

# GASES

## THE PROPERTIES OF GASES

Only 4 quantities are needed to define the **state of a gas**:

- the *quantity* of the gas,  $n$  (in moles)
- the *temperature* of the gas,  $T$  (in **KELVINS**)
- the *volume* of the gas,  $V$  (in liters)
- the *pressure* of the gas,  $P$  (in atmospheres)

A gas uniformly fills any container, is easily compressed & mixes completely with any other gas.

## GAS PRESSURE:

A measure of the force that a gas exerts on its container. Force is the physical quantity that interferes with inertia. Gravity is the force responsible for weight.

Force = mass x acceleration; Newton's 2<sup>nd</sup> Law.

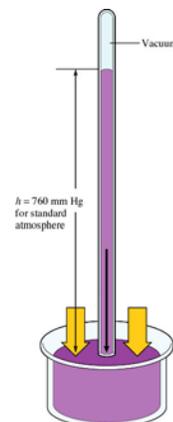
$$N = \text{kg} \times \text{m/s}^2$$

**Pressure**-- Force/ unit area;  $\text{N/m}^2$

**Barometer**--invented by Evangelista Torricelli in 1643; uses the height of a column of mercury to measure gas pressure (especially atmospheric)

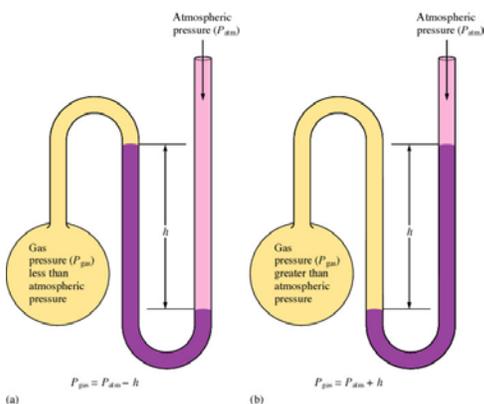
1 mm of Hg = 1 torr

$$760.00 \text{ mm Hg} = 760.00 \text{ torr} = 1.00 \text{ atm} = 101.325 \text{ kPa} \approx 10^5 \text{ Pa}$$



Early barometer

At sea level all of the above define STANDARD PRESSURE. The SI unit of pressure is the Pascal (Blaise Pascal);  $1 \text{ Pa} = 1 \text{ N} / \text{m}^2$



The **manometer**—a device for measuring the pressure of a gas in a container. The pressure of the gas is given by  $h$  [the difference in mercury levels] in units of torr (equivalent to mm Hg).

- Gas pressure = atmospheric pressure -  $h$
- Gas pressure = atmospheric pressure +  $h$

### Exercise 1 Pressure Conversions

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

$$6.4 \times 10^{-2} \text{ atm}$$
$$6.5 \times 10^3 \text{ Pa}$$

### Exercise

Rank the following pressures in decreasing order of magnitude (largest first, smallest last): 75 kPa, 300. torr, 0.60 atm and 350. mm Hg.

### GAS LAWS: THE EXPERIMENTAL BASIS

- **BOYLE'S LAW:** "father of chemistry"--*the volume of a confined gas is inversely proportional to the pressure exerted on the gas. ALL GASES BEHAVE IN THIS MANNER!*
- Robert Boyle was an Irish chemist. He studied P V relationships using a J-tube set up in the multi-story entryway of his home.

