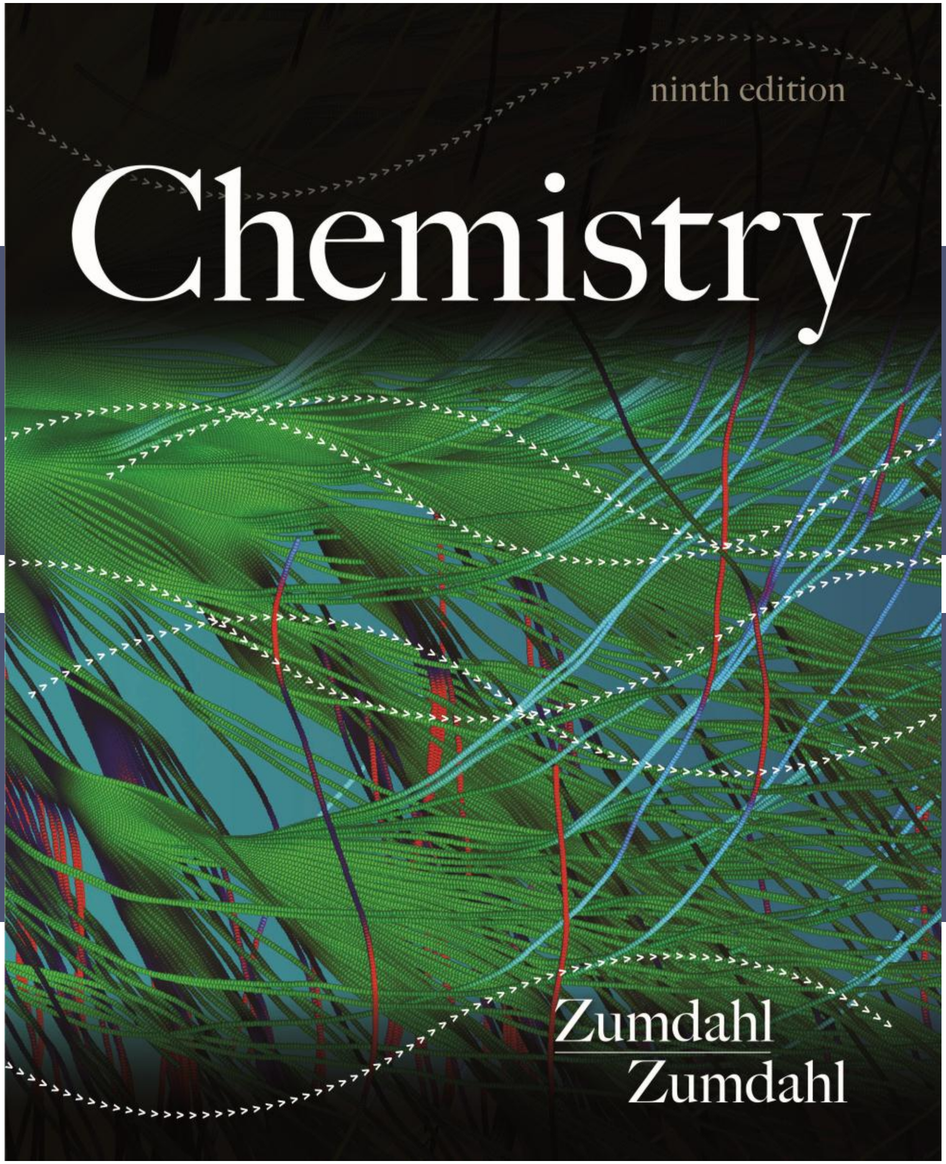

Chapter 5

Gases

Chapter 5
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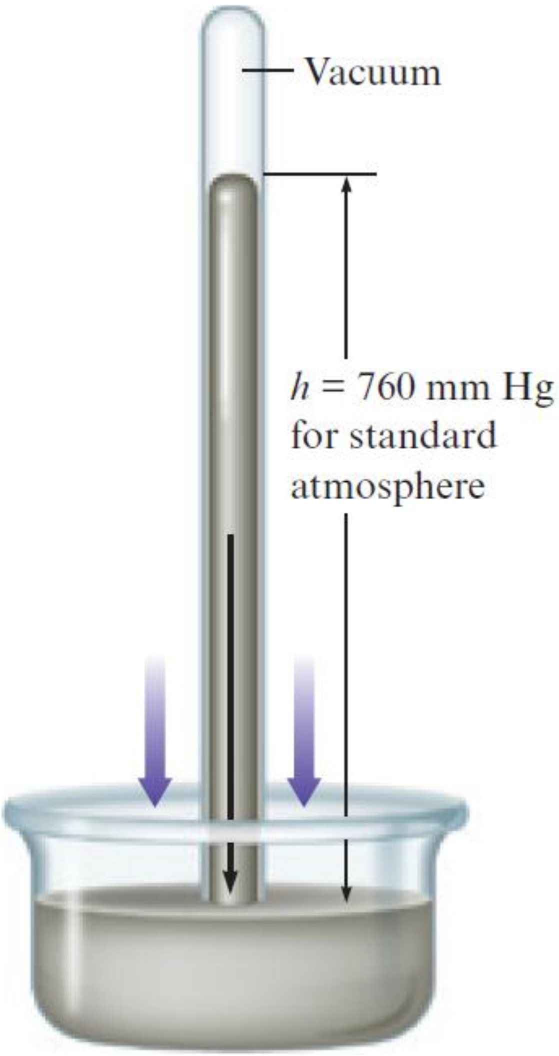
Properties of Gases

- Uniformly fill any container
- Easily compressible
- Completely mix with other gases
- Exert pressure on their surroundings

Figure 5.1 - The Collapsing Can Experiment

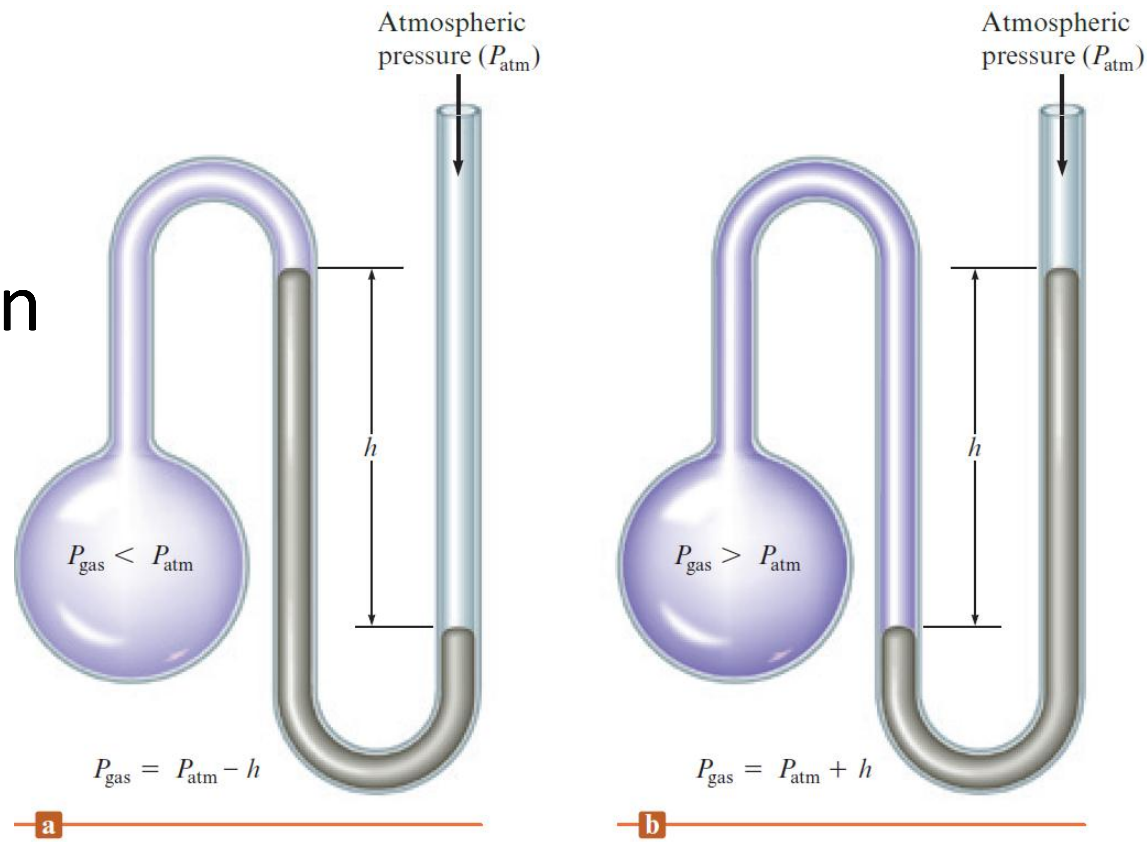
Barometer

- Device used to measure atmospheric pressure
  - Atmospheric pressure results from the weight of the air
    - Variations are attributed to change in weather conditions and altitudes
- Contains a glass tube filled with liquid mercury that is inverted in a dish containing mercury



Manometer

- Device used for measuring the pressure of a gas in a container



Units of Pressure

- Based on the height of the mercury column that a gas can support
- **mm Hg** (millimeter of mercury)
  - **Torr** and mm Hg are used interchangeably
- **Standard atmosphere** (atm)

1 standard atmosphere = 1 atm = 760 mm Hg = 760 torr

Units of Pressure (Continued)

- Defined as force per unit area

$$\text{Pressure} = \frac{\text{force}}{\text{area}}$$

  - SI system
    - Newton/m^2 = 1 **pascal** (Pa)
    - 1 standard atmosphere = 101,325 Pa
    - 1 atmosphere ≈ 10^5 Pa

Interactive Example 5.1 - Pressure Conversions

- The pressure of a gas is measured as 49 torr
  - Represent this pressure in both atmospheres and pascals

Interactive Example 5.1 - Solution

$$49\ \cancel{\text{torr}} \times \frac{1\ \text{atm}}{760\ \cancel{\text{torr}}} = 6.4 \times 10^{-2}\ \text{atm}$$

$$6.4 \times 10^{-2}\ \cancel{\text{atm}} \times \frac{101{,}325\ \text{Pa}}{1\ \cancel{\text{atm}}} = 6.5 \times 10^{3}\ \text{Pa}$$

Boyle's Law

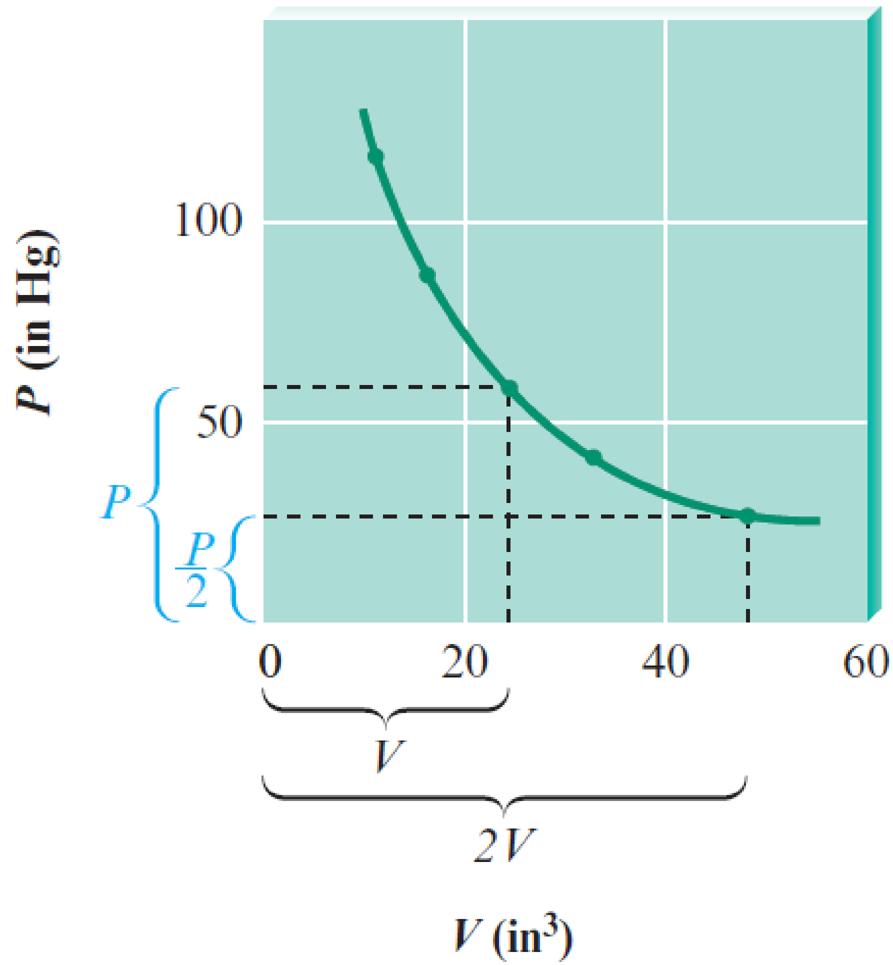
- Robert Boyle studied the relationship between the pressure of trapped gas and its volume

$$PV = k$$

  - *k* - Constant for a given sample of air at a specific temperature
  - Deduced that pressure and volume are inversely related
- **Ideal gas**: Gas that strictly obeys Boyle's law

Plot of Boyle's Law

- A plot of *P* versus *V* shows that the volume doubles as the pressure is halved
  - Resulting curve is a hyperbola



Plot of Boyle's Law (Continued)

