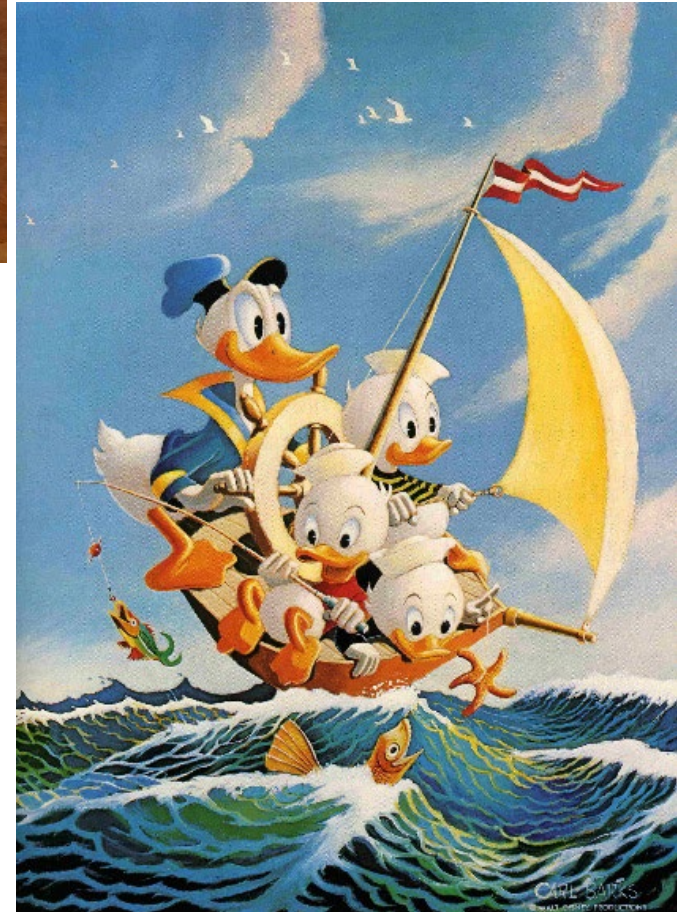


PROPERTIES OF GASES



DALTON'S LAW OF PARTIAL PRESSURES

The total pressure of a mixture of gases equals the partial pressures of each of the constituent gases. Furthermore, a mixture of gases that do not react with one another behaves like a single pure gas.

$$P_{\text{total}} = P_{\text{gas A}} + P_{\text{gas B}} + P_{\text{gas C}} \dots$$

If the total pressure of a system was 2.5 atm, what is the partial pressure of carbon monoxide if the gas mixture also contained 0.4 atm O₂ and 1.48 atm of N₂?

Mole Fraction

The easiest way to express the relation between the total pressure of a mixture and the partial pressures of its components is to introduce the mole fraction, χ_J , of each component. The mole fraction is a dimensionless number that expresses the ratio of the number of moles of one component (A) to the total number of moles in the mixture. That is,

$$\chi_A = n_A / (n_A + n_B + \dots)$$

The partial pressure of a gas is then related to the total pressure by the mole fraction as follows:

$$P_A = \chi_A P_{\text{Total}}$$

This is known as vapor pressure lowering.

DALTON'S LAW OF PARTIAL PRESSURES-2

EXAMPLE:

A 250.0 mL sample of a gas mixture was analyzed and found to contain 2.00 g of N₂ and 1.75 g of O₃ at 55.0°C. What is the total pressure of the mixture and the partial pressure of each component?