- $P \propto 1/V$  plot = straight line
- pressure and volume are **inversely** proportional
- Volume  $\uparrow$  pressure  $\downarrow$  **at constant temperature**, the converse is also true
- *for a given quantity of a gas at constant temperature, the product of pressure and volume is a constant.*

$$PV = k$$

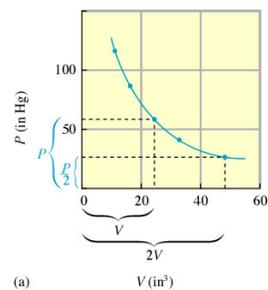
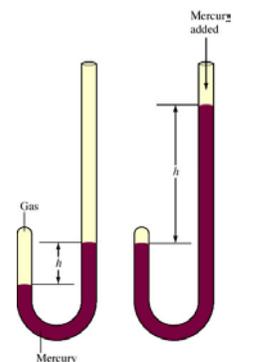
$$\text{Therefore, } V = \frac{k}{P} = k \frac{1}{P}$$

which is the equation for a straight line of the type

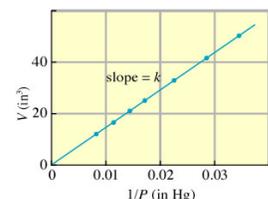
$$y = mx + b \quad \text{where } m = \text{slope, and } b \text{ the } y\text{-intercept}$$

In this case,  $y = V$ ,  $x = 1/P$  and  $b = 0$ . Check out the plot on the right (b). The data Boyle collected is graphed on (a) above.

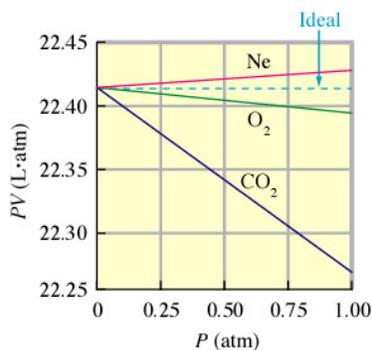
- $P_1V_1 = P_2V_2$  is the easiest form of Boyle's law to **memorize**
- Boyle's Law has been tested for over three centuries. It holds true **only at low pressures**.



(a)



(b)



A plot of PV versus P for several gases at pressures below 1atm is pictured at left.

An **ideal** gas is expected to have a constant value of PV, as shown by the dotted line. CO<sub>2</sub> shows the largest change in PV, and this change is actually quite small: PV changes from about 22.39 L·atm at 0.25 atm to 22.26 L·atm at 1.00 atm. Thus Boyle's Law is a good approximation at these relatively low pressures.

**Exercise 2 Boyle's Law I**

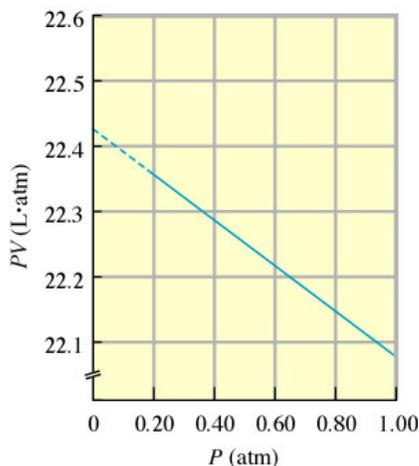
Sulfur dioxide ( $\text{SO}_2$ ), a gas that plays a central role in the formation of acid rain, is found in the exhaust of automobiles and power plants. Consider a 1.53- L sample of gaseous  $\text{SO}_2$  at a pressure of  $5.6 \times 10^3$  Pa. If the pressure is changed to  $1.5 \times 10^4$  Pa at a constant temperature, what will be the new volume of the gas ?

**0.57 L****Exercise 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 mol  $\text{NH}_3$  gas at a temperature of  $0^\circ\text{C}$ . Using the results listed below, calculate the Boyle's law constant for  $\text{NH}_3$  at the various pressures.

<i>Experiment</i>	<i>Pressure (atm)</i>	<i>Volume (L)</i>
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

**experiment 1 is 22.37**  
**experiment 2 is 22.32**  
**experiment 3 is 22.31**  
**experiment 4 is 22.25**  
**experiment 5 is 22.16**  
**experiment 6 is 22.08**



PLOT the values of PV for the six experiments above.

