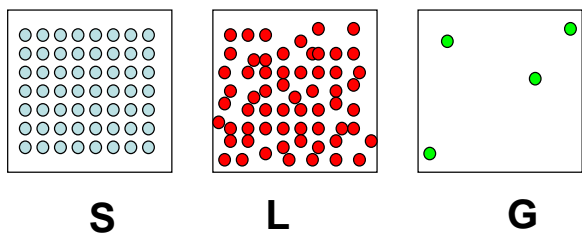


Chapter 13: The Behavior of Gases

I. First Concepts

a. The 3 states of matter most important to us: solids, liquids, and gases.



b. Real Gases and Ideal Gases

- i. Real gases exist, ideal gases do not
- ii. Under conditions of high temperature and low pressure, an ideal gas approximates the behavior of a real gas
 1. This is because at high T and low P, the particles are far apart, and therefore do not interact as strongly with one another.

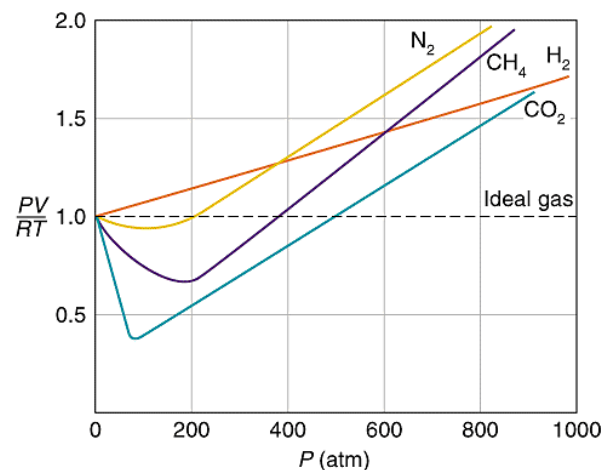


Figure 1. This graph shows how real gas's behavior deviates from ideal behavior as pressure increases. If a gas were to behave perfectly ideally, then the ratio PV/RT would equal exactly 1 for one mole of gas (dashed line).

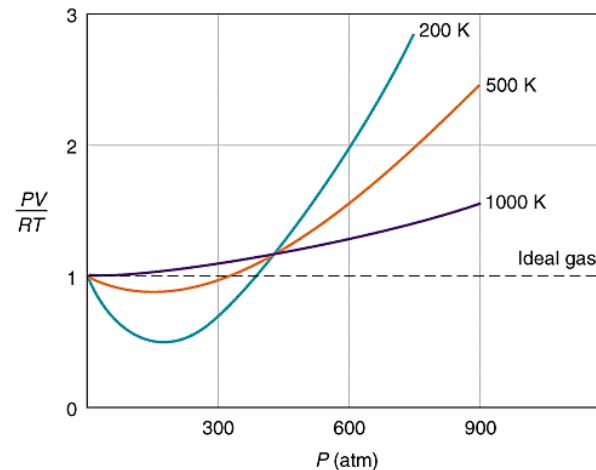


Figure 2. This graph shows how a real gas's behavior deviates from ideal behavior as pressure increases. Each curve represents the behavior of the gas at a different temperature. If a gas were to

behave perfectly ideally, then the ratio PV/RT would equal exactly 1 for one mole of gas (dashed line).

iii. Assumptions of the ideal gas model

1. The particles are tiny point masses of zero volume
2. The particles do not interact with each other (do not attract or repel one another)
3. The particles collide with each other and the walls of the container with perfect elasticity.
 - a. An elastic collision is one in which the particle loses no energy of motion. (This is impossible).
 - b. In reality, all collisions are inelastic. That is, when a basketball is dropped and bounces off the ground, the ball does not bounce back up to the height from which it was dropped. Some of the energy is lost to friction and vibration.

c. Force and Pressure

i. Force is any push or pull on an object

1. $F = \text{mass} \times \text{acceleration} = m \times a$
2. The units of force:

$$= \text{kg} \times \frac{\text{m}}{\text{s}^2} = \frac{\text{kgm}}{\text{s}^2} = \text{N (newtons)}$$

ii. Pressure is the force exerted on an object per unit area

$$P = \frac{F}{A} = \frac{\text{force}}{\text{area}} = \frac{\text{N}}{\text{m}^2} = \text{Pa} = \text{pascal}$$

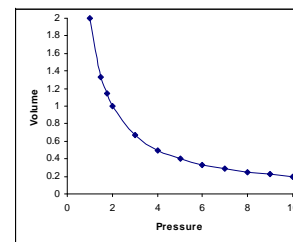
- 1.
2. The pressure of the gases of the atmosphere that is pushing on each one of us and on all objects is called **atmospheric pressure**.
3. Standard atmospheric pressure = the pressure that is exerted by the atmosphere at sea level on a pleasant day of calm weather.
4. Standard atmospheric pressure =
 - a. 101.3 kPa
 - b. 760 mm Hg = 760 torr
 - c. 1.00 atm
5. A **barometer** is an instrument used to measure gas pressure. It is a tube of mercury that is inverted in a bath of mercury. On a nice day at sea level, the atmosphere will support a column of mercury to a height of 760 mm.

d. Converting units of pressure.