Video
in
canvas

$$2.00\text{g N}_2 \left(\frac{1 \text{ mol}}{28\text{g}} \right) = 0.071429 \text{ mol N}_2$$

$$1.75\text{g O}_3 \left(\frac{1 \text{ mol}}{48\text{g}} \right) = 0.036458 \text{ mol O}_3$$

$$P_{\text{N}_2} = nRT/V = (0.071429 \text{ mol})(0.082 \text{ L atm} / \text{mol K})(55+273\text{K}) / 0.250 \text{ L}$$

$$P_{\text{N}_2} = 7.685 \text{ atm}$$

$$P_{\text{O}_3} = nRT/V = (0.036458 \text{ mol})(0.082 \text{ L atm} / \text{mol K})(55+273\text{K}) / 0.250 \text{ L}$$

$$P_{\text{O}_3} = 3.9223 \text{ atm}$$

$$P_{\text{T}} = P_{\text{N}_2} + P_{\text{O}_3} = 11.607 \text{ atm}$$

or

$$X_{\text{N}_2} = 0.071429 / (0.071429 + 0.036458)$$

$$X_{\text{N}_2} = 0.66207$$

$$P_{\text{N}_2} = (0.66207) (11.6 \text{ atm}) = 7.68 \text{ atm!}$$

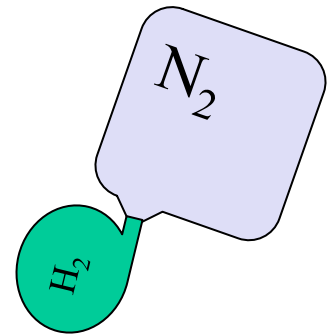
$$P_{\text{T}} = n_{\text{T}} RT/V = (0.07143+0.03646)(0.082 \text{ L atm} / \text{mol K})(55+273\text{K})/0.250 \text{ L}$$

P_T = 11.6 atm then use the mole fraction to find the partial pressures

LECTURE Exercise 2

A 4.0-L vessel containing N_2 at STP and a 2.0-L vessel containing H_2 at STP are connected by a valve. If the valve is opened allowing the two gases to mix, what is the mole fraction of hydrogen in the mixture?

- (a) 0.034
- (b) 0.33
- (c) 0.50
- (d) 0.67
- (e) 0.96



Sample Exercise 10.11 Relating Mole Fractions and Partial Pressures

A study of the effects of certain gases on plant growth requires a synthetic atmosphere composed of 1.5 mol % CO_2 , 18.0 mol % O_2 and 80.5 mol % Ar.

(a) Calculate the partial pressure of O_2 in the mixture if the total pressure of the atmosphere is to be 745 torr.

(b) If this atmosphere is to be held in a 121-L space at 295 K, how many moles of O_2 are needed?

Sample Exercise 10.11 Relating Mole Fractions and Partial Pressures

7

A study of the effects of certain gases on plant growth requires a synthetic atmosphere composed of 1.5 mol % CO₂, 18.0 mol % O₂ and 80.5 mol % Ar. (a) Calculate the partial pressure of O₂ in the mixture if the total pressure of the atmosphere is to be 745 torr. (b) If this atmosphere is to be held in a 121-L space at 295 K, how many moles of O₂ are needed?

Solve

(a) The mole percent is the mole fraction times 100. Therefore, the mole fraction of O₂ is 0.180.

$$P_{\text{O}_2} = (0.180)(745 \text{ torr}) = 134 \text{ torr}$$

(b) Tabulating the given variables and converting to appropriate units, we have:

$$P_{\text{O}_2} = (134 \cancel{\text{torr}}) \left(\frac{1 \text{ atm}}{760 \cancel{\text{torr}}} \right) = 0.176 \text{ atm}$$

$$V = 121 \text{ L}$$

$$n_{\text{O}_2} = ?$$

$$R = 0.08206 \frac{\text{L-atm}}{\text{mol-K}}$$

$$T = 295 \text{ K}$$

Solving the ideal-gas equation for n_{O_2} , we have:

$$\begin{aligned} n_{\text{O}_2} &= P_{\text{O}_2} \left(\frac{V}{RT} \right) \\ &= (0.176 \cancel{\text{atm}}) \frac{121 \cancel{\text{L}}}{(0.08206 \cancel{\text{L-atm/mol-K}})(295 \cancel{\text{K}})} \end{aligned}$$

$$n_{\text{O}_2} = 0.880 \text{ mol}$$

IDEAL GAS LAW APPLICATION

A compound has an empirical formula CHCl . If a 256-mL flask contains 0.800 g of the gas that was collected over water.