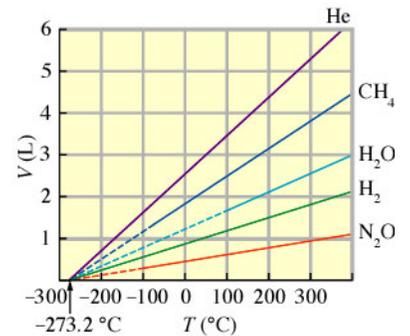
Extrapolate it back to see what PV equals at 0.00 atm pressure.

Compare it to the PV vs. P graph on page 2 of these notes.

What is the y-intercept for all of these gases?

Remember, gases behave most ideally at low pressures. You can't get a pressure lower than 0.00 atm!

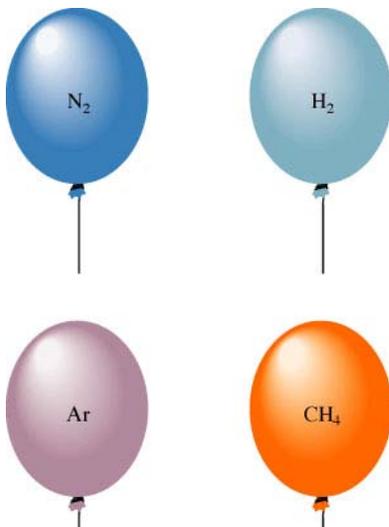
- **CHARLES' LAW:** *If a given quantity of gas is held at a constant pressure, then its volume is directly proportional to the absolute temperature. Must use KELVIN*
- Jacques Charles was a French physicist and the first person to fill a hot "air" balloon with hydrogen gas and made the first solo balloon flight!
  - $V \propto T$  plot = straight line
  - $V_1 T_2 = V_2 T_1$
  - Temperature  $\uparrow$  Volume  $\uparrow$  **at constant pressure**
  - This figure shows the plots of V vs. T (Celcius) for several gases. The solid lines represent experimental measurements on gases. The dashed lines represent extrapolation of the data into regions where these gases would become liquids or solids. Note that the samples of the various gases contain different numbers of moles.
  - What is the temperature when the Volume extrapolates to zero?



#### Exercise 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 ?

2.79 L



These balloons each hold 1.0 L of gas at 25°C and 1 atm. Each balloon contains 0.041 mol of gas, or  $2.5 \times 10^{22}$  molecules.

- **GAY-LUSSAC'S LAW of combining volumes:** volumes of gases always combine with one another in the ratio of small whole numbers, as long as volumes are measured at the same T and P.
  - $P_1 T_2 = P_2 T_1$
  - Avogadro's hypothesis: *equal volumes of gases under the same conditions of temperature and pressure contain equal numbers of molecules.*
- **AVOGADRO'S LAW:** *The volume of a gas, at a given temperature and pressure, is directly proportional to the quantity of gas.*
  - $V \propto n$
  - $n \uparrow$  Volume  $\uparrow$  **at constant T & P**

HERE'S AN EASY WAY TO MEMORIZE ALL OF THIS!

Start with the combined gas law:

$$P_1 V_1 T_2 = P_2 V_2 T_1$$

Memorize it.

Next, put the guy's names in alphabetical order.

Boyle's uses the first 2 variables, Charles' the second 2 variables & Gay-Lussac's the remaining combination of variables. What ever doesn't appear in the formula, is being held CONSTANT!

**Exercise 5 Avogadro's Law**

Suppose we have a 12.2-L sample containing 0.50 mol oxygen gas (O<sub>2</sub>) at a pressure of 1 atm and a temperature of 25°C. If all this O<sub>2</sub> were converted to ozone (O<sub>3</sub>) at the same temperature and pressure, what would be the volume of the ozone ?

**8.1 L****THE IDEAL GAS LAW**

Four quantities describe the state of a gas: pressure, volume, temperature, and # of moles (quantity).  
Combine all 3 laws:

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

Replace the % with a constant, R, and you get:

$$PV = nRT$$

The ideal gas law! *It is an equation of state.* R = 0.8206 L· atm/mol · K

Useful only at low Pressures and high temperatures! Guaranteed points on the AP Exam!