- Boyle's law can be rearranged to mirror the straight line equation

$$V = \frac{k}{P} = k\frac{1}{P} \qquad\qquad y = mx + b$$

  - $y = V$
  - $x = 1/P$
  - $m = k$
  - $b = 0$

Figure 5.5 - Linear Plot of Boyle's Law

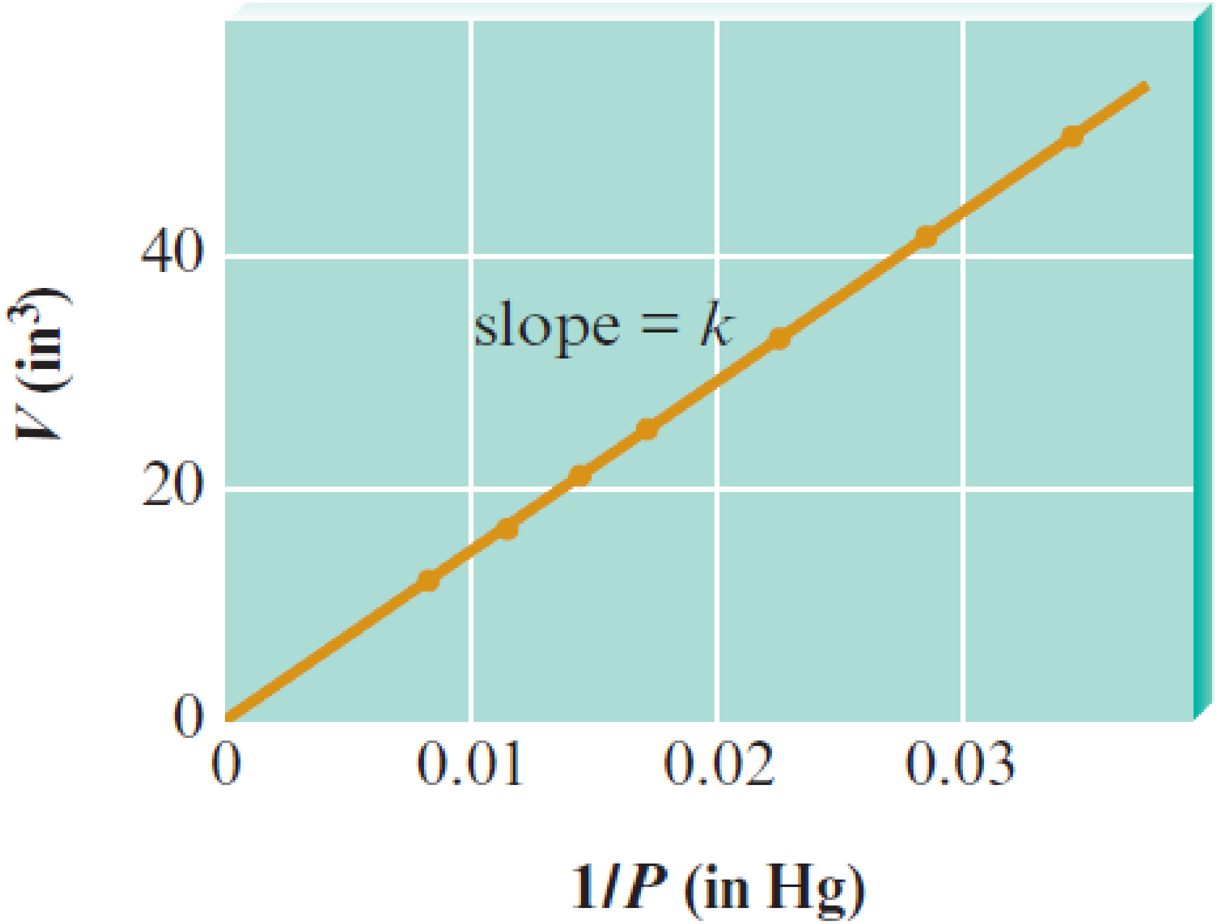
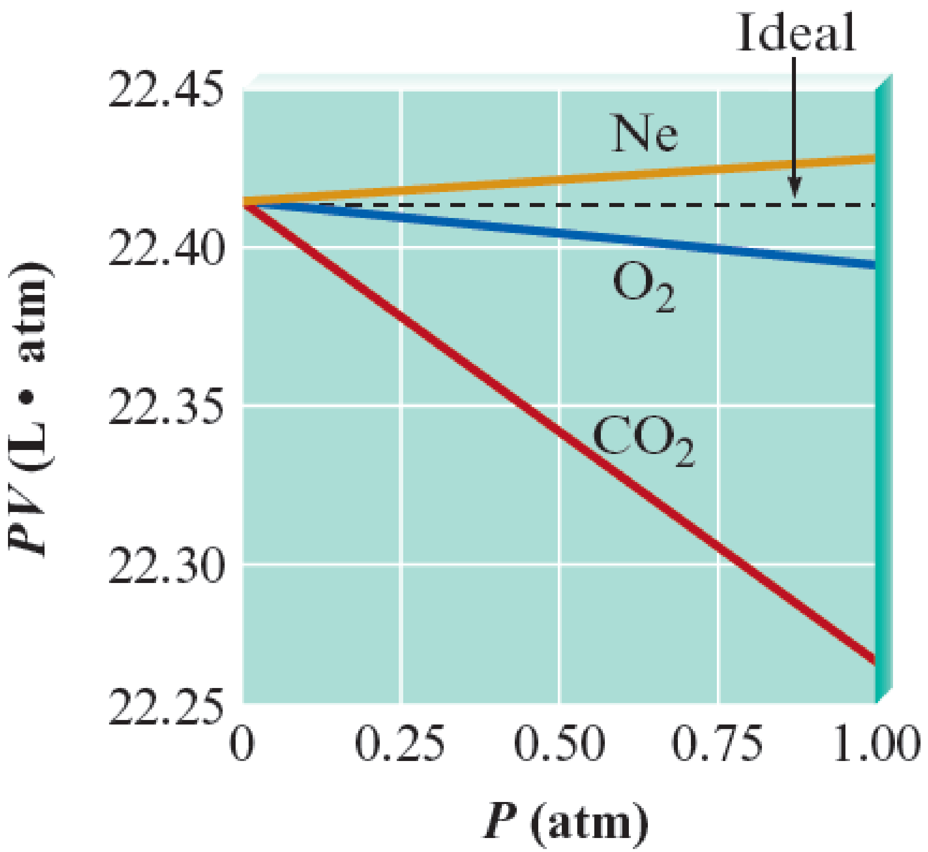


Figure 5.6 - A Plot of *PV* versus *P* for Several Gases at Pressures below 1 atm



Interactive Example 5.2 - Boyle’s Law I

- Sulfur dioxide (SO_2), a gas that plays a central role in the formation of acid rain, is found in the exhaust of automobiles and power plants
  - Consider a 1.53-L sample of gaseous SO_2 at a pressure of 5.6×10^3 Pa
    - If the pressure is changed to 1.5×10^4 Pa at a constant temperature, what will be the new volume of the gas?

Interactive Example 5.2 - Solution

- Where are we going?
  - To calculate the new volume of gas
- What do we know?
  - $P_1 = 5.6 \times 10^3$ Pa　　$P_2 = 1.5 \times 10^4$ Pa
  - $V_1 = 1.53$ L　　$V_2 = ?$
- What information do we need?
  - Boyle's law

$$PV = k$$

Interactive Example 5.2 - Solution (Continued 1)

- How do we get there?
  - What is Boyle's law (in a form useful with our knowns)?

$$P_1V_1 = P_2V_2$$

  - What is V_2?

$$V_2 = \frac{P_1V_1}{P_2} = \frac{5.6 \times 10^3 \ \cancel{\text{Pa}} \times 1.53 \text{ L}}{1.5 \times 10^4 \ \cancel{\text{Pa}}} = 0.57 \text{ L}$$

    - The new volume will be 0.57 L

Interactive Example 5.2 - Solution (Continued 2)

- Reality check
  - The new volume (0.57 L) is smaller than the original volume
  - As pressure increases, the volume should decrease, so our answer is reasonable

Exercise

- A particular balloon is designed by its manufacturer to be inflated to a volume of no more than 2.5 L
  - If the balloon is filled with 2.0 L helium at sea level, is released, and rises to an altitude at which the atmospheric pressure is only 500 mm Hg, will the balloon burst? (Assume temperature is constant)

The balloon will burst

Example 5.3 - Boyle's Law II

- In a study to see how closely gaseous ammonia obeys Boyle's law, several volume measurements were made at various pressures, using 1.0 mole of NH_3 gas at a temperature of 0° C
  - Using the results listed below, calculate the Boyle's law constant for NH_3 at the various pressures

Example 5.3 - Boyle's Law II (Continued)

Experiment	Pressure (atm)	Volume (L)
1	0.1300	172.1
2	0.2500	89.28
3	0.3000	74.35
4	0.5000	44.49
5	0.7500	29.55
6	1.000	22.08

Example 5.3 - Solution

- To determine how closely NH_3 gas follows Boyle's law under these conditions, we calculate the value of *k* (in L · atm) for each set of values

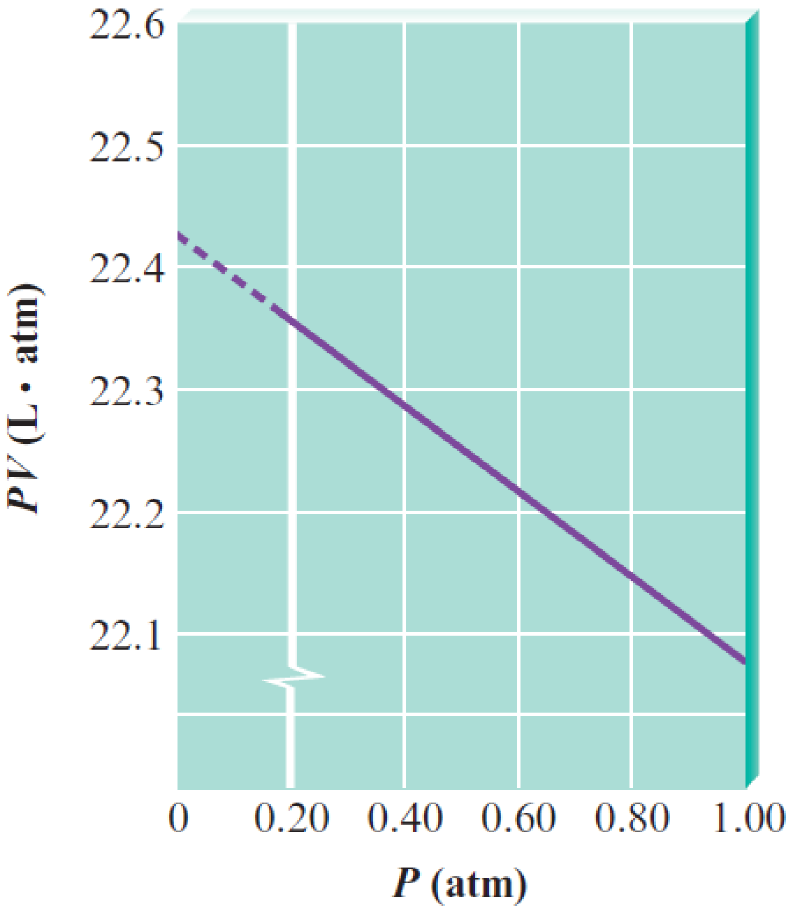
Experiment	1	2	3	4	5	6
$k = PV$	22.37	22.32	22.31	22.25	22.16	22.08

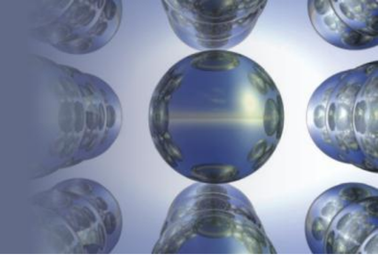
Example 5.3 - Solution (Continued 1)

- Although the deviations from true Boyle's law behavior are quite small at these low pressures, note that the value of *k* changes regularly in one direction as the pressure is increased
  - To calculate the ideal value of *k* for NH_3, we can plot *PV* versus *P*, and extrapolate back to zero pressure, where, for reasons we will discuss later, a gas behaves most ideally

Example 5.3 - Solution (Continued 2)

- The value of *k* obtained by this extrapolation is 22.41 L· atm
  - Notice that this is the same value obtained from similar plots for the gases CO_2, O_2, and Ne at 0° C

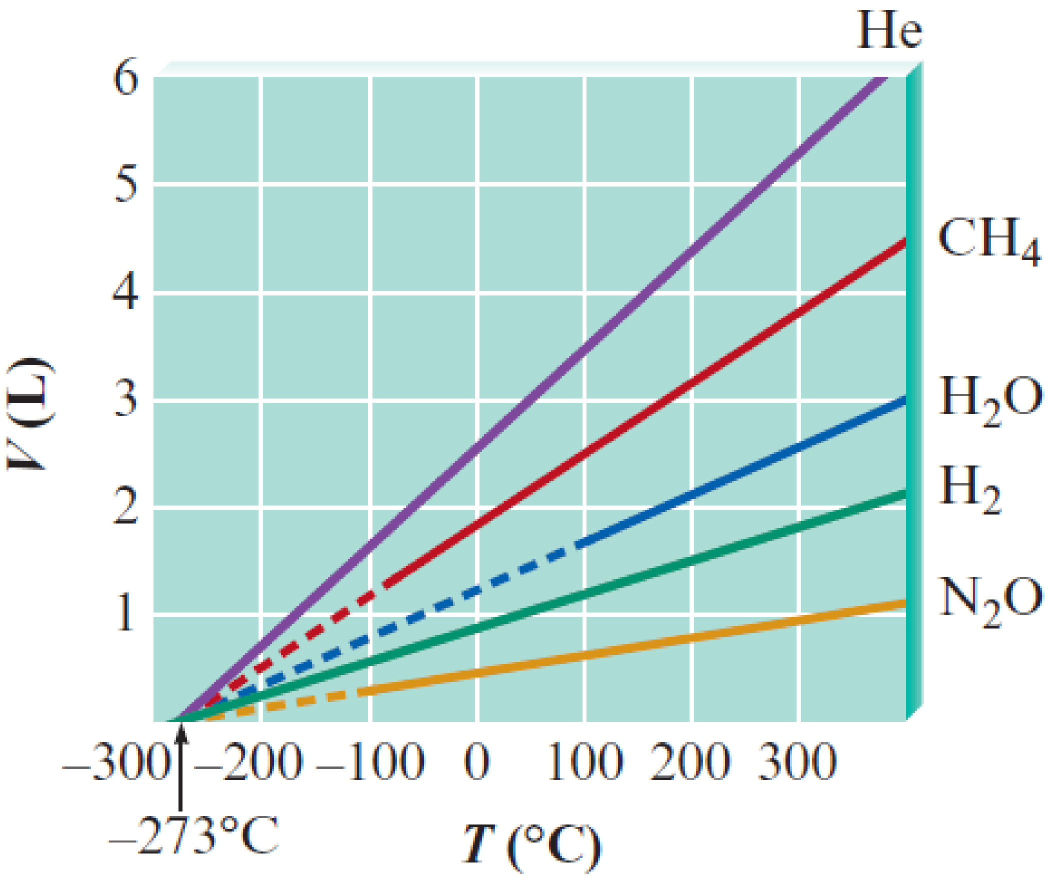




Jacques Charles

- French physicist who made the first solo balloon flight
- Determined that the volume of a gas at constant pressure increases linearly with the temperature of the gas
  - Plot of volume of gas (at constant pressure) versus temperature (in °C) gives a straight line

Figure 5.8 - Plots of *V* versus *T* (° C) for Several Gases



Charles's Law

- Volume of each gas is directly proportional to the temperature
  - Volume extrapolates to zero when the temperature is equal to 0 K (**absolute zero**)

$$V = bT$$

    - *T* - Temperature in kelvins
    - *b* - Proportionality constant

Critical Thinking

- According to Charles's law, the volume of a gas is directly related to its temperature in kelvins at constant pressure and number of moles
  - What if the volume of a gas was directly related to its temperature in degrees Celsius at constant pressure and number of moles?
    - What differences would you notice?

Interactive Example 5.4 - Charles's Law

- A sample of gas at 15°C and 1 atm has a volume of 2.58 L
  - What volume will this gas occupy at 38°C and 1 atm?

Interactive Example 5.4 - Solution

- Where are we going?
  - To calculate the new volume of gas
- What do we know?
  - $T_1 = 15^\circ\ \text{C} + 273 = 288\ \text{K}$
  - $T_2 = 38^\circ\ \text{C} + 273 = 311\ \text{K}$
  - $V_1 = 2.58\ \text{L}$
  - $V_2 = ?$

Interactive Example 5.4 - Solution (Continued 1)

- What information do we need?
  - Charles's law

$$\frac{V}{T} = b$$

- How do we get there?
  - What is Charles's law (in a form useful with our knowns)?