Determine the molecular formula if the temperature was 45°C and the barometric pressure was 712 torr.

DALTON'S LAW OF PARTIAL PRESSURE & STOICHIOMETRY

EXAMPLE:

In an experiment similar to lab #6 at LACC, 0.2898 g of aluminum metal is dissolved in 35 mL of 3.0 M sulfuric acid. The product gas was collected over water. What volume of gas was produced if the barometer in the room read 752.83 mmHg and 19.8 °C?

DALTON'S LAW OF PARTIAL PRESSURE & STOICHIOMETRY -2

EXAMPLE:

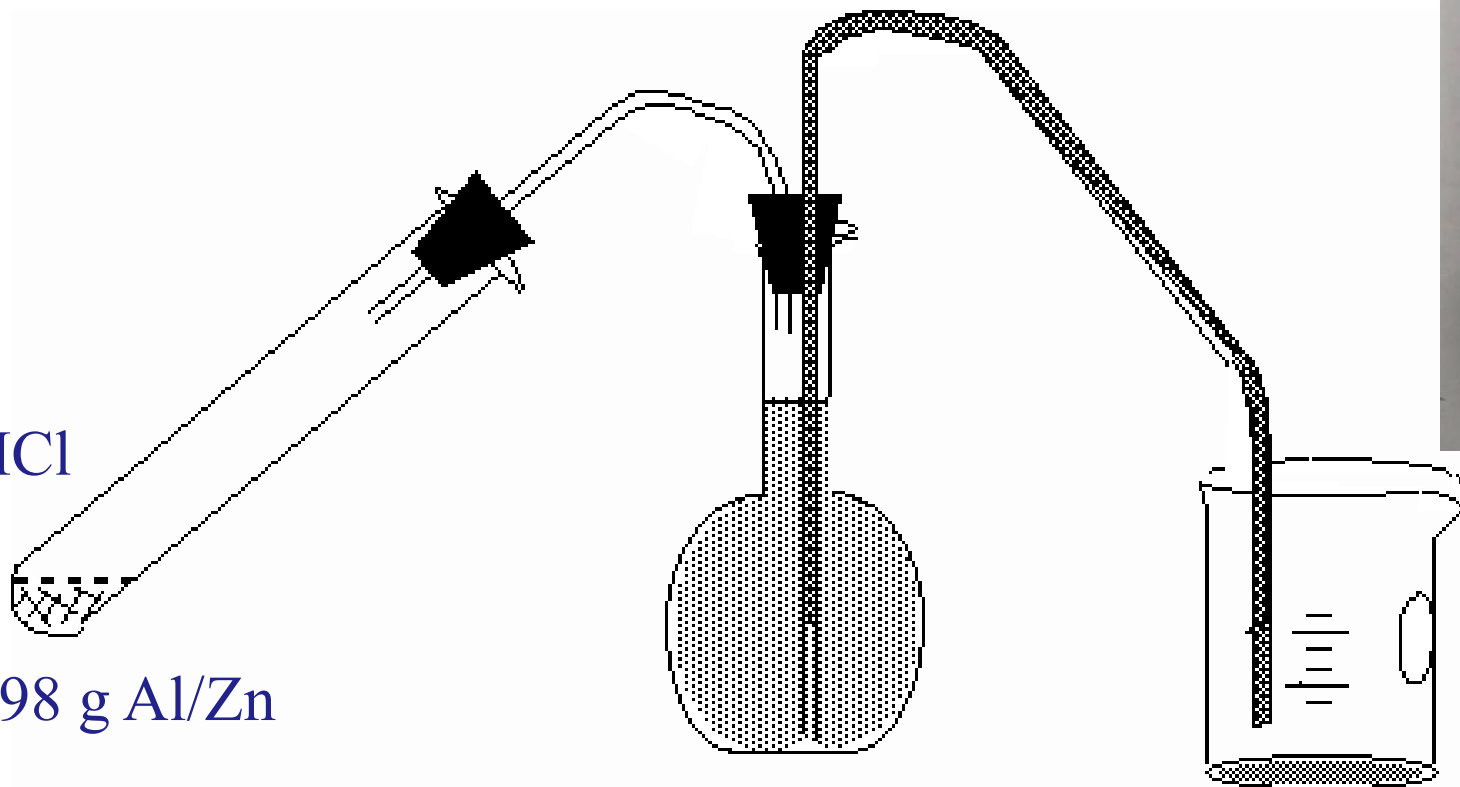
Since collected over water

$$P_{\text{gas}} = P_{\text{bar}} - P_{\text{H}_2\text{O}}$$

There is a mixture of hydrogen gas and water vapor

3.0M HCl

0.2898 g Al/Zn



The amount of water displaced = volume of gas produced



DALTON'S LAW OF PARTIAL PRESSURE & STOICHIOMETRY -4

ANSWER



