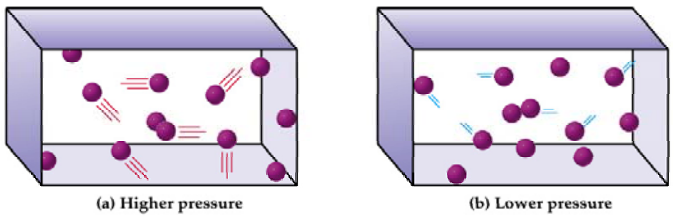
Chapter 10 Gases

Forestville Central School

Properties of Gases

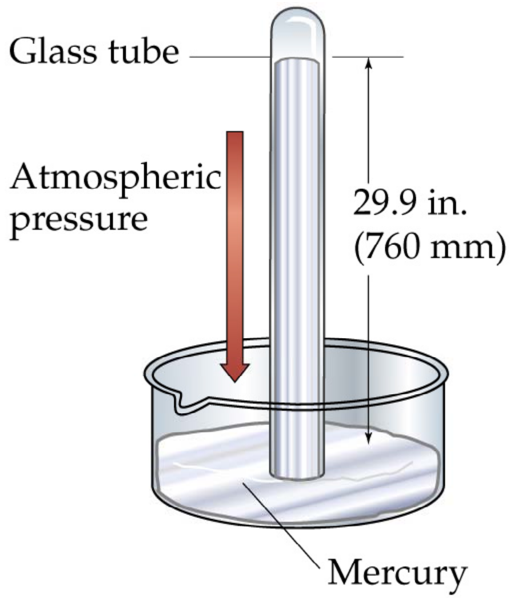
- Properties of Gases:
  1. Gases have an indefinite shape.
  2. Gases can expand.
  3. Gases can compress.
  4. Gases have low densities.
  5. Gases diffuse uniformly throughout their containers to form homogeneous mixtures.

Atmospheric Pressure



- Gas pressure is the result of constantly moving molecules striking the inside surface of its container.
- Depends on:
  - Number of collisions.
  - Energy of the molecules.

Atmospheric Pressure

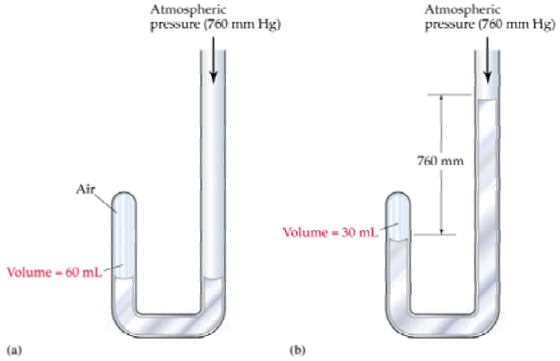


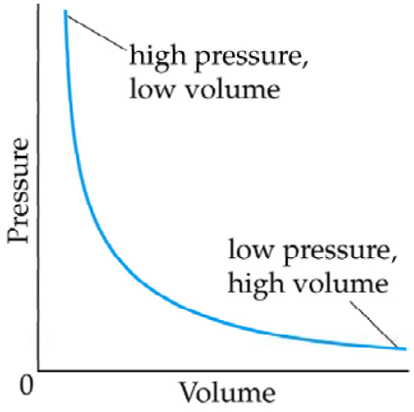
- Atmospheric pressure:
  - Result of air molecules striking various surfaces in the environment.
- 1 atm = 760 mm Hg
  = 760 Torr

Variables Affecting Gas Pressure

- In a gaseous system, there are three ways to change the pressure.
  - Increase or decrease the volume of the container.
  - Increase or decrease the temperature of the gas.
  - Increase or decrease the number of molecules in the container.

