Unlike liquids and solids, gases
  - expand to fill their containers;
  - are highly compressible;
  - have extremely low densities.

Pressure

- Pressure is the amount of force applied to an area.

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

- Atmospheric pressure is the weight of air per unit of area.



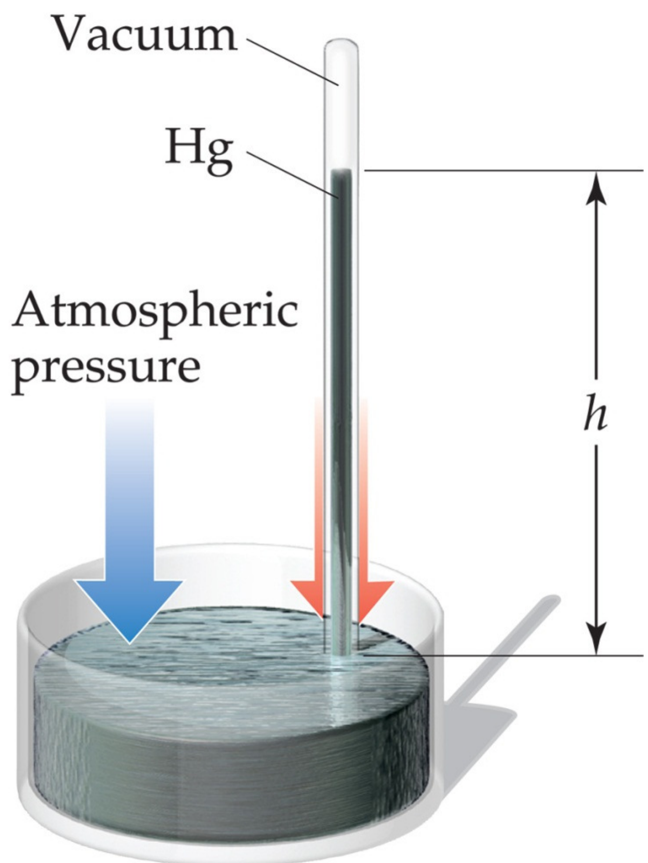
Units of Pressure

- Pascals
  - $1 \text{ Pa} = 1 \text{ N/m}^2$
- Bar
  - $1 \text{ bar} = 10^5 \text{ Pa} = 100 \text{ kPa}$

Units of Pressure

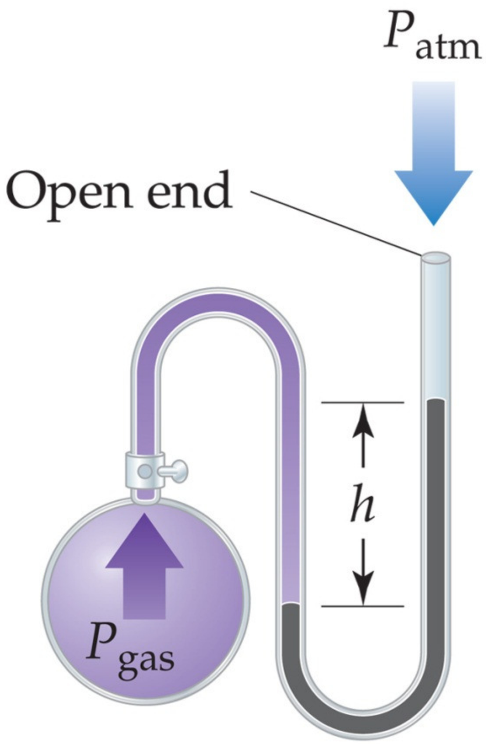
- mm Hg or torr
  - These units are literally the difference in the heights measured in mm (h) of two connected columns of mercury.

- Atmosphere
  - 1.00 atm = 760 torr



Gases

Manometer



$$P_{gas} = P_{atm} + P_h$$

This device is used to measure the difference in pressure between atmospheric pressure and that of a gas in a vessel.

Standard Pressure

- Normal atmospheric pressure at sea level is referred to as **standard pressure**.

- It is equal to
  - 1.00 atm
  - 760 torr (760 mm Hg)
  - 101.325 kPa

Sample Exercise 10.1 Converting Units of Pressure

(a) Convert 0.357 atm to torr. **(b)** Convert 6.6×10^{-2} torr to atm. **(c)** Convert 147.2 kPa to torr.