0.2892g

35mL

V=?

27g/mol

3.0 M

$P_{\text{bar}} = 752.83 \text{ mmHg}$

$T = 19.8^\circ\text{C}$

Limiting Reactant Problem!

$$0.2892 \text{ g Al} (1 \text{ mol}/27 \text{ g/mol}) (3 \text{ mol H}_2 / 2 \text{ mol Al}) = n_{\text{H}_2} \quad \mathbf{0.01607 \text{ mol}}$$

$$0.035 \text{ L H}_2\text{SO}_4 (3.0 \text{ mol/L}) (3 \text{ mol H}_2 / 3 \text{ mol H}_2\text{SO}_4) = n_{\text{H}_2} \quad \mathbf{0.1050 \text{ mol}}$$

“over water” means $P_{\text{H}_2} = P_{\text{bar}} - P_{\text{H}_2\text{O}}$

$$P_{\text{H}_2} = 752.83 - \mathbf{17.54} = \mathbf{735.29 \text{ torr}} (1 \text{ atm}/760 \text{ torr}) = \mathbf{0.96749 \text{ atm}}$$

$$V = \frac{nRT}{P} \quad V = \frac{(\mathbf{0.016067 \text{ mol}})(\mathbf{0.082 \text{ L-atm/mol-K}})(\mathbf{19.8+273.15})}{\mathbf{0.96749 \text{ atm}}}$$