Boyle's Law

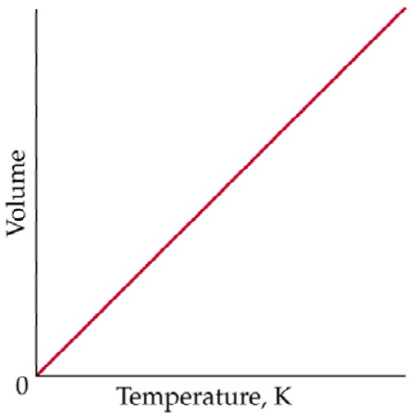




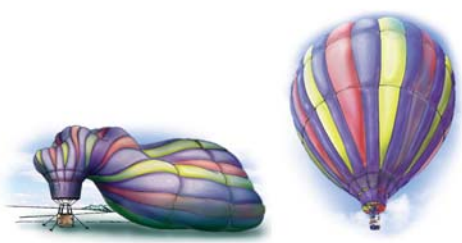
- Boyle's Law:
  - The volume of a gas is inversely proportional to the pressure when the temperature remains constant.

$$P_1V_1 = P_2V_2$$

Charles' Law



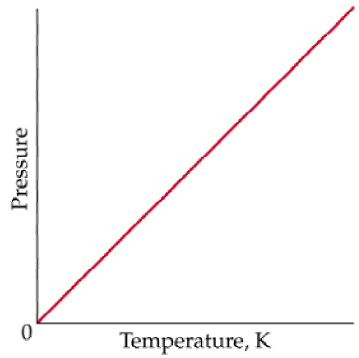
Bettelheim, Brown, March, Introduction to General, Organic and Biochemistry, 6/e
and Introduction to Organic and Biochemistry, 4/e
Figure 5.5



- Charles' Law:
  - The volume of a gas is directly proportional to the absolute temperature if the pressure remains constant.

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

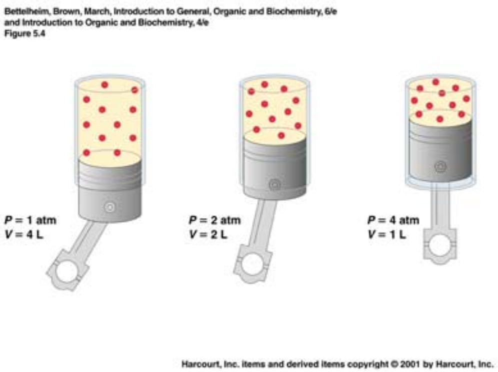
Gay-Lussac's Law

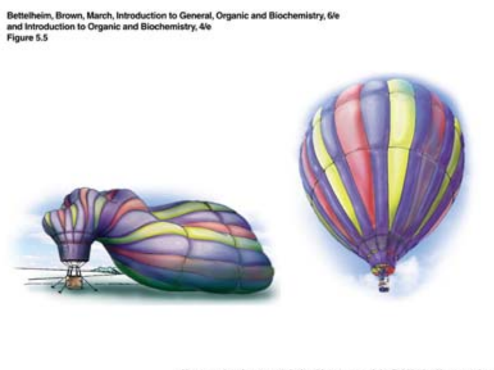


- Gay-Lussac's Law:
  - The pressure of a gas is directly proportional to its absolute temperature if the volume remains constant.

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

Combined Gas Law



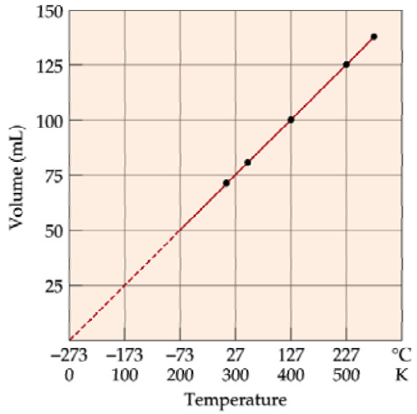


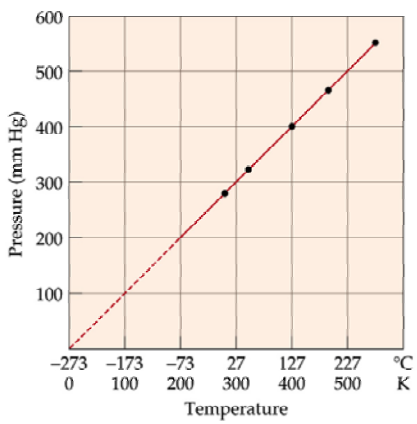
- Combined Gas Law:

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

- Standard Temperature and Pressure (STP):
  - 1 atm, 0^0C (273.15K)

Ideal Gas Behavior





- Absolute zero:
  - The temperature at which the pressure and volume of a gas theoretically reach zero.
  - -273.15^{o}C, 0K

Ideal Gas Equation

- Ideal gases follow the ideal gas equation
- An gas is considered “ideal” when molecular volume is ignored in volume calculations and gas molecules do not attract each other

$$PV = nRT$$

Further Application of Gas laws

The ideal-gas equation can be manipulated to solve a variety of different types of problems. In order to determine the density of a gas, we rearrange the equation to

$$PV = nRT$$

Can you Do this??