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Interactive Example 5.4 - Solution (Continued 2)

- What is V_2?

$$V_2 = \left(\frac{T_2}{T_1}\right) V_1 = \left(\frac{311\ \cancel{K}}{288\ \cancel{K}}\right) 2.58\ \text{L}$$

$$V_2 = 2.79\ \text{L}$$

- Reality check
  - The new volume is greater than the original volume, which makes physical sense because the gas will expand as it is heated

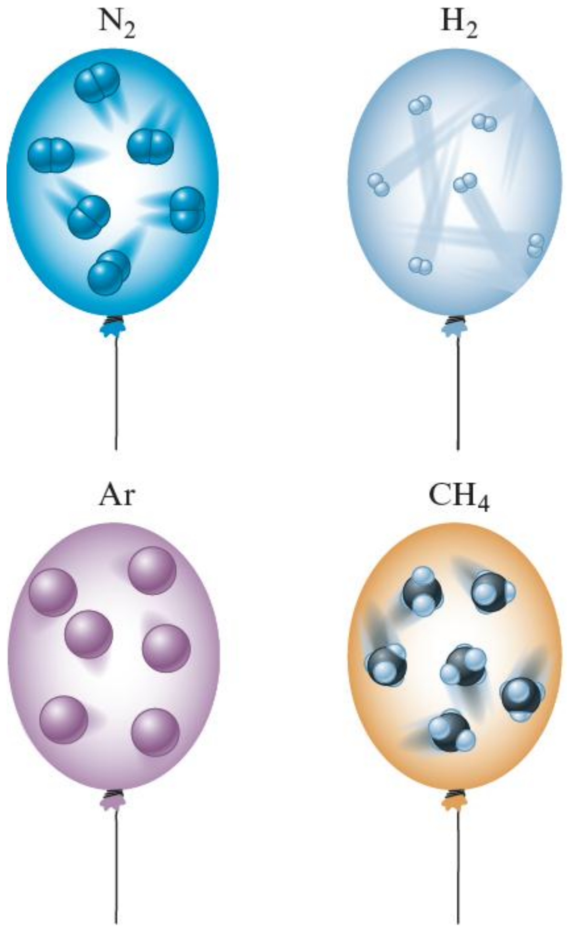
Avogadro's Law

- For a gas at constant pressure and temperature, the volume is directly proportional to the number of moles of gas

$$V = an$$

  - *V* - Volume of gas
  - *n* - Number of moles of gas particles
  - *a* - Proportionality constant

Figure 5.10 - Representation of Avogadro’s Law



These balloons each hold 1.0 L gas at 25°C and 1 atm, and each balloon contains 0.041 mole of gas, or 2.5×10^{22} molecules

Interactive Example 5.5 - Avogadro's Law

- Suppose we have a 12.2-L sample containing 0.50 mole of oxygen gas (O_2) at a pressure of 1 atm and a temperature of 25° C
  - If all this O_2 were converted to ozone (O_3) at the same temperature and pressure, what would be the volume of the ozone?

Interactive Example 5.5 - Solution

- Where are we going?
  - To calculate the volume of the ozone produced by 0.50 mole of oxygen
- What do we know?
  - $n_1 = 0.50 \text{ mol } O_2$
  - $n_2 = ? \text{ mol } O_3$
  - $V_1 = 12.2 \text{ L } O_2$
  - $V_2 = ? \text{ L } O_3$

Interactive Example 5.5 - Solution (Continued 1)

- What information do we need?
  - Balanced equation
  - Moles of O_3
  - Avogadro's law: $V = an$
- How do we get there?
  - How many moles of O_3 are produced by 0.50 mole of O_2?

Interactive Example 5.5 - Solution (Continued 2)

- What is the balanced equation?

$$3O_2(g) \rightarrow 2O_3(g)$$

- What is the mole ratio between O_3 and O_2?

$$\frac{2 \text{ mol } O_3}{3 \text{ mol } O_2}$$

- Now we can calculate the moles of O_3 formed

$$0.50 \cancel{\text{ mol } O_2} \times \frac{2 \text{ mol } O_3}{3 \cancel{\text{ mol } O_2}} = 0.33 \text{ mol } O_3$$

Interactive Example 5.5 - Solution (Continued 3)

- What is the volume of O_3 produced?
  - Avogadro's law states that $V = an$, which can be rearranged to give

$$\frac{V}{n} = a$$

  - Since a is a constant, an alternative representation is

$$\frac{V_1}{n_1} = a = \frac{V_2}{n_2}$$

  - V_1 is the volume of n_1 moles of O_2 gas and V_2 is the volume of n_2 moles of O_3 gas

Interactive Example 5.5 - Solution (Continued 4)

- In this case we have
  - n_1 = 0.50 mol n_2 = 0.33 mol
  - V_1 = 12.2 L V_2 = ?
- Solving for V_2 gives

$$V_2 = \left(\frac{n_2}{n_1}\right)V_1 = \left(\frac{0.33\ \cancel{\text{mol}}}{0.50\ \cancel{\text{mol}}}\right)12.2\text{ L} = 8.1\text{ L}$$

- Reality check - Note that the volume decreases, as it should, since fewer moles of gas molecules will be present after O_2 is converted to O_3

Exercise

- An 11.2-L sample of gas is determined to contain 0.50 mole of N_2
  - At the same temperature and pressure, how many moles of gas would there be in a 20-L sample?

0.89 mol

Ideal Gas Law

- The gas laws of Boyle, Charles, and Avogadro can be combined to give the ideal gas law

$$V = R\left(\frac{Tn}{P}\right)$$

  - **Universal gas constant**: Combined proportionality constant (R)
    - R = 0.08206 L · atm/K · mol when pressure is expressed in atmospheres and the volume in liters

Ideal Gas Law (Continued)

- Can be rearranged to:

$$PV = nRT$$

- Equation of state for a gas, where the state of the gas is its condition at a given time
  - Based on experimental measurements of the properties of gases
  - Any gas that obeys this law is said to be behaving ideally
    - An ideal gas is a hypothetical substance

Gas Law Problems

- Types
  - Boyle's law problems
  - Charles's law problems
  - Avogadro's law problems
- Ideal gas law can be applied to any problem
  - Place the variables that change on one side of the equal sign and the constants on the other side

Interactive Example 5.7 - Ideal Gas Law II

- Suppose we have a sample of ammonia gas with a volume of 7.0 mL at a pressure of 1.68 atm
  - The gas is compressed to a volume of 2.7 mL at a constant temperature
  - Use the ideal gas law to calculate the final pressure

Interactive Example 5.7 - Solution

- Where are we going?
  - To use the ideal gas equation to determine the final pressure
- What do we know?
  - P_1 = 1.68 atm
  - P_2 = ?
  - V_1 = 7.0 mL
  - V_2 = 2.7 mL

Interactive Example 5.7 - Solution (Continued 1)

- What information do we need?
  - Ideal gas law

$$PV = nRT$$

  - R = 0.08206 L · atm/K · mol
- How do we get there?
  - What are the variables that change?

$$P, V$$

Interactive Example 5.7 - Solution (Continued 2)

- What are the variables that remain constant?

$$n, R, T$$

- Write the ideal gas law, collecting the change variables on one side of the equal sign and the variables that do not change on the other

$$\underbrace{PV}_{\text{Change}} = \underbrace{nRT}_{\text{Remain constant}}$$

Interactive Example 5.7 - Solution (Continued 3)

- Since n and T remain the same in this case, we can write $P_1V_1 = nRT$ and $P_2V_2 = nRT$, and combining these gives

$$P_1V_1 = nRT = P_2V_2 \quad \text{or} \quad P_1V_1 = P_2V_2$$

  - $P_1 = 1.68$ atm $\quad$ $V_1 = 7.0$ mL $\quad$ $V_2 = 2.7$ mL
- Solving for P_2 gives

$$P_2 = \left(\frac{V_1}{V_2}\right)P_1 = \left(\frac{7.0\ \cancel{\text{mL}}}{2.7\ \cancel{\text{mL}}}\right)1.68\ \text{atm} = 4.4\ \text{atm}$$

Interactive Example 5.7 - Solution (Continued 4)

- Reality check - The volume decreased (at constant temperature), so the pressure should increase
  - Note that the calculated final pressure is 4.4 atm
  - Most gases do not behave ideally above 1 atm
    - If the pressure of this gas sample was measured, the observed pressure would differ slightly from 4.4 atm

Interactive Example 5.9 - Ideal Gas Law IV

- A sample of diborane gas (B_2H_6), a substance that bursts into flame when exposed to air, has a pressure of 345 torr at a temperature of –15°C and a volume of 3.48 L
  - If conditions are changed so that the temperature is 36°C and the pressure is 468 torr, what will be the volume of the sample?

Interactive Example 5.9 - Solution

- Where are we going?
  - To use the ideal gas equation to determine the final volume
- What do we know?
  - $T_1 = 15^\circ\text{ C} + 273 = 258\text{ K}$ $\quad$ $T_2 = 36^\circ\text{ C} + 273 = 309\text{ K}$
  - $V_1 = 3.48\text{ L}$ $\quad$ $V_2 = ?$
  - $P_1 = 345\text{ torr}$ $\quad$ $P_2 = 468\text{ torr}$

Interactive Example 5.9 - Solution (Continued 1)

- What information do we need?
  - Ideal gas law

$$PV = nRT$$

  - R = 0.08206 L · atm/K · mol
- How do we get there?
  - What are the variables that change?

$$P, V, T$$

Interactive Example 5.9 - Solution (Continued 2)

- What are the variables that remain constant?

$$n, R$$

- Write the ideal gas law, collecting the change variables on one side of the equal sign and the variables that do not change on the other

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

Interactive Example 5.9 - Solution (Continued 3)

$$\frac{P_1V_1}{T_1} = nR = \frac{P_2V_2}{T_2} \quad \text{or} \quad \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

- Solving for V_2

$$V_2 = \frac{T_2P_1V_1}{T_1P_2} = \frac{(309\ \cancel{\text{K}})(345\ \cancel{\text{torr}})(3.48\ \text{L})}{(258\ \cancel{\text{K}})(468\ \cancel{\text{torr}})}$$

$$V_2 = 3.07\ \text{L}$$

Interactive Example 5.10 - Ideal Gas Law V

- A sample containing 0.35 mole of argon gas at a temperature of 13°C and a pressure of 568 torr is heated to 56°C and a pressure of 897 torr
  - Calculate the change in volume that occurs

Interactive Example 5.10 - Solution

- Where are we going?
  - To use the ideal gas equation to determine the final volume
- What do we know?

State 1	State 2
$n_1 = 0.35$ mol	$n_2 = 0.35$ mol
$P_1 = 568 \cancel{\text{torr}} \times \dfrac{1 \text{ atm}}{760 \cancel{\text{torr}}} = 0.747$ atm	$P_2 = 897 \cancel{\text{torr}} \times \dfrac{1 \text{ atm}}{760 \cancel{\text{torr}}} = 1.18$ atm
$T_1 = 13°\text{C} + 273 = 286$ K	$T_2 = 56°\text{C} + 273 = 329$ K

Interactive Example 5.10 - Solution (Continued 1)

- What information do we need?
  - Ideal gas law

$$PV = nRT$$

  - R = 0.08206 L · atm/K · mol
  - V_1 and V_2

Interactive Example 5.10 - Solution (Continued 2)

- How do we get there?
  - What is V_1?

$$V_1 = \frac{n_1 R T_1}{P_1}$$

$$V_1 = \frac{(0.35\ \cancel{\text{mol}})(0.08206\ \text{L} \cdot \cancel{\text{atm}}/\cancel{\text{K}} \cdot \cancel{\text{mol}})(286\ \cancel{\text{K}})}{(0.747\ \cancel{\text{atm}})}$$

$$V_1 = 11\ \text{L}$$

Interactive Example 5.10 - Solution (Continued 3)

- What is V_2?

$$V_2 = \frac{n_2RT_2}{P_2}$$

$$V_2 = \frac{(0.35\ \cancel{\text{mol}})(0.08206\ \text{L} \cdot \cancel{\text{atm}}/\cancel{\text{K}} \cdot \cancel{\text{mol}})(329\ \cancel{\text{K}})}{(1.18\ \cancel{\text{atm}})}$$

$$V_2 = 8.0\ \text{L}$$

Interactive Example 5.10 - Solution (Continued 4)

- What is the change in volume Δ*V*?

$$\Delta V = V_2 - V_1 = 8.0 \text{ L} - 11 \text{ L} = -3 \text{ L}$$

- Conclusion
  - The change in volume is negative because the volume decreases

Molar Volume of an Ideal Gas

- **Molar volume**: For 1 mole of an ideal gas at 0° C and 1 atm, the volume of the gas is 22.42 L

$$V = \frac{nRT}{P} = \frac{(1.000\ \cancel{\text{mol}})(0.08206\ \text{L} \cdot \cancel{\text{atm}} / \cancel{\text{K}} \cdot \cancel{\text{mol}})(273.2\ \cancel{\text{K}})}{1.000\ \cancel{\text{atm}}}$$

$$= 22.42\ \text{L}$$

  - **Standard temperature and pressure (STP)**: Conditions 0° C and 1 atm

Critical Thinking

- What if STP was defined as normal room temperature (22° C) and 1 atm?
  - How would this affect the molar volume of an ideal gas?
    - Include an explanation and a number

Interactive Example 5.12 - Gas Stoichiometry II

- Quicklime (CaO) is produced by the thermal decomposition of calcium carbonate ($CaCO_3$)
  - Calculate the volume of CO_2 at STP produced from the decomposition of 152 g $CaCO_3$ by the reaction

$$\mathrm{CaCO_3}(s) \rightarrow \mathrm{CaO}(s) + \mathrm{CO_2}(g)$$

Interactive Example 5.12 - Solution

- Where are we going?
  - To use stoichiometry to determine the volume of CO_2 produced
- What do we know?
  - $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
- What information do we need?
  - Molar volume of a gas at STP is 22.42 L

Interactive Example 5.12 - Solution (Continued 1)

- How do we get there?
  - What is the balanced equation?

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

  - What are the moles of $CaCO_3$ (100.09 g/mol)?

$$152\ \cancel{\text{g CaCO}_3} \times \frac{1\ \text{mol CaCO}_3}{100.09\ \cancel{\text{g CaCO}_3}} = 1.52\ \text{mol CaCO}_3$$

Interactive Example 5.12 - Solution (Continued 2)

- What is the mole ratio between CO_2 and $CaCO_3$ in the balanced equation?

$$\frac{1 \text{ mol } CO_2}{1 \text{ mol } CaCO_3}$$

- What are the moles of CO_2?
  - 1.52 moles of CO_2, which is the same as the moles of $CaCO_3$ because the mole ratio is 1

Interactive Example 5.12 - Solution (Continued 3)

- What is the volume of CO_2 produced?
  - We can compute this by using the molar volume since the sample is at STP:

$$1.52\ \cancel{\text{mol CO}_2} \times \frac{22.42\ \text{L CO}_2}{1\ \cancel{\text{mol CO}_2}} = 34.1\ \text{L CO}_2$$

- Thus, the decomposition of 152 g $CaCO_3$ produces 34.1 L CO_2 at STP

Interactive Example 5.13 - Gas Stoichiometry III

- A sample of methane gas having a volume of 2.80 L at 25° C and 1.65 atm was mixed with a sample of oxygen gas having a volume of 35.0 L at 31° C and 1.25 atm, and the mixture was then ignited to form carbon dioxide and water
  - Calculate the volume of CO_2 formed at a pressure of 2.50 atm and a temperature of 125° C

Interactive Example 5.13 - Solution

- Where are we going?
  - To determine the volume of CO_2 produced
- What do we know?