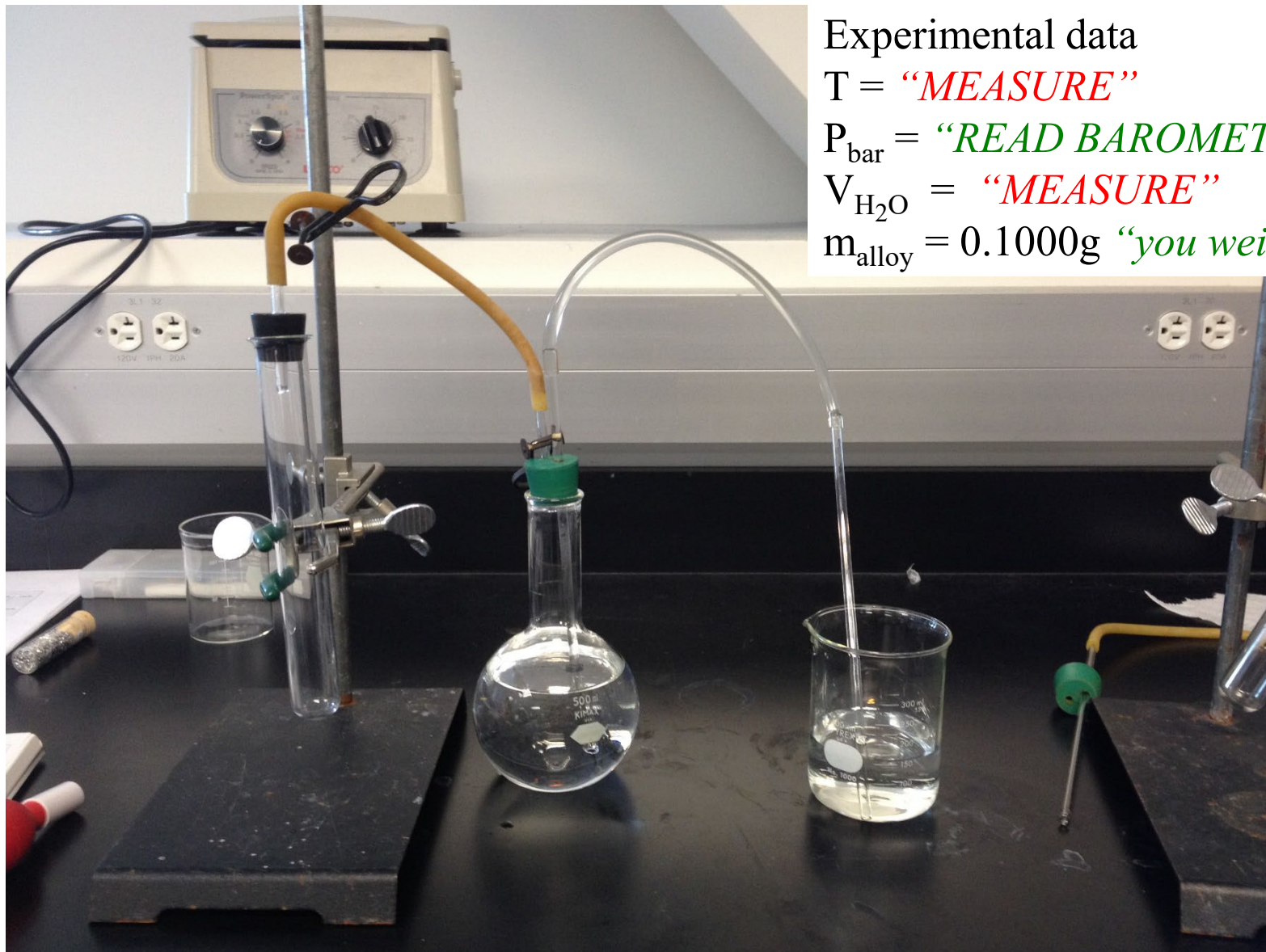
265 mL

DALTON'S LAW OF PARTIAL PRESSURE & STOICHIOMETRY -LECTURE

EXAMPLE:

In an experiment similar to lab #6 at LACC, 0.5000 g of aluminum metal is dissolved in 45.0 mL of 3.00 M sulfuric acid. The product gas was collected over water. What volume of gas was produced if the barometer in the room read 758.83 mmHg and 20.8 °C?

The set up for experiment 6



Experimental data

$T = \text{"MEASURE"}$

$P_{\text{bar}} = \text{"READ BAROMETER"}$

$V_{\text{H}_2\text{O}} = \text{"MEASURE"}$

$m_{\text{alloy}} = 0.1000\text{g}$ *"you weigh"*

The set up for experiment 6 -2

Experimental data

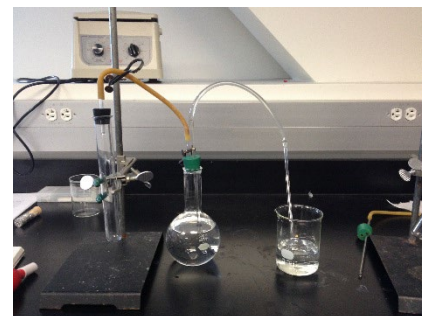
$T = 19.6^{\circ}\text{C}$

$P_{\text{bar}} = 758.9 \text{ mmHg}$

$V_{\text{H}_2\text{O}} = 115.2 \text{ mL}$

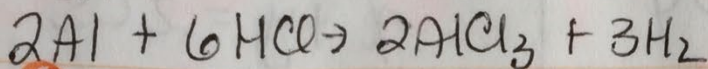
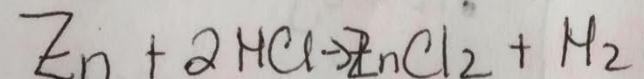
$m_{\text{alloy}} = 0.1000 \text{ g}$