Can be rearranged to find molar mass and Density!

$$D = \frac{P\,\mathcal{M}}{R\,T}$$

For Example

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

Example #2

- We can use stoichiometry and the gas laws for a wide variety of uses

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

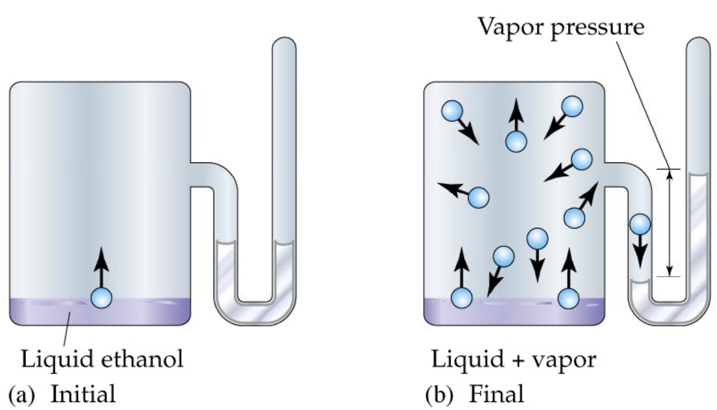
$$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$$

- If the 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°C, how many grams of NaN_3 must be decomposed?

AP Chemistry

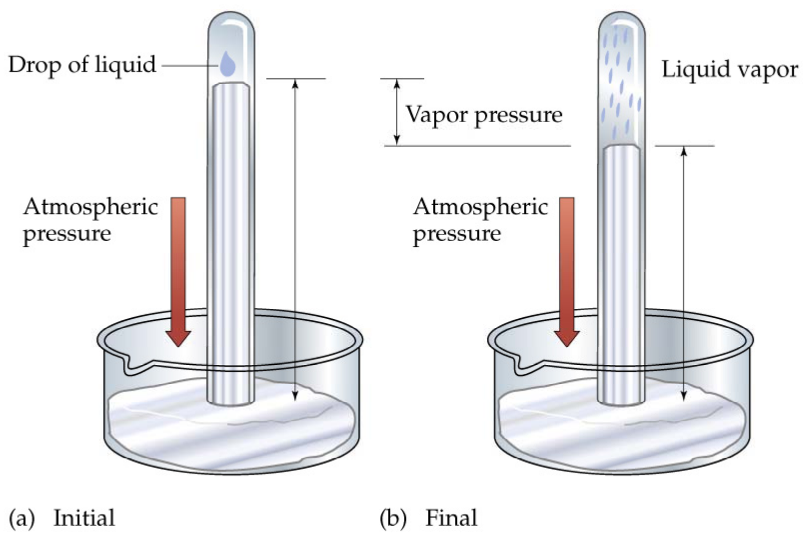
- <u>Chapter 10 Lecture Notes con't</u>
  - 10.6 Dalton's Law
  - 10.7 Kinetic Molecular Theory
  - 10.8 Molecular Effusion and Diffusion
  - 10.9 Real Gas Deviations

The Vapor Pressure Concept



- Vapor Pressure:
  - The pressure exerted by molecules in the vapor above a liquid when the rate of the evaporation and condensation are equal.

The Vapor Pressure Concept



- Vapor pressure increases as temperature increases.
  - Liquid molecules evaporate faster and vapor molecules have more kinetic energy.

10.6 Dalton's Law

- **Dalton's law of partial pressures** states that the total pressure (*Pt*) exerted by a mixture of gases is the sum of the pressures that would be exerted by each individual gas were it the only gas present.