Sample Exercise 10.1 Converting Units of Pressure

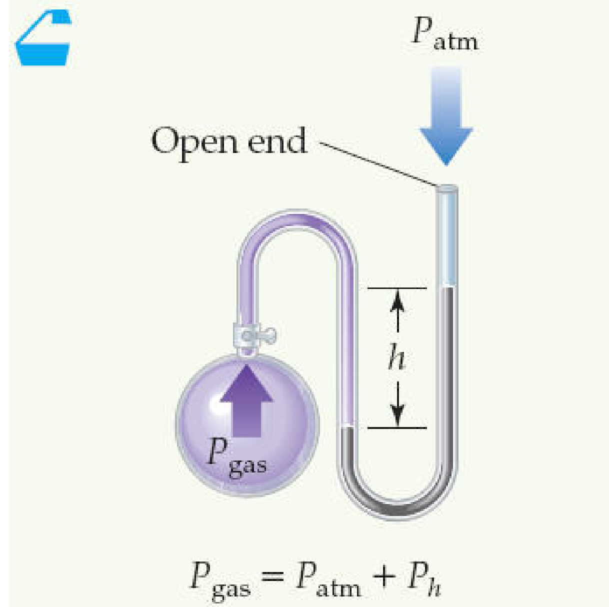
Practice Exercise

(a) In countries that use the metric system, such as Canada, atmospheric pressure in weather reports is given in units of kPa. Convert a pressure of 745 torr to kPa. **(b)** An English unit of pressure sometimes used in engineering is pounds per square inch ($lb/in.^2$), or psi: 1 atm = 14.7 $lb/in.^2$. If a pressure is reported as 91.5 psi, express the measurement in atmospheres.

Answer: **(a)** 99.3 kPa, **(b)** 6.22 atm

Sample Exercise 10.2 Using a Manometer to Measure Gas Pressure

On a certain day the barometer in a laboratory indicates that the atmospheric pressure is 764.7 torr. A sample of gas is placed in a flask attached to an open-end mercury manometer, shown in Figure 10.3. A meter stick is used to measure the height of the mercury above the bottom of the manometer. The level of mercury in the open-end arm of the manometer has a height of 136.4 mm, and the mercury in the arm that is in contact with the gas has a height of 103.8 mm. What is the pressure of the gas **(a)** in atmospheres,
(b) in kPa?



▲ Figure 10.3 **A mercury manometer.** This device is sometimes employed in the laboratory to measure gas pressures near atmospheric pressure.

Sample Exercise 10.2 Using a Manometer to Measure Gas Pressure

Sample Exercise 10.2 Using a Manometer to Measure Gas Pressure

Practice Exercise

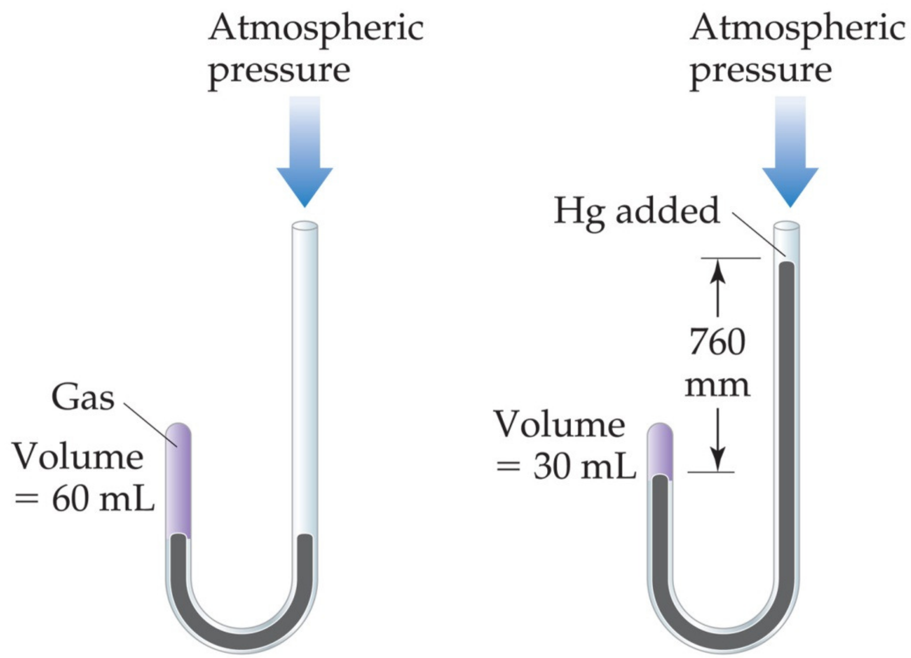
Convert a pressure of 0.975 atm into Pa and kPa

.