	CH_4	O_2	CO_2
P	1.65 atm	1.25 atm	2.50 atm
V	2.80 L	35.0 L	?
T	25°C + 273 = 298 K	31°C + 273 = 304 K	125°C + 273 = 398 K

Interactive Example 5.13 - Solution (Continued 1)

- What information do we need?
  - Balanced chemical equation for the reaction
  - Ideal gas law

$$PV = nRT$$

  - R = 0.08206 L · atm/K · mol

Interactive Example 5.13 - Solution (Continued 2)

- How do we get there?
  - What is the balanced equation?
    - From the description of the reaction, the unbalanced equation is

$$CH_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$

    - This can be balanced to give

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

Interactive Example 5.13 - Solution (Continued 3)

- What is the limiting reactant?
  - We can determine this by using the ideal gas law to determine the moles for each reactant

$$n_{CH_4} = \frac{PV}{RT} = \frac{(1.65\ \cancel{\text{atm}})(2.80\ \cancel{\text{L}})}{(0.08206\ \cancel{\text{L}} \cdot \cancel{\text{atm}}/\cancel{\text{K}} \cdot \text{mol})(298\ \cancel{\text{K}})} = 0.189\ \text{mol}$$

$$n_{O_2} = \frac{PV}{RT} = \frac{(1.25\ \cancel{\text{atm}})(35.0\ \cancel{\text{L}})}{(0.08206\ \cancel{\text{L}} \cdot \cancel{\text{atm}}/\cancel{\text{K}} \cdot \text{mol})(304\ \cancel{\text{K}})} = 1.75\ \text{mol}$$

Interactive Example 5.13 - Solution (Continued 4)

- In the balanced equation for the combustion reaction, 1 mole of CH_4 requires 2 moles of O_2
- The moles of O_2 required by 0.189 mole of CH_4 can be calculated as follows:

$$0.189\ \cancel{\text{mol CH}_4} \times \frac{2\ \text{mol O}_2}{1\ \cancel{\text{mol CH}_4}} = 0.378\ \text{mol O}_2$$

- The limiting reactant is CH_4

Interactive Example 5.13 - Solution (Continued 5)

- What are the moles of CO_2?
  - Since CH_4 is limiting, we use the moles of CH_4 to determine the moles of CO_2 produced

$$0.189\ \cancel{\text{mol CH}_4} \times \frac{1\ \text{mol CO}_2}{1\ \cancel{\text{mol CH}_4}} = 0.189\ \text{mol CO}_2$$

- What is the volume of CO_2 produced?
  - Since the conditions stated are not STP, we must use the ideal gas law to calculate the volume

Interactive Example 5.13 - Solution (Continued 6)

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

- In this case, n = 0.189 mol, *T* = 125° C + 273 = 398 K, *P* = 2.50 atm, and *R* = 0.08206 L · atm/K · mol

$$V = \frac{(0.189\ \cancel{\text{mol}})(0.08206\ \text{L} \cdot \cancel{\text{atm}} / \cancel{\text{K}} \cdot \cancel{\text{mol}})(398\ \cancel{\text{K}})}{2.50\ \cancel{\text{atm}}} = 2.47\ \text{L}$$

  - This represents the volume of CO_2 produced under these conditions

Molar Mass of a Gas

- Ideal gas law is essential for the calculation of molar mass of a gas from its measured density

$$n = \frac{\text{grams of gas}}{\text{molar mass}} = \frac{\text{mass}}{\text{molar mass}} = \frac{m}{\text{molar mass}}$$

  - Substituting into the ideal gas equation gives:

$$P = \frac{nRT}{V} = \frac{(m/\text{molar mass})RT}{V} = \frac{m(RT)}{V(\text{molar mass})}$$

Molar Mass of a Gas (Continued)

- Density, *d*, is equal to *m*/*V*

$$P = \frac{dRT}{\text{molar mass}}$$

Or

$$\text{Molar mass} = \frac{dRT}{P}$$

Interactive Example 5.14 - Gas Density/Molar Mass

- The density of a gas was measured at 1.50 atm and 27° C and found to be 1.95 g/L
  - Calculate the molar mass of the gas

Interactive Example 5.14 - Solution

- Where are we going?
  - To determine the molar mass of the gas
- What do we know?
  - P = 1.50 atm
  - T = 27° C + 273 = 300 K
  - d = 1.95 g/L

Interactive Example 5.14 - Solution (Continued)

- What information do we need?
  - $\text{Molar mass} = \dfrac{dRT}{P}$
  - $R = 0.08206\ \text{L} \cdot \text{atm/K} \cdot \text{mol}$
- How do we get there?

$$\text{Molar mass} = \frac{dRT}{P} = \frac{\left(1.95\,\dfrac{\text{g}}{\cancel{\text{L}}}\right)\left(0.08206\,\dfrac{\cancel{\text{L}} \cdot \cancel{\text{atm}}}{\cancel{\text{K}} \cdot \text{mol}}\right)\left(300\ \cancel{\text{K}}\right)}{1.50\ \cancel{\text{atm}}} = 32.0\ \text{g/mol}$$

- Reality check - These are the units expected for molar mass

Law of Partial Pressures - John Dalton

- For a mixture of gases in a container, the total pressure exerted is the sum of the partial pressures

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

  - **Partial pressure**: Pressure that a gas would exert if it were alone in a container
    - Represented by symbols P_1, P_2, and P_3

Law of Partial Pressures - John Dalton (Continued 1)

- Assume that all gases behave ideally
  - Their partial pressures can be calculated from the ideal gas law

$$P_1 = \frac{n_1RT}{V}, \quad P_2 = \frac{n_2RT}{V}, \quad P_3 = \frac{n_3RT}{V}, \ \ldots$$

  - Total pressure of the mixture P_{TOTAL}

$$P_{\text{TOTAL}} = P_1 + P_2 + P_3 + \ldots = \frac{n_1RT}{V} + \frac{n_2RT}{V} + \frac{n_3RT}{V} + \ldots$$

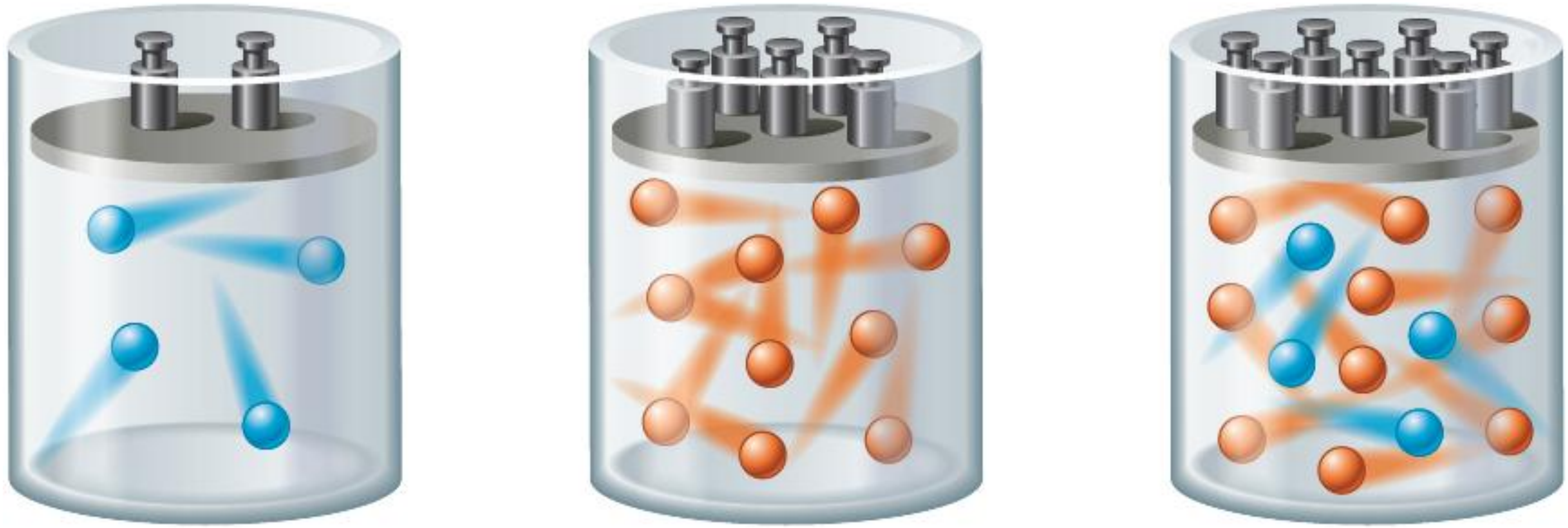
Law of Partial Pressures - John Dalton (Continued 2)

$$P_{\text{TOTAL}} = (n_1 + n_2 + n_3 + \ldots)\left(\frac{RT}{V}\right)$$

$$P_{\text{TOTAL}} = n_{\text{TOTAL}}\left(\frac{RT}{V}\right)$$

- n_{TOTAL} - Sum of the number of moles of the various gases

Figure 5.12 - Schematic Diagram of Dalton's Law of Partial Pressures



Characteristics of an Ideal Gas

- Since the pressure exerted by an ideal gas is unaffected by its identity:
  - The volume of individual gas particles must not be important
  - The forces among the particles must not be important

Interactive Example 5.15 - Dalton's Law I

- Mixtures of helium and oxygen can be used in scuba diving tanks to help prevent "the bends"
  - For a particular dive, 46 L He at 25° C and 1.0 atm and 12 L O_2 at 25° C and 1.0 atm were pumped into a tank with a volume of 5.0 L
    - Calculate the partial pressure of each gas and the total pressure in the tank at 25° C

Interactive Example 5.15 - Solution

- Where are we going?
  - To determine the partial pressure of each gas
  - To determine the total pressure in the tank at 25° C
- What do we know?

	He	O_2	Tank
P	1.00 atm	1.00 atm	? atm
V	46 L	12 L	5.0 L
T	25°C + 273 = 298 K	25°C + 273 = 298 K	25°C + 273 = 298 K

Interactive Example 5.15 - Solution (Continued 1)

- What information do we need?
  - Ideal gas law

$$PV = nRT$$

  - R = 0.08206 L · atm/K · mol
- How do we get there?
  - How many moles are present for each gas?

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

Interactive Example 5.15 - Solution (Continued 2)

$$n_{\mathrm{He}} = \frac{(1.0\ \cancel{\mathrm{atm}})(46\ \cancel{\mathrm{L}})}{(0.08206\ \cancel{\mathrm{L}} \cdot \cancel{\mathrm{atm}} / \cancel{\mathrm{K}} \cdot \mathrm{mol})(298\ \cancel{\mathrm{K}})} = 1.9\ \mathrm{mol}$$

$$n_{\mathrm{O_2}} = \frac{(1.0\ \cancel{\mathrm{atm}})(12\ \cancel{\mathrm{L}})}{(0.08206\ \cancel{\mathrm{L}} \cdot \cancel{\mathrm{atm}} / \cancel{\mathrm{K}} \cdot \mathrm{mol})(298\ \cancel{\mathrm{K}})} = 0.49\ \mathrm{mol}$$

Interactive Example 5.15 - Solution (Continued 3)

- What is the partial pressure for each gas in the tank?
  - The tank containing the mixture has a volume of 5.0 L, and the temperature is 25° C
  - We can use these data and the ideal gas law to calculate the partial pressure of each gas

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

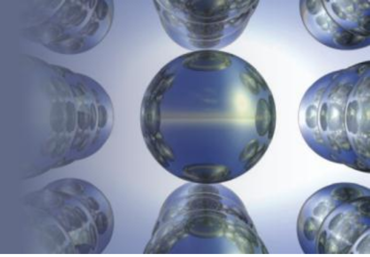
$$P_{He} = \frac{(1.9\ \cancel{\text{mol}})(0.08206\ \cancel{\text{L}} \cdot \text{atm}/\cancel{\text{K}} \cdot \cancel{\text{mol}})(298\ \cancel{\text{K}})}{5.0\ \cancel{\text{L}}} = 9.3\ \text{atm}$$

Interactive Example 5.15 - Solution (Continued 4)

$$P_{O_2} = \frac{(0.49\ \cancel{mol})(0.08206\ \cancel{L} \cdot atm/\cancel{K} \cdot \cancel{mol})(298\ \cancel{K})}{5.0\ \cancel{L}} = 2.4\ atm$$

- What is the total pressure of the mixture of gases in the tank?
  - The total pressure is the sum of the partial pressures

$$P_{TOTAL} = P_{He} + P_{O_2} = 9.3\ atm + 2.4\ atm = 11.7\ atm$$



Mole Fraction (χ)

- Ratio of number of moles of a given component in a mixture to the total number of moles in the mixture
- Example
  - For a given component in a mixture, χ_1 is calculated as follows:

$$\chi_1 = \frac{n_1}{n_{\text{TOTAL}}} = \frac{n_1}{n_1 + n_2 + n_3 + \ldots}$$

Mole Fraction (χ) in Terms of Pressure

- Number of moles of a gas is directly proportional to the pressure of the gas, since

$$n = P\left(\frac{V}{RT}\right)$$

- Representation of mole fraction in terms of pressure

$$\chi_1 = \frac{n_1}{n_{\text{TOTAL}}} = \frac{P_1\left(V/RT\right)}{P_1\left(V/RT\right) + P_2\left(V/RT\right) + P_3\left(V/RT\right) + \ldots}$$

Mole Fraction (χ) in Terms of Pressure (Continued)

$$\chi_1 = \frac{(V/RT)P_1}{(V/RT)(P_1 + P_2 + P_3 + \ldots)} = \frac{P_1}{P_1 + P_2 + P_3 + \ldots} = \frac{P_1}{P_{\text{TOTAL}}}$$

- Mole fraction of each component in a mixture of ideal gases is directly related to its partial pressure

$$\chi_2 = \frac{n_2}{n_{\text{TOTAL}}} = \frac{P_2}{P_{\text{TOTAL}}}$$

Interactive Example 5.16 - Dalton's Law II

- The partial pressure of oxygen was observed to be 156 torr in air with a total atmospheric pressure of 743 torr
  - Calculate the mole fraction of O_2 present

Interactive Example 5.16 - Solution

- Where are we going?
  - To determine the mole fraction of O_2
- What do we know?
  - P_{O_2} = 156 torr
  - P_{TOTAL} = 743 torr

Interactive Example 5.16 - Solution (Continued)

- How do we get there?
  - The mole fraction of O_2 can be calculated from the equation

$$\chi_{O_2} = \frac{P_{O_2}}{P_{TOTAL}} = \frac{156\ \cancel{\text{torr}}}{743\ \cancel{\text{torr}}} = 0.210$$

    - Note that the mole fraction has no units

Rearranging the Expression of Mole Fraction

$$\chi_1 = \frac{P_1}{P_{\mathrm{TOTAL}}}$$

- This can be rearranged to give

$$P_1 = \chi_1 + P_{\mathrm{TOTAL}}$$

- The partial pressure of a particular component of a gaseous mixture is the mole fraction of that component times the total pressure

Vapor Pressure of Water

- When the rate of escape of a gas in a container equals the rate of return:
  - Number of water molecules in the vapor state remain constant
  - Pressure of water vapor remains constant

Figure 5.13 - The Production of Oxygen by Thermal Decomposition of Potassium Chlorate

Interactive Example 5.18 - Gas Collection over Water

- A sample of solid potassium chlorate ($KClO_3$) was heated in a test tube and decomposed by the following reaction:

$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

  - The oxygen produced was collected by displacement of water at 22° C at a total pressure of 754 torr
  - The volume of the gas collected was 0.650 L, and the vapor pressure of water at 22° C is 21 torr

Interactive Example 5.18 - Gas Collection over Water
(Continued)

- Calculate the partial pressure of O_2 in the gas collected and the mass of $KClO_3$ in the sample that was decomposed

Interactive Example 5.18 - Solution

- Where are we going?
  - To determine the partial pressure of O_2 in the gas collected
  - Calculate the mass of $KClO_3$ in the original sample
- What do we know?