II. Demonstrating the effects of pressure with a vacuum pump and bell jar

III. Boyle's law

- a. The relationship between P and V is inverse (at constant T).
- b. If P increases, then V decreases and vice-versa.
- c. The graph is hyperbolic (makes a hyperbola). "Curved" is an OK description for our purposes.

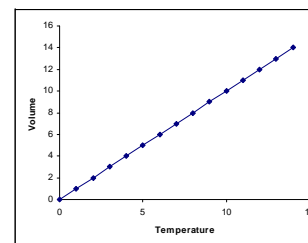


d. $PV = k$

e. $P_1V_1 = P_2V_2$

IV. Charles's Law

- a. The relationship between T and V is direct (at constant P).
- b. If T increases, then V increases. If T decreases, then V decreases.
- c. The graph is linear with a positive slope.



d. $V / T = k$

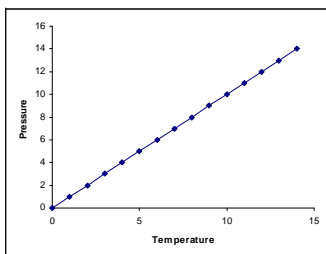
$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

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

f. Temperature must always be in Kelvins

V. Gay-Lussac's Law

- a. The relationship between T and P is direct (at constant V).
- b. If T increases, then P increases. If T decreases, then P decreases.
- c. The graph is linear with a positive slope.



d. $P/T = k$

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

e.

f. Temperature must always be in Kelvins

VI. The Combined Gas Law

a. This combines the first three gas laws

b. $PV/T = k$

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

c.

d. Major helpful tip: Combined gas law problems often require you to find the new volume of a gas at STP. This means that the new pressure (P_2) is standard pressure, and the new temperature (T_2) is standard temperature.

Std P = 1 atm
Std T = 0°C = 273 K

e. Temperature must always be in Kelvins

VII. The Ideal Gas Law

a. This summarizes all gas behavior, including the effect of adding more moles of gas

b. Symbol for moles = "n"

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$$

c.

d. Often, this equation is used in this format:

$$PV = nRT$$

e. R = the ideal gas constant. You don't need to memorize the value of R , but you probably will end up doing so as a result of doing the homework.

$$R = 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{K} \cdot \text{mol}}$$

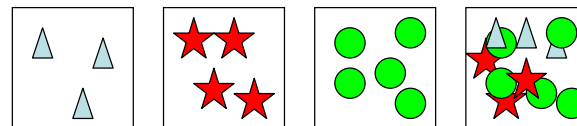
f.

VIII. Dalton's Law of Partial Pressures

a. Dalton's Law of Partial Pressures: The pressure of a mixture of gases is equal to the sum of the partial pressures of each of the individual gases in the mixture.

b. $P_{\text{total}} = P_1 + P_2 + P_3 + \dots + P_n$, where n = the total number of gases in the mixture

c. Suppose that three different gases in three separate but identically-sized containers are all combined into one container. The resulting pressure of this mixture of gases must equal the sum of the individual pressures.



$$P_A + P_B + P_C = P_{\text{total}}$$

3 atm 4 atm 5 atm 12 atm

d. Vocabulary note: " $P_a = 3\text{atm}$ " is pronounced, "the partial pressure of gas 'A' is 3 atmospheres." The **partial pressure** of a gas is the pressure that the gas in some mixture would exert if it were the only gas present in the container.

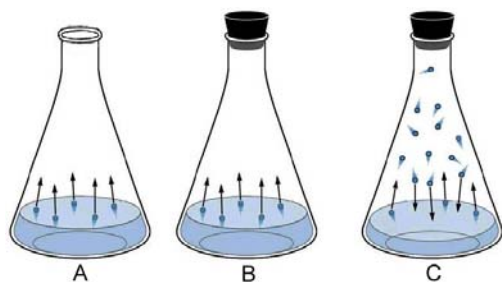
e. This is an important principle. It is important to notice that the sizes of the particles and the identities of the gases involved are unimportant. Dalton's Law holds true for any mixture of gases. This observation is closely related to the work of an Italian scientist, Amadeo Avogadro.

f. Avogadro's hypothesis

i. Equal volumes of any gases must contain the same number of particles (as long as the gases are at the same temp and pressure).