Answer: 9.88×10^4 Pa and 98.8 kPa

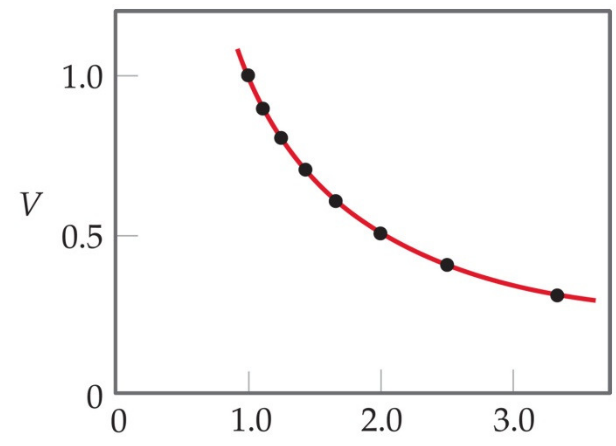
Boyle's Law

The volume of a fixed quantity of gas at constant temperature is inversely proportional to the pressure.



As *P* and *V* are inversely proportional

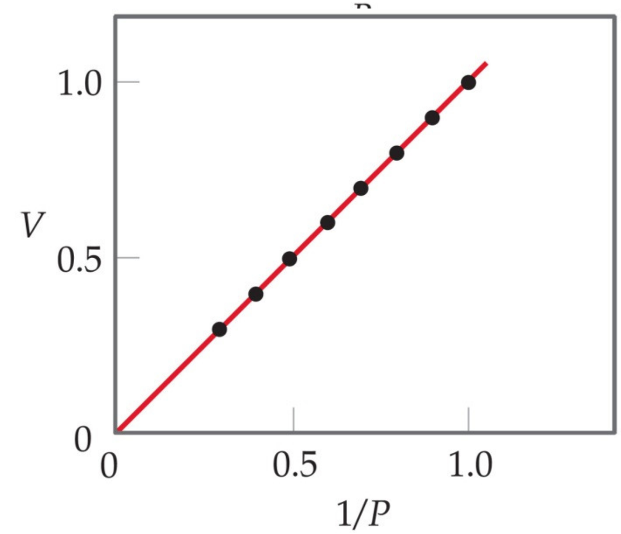
A plot of *V* versus *P* results in a curve.



Since $PV = k$

$V = k\,(1/P)$

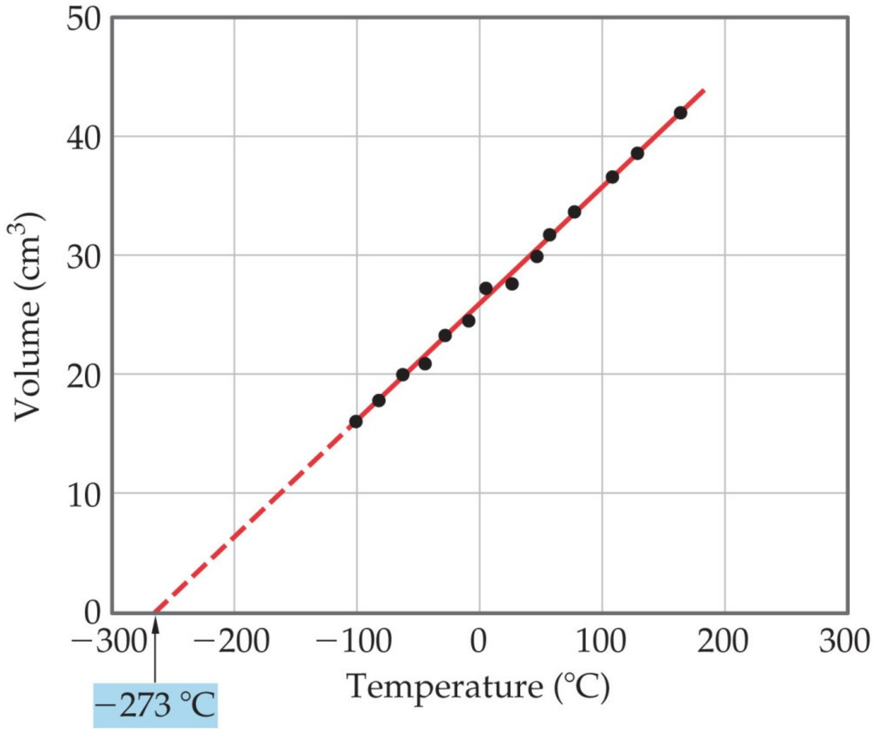
This means a plot of *V* versus 1/*P* will be a straight line.