	Gas Collected	Water Vapor
P	754 torr	21 torr
V	0.650 L	
T	22°C + 273 = 295 K	22°C + 273 = 295 K

Interactive Example 5.18 - Solution (Continued 1)

- How do we get there?
  - What is the partial pressure of O_2?

$$P_{\text{TOTAL}} = P_{\text{O}_2} + P_{\text{H}_2\text{O}} = P_{\text{O}_2} + 21 \text{ torr} = 754 \text{ torr}$$

$$P_{\text{O}_2} = 754 \text{ torr} - 21 \text{ torr} = 733 \text{ torr}$$

  - What is the number of moles of O_2?
    - Now we use the ideal gas law to find the number of moles of O_2

$$n_{\text{O}_2} = \frac{P_{\text{O}_2} V}{RT}$$

Interactive Example 5.18 - Solution (Continued 2)

- In this case, the partial pressure of the O_2 is

$$P_{O_2} = 733 \text{ torr} = \frac{733 \cancel{\text{ torr}}}{760 \cancel{\text{ torr}}/\text{atm}} = 0.964 \text{ atm}$$

- To find the moles of O_2 produced, we use

$$V = 0.650 \text{ L}$$

$$T = 22^\circ \text{ C} + 273 = 295 \text{ K}$$

$$R = 0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol}$$

$$n_{O_2} = \frac{(0.964 \cancel{\text{ atm}})(0.650 \cancel{\text{ L}})}{(0.08206 \cancel{\text{ L}} \cdot \cancel{\text{atm}}/\cancel{\text{K}} \cdot \text{mol})(295 \cancel{\text{ K}})} = 2.59 \times 10^{-2} \text{ mol}$$

Interactive Example 5.18 - Solution (Continued 3)

- How many moles of $KClO_3$ are required to produce this amount of O_2?
  - What is the balanced equation?

$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

  - What is the mole ratio between $KClO_3$ and O_2 in the balanced equation?

$$\frac{2 \text{ mol } KClO_3}{3 \text{ mol } O_2}$$

Interactive Example 5.18 - Solution (Continued 4)

- What are the moles of $KClO_3$?

$$2.59 \times 10^{-2} \ \cancel{\text{mol O}_2} \times \frac{2 \text{ mol KClO}_3}{3 \ \cancel{\text{mol O}_2}} = 1.73 \times 10^{-2} \text{ mol KClO}_3$$

- What is the mass of $KClO_3$ (molar mass 122.6 g/mol) in the original sample?

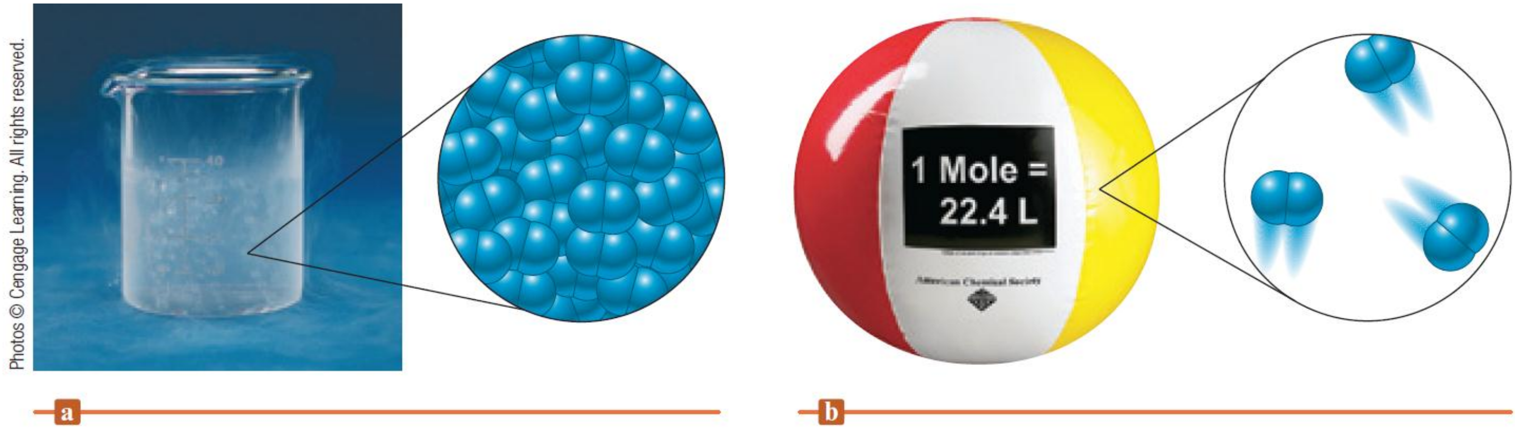
$$1.73 \times 10^{-2} \ \cancel{\text{mol KClO}_3} \times \frac{122.6 \text{ g KClO}_3}{\cancel{\text{mol KClO}_3}} = 2.12 \text{ g KClO}_3$$

The Kinetic Molecular Theory (KMT)

- Attempts to explain the properties of an ideal gas
  - Real gases do not conform to the postulates of the KMT
- Based on speculations about the behavior of individual gas particles

Postulates of the Kinetic Molecular Theory

1. Particles are so small compared with the distances between them that the volume of the individual particles can be assumed to be negligible (zero)



Postulates of the Kinetic Molecular Theory (Continued 1)

2. Particles are in constant motion
   - Collisions of the particles with the walls of the container are the cause of the pressure exerted by the gas
3. Particles are assumed to exert no forces on each other
   - Particles are assumed neither to attract nor to repel each other

Postulates of the Kinetic Molecular Theory (Continued 2)

4. Average kinetic energy of a collection of gas particles is assumed to be directly proportional to the Kelvin temperature of the gas

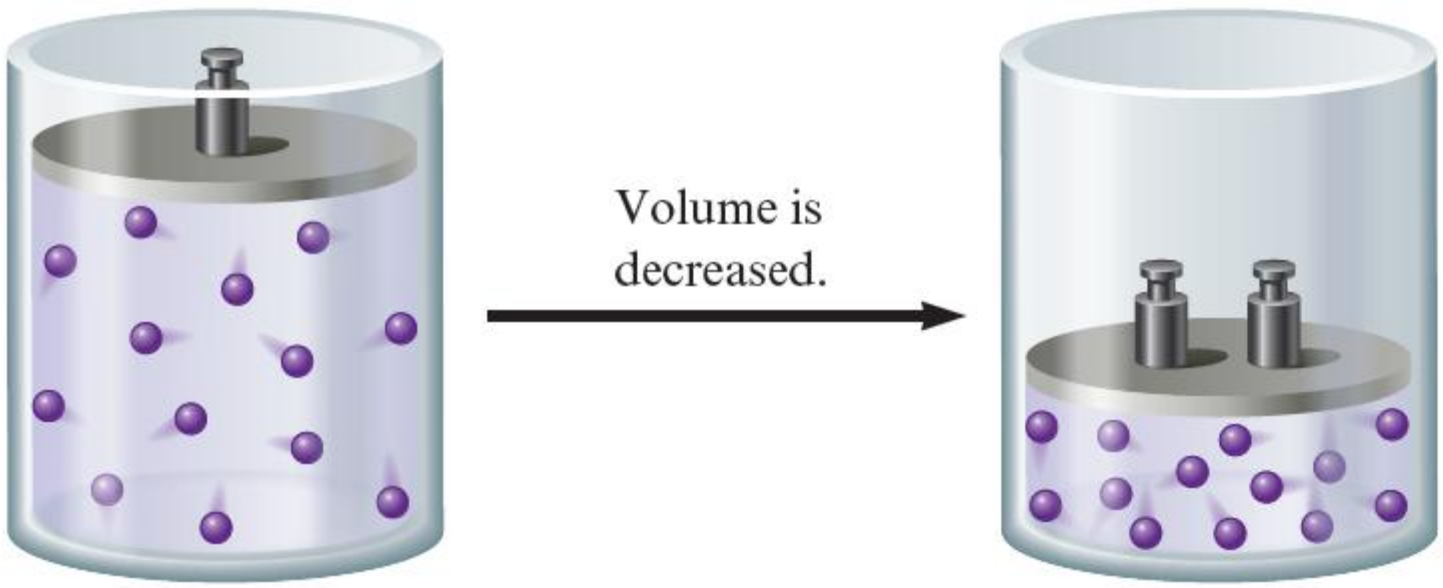
Pressure and Volume (Boyle's Law)

- For a given sample of a gas at any given temperature, if the volume decreases, the pressure increases

$$P = \underbrace{(nRT)}_{\text{Constant}}\frac{1}{V}$$

  - According to KMT, decrease in volume implies that the gas particles would hit the wall more often
    - Increases pressure

Figure 5.16 - Effects of Decreasing the Volume of a Sample of Gas at Constant Temperature



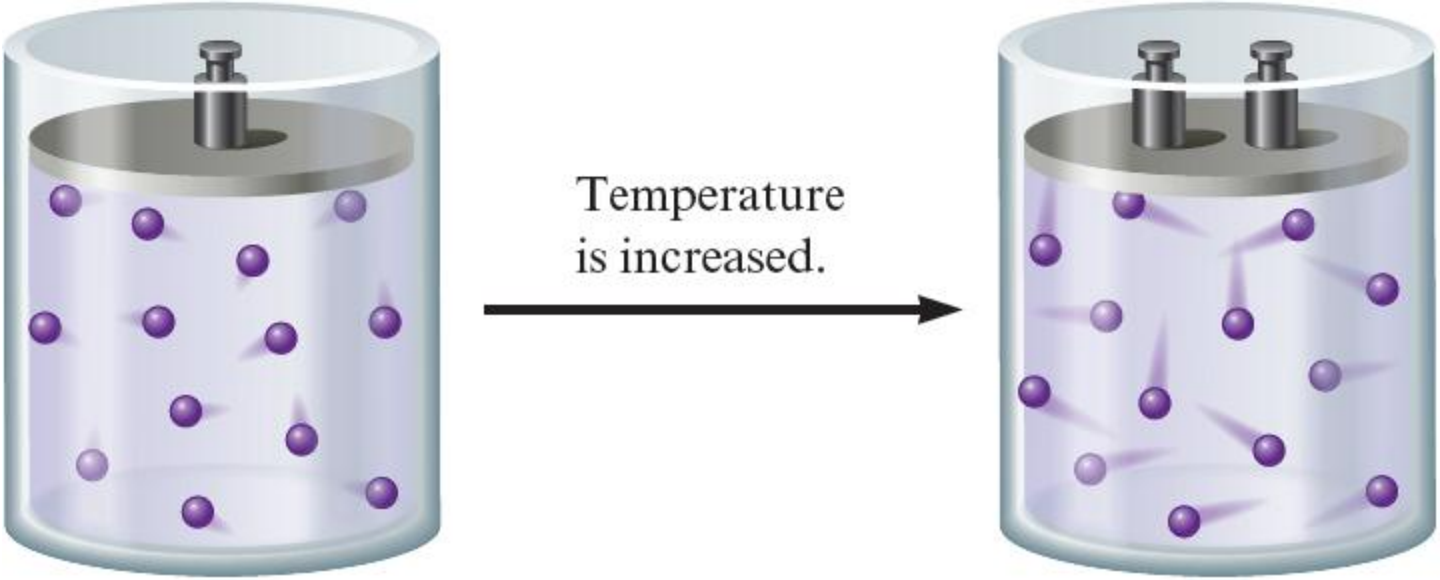
Pressure and Temperature

- According to the ideal gas law, for a given sample of an ideal gas at constant volume, the pressure is directly proportional to the temperature

$$P = \underbrace{\left(\frac{nR}{V}\right)}_{\text{Constant}} T$$

  - According to KMT, as the temperature increases, the speed of the gas particles increases, and they hit the wall with greater force resulting in increased pressure

Figure 5.16 - Effects of Increasing the Temperature of a Sample of Gas at Constant Volume



Critical Thinking

- You have learned the postulates of the KMT
  - What if we could not assume the third postulate to be true?
    - How would this affect the measured pressure of a gas?

Volume and Temperature (Charles's Law)

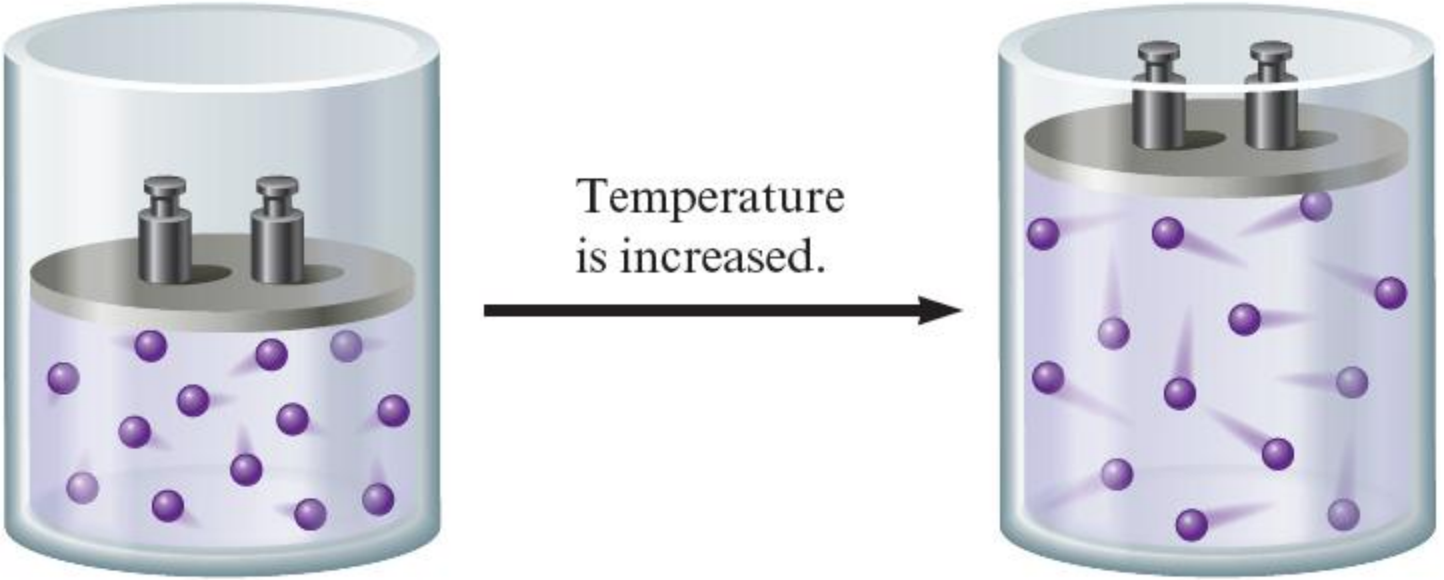
- According to the ideal gas law, for a sample of a gas at constant pressure, the volume of the gas is directly proportional to the temperature in Kelvins

$$V = \underbrace{\left(\frac{nR}{P}\right)}_{\text{Constant}} T$$

  - According to the KMT, at higher temperature, the speed of gas molecules increases
    - Hit the walls with more force