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

- The ratio of partial pressure of a particular component of a gaseous mixture to the total pressure exerted by the gas mixture is the **mole fraction**. Mole fraction, denoted *X*, is a measure of a gas's concentration in a mixture.

$$X_1 = \frac{P_1}{P_t} = \frac{n_1}{n_2}$$

- Knowing the mole fraction of a component and the total pressure of a mixture, we can calculate the partial pressure of the component.

$$\mathrm{P}_1 = X_1\mathrm{P}_t$$

For Example

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?

b) What is the total pressure inside the vessel?

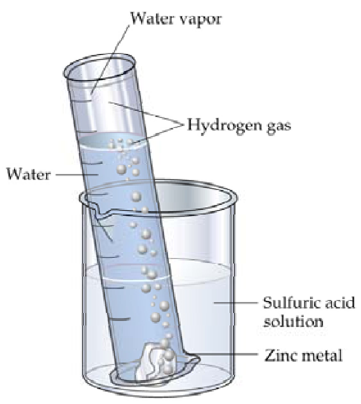
Example #2

- 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 120 L space at 295 K, how many moles of O_2 are needed?

Dalton's Law

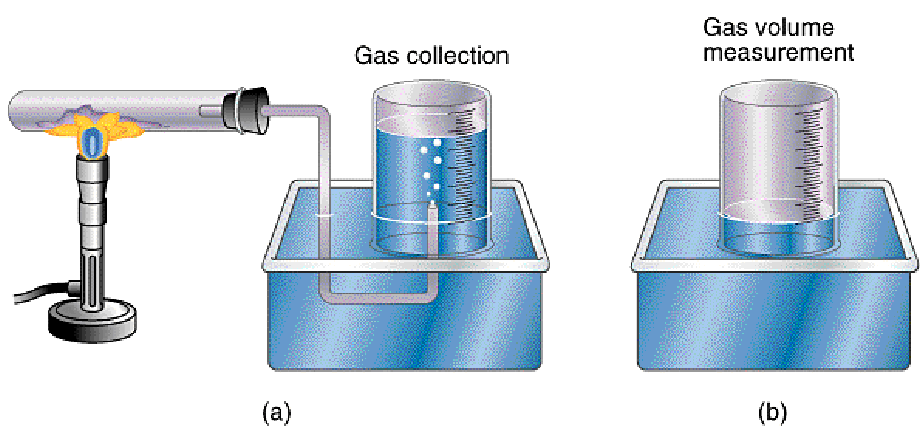


- Dalton's Law of Partial Pressures:
  - Total pressure of a gaseous mixture is equal to the sum of the individual pressures of each gas.
  - Partial Pressure: Pressure exerted by each gas.

Collecting Gas Over Water

- In this experiment, we will collect a gas over water and stoichiometrically determine how much reactant was present before decomposition.

$$P_{total} = P_{gas} + PH_2O$$

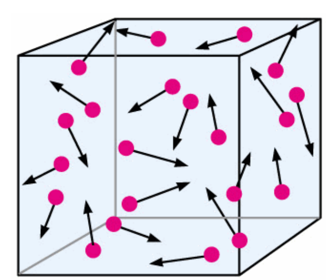


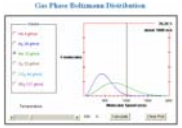
Kinetic Molecular Theory

Kinetic-molecular theory is a way to explain *why* gases obey the gas laws, and it is summarized by the following statements.

1. Gases consist of large numbers of molecules (atoms) that are in continuous, random motion.
2. The volume of all the molecules of the gas is negligible compared to the total volume in which the gas is contained.
3. Attractive and repulsive forces between gas molecules are negligible.
4. Energy can be transferred between molecules during collisions, but the *average* kinetic energy of the molecules does not change with time, as long as the temperature of the gas remains constant. In other words, the collisions are perfectly elastic.
5. The average kinetic energy of the molecules is proportional to the absolute temperature. At any given temperature the molecules of all gases have the same kinetic energy.