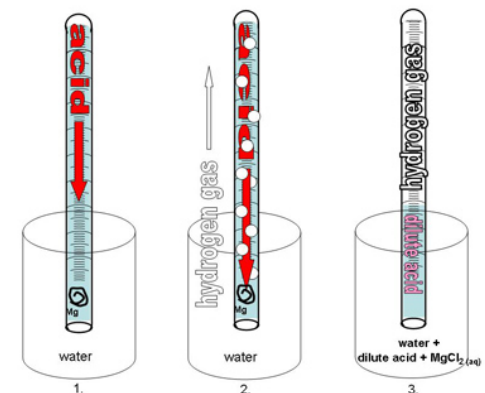
ii. This is a remarkable statement, because it holds true even when the gas particles are not the same size. For instance, 1 liter of H_2 gas (2.02 g/mol) and 1 liter of UF_6 gas (352 g/mol) both contain the same number of particles.

- iii. This counterintuitive statement makes sense if one takes into account the fact that the distances between the particles are so great that the volumes of the individual particles are negligible.
- g. Avogadro's number
 - i. Amadeo Avogadro didn't actually determine Avogadro's number ($N_A = 6.022 \times 10^{23}$). It was determined 9 years after he was dead. However, the number is named in his honor because Avogadro's ideas and work paved the way to the determination of the size of one mole.
- h. Collecting gas over water
 - i. An important application of Dalton's Law is the computation of the pressure of a pure gas when the gas is collected over water.
 - ii. When gas is collected over water, the water evaporates to a small extent. Eventually the water will reach equilibrium. This means that the water will continue to evaporate, but the net rate of evaporation reaches zero. When the rate of evaporation is exactly equal to the reverse process of condensation, water-water vapor equilibrium is achieved.
 - iii. When this equilibrium is achieved, the water vapor present, like any gas, will exert a pressure.
 - iv. This pressure of water vapor by itself, **the partial pressure of water vapor**, must be subtracted from the total pressure of the gas collected in order to determine the pressure of the pure gas, also called the "dry gas pressure."
 - v. Examine the diagram of three flasks.



The water in flask "A" will evaporate, and because the container is open to the air, its net evaporation will continue also. Container "B" is closed, but the water will still evaporate. "In diagram "C", we see that the water is still evaporating, but the net evaporation has stopped. The net evaporation has a rate of zero because the rate of evaporation is equaled by the rate of condensation. That is, every time a water molecule evaporates,

- another water vapor molecule condense. This state of balance is called "equilibrium." (In AP Chemistry you will learn that this is actually a dynamic equilibrium, as opposed to a static equilibrium.)
- vi. Examine the diagram below of the lab set-up for generating hydrogen gas and collecting over water.



- vii. Once the hydrogen gas is collected over water, there is also a certain amount of water vapor collected along with the hydrogen. Therefore, we must look up the vapor pressure of water at the temperature of the lab that day, and subtract that pressure away from the total pressure of the gas collected.

$$P_{\text{total}} = P_{\text{hydrogen}} + P_{\text{water vapor}}$$
 This information can be looked up in a table that is located in your book. Here is a partial table below.

Temp (°C)	Saturated Vapor Pressure (mmHg)
-10	2.15
0	4.58
5	6.54
10	9.21
11	9.84
12	10.52

13	11.23
14	11.99
15	12.79
20	17.54
25	23.76
30	31.8
37	47.07

IX. Graham's Law of Effusion

- All substances sitting in our classroom are at the same temperature, except the human beings and probably the electrical devices. When hot and cold objects are in contact for a long time, they reach *thermal equilibrium*, that is, they will reach the same temperature.
- Thus, the air in the classroom has the same temperature as the desks and the chairs and the walls.
- Temperature – how hot or cold an object is – is a measure of the average kinetic energy of the particles of a substance. (Kinetic energy is the energy of motion.) The particles of all substances at the same temperature have the same average kinetic energy.
- The K.E. of an object is determined by how big the object is and how fast the object is moving. (Which has more KE: a baseball moving at 50 mph or a truck moving at 50 mph? Which has more KE: a baseball moving at 50 mph or a baseball moving at 100 mph?)
KE = $\frac{1}{2} \times m \times v^2$ (v = velocity, which for our purposes is the same thing as speed)

$$KE = \frac{1}{2}mv^2$$

- This is the kinetic energy of a particle of a gas
- Let us consider two different gases, Gas "A" and gas "B", that have different masses. If both gases have the same temperature, then the particles of both gases must have the same average KE. What does this tell us about the speeds of the particles of the two different gases?