Volume and Temperature (Charles's Law) (Continued)

- Pressure can be kept constant only by increasing the volume of the container



Volume and Number of Moles (Avogadro's Law)

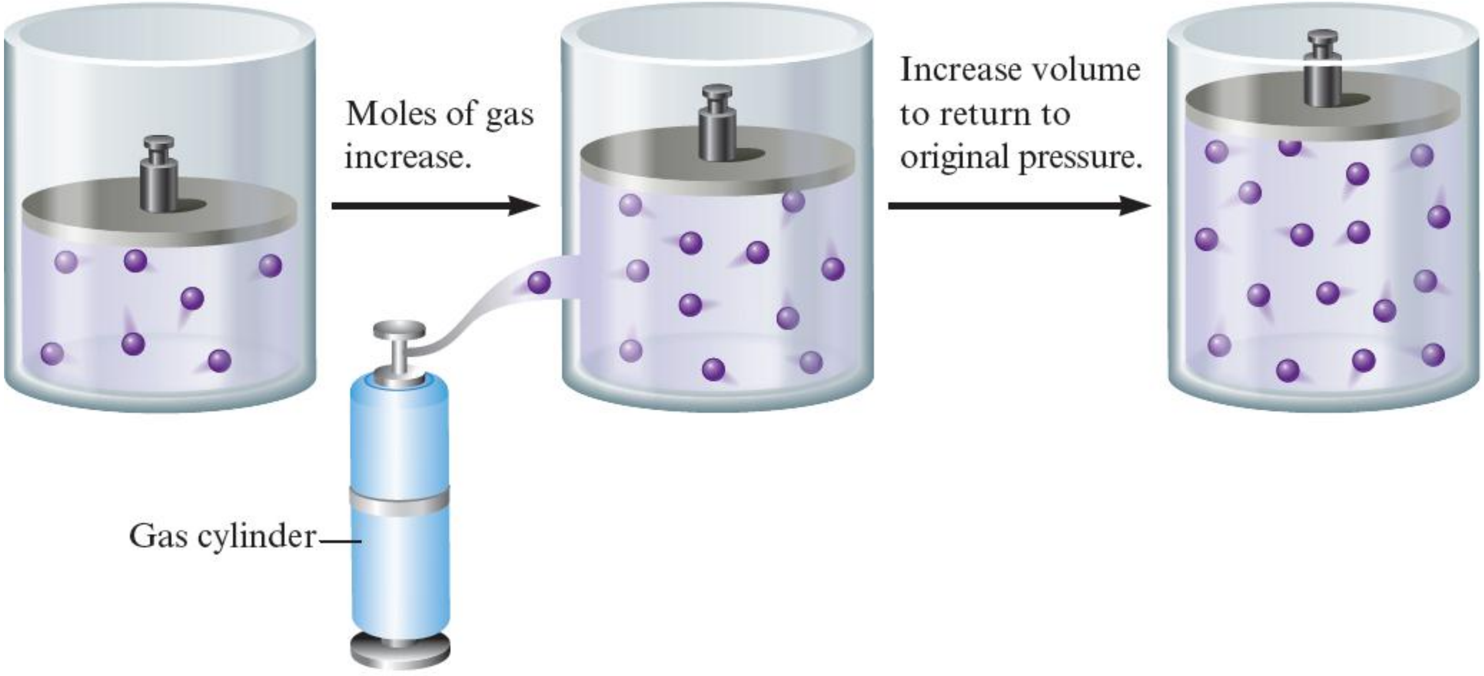
- According to the ideal gas law, the volume of a gas at constant temperature and pressure directly depends on the number of gas particles present

$$V = \underbrace{\left(\frac{RT}{P}\right)}_{\text{Constant}} n$$

  - According to the KMT, increase in number of gas particles at the same temperature would cause the pressure to increase if the volume were held constant

Volume and Number of Moles (Avogadro's Law)
(Continued)

- Pressure can return to its original value if volume is increased



Mixture of Gases (Dalton's Law)

- Total pressure exerted by a mixture of gases is the sum of the pressures of the individual gases
- The KMT assumes that:
  - All gas particles are independent of one another
  - Volume of individual particles are not important

Deriving the Ideal Gas Law

- Pressure can be expressed differently by applying the definitions of velocity, momentum, force, and pressure to the collection of particles in an ideal gas

$$P = \frac{2}{3}\left[\frac{nN_{\mathrm{A}}\left(\frac{1}{2}\overline{mu^2}\right)}{V}\right]$$

Deriving the Ideal Gas Law (Continued 1)

- Here,
  - P - Pressure of the gas
  - n - Number of moles of gas
  - N_A - Avogadro's number
  - m - Mass of each particle
  - $\overline{u^2}$ - Average of the square of the velocities of the particles
  - V - Volume of the container
  - $\frac{1}{2}m\overline{u^2}$ - Average kinetic energy of a gas particle

Deriving the Ideal Gas Law (Continued 2)

- Average kinetic energy for a mole of gas particles can be ascertained by multiplying the average kinetic energy of an individual particle by N_A

$$(\mathrm{KE})_{\mathrm{avg}} = N_{\mathrm{A}}\left(\frac{1}{2}m\overline{u^2}\right)$$

- The expression for pressure can be rewritten as:

$$P = \frac{2}{3}\left[\frac{n(\mathrm{KE})_{\mathrm{avg}}}{V}\right] \quad \text{or} \quad \frac{PV}{n} = \frac{2}{3}(\mathrm{KE})_{\mathrm{avg}}$$

Deriving the Ideal Gas Law (Continued 3)

- Since $(\mathrm{KE})_{\mathrm{avg}} \propto T$:

$$\frac{PV}{n} = \frac{2}{3}(\mathrm{KE})_{\mathrm{avg}} \propto T \quad \text{or} \quad \frac{PV}{n} \propto T$$

  - The agreement between the ideal gas law and the postulates of the KMT prove the validity of the KMT model

The Meaning of Temperature

- According to the KMT, the Kelvin temperature indicates the average kinetic energy of gas particles
  - This relationship between temperature and average kinetic energy can be obtained when the following equations are combined:

$$\frac{PV}{n} = RT = \frac{2}{3}(\mathrm{KE})_{\mathrm{avg}}$$

The Meaning of Temperature (Continued)

$$(\mathrm{KE})_{\mathrm{avg}} = \frac{3}{2}RT$$

- This shows that the Kelvin temperature is an index of the random motions of particles of a gas
  - As temperature increases, the motion of the particles becomes greater

Root Mean Square Velocity

- Refers to the square root of $\overline{u^2}$
- Symbolized by u_{rms}

$$u_{\text{rms}} = \sqrt{\overline{u^2}}$$

  - The expression for u_{rms} can also be attained from the following equations:

$$(\text{KE})_{\text{avg}} = N_{\text{A}}\left(\frac{1}{2}\overline{mu^2}\right) \text{ and } (\text{KE})_{\text{avg}} = \frac{3}{2}RT$$

Root Mean Square Velocity (Continued 1)

- Combination of the equations gives

$$N_{\mathrm{A}}\left(\frac{1}{2}\overline{mu^2}\right)=\frac{3}{2}RT \quad \text{or} \quad \overline{u^2}=\frac{3RT}{N_{\mathrm{A}}m}$$

- Taking square root on both sides, we get:

$$\sqrt{\overline{u^2}}=u_{\mathrm{rms}}=\sqrt{\frac{3RT}{N_{\mathrm{A}}m}}$$

  - m - Mass (kg) of a single gas particle
  - N_{A} - Number of particles in a mole

Root Mean Square Velocity (Continued 2)

- Substituting M for $N_A m$ in the equation for u_{rms}, we obtain:

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

- M - Mass of a mole of gas particles (kg)
- R = 8.3145 J/K·mol
  - J = **joule** = kg · m^2/s^2

Interactive Example 5.19 - Root Mean Square Velocity

- Calculate the root mean square velocity for the atoms in a sample of helium gas at 25° C

Interactive Example 5.19 - Solution

- Where are we going?
  - To determine the root mean square velocity for the atoms of He
- What do we know?
  - $T = 25^\circ\text{C} + 273 = 298\text{ K}$
  - $R = 8.3145\text{ J/K}\cdot\text{mol}$
- What information do we need?
  - Root mean square velocity is $u_{\text{rms}} = \sqrt{\dfrac{3RT}{M}}$

Interactive Example 5.19 - Solution (Continued 1)

- How do we get there?
  - What is the mass of a mole of He in kilograms?

$$M = 4.00\ \frac{\cancel{\text{g}}}{\text{mol}} \times \frac{1\ \text{kg}}{1000\ \cancel{\text{g}}} = 4.00 \times 10^{-3}\ \text{kg/mol}$$

  - What is the root mean square velocity for the atoms of He?

$$u_{\text{rms}} = \sqrt{\frac{3\left(8.3145\dfrac{\text{J}}{\cancel{\text{K}}\cdot\cancel{\text{mol}}}\right)\left(298\ \cancel{\text{K}}\right)}{4.00 \times 10^{-3}\ \dfrac{\text{kg}}{\cancel{\text{mol}}}}} = \sqrt{1.86 \times 10^{6}\ \frac{\text{J}}{\text{kg}}}$$

Interactive Example 5.19 - Solution (Continued 2)

- Since the units of J are kg · m²/s², this expression gives:

$$u_{\text{rms}} = \sqrt{1.86 \times 10^6 \frac{\cancel{\text{kg}} \cdot \text{m}^2}{\cancel{\text{kg}} \cdot \text{s}^2}} = 1.36 \times 10^3 \text{ m}/s$$

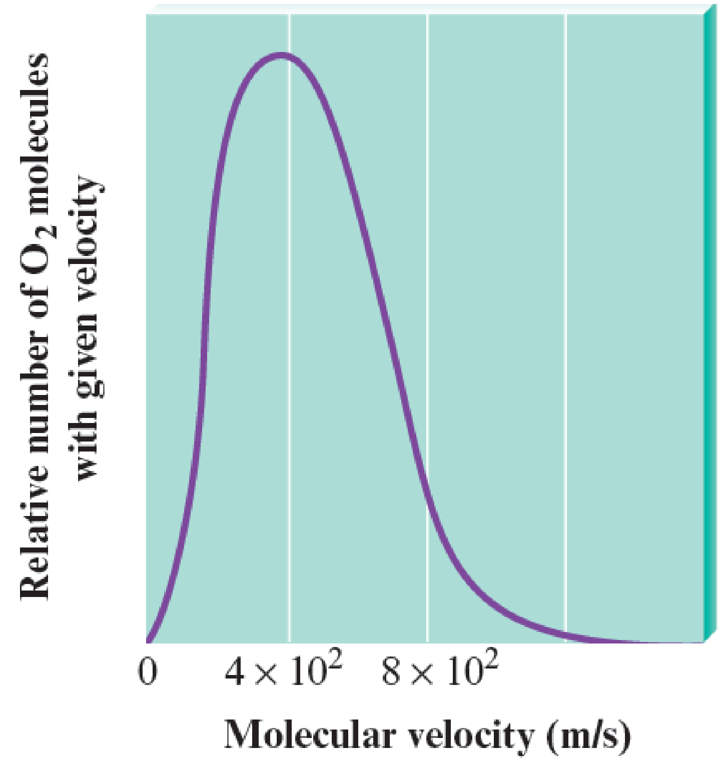
- Reality check - The resulting units are appropriate for velocity

Range of Velocities in Gas Particles

- Real gases have large number of collisions between particles
  - Path of a gas particle is erratic
- Mean free path
  - Average distance travelled by a particle between collisions in a gas sample
  - 1×10^{-7} m for O_2 at STP

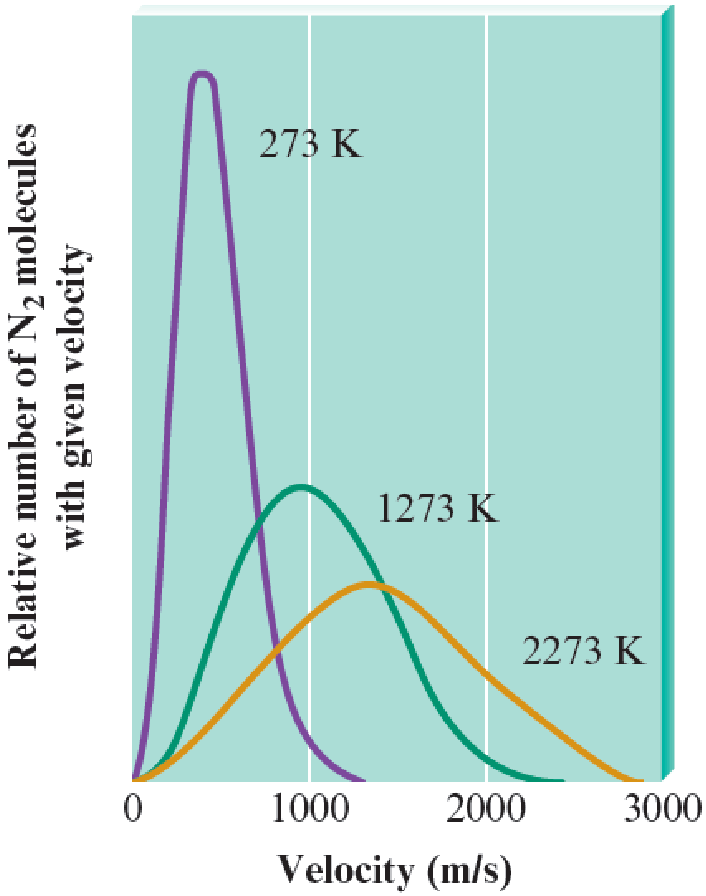
Effect of Collisions among Gas Particles

- When particles collide and exchange kinetic energy, a large range of velocities is produced
  - This plot depicts the relative number of O_2 molecules that have a given velocity at STP