These next exercises can all be solved with the ideal gas law, BUT, you can use another if you like!

**Exercise 6 Ideal Gas Law I**

A sample of hydrogen gas (H<sub>2</sub>) has a volume of 8.56 L at a temperature of 0°C and a pressure of 1.5 atm. Calculate the moles of H<sub>2</sub> molecules present in this gas sample.

**0.57 mol****Exercise 7 Ideal Gas Law II**

Suppose we have a sample of ammonia gas with a volume of 3.5 L at a pressure of 1.68 atm. The gas is compressed to a volume of 1.35 L at a constant temperature. Use the ideal gas law to calculate the final pressure.

**4.4 atm**

**Exercise 8                    Ideal Gas Law III**

A sample of methane gas that has a volume of 3.8 L at 5°C is heated to 86°C at constant pressure. Calculate its new volume.

**4.9 L****Exercise 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 ?

**3.07 L****Exercise 10                  Ideal Gas Law V**

A sample containing 0.35 mol 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.

**decreases by 3 L****GAS STOICHIOMETRY**

Use  $PV = nRT$  to solve for the volume of one mole of gas at STP:

Look familiar? This is the **molar volume** of a gas at STP. Work stoichiometry problems using your favorite method, dimensional analysis, mole map, the table way...just work FAST! Use the ideal gas law to convert quantities that are NOT at STP.

**Exercise 11            Gas Stoichiometry I**

A sample of nitrogen gas has a volume of 1.75 L at STP. How many moles of N<sub>2</sub> are present ?

**7.81 x 10<sup>-2</sup> mol N<sub>2</sub>**

**Exercise 12            Gas Stoichiometry II**

Quicklime (CaO) is produced by the thermal decomposition of calcium carbonate (CaCO<sub>3</sub>). Calculate the volume of CO<sub>2</sub> at STP produced from the decomposition of 152 g CaCO<sub>3</sub> by the reaction



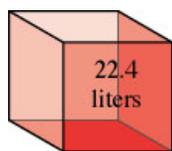
**34.1 L CO<sub>2</sub> at STP**

**Exercise 13            Gas Stoichiometry III**

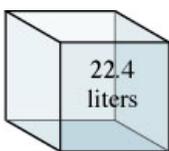
A sample of methane gas having a volume of 2.80 L at 25°C and 1.65 atm was mixed with a sample of oxygen gas having a volume of 35.0 L at 31°C and 1.25 atm. The mixture was then ignited to form carbon dioxide and water. Calculate the volume of CO<sub>2</sub> formed at a pressure of 2.50 atm and a temperature of 125°C.

**2.47 L**

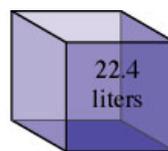
## THE DENSITY OF GASES:



22.4  
liters  
32 g of O<sub>2</sub>  
(1 mole)



22.4  
liters  
2 g of H<sub>2</sub>  
(1 mole)



22.4  
liters  
17 g of NH<sub>3</sub>  
(1 mole)

$$d = \frac{m}{V} = \frac{P(\text{FW})}{RT} \quad \{\text{for ONE mole of gas}\} = \frac{\text{FW}}{22.4 \text{ L}} \quad \text{AND} \quad \text{Molar Mass} = \text{FW} = \frac{dRT}{P}$$

“Molecular Mass kitty cat”—all good cats put dirt [dRT] over their pee [P]. Corny, but you’ll thank me later!

Just remember that densities of gases are reported in g/L NOT g/mL.

What is the approximate molar mass of air? \_\_\_\_\_

The density of air is approx. \_\_\_\_\_ g/L.