Gases

Charles's Law

- The volume of a fixed amount of gas at constant pressure is directly proportional to its absolute temperature.
- i.e., $\frac{V}{T} = k$

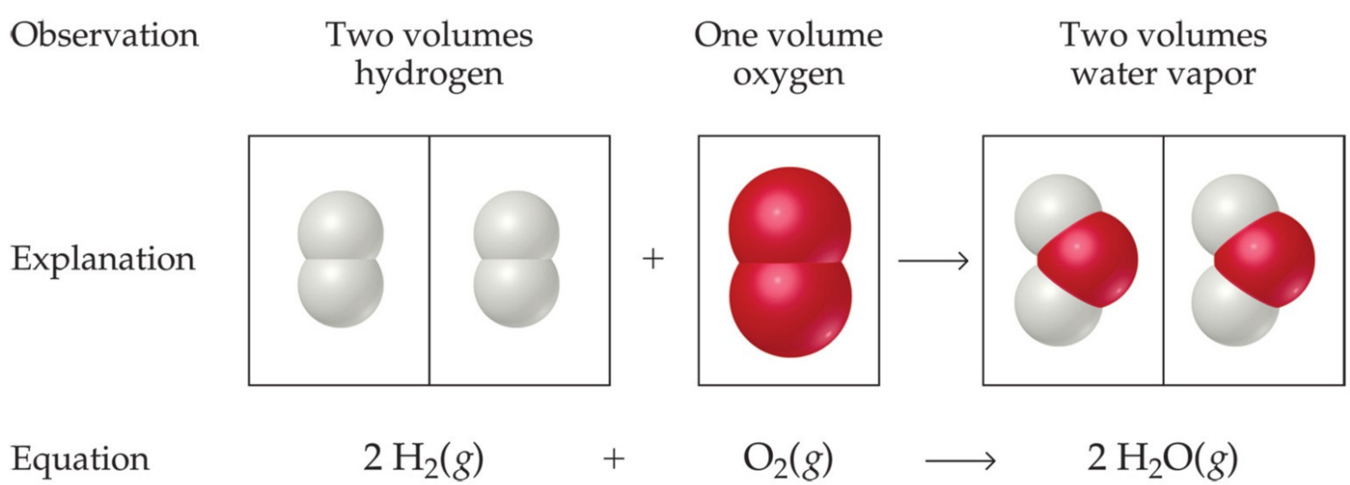


A plot of *V* versus *T* will be a straight line.

Gases

Avogadro's Law

- The volume of a gas at constant temperature and pressure is directly proportional to the number of moles of the gas.
- Mathematically, this means $V = kn$



Ideal-Gas Equation

- So far we've seen that

$$V \propto 1/P \quad \text{(Boyle's law)}$$
$$V \propto T \quad \text{(Charles's law)}$$
$$V \propto n \quad \text{(Avogadro's law)}$$

- Combining these, we get

$$V \propto \frac{nT}{P}$$

Sample Exercise 10.3 Evaluating the Effects of Changes in *P*, *V*, *n* and *T* on a Gas

Suppose we have a gas confined to a cylinder as shown in Figure 10.12. Consider the following changes: **(a)** Heat the gas from 298 K to 360 K, while maintaining the piston in the position shown in the drawing. **(b)** Move the piston to reduce the volume of gas from 1 L to 0.5 L. **(c)** Inject additional gas through the gas inlet valve. Indicate how each of these changes will affect the average distance between molecules, the pressure of the gas, and the number of moles of gas present in the cylinder

Practice Exercise

What happens to the density of a gas as **(a)** the gas is heated in a constant-volume container; **(b)** the gas is compressed at constant temperature; **(c)** additional gas is added to a constant-volume container?

Answer: **(a)** no change, **(b)** increase, **(c)** increase

Ideal-Gas Equation

The constant of proportionality is known as *R*, the gas constant.

Units	Numerical Value
L-atm/mol-K	0.08206
J/mol-K*	8.314
cal/mol-K	1.987
m^3-Pa/mol-K*	8.314
L-torr/mol-K	62.36

*SI unit

Ideal-Gas Equation

The relationship

$$V \propto \frac{nT}{P}$$

then becomes

$$V = R\frac{nT}{P}$$

or

$$PV = nRT$$

Densities of Gases

If we divide both sides of the ideal-gas equation by V and by RT, we get

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

Gases

Sample Exercise 10.4 Using the Ideal-Gas equation

Calcium carbonate, $CaCO_3(s)$, decomposes upon heating to give $CaO(s)$ and $CO_2(g)$. A sample of $CaCO_3$ is decomposed, and the carbon dioxide is collected in a 250-mL flask. After the 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?