Effect of Temperature on Velocity Distribution

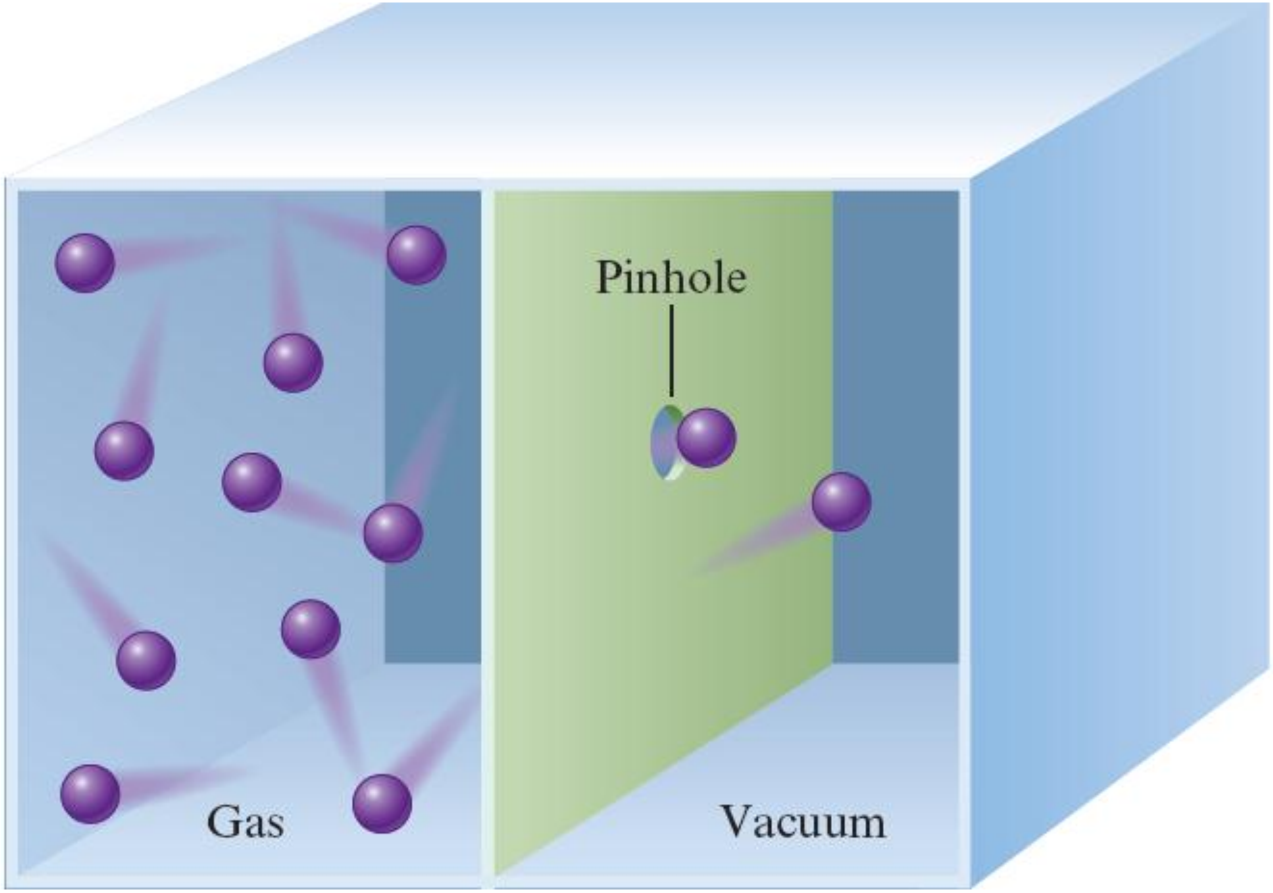
- As the temperature increases, the range of velocities becomes larger
  - Peak of the curve reflects the most probable velocity



Effusion and Diffusion - An Introduction

- **Diffusion**: Describes the mixing of gases
  - Rate of diffusion is the rate of mixing of gases
- **Effusion**: Describes the passage of a gas through a tiny orifice into an evacuated chamber
  - Rate of effusion measures the speed at which the gas is transferred into the chamber
    - Inversely proportional to the square root of the mass of the gas particles

Figure 5.22 - The Effusion of a Gas into an Evacuated Chamber



Graham's Law of Effusion

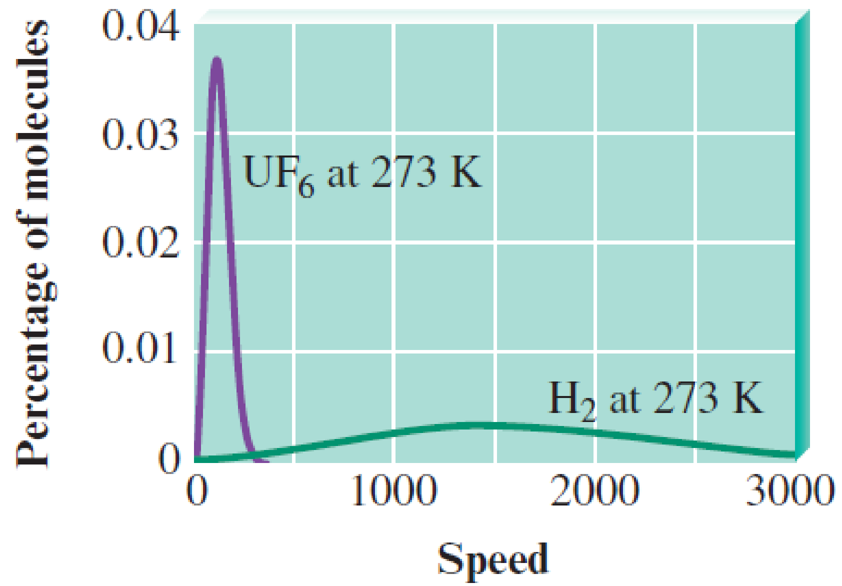
- Relative rates of effusion of two gases at the same *T* and *P* are given by the inverse ratio of the square roots of the masses of the gas particles

$$\frac{\text{Rate of effusion for gas 1}}{\text{Rate of effusion for gas 2}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

  - M_1 and M_2 - Molar masses of the gases

Interactive Example 5.20 - Effusion Rates

- Calculate the ratio of the effusion rates of hydrogen gas (H_2) and uranium hexafluoride (UF_6), a gas used in the enrichment process to produce fuel for nuclear reactors



Relative molecular speed distribution of H_2 and UF_6

Interactive Example 5.20 - Solution

- First we need to compute the molar masses
  - Molar mass of H_2 = 2.016 g/mol, and molar mass of UF_6 = 352.02 g/mol
    - Using Graham's law, we have:

$$\frac{\text{Rate of effusion for H}_2}{\text{Rate of effusion for UF}_6} = \frac{\sqrt{M_{\text{UF}_6}}}{\sqrt{M_{\text{H}_2}}} = \sqrt{\frac{352.02}{2.016}} = 13.2$$

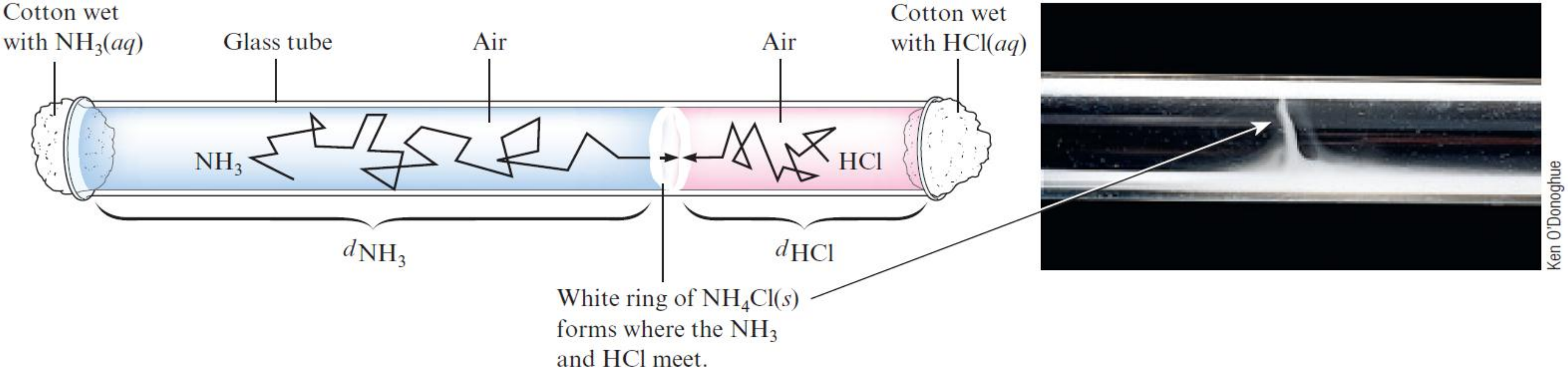
- The effusion rate of the very light H_2 molecules is about 13 times that of the massive UF_6 molecules

Prediction of the Relative Effusion Rates of Gases by the KMT

- The kinetic molecular model fits the experimental results for the effusion of gases
- Prediction for two gases at the same pressure and temperature

$$\frac{\text{Effusion rate for gas 1}}{\text{Effusion rate for gas 2}} = \frac{u_{\text{rms}} \text{ for gas 1}}{u_{\text{rms}} \text{ for gas 2}} = \frac{\sqrt{\dfrac{3RT}{M_1}}}{\sqrt{\dfrac{3RT}{M_2}}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

Figure 5.24 - Demonstration of the Relative Diffusion Rates of NH_3 and HCl Molecules



Ideal Gas Behavior

- Exhibited by real gases under certain conditions of:
  - Low pressure
  - High temperature

Figure 5.25 - Plots of *PV/nRT* versus *P* for Several Gases (200 K)

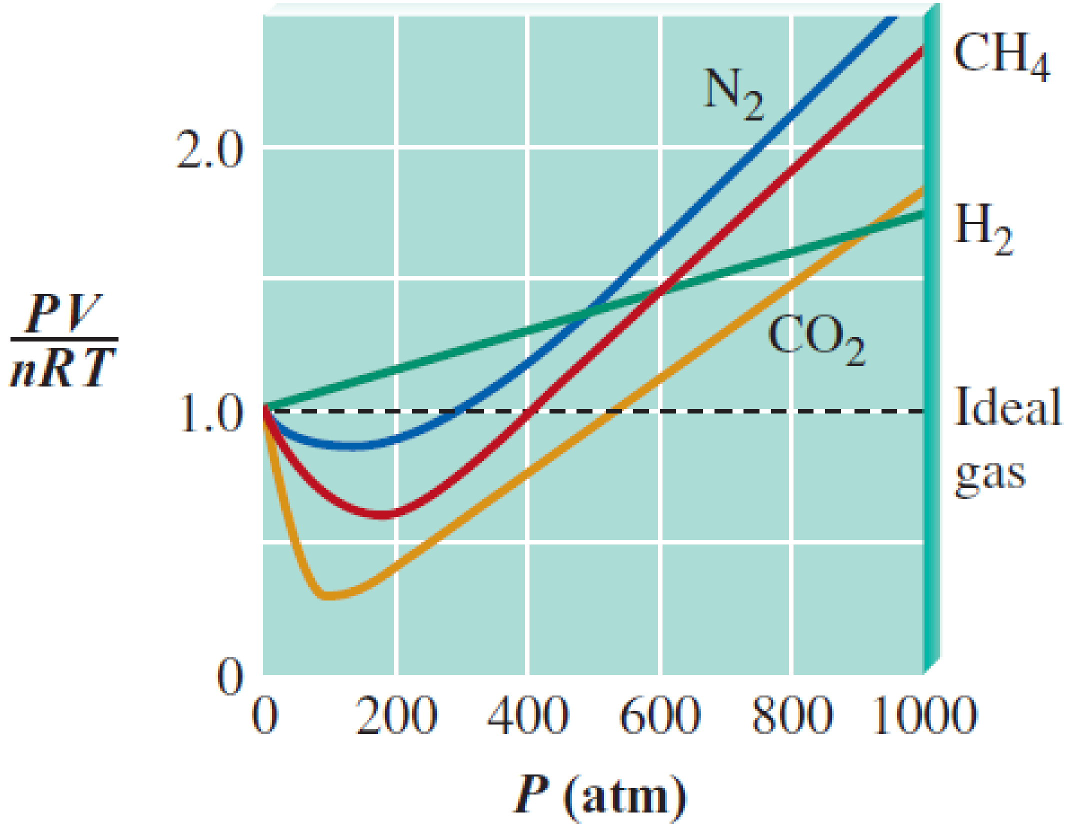
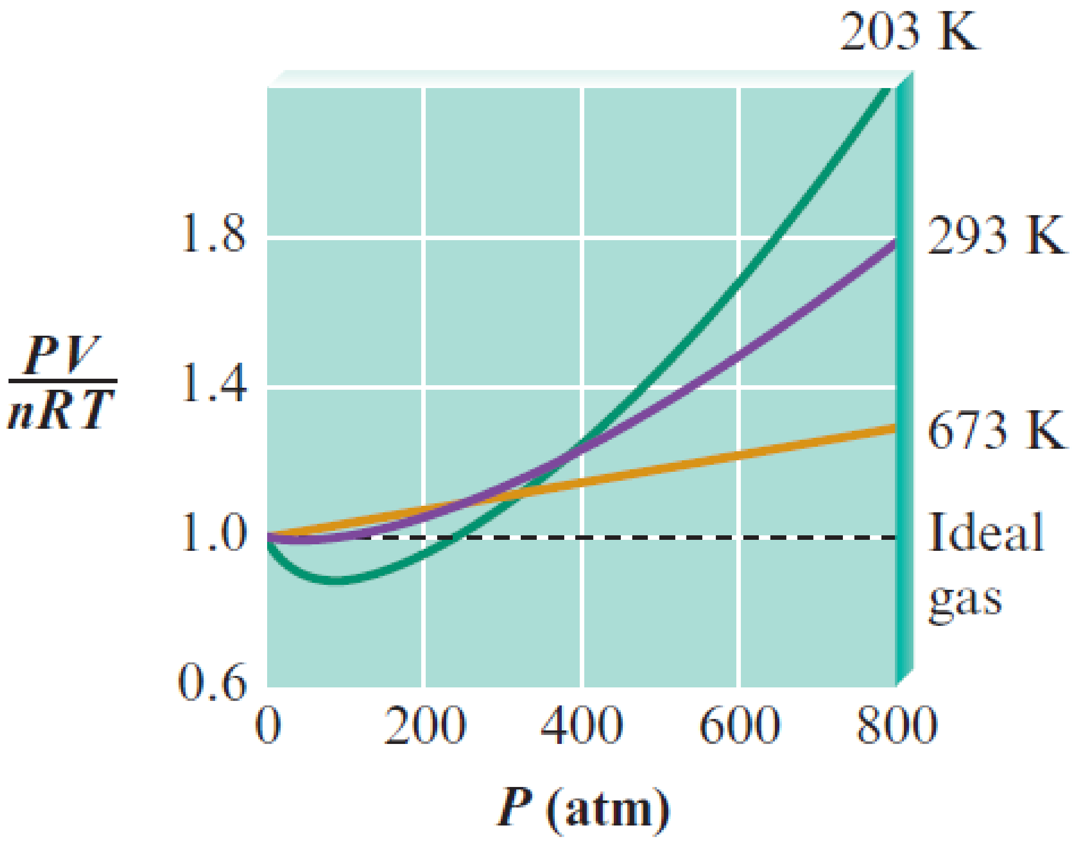


Figure 5.26 - Plots of *PV/nRT* versus *P* for Nitrogen Gas at Three Temperatures



van der Waals Equation

- Actual volume available to a given gas molecule can be calculated as follows:

$$V - nb$$

  - *V* - Volume of the container
  - *nb* - Correction factor for the volume of the molecules
    - *n* - Number of moles of gas
    - *b* - Empirical constant

van der Waals Equation (Continued 1)

- The ideal gas equation can be modified as follows:

$$P' = \frac{nRT}{V - nb}$$

- When gas particles come close together, attractive forces occur
    - Cause the particles to hit the walls very slightly and less often

van der Waals Equation (Continued 2)

- Size of the correction factor depends on the concentration of gas molecules
  - Higher the concentration, the more likely that the particles will attract each other

$$P_{obs} = P' - a\left(\frac{n}{V}\right)^2$$

    - *a* - Proportionality constant whose value can be determined by observing the actual behavior of the gas

van der Waals Equation (Continued 3)

$$\underbrace{\underbrace{\left[P_{\text{obs}} + a\left(\frac{n}{V}\right)^2\right]}_{\text{Corrected pressure}}}_{P_{\text{ideal}}} \times \underbrace{\underbrace{(V - nb)}_{\text{Corrected volume}}}_{V_{\text{ideal}}} = nRT$$

- Values of *a* are *b* vary until the best fit for the observed pressure is obtained

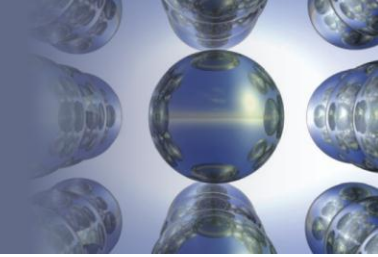


Table 5.3 - Values of the van der Waals Constants for some Common Gases

Gas	$a\left(\frac{\text{atm} \cdot \text{L}^2}{\text{mol}^2}\right)$	$b\left(\frac{\text{L}}{\text{mol}}\right)$
He	0.0341	0.0237
Ne	0.211	0.0171
Ar	1.35	0.0322
Kr	2.32	0.0398
Xe	4.19	0.0511
H_2	0.244	0.0266
N_2	1.39	0.0391
O_2	1.36	0.0318
Cl_2	6.49	0.0562
CO_2	3.59	0.0427
CH_4	2.25	0.0428
NH_3	4.17	0.0371
H_2O	5.46	0.0305

Critical Thinking

- You have learned that no gases behave perfectly ideally, but under conditions of high temperature and low pressure (high volume), gases behave more ideally
  - What if all gases always behaved perfectly ideally?
    - How would the world be different?