List 3 gases that float in air:

List 3 gases that sink in air:

### Exercise 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.

**32.0 g/mol**

## GAS MIXTURES AND PARTIAL PRESSURES

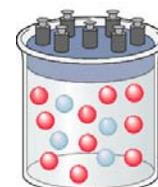
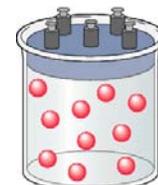
*The pressure of a mixture of gases is the sum of the pressures of the different components of the mixture:*

$$P_{\text{total}} = P_1 + P_2 + \dots + P_n$$

John Dalton=s Law of Partial Pressures also uses the concept of mole fraction, X

$$X_A = \frac{\text{moles of A}}{\text{moles A} + \text{moles B} + \text{moles C} + \dots}$$

so now,  $P_A = X_A P_{\text{total}}$



The partial pressure of each gas in a mixture of gases in a container depends on the number of moles of that gas. The total pressure is the SUM of the partial pressures and depends on the total moles of gas particles present, no matter what they are!

**Exercise 15 Dalton's Law I**

Mixtures of helium and oxygen are 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<sub>2</sub> 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.

$$\begin{aligned}P_{\text{He}} &= 9.3 \text{ atm} \\P_{\text{O}_2} &= 2.4 \text{ atm} \\P_{\text{TOTAL}} &= 11.7 \text{ atm}\end{aligned}$$

**Exercise 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<sub>2</sub> present.

0.210

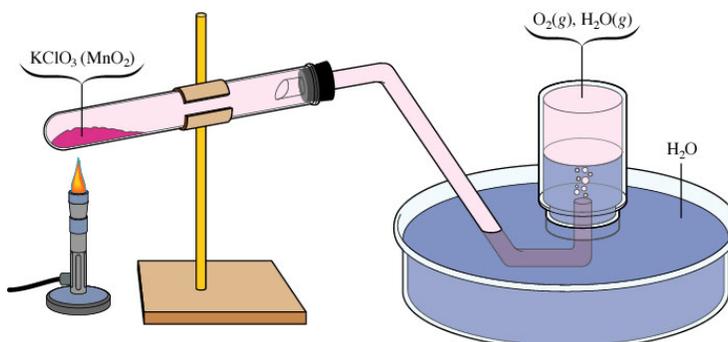
**Exercise 17 Dalton's Law III**

The mole fraction of nitrogen in the air is 0.7808. Calculate the partial pressure of N<sub>2</sub> in air when the atmospheric pressure is 760. torr.

593 torr

**WATER DISPLACEMENT**

It is common to collect a gas by water displacement which means some of the pressure is due to water vapor collected as the gas was passing through! You must correct for this. You look up the partial pressure due to water vapor by knowing the temperature.

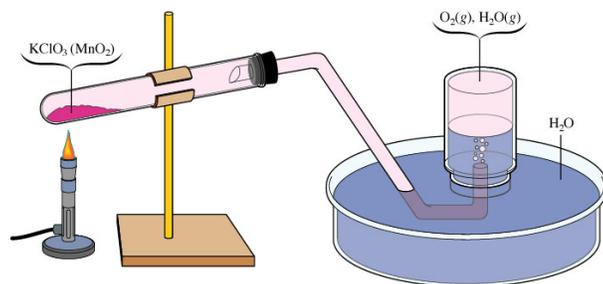


## Exercise 8 Gas Collection over Water

A sample of solid potassium chlorate ( $\text{KClO}_3$ ) was heated in a test tube (see the figure above) and decomposed by the following reaction:



The oxygen produced was collected by displacement of water at  $22^\circ\text{C}$  at a total pressure of 754 torr. The volume of the gas collected was 0.650 L, and the vapor pressure of water at  $22^\circ\text{C}$  is 21 torr. Calculate the partial pressure of  $\text{O}_2$  in the gas collected and the mass of  $\text{KClO}_3$  in the sample that was decomposed.