See CD video





See CD video

Kinetic Energies of Molecules

- Average KE is proportional to Absolute Temp
  - Average KE = Average speeds
  - KE is same for all gases in a mixture at same temperature
  - Since this is only an average, some are moving faster than others..*some slower.*

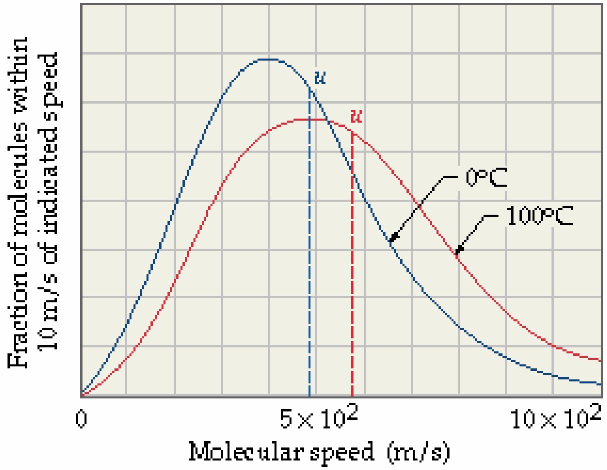


Figure 10.17. Distribution of molecular speeds for nitrogen at 0° C (blue line) and 100° C.

Notice the higher
The temp, the more
Molecules moving
At a greater speed

For Example

- A sample of O_2 Gas initially at STP is compressed to a smaller volume at constant temperature. What effect does this have on:

- The average kinetic energy of O_2 molecules?

- The average speed of O_2 molecules?

- The total number of collisions of O_2 molecules with the walls of the container per unit time?

Molecular Diffusion

- Diffusion is the rate a gas molecule spreads out throughout a substance

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

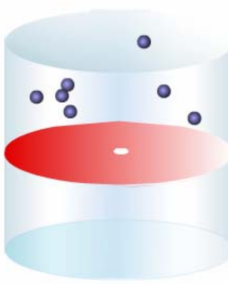
- Where R = 8.314 kg•m^2/s^2•mol •/K
- We can find the average speed of diffusion at a given temperature by the above equation

For Example

- Calculate the rms speed, u, of an N_2 molecule at 25°C. Where R = 8.314 J/mol-K. (Where R = 8.314 kg•m^2/s^2•mol •/K)

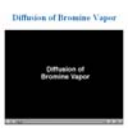
Graham’s Law

See
Cd Video



- **Effusion** is the escape of a gas from a container through a small opening.
  - A consequence of the fact that molecular speed at a particular temperature depends on molecular mass is that lighter molecules undergo diffusion and effusion at faster rates than do heavier molecules. This is summarized by **Graham's law**:

$$\frac{r_1}{r_2} = \sqrt{\frac{\mathcal{M}_2}{\mathcal{M}_1}}$$



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Cd Video

  - where r_1 and r_2 are the *rates* of effusion of two different gases under the same conditions.
  - This explains why balloons filled with helium deflate more rapidly than those filled with air. The helium atoms, being lighter, escape through the tiny openings in the porous rubber faster than the heavier nitrogen and oxygen molecules that make up air.

For Example

- Calculate the ratio of the effusion rates of N_2 and O_2

10.9 Real Gas Deviations

- Although the differences in behavior between real and ideal gases are usually small, it is worthwhile to consider the small differences.
- Deviations from ideal behavior result from the error in assuming that
  - (1) gas molecules occupy no volume, and
  - (2) gas molecules exhibit no intermolecular forces. At very low pressures and very high temperatures, these assumptions are reasonably valid.
- The van der Waals equation corrects for these differences.

$$P = \frac{nRT}{V - nb} - \frac{n^2 a}{V^2}$$

Correction for nonzero volume of molecules

Correction for molecular attractions

10.9 practice problem

What pressure is exerted by 2.00 moles of oxygen gas in a volume of 5.00 L at a temperature of 50° °C?

TABLE 10.3 Van der Waals Constants for Gas Molecules		
Substance (L/mol)	*a* (L^2-atm/mol^2)	*b*
He	0.0341	0.02370
Ne	0.211	0.0171
Ar	1.34	0.0322
Kr	2.32	0.0398
Xe	4.19	0.0510
H_2	0.244	0.0266
N_2	1.39	0.0391
O_2	1.36	0.0318
Cl_2	6.49	0.0562
H_2O	5.46	0.0305
CH_4	2.25	0.0428
CO_2	3.59	0.0427
CCl_4	20.4	0.1383