$$KE \text{ of gas A} = \frac{1}{2}m_A v_A^2 \quad KE \text{ of gas B} = \frac{1}{2}m_B v_B^2$$

- If gas A and gas B are at the same temperature, then their KEs are equal.

$$KE \text{ of gas A} = KE \text{ of gas B}$$

$$\frac{1}{2}m_A v_A^2 = \frac{1}{2}m_B v_B^2$$

$$m_A v_A^2 = m_B v_B^2$$

$$\frac{m_A v_A^2}{m_B} = \frac{m_B v_B^2}{m_B}$$

$$\frac{m_A v_A^2}{m_B} = v_B^2$$

$$\frac{m_A v_A^2}{m_B v_A^2} = \frac{v_B^2}{v_A^2}$$

$$\frac{m_A}{m_B} = \frac{v_B^2}{v_A^2}$$

$$\sqrt{\frac{m_A}{m_B}} = \sqrt{\frac{v_B^2}{v_A^2}}$$

$$\sqrt{\frac{m_A}{m_B}} = \frac{\sqrt{v_B^2}}{\sqrt{v_A^2}} = \frac{v_B}{v_A} \quad (\text{we will ignore the negative root!})$$

$$*** \sqrt{\frac{m_A}{m_B}} = \frac{v_B}{v_A} *** \text{ Graham's Law of Effusion}$$

- Graham's Law of Effusion:** The ratio of the velocities of the particles of two different gases is inversely proportional to the square root of the ratio of the masses of the particles.
- Example Problem:** Two different balloons have been filled with two different gases: one contains helium (He) and the other contains krypton (Kr). If the balloons both develop leaks of the same size, which gas will effuse (leak out of the balloon) more quickly? How many times more quickly will it leak?

Solution: Do not try to solve for v_a or v_b . **Solve for the ratio " v_b/v_a "!**

Assign each gas to a letter. Put the lighter gas on top (Gas B).

He = 4.00 g/mol, Kr = 83.8 g/mol.

Gas B = He (it's lighter), Gas A = Kr (it's heavier).

$$\sqrt{\frac{m_A}{m_B}} = \frac{v_B}{v_A}$$

$$\sqrt{\frac{m_{Kr}}{m_{He}}} = \frac{v_{He}}{v_{Kr}}$$

$$\sqrt{\frac{83.8}{4.00}} = \frac{V_{\text{He}}}{V_{\text{Kr}}}$$

$$\sqrt{20.95} = \frac{V_{\text{He}}}{V_{\text{Kr}}}$$

$$4.58 = \frac{V_{\text{He}}}{V_{\text{Kr}}}$$

Answer: He particles are lighter than Kr particles, thus the helium will effuse from its balloon more quickly than the krypton will effuse from the balloon full of Kr. The He gas will effuse 4.58 times faster than the Kr gas will effuse.

X. Gas Density and Determining Molecular Mass of Gases

$$D = \frac{m}{V} = \frac{\text{number of grams}}{\text{number of liters}}$$

- Previously, we used units of g/ml (g/cm^3) to express density. However, because gases are typically thousands of times less dense than liquids and solids, it is more convenient to use g/L to express their densities.
- Remember: density of a substance does not depend on the amount of the substance. (Does mass depend on the amount of the substance? Yes. Does color? No. Density is more like color, which is an intrinsic property.) Therefore, when calculating the density of gases, we can assume any amount of the gas that is convenient for us. To make things easy on yourself, **always assume one mole of gas for your calculations.** (You can assume any amount that you want, and the problem will still work. It will take longer to solve, though.)
- Why assume one mole of gas? Here's why: because the **mass** of one mole of any gas can be looked up on the periodic table. (Remember: the "molar mass" of a substance is the "mass of one mole.") Additionally, the **volume** of a mole of any gas is 22.4 liters (at STP). The "22.4 L/mol" value is only valid for gases at STP, but I will only test you on gas densities at STP for the sake of simplicity. (It should be noted, however, that you could determine the gas density at any conditions using the combined gas law.)
- Example problem:
Find the density of H_2 gas at STP.
Solution:

$$D = \frac{m}{V} = \frac{\text{mass of one mole}}{\text{volume of one mole at STP}}$$

$$D = \frac{2.02 \text{ g (from periodic table)}}{22.4 \text{ L}}$$

$$D = 0.0902 \text{ g/L}$$

XI. Gas Stoichiometry