Sample Exercise 10.4 Using the Ideal-Gas equation

Practice Exercise

Tennis balls are usually filled with air or N_2 gas to a pressure above atmospheric pressure to increase their "bounce." If a particular tennis ball has a volume of 144 cm^3 and contains 0.33 g of N_2 gas, what is the pressure inside the ball at 24 °C?

Answer: 2.0 atm

Sample Exercise 10.5 Calculating the Effect of Temperature Changes on Pressure

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

Sample Exercise 10.5 Calculating the Effect of Temperature Changes on Pressure

Practice Exercise

A large natural-gas storage tank is arranged so that the pressure is maintained at 2.20 atm. On a cold day in December when the temperature is –15 °C (4 °F), the volume of gas in the tank is 3.25×10^3 m^3. What is the volume of the same quantity of gas on a warm July day when the temperature is 31 °C (88 °F)?

Answer: 3.83×10^3 m^3

Sample Exercise 10.6 Calculating the Effect of Changing *P* and *T* on the Volume of a Gas

An inflated balloon has a volume of 6.0 L at sea level (1.0 atm) and is allowed to ascend in altitude 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.

Sample Exercise 10.6 Calculating the Effect of Changing *P* and *T* on the Volume of a Gas

Practice Exercise

A 0.50-mol sample of oxygen gas is confined at 0 °C in a cylinder with a movable piston, such as that shown in Figure 10.12. The gas has an initial pressure of 1.0 atm. The piston then compresses the gas so that its final volume is half the initial volume. The final pressure of the gas is 2.2 atm. What is the final temperature of the gas in degrees Celsius?

Answer: 27 °C

Densities of Gases

- We know that
  - moles × molecular mass = mass

$$n \times \mathcal{M} = m$$

- So multiplying both sides by the molecular mass ($\mathcal{M}$) gives

$$\frac{m}{V} = \frac{P\mathcal{M}}{RT}$$

Gases

Densities of Gases

- Mass ÷ volume = density
- So,

$$d = \frac{m}{V} = \frac{P\mathcal{M}}{RT}$$

Note: One only needs to know the molecular mass, the pressure, and the temperature to calculate the density of a gas.

Molecular Mass

We can manipulate the density equation to enable us to find the molecular mass of a gas:

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

Becomes

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

Gases

Sample Exercise 10.7 Calculating Gas Density

What is the density of carbon tetrachloride vapor at 714 torr and 125 °C?

Sample Exercise 10.7 Calculating Gas Density

Practice Exercise

The mean molar mass of the atmosphere at the surface of Titan, Saturn's largest moon, is 28.6 g/mol. The surface temperature is 95 K, and the pressure is 1.6 atm. Assuming ideal behavior, calculate the density of Titan's atmosphere.

Answer: 5.9 g/L

Sample Exercise 10.8 Calculating the Molar Mass of a Gas

A series of measurements are made to determine the molar mass of an unknown gas. First, a large flask is evacuated and found to weigh 134.567 g. It is then filled with the gas to a pressure of 735 torr at 31 °C and reweighed. Its mass is now 137.456 g. Finally, the flask is filled with water at 31 °C and found to weigh 1067.9 g. (The density of the water at this temperature is 0.997 g/mL.) Assume that the ideal-gas equation applies, and calculate the molar mass of the unknown gas.

Sample Exercise 10.8 Calculating the Molar Mass of a Gas

Practice Exercise

Calculate the average molar mass of dry air if it has a density of 1.17 g/L at 21 °C and 740.0 torr.

Answer: 29.0 g/mol.

Sample Exercise 10.9 Relating the Volume of a Gas to the Amount of Another Substance in a Reaction

The safety air bags in automobiles are inflated by nitrogen gas generated by the rapid decomposition of sodium azide, NaN_3:

$$2\ NaN_3(s) \rightarrow 2\ Na(s) + 3\ N_2(g)$$

If an air bag has a volume of 36 L and is to be filled with nitrogen gas at a pressure of 1.15 atm at a temperature of 26.0 °C, how many grams of NaN_3 must be decomposed?