Behavior of Real Gases - Conclusions

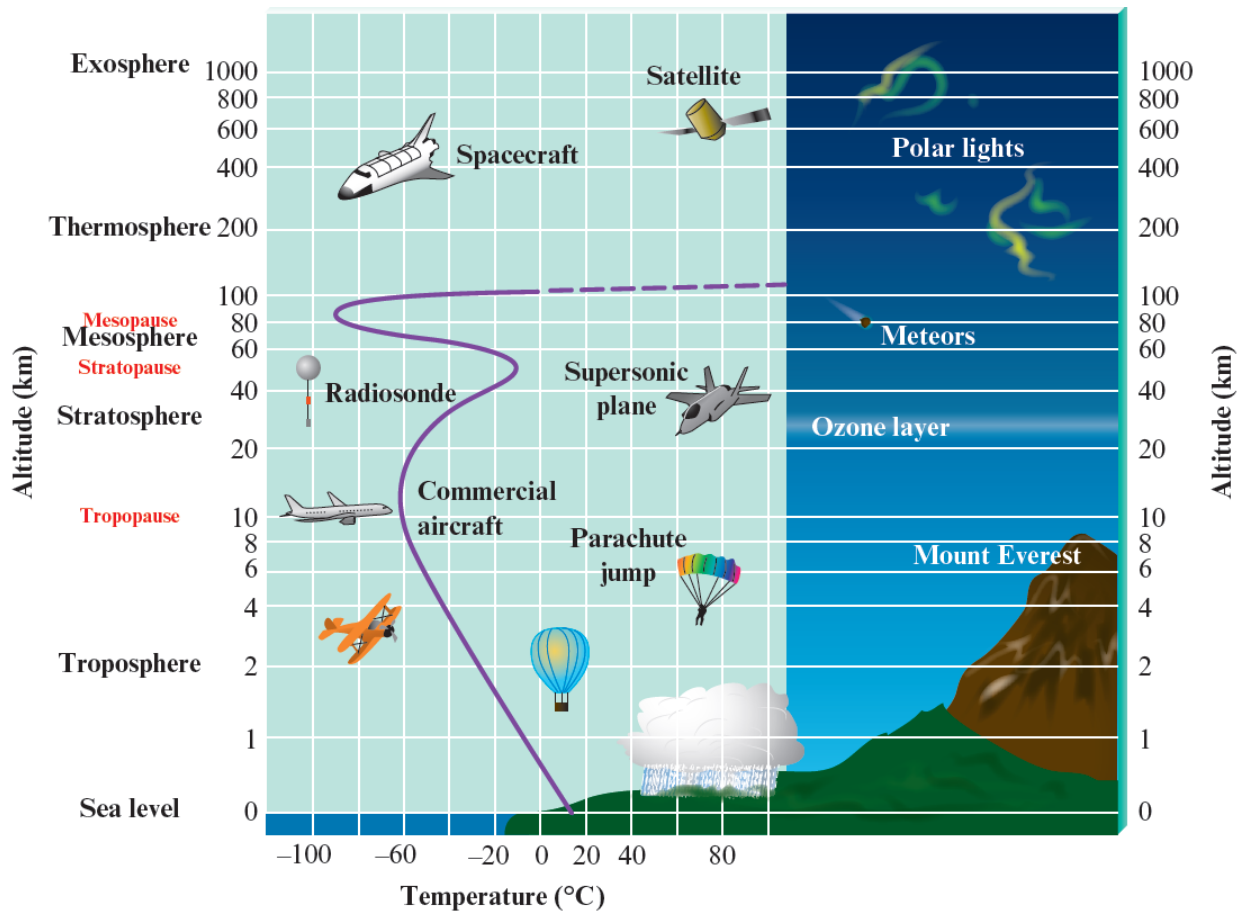
- For a real gas, the actual observed pressure is lower than the pressure expected for an ideal gas
  - Caused due to intermolecular attractions that occur in real gases, which increase in the following order:

$$H_2 < N_2 < CH_4 < CO_2$$

The Atmosphere

- Surrounds the earth's surface
- Contains essential gases such as N_2, O_2, H_2O, and CO_2
- Composition is not constant due to gravitational effects
  - Heavy molecules stay closer to the earth’s surface
  - Light molecules migrate to higher altitudes

Figure 5.30 - The Variation of Temperature with Altitude



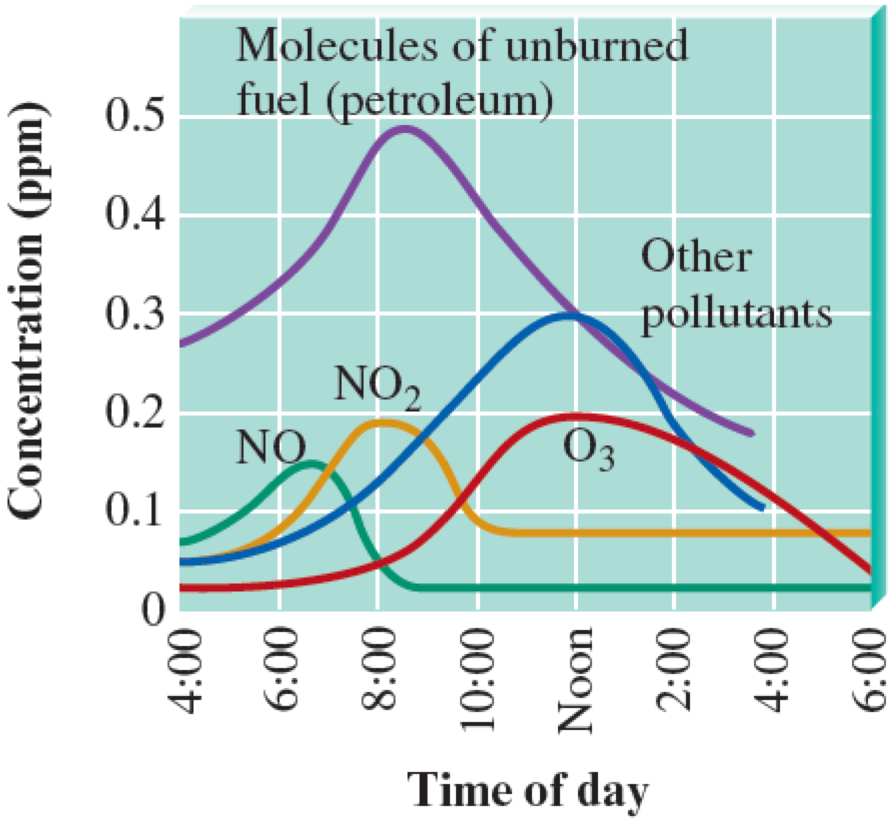
Chemistry in the Atmosphere - An Introduction

- Higher levels of atmosphere
  - Chemistry is affected by high-energy radiation and particles from the sun and other sources in space
  - Ozone - Prevents ultraviolet radiation from reaching the earth
- Troposphere - Layer that is closest to the earth's surface
  - Chemistry is highly affected by human activities

Air Pollution

- Sources
  - Transportation
  - Production of electricity
- Combustion of petroleum produces CO, CO_2, NO, and NO_2
  - When the mixture is trapped close to the ground in stagnant air, reactions occur to produce chemicals that harm living systems

Figure 5.31 - Concentration for Some Smog Components versus Time of Day



Pollution Due to Transportation

- Caused by the presence of nitrogen oxides in the air
  - Leads to the production of **photochemical smog**

$$NO_2(g) \rightarrow NO(g) + O(g)$$

$$O(g) + O_2(g) \rightarrow O_3(g)$$

$$NO(g) + \frac{1}{2}O_2(g) \rightarrow NO_2(g)$$

$$\text{Net reaction: } \frac{3}{2}O_2(g) \rightarrow O_3(g)$$

Pollution Due to the Production of Electricity

- Caused due to the presence of sulfur in coal, which, when burned, produces SO_2

$$\mathrm{S}(\text{in coal}) + \mathrm{O_2}(g) \rightarrow \mathrm{SO_2}(g)$$

  - After further oxidation, SO_2 is converted to SO_3 in air

$$2\mathrm{SO_2}(g) + \mathrm{O_2}(g) \rightarrow 2\mathrm{SO_3}(g)$$

Pollution due to the Production of Electricity (Continued)

- SO_3 can combine with water droplets to form sulfuric acid

$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$$

- Sulfuric acid is corrosive to living beings and building materials
  - Can lead to **acid rains**

Sulfur Dioxide Pollution

- Complicated due to the energy crisis
  - Lower petroleum supplies would mean a shift to the usage of high-sulfur coal
- High-sulfur coal can be used without harming the air quality by removing the SO_2 from the exhaust gas by a system called the scrubber
  - Involves the decomposition of $CaCO_3$ to lime and carbon dioxide

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

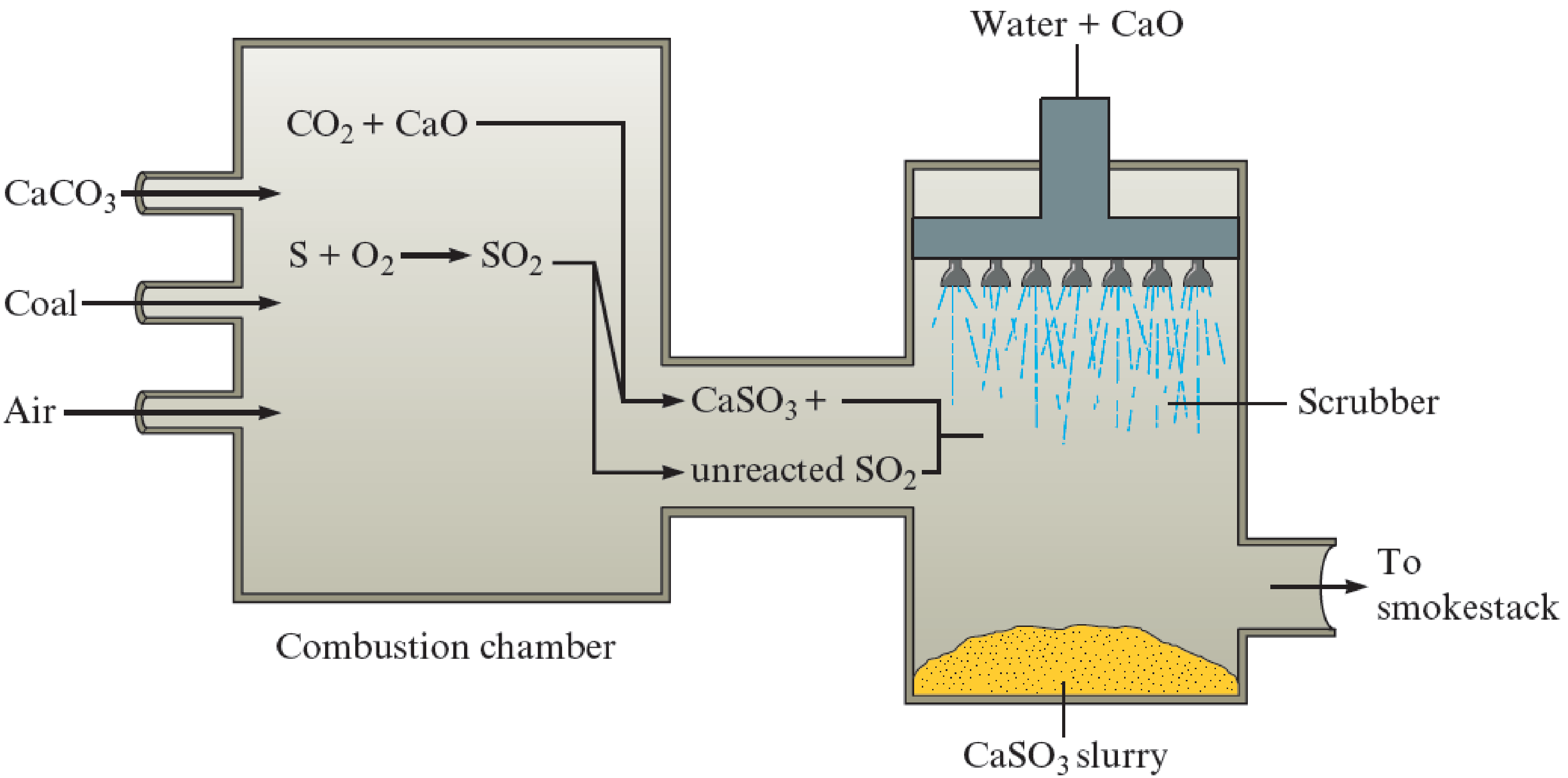
Sulfur Dioxide Pollution (Continued)

- Lime combines with the sulfur dioxide to form calcium sulfite

$$CaO(s) + SO_2(g) \rightarrow CaSO_3(s)$$

- An aqueous suspension of lime is injected into the exhaust gases to produce a slurry
  - Helps remove calcium sulfite and any remaining unreacted SO_2

Figure 5.33 - A Schematic Diagram of a Scrubber



Drawbacks of Scrubbing

- Complicated and expensive
- Consumes huge amounts of energy
- Calcium sulfite that is produced is buried in landfills
  - No use has been found for calcium sulfite