**Partial pressure of  $\text{O}_2 = 733$  torr  
2.12 g  $\text{KClO}_3$**

## KINETIC MOLECULAR THEORY OF GASES

Assumptions of the MODEL:

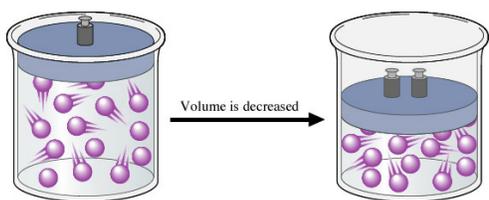
1. All particles are in constant, random, motion.
2. All collisions between particles are perfectly elastic.
3. The volume of the particles in a gas is negligible
4. The average kinetic energy of the molecules is its Kelvin temperature.

This neglects any intermolecular forces as well.

Gases expand to fill their container, solids/liquids do not.

Gases are compressible, solids/liquids are not appreciably compressible.

This helps explain:

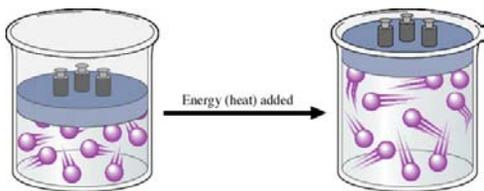


Boyle's Law – P & V

If the volume is decreased that means that the gas particles will hit the wall more often, thus increasing pressure

$$P = (nRT) \frac{1}{V}$$

↑  
Constant



### Charles' Law – V & T

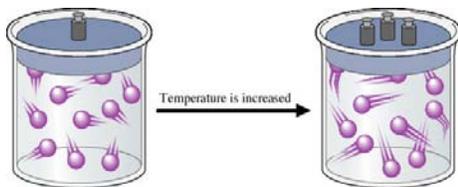
When a gas is heated, the speed of its particles increase and thus hitting the walls more often and with more force. The only way to keep the P constant is to increase the volume of the container.

$$V = \left( \frac{nR}{P} \right) T$$



Constant

### Gay-Lussac's Law – P & T



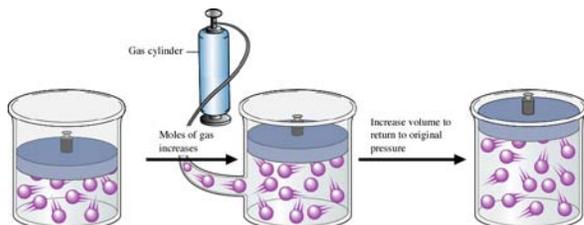
When the temperature of a gas increases, the speeds of its particles increase, the particles are hitting the wall with greater force and greater frequency. Since the volume remains the same this would result in increased gas pressure.

$$P = \left( \frac{nR}{V} \right) T$$



Constant

### Avogadro's Law – V & n



An increase in the number of particles at the same temperature would cause the pressure to increase if the volume were held constant. The only way to keep constant P is to vary the V.

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



Constant

Dalton's Law – The P exerted by a mixture of gases is the SUM of the partial pressures since gas particles are independent of each other and the volumes of the individual particles DO NOT matter.

### DISTRIBUTION OF MOLECULAR SPEEDS

Plot # of gas molecules having various speeds  $v$ . the speed and you get a curve. Changing the temperature affects the shape of the curve NOT the area beneath it. Change the # of molecules and all bets are off!

Maxwell's equation:

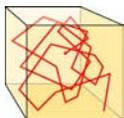
$$\sqrt{u^2} = u_{rms} = \sqrt{\frac{3RT}{FW}}$$

Use 8.314510 J/K·mol for this equation—the “energy” R since it's really kinetic energy that is at work here!

### Exercise 19      Root Mean Square Velocity

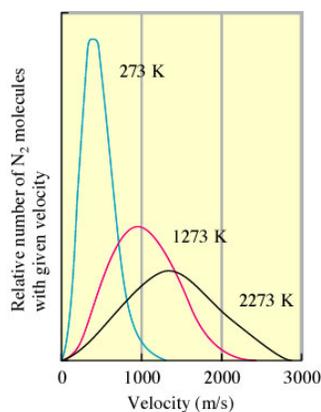
Calculate the root mean square velocity for the atoms in a sample of helium gas at 25°C.

$$1.36 \times 10^3 \text{ m/s}$$



If we could monitor the path of a single molecule it would be very erratic.