Sample Exercise 10.9 Relating the Volume of a Gas to the Amount of Another Substance in a Reaction

Practice Exercise

In the first step in the industrial process for making nitric acid, ammonia reacts with oxygen in the presence of a suitable catalyst to form nitric oxide and water vapor:

$$4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g)$$

How many liters of NH_3(g) at 850 ° C and 5.00 atm are required to react with 1.00 mol of O_2(g) in this reaction?

Answer: 14.8 L

Dalton's Law of Partial Pressures

- The total pressure of a mixture of gases equals the sum of the pressures that each would exert if it were present alone.
- In other words,

$$P_{\text{total}} = P_1 + P_2 + P_3 + \ldots$$

Sample Exercise 10.10 Applying Dalton's Law to the Partial Pressures

A gaseous mixture made from 6.00 g O_2 and 9.00 g CH_4 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?

Sample Exercise 10.10 Applying Dalton's Law to the Partial Pressures

Practice Exercise

What is the total pressure exerted by a mixture of 2.00 g of H_2 and 8.00 g of N_2 at 273 K in a 10.0-L vessel?

Answer: 2.86 atm

Sample Exercise 10.11 Relating Mole Fractions to Partial Pressures

A study of the effects of certain gases on plant growth requires a synthetic atmosphere composed of 1.5 mol percent CO_2, 18.0 mol percent O_2, and 80.5 mol percent Ar. **(a)** Calculate the partial pressure of O_2 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 O_2 are needed?

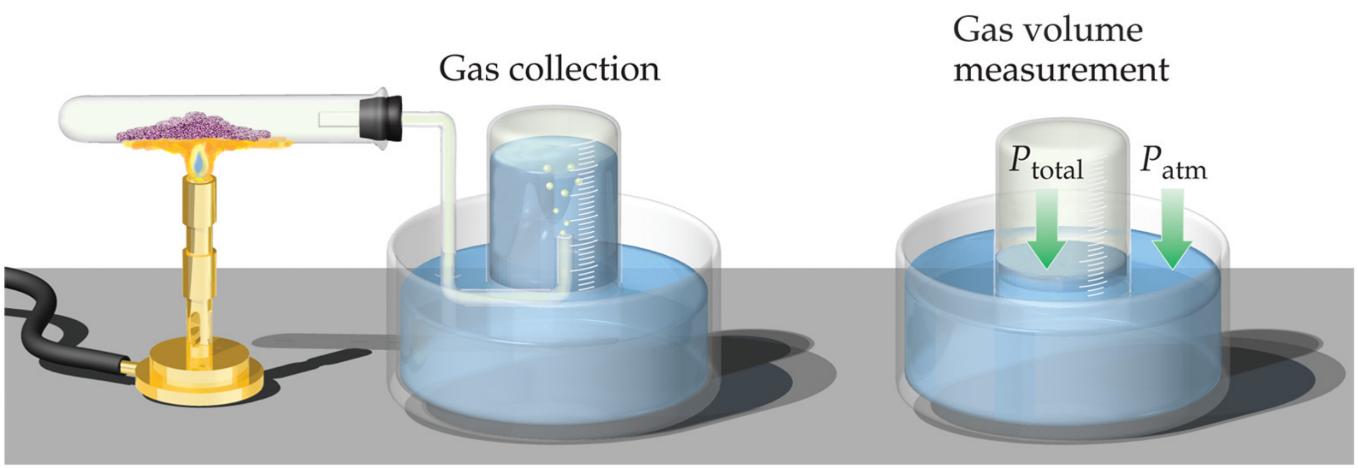
Sample Exercise 10.11 Relating Mole Fractions to Partial Pressures

Practice Exercise

From data gathered by *Voyager 1*, scientists have estimated the composition of the atmosphere of Titan, Saturn's largest moon. The total pressure on the surface of Titan is 1220 torr. The atmosphere consists of 82 mol percent N_2, 12 mol percent Ar, and 6.0 mol percent CH_4. Calculate the partial pressure of each of these gases in Titan's atmosphere.

Answer: 1.0×10^3 torr N_2, 1.5×10^2 torr Ar, and 73 torr CH_4

Partial Pressures



- When one collects a gas over water, there is water vapor mixed in with the gas.
- To find only the pressure of the desired gas, one must subtract the vapor pressure of water from the total pressure.

Sample Exercise 10.12 Calculating the Amount of Gas Collected over Water

A sample of $KClO_3$ is partially decomposed (Equation 10.16), producing O_2 gas that is collected over water as in Figure 10.16. The volume of gas collected is 0.250 L at 26 °C and 765 torr total pressure. **(a)** How many moles of O_2 are collected? **(b)** How many grams of $KClO_3$ were decomposed?

Sample Exercise 10.12 Calculating the Amount of Gas Collected over Water

Practice Exercise

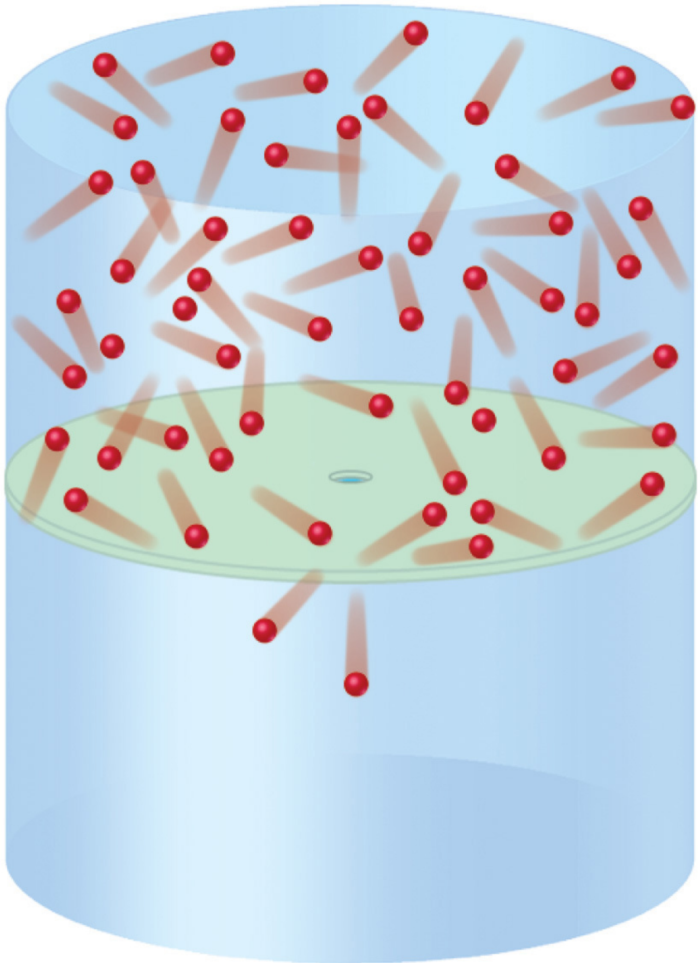
Ammonium nitrite, NH_4NO_2, decomposes upon heating to form N_2 gas:

$$NH_4NO_2(s) \longrightarrow N_2(g) + 2\,H_2O(l)$$

When a sample of NH_4NO_2 is decomposed in a test tube, as in Figure 10.16, 511mL of N_2 gas is collected over water at 26 °C and 745 torr total pressure. How many grams of NH_4NO_2 were decomposed?