**Mean free path**—the average distance a particle travels between collisions. It's on the order of a tenth of a micrometer. WAY SMALL!



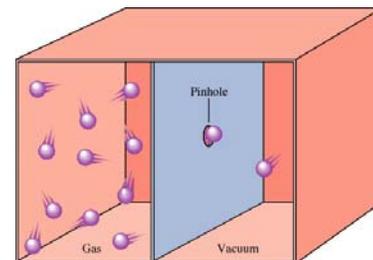
Examine the effect of temperature on the numbers of molecules with a given velocity as it relates to temperature. **HEAT 'EM UP, SPEED 'EM UP!!**

Drop a vertical line from the peak of each of the three bell shaped curves—that point on the x-axis represents the AVERAGE velocity of the sample at that temperature. Note how the bells are “squashed” as the temperature increases. You may see graphs like this on the AP exam where you have to identify the highest temperature based on the shape of the graph!

### GRAHAM'S LAW OF DIFFUSION AND EFFUSION

Effusion is closely related to diffusion. **Diffusion** is the term used to describe the mixing of gases. The *rate* of diffusion is the *rate* of the mixing.

**Effusion** is the term used to describe the passage of a gas through a tiny orifice into an evacuated chamber as shown on the right. The rate of effusion measures the speed at which the gas is transferred into the chamber.



*The rates of effusion of two gases are inversely proportional to the square roots of their molar masses at the same temperature and pressure.*

$$\frac{\text{Rate of effusion of gas 1}}{\text{Rate of effusion of gas 2}} = \sqrt{\frac{FW_2}{FW_1}}$$

REMEMBER *rate* is a change in a quantity over time, NOT just the time!

### Exercise 20 Effusion Rates

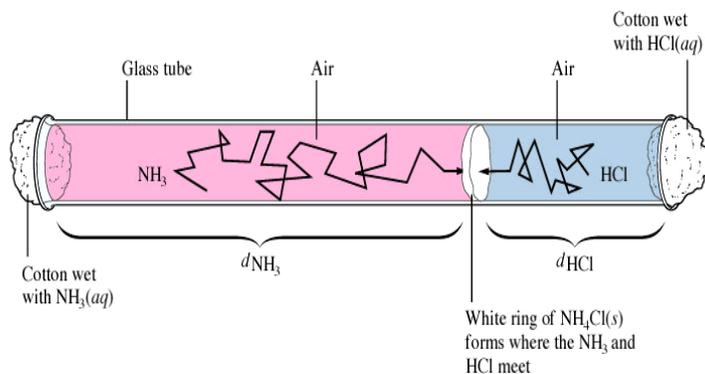
Calculate the ratio of the effusion rates of hydrogen gas ( $\text{H}_2$ ) and uranium hexafluoride ( $\text{UF}_6$ ), a gas used in the enrichment process to produce fuel for nuclear reactors.

13.2

### Exercise

A pure sample of methane is found to effuse through a porous barrier in 1.50 minutes. Under the same conditions, an equal number of molecules of an unknown gas effuses through the barrier in 4.73 minutes. What is the molar mass of the unknown gas?

### Diffusion



$$\frac{\text{Distance traveled by NH}_3}{\text{Distance traveled by HCl}} = \frac{u_{\text{rms for NH}_3}}{u_{\text{rms for HCl}}} = \sqrt{\frac{FW_{\text{HCl}}}{FW_{\text{NH}_3}}} = \sqrt{\frac{36.5}{17}} = 1.5$$

The observed ratio is LESS than a 1.5 distance ratio—why?

This diffusion is slow considering the molecular velocities are 450 and 660 meters per second—which one is which?