Answer: 1.26 g

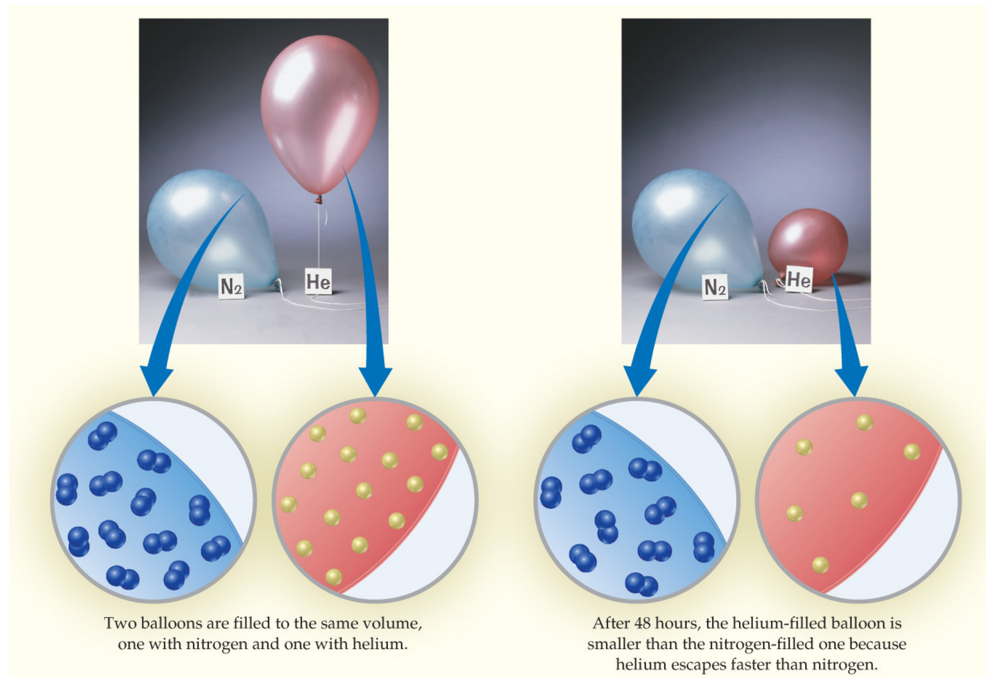
Effusion



Effusion is the escape of gas molecules through a tiny hole into an evacuated space.

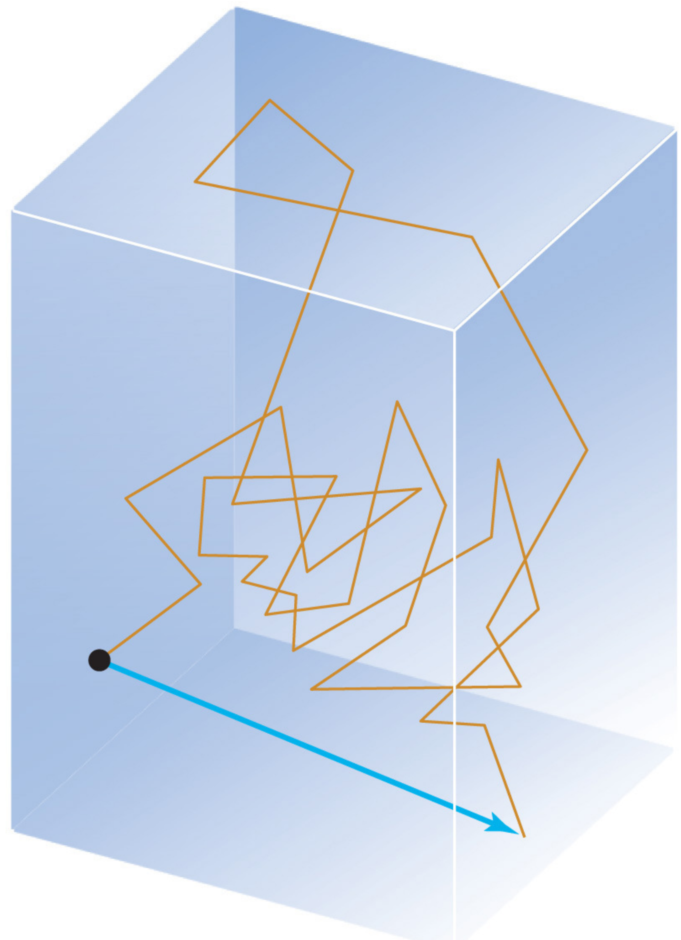
Effusion

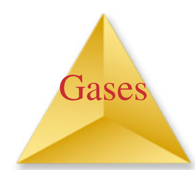
The difference in the rates of effusion for helium and nitrogen, for example, explains a helium balloon would deflate faster.



Diffusion

Diffusion is the spread of one substance throughout a space or throughout a second substance.





Graham's Law

$$KE_1 = KE_2$$

$$1/2\ m_1 v_1^2 = 1/2\ m_2 v_2^2$$

$$\frac{m_1}{m_2} = \frac{v_2^2}{v_1^2}$$

$$\frac{\sqrt{m_1}}{\sqrt{m_2}} = \frac{\sqrt{v_2^2}}{\sqrt{v_1^2}} = \frac{v_2}{v_1}$$

Sample Exercise 10.14 Calculating a Root-Mean-Square Speed

Calculate the rms speed, u, of an N_2 molecule at 25 °C.

Practice Exercise

What is the rms speed of an He atom at 25 °C?

Answer: 1.36×10^3 m/s

Sample Exercise 10.15 Applying Graham's Law

An unknown gas composed of homonuclear diatomic molecules effuses at a rate that is only 0.355 times that of O_2 at the same temperature. Calculate the molar mass of the unknown, and identify it.

Sample Exercise 10.15 Applying Graham's Law

Practice Exercise

Calculate the ratio of the effusion rates of N_2 and O_2, r_{N_2}/r_{O_2}.

Answer: $r_{N_2}/r_{O_2} = 1.07$

Next

- Physical Properties of Solution