Create a graph for 1.00 g sample using 0% Al, 0% Zn, and 50% of each. Then use the data to the left to determine the percentages of your alloy.



graphing

EXPT 6



Alloy = mixture

$$x\text{Zn} + y\text{Al} = 100\%$$

$$\textcircled{1} n_{\text{TOTAL}} = x n_{\text{H}_2} + y n_{\text{H}_2} \quad (x = \text{mole fraction}) = \frac{n_{\text{H}_2}}{n_{\text{T}}}$$

$$\textcircled{2} n_{\text{H}_2, \text{total}} = \left[n_{\text{Al}} \times \frac{3 \text{ mol H}_2}{2 \text{ mol Al}} \right] + \left[n_{\text{Zn}} \times \frac{1 \text{ mol H}_2}{1 \text{ mol Zn}} \right]$$

$$\textcircled{3} n_{\text{Al}} \left(\frac{3 \text{ H}_2}{2 \text{ Al}} \right) = n_{\text{H}_2, \text{from Al}}$$

Note

$$\textcircled{2} n_{\text{H}_2, \text{total}} = n_{\text{H}_2, \text{from Al}} + n_{\text{H}_2, \text{from Zn}}$$

$$\textcircled{4} y = mx + b$$

$y = n_{\text{H}_2, \text{total}} \quad x =$



LECTURE QUIZ #19B problems: reactions

1. A syringe containing 50.0 mL of vacuum weighs 75.212 g. The same syringe containing 50.0 mL of gaseous butane at a pressure of 0.923 atm and a temperature of 24 °C weighs 75.322 g. What is the molar mass of butane?

2c. What is the pressure (in kPa) in a 35.0 L balloon at 25.0 °C filled with pure hydrogen gas produced by the reaction of 34.11 g of CaH_2 with water?

DENSITY OF A GAS

The density of a gas at STP can be calculated by

$$d_{\text{STP}} = \text{molar mass/molar volume} = \text{MM}/V_m$$

The density of a gas **not** at STP can be calculated by

$$d = (\text{MM}) P / R T$$

or

$$dRT/P = \text{MM}$$

DENSITY OF A GAS -2

1) Calculate the density of hydrogen sulfite gas at STP.

$$d_{\text{(STP)}} = \text{MM}/V_{\text{m}} = (82 \text{ g/mol}) / 22.4 \text{ L/mol} = 3.66 \text{ g/L}$$

2) Identify an unknown homonuclear diatomic gas that was found to have a density of 3.165 g/L at STP.

DENSITY OF A GAS -3

$$d = (MM) P / R T$$

1. Calculate the density of hydrogen sulfite gas at 587 torr and 56.9 °C.

$$d = (82 \text{ g/mol} * 0.772 \text{ atm}) / (0.082 \text{ L-atm/mol-K} * 330 \text{ K}) = 2.34 \text{ g/L}$$

2. Identify an unknown homonuclear diatomic gas that was found to have a density of 1.950 g/L at 2.57 atm and 177°C.

Sample Exercise 10.8 Calculating the Molar Mass of a Gas

A large evacuated flask initially has a mass of 134.567 g. When the flask is filled with a gas of unknown molar mass to a pressure of 735 torr at 31 °C, its mass is 137.456 g. When the flask is evacuated again and then filled with water at 31 °C, its mass is 1067.9 g. (The density of water at this temperature is 0.997 g/mL.) Assuming the ideal-gas equation applies, calculate the molar mass of the gas.

1) Draw it out – label – and look for relationships

2) .
$$\mathcal{M} = \frac{dRT}{P}$$

Sample Exercise 10.8 Calculating the Molar Mass of a Gas

A large evacuated flask initially has a mass of 134.567 g. When the flask is filled with a gas of unknown molar mass to a pressure of 735 torr at 31 °C, its mass is 137.456 g. When the flask is evacuated again and then filled with water at 31 °C, its mass is 1067.9 g. (The density of water at this temperature is 0.997 g/mL.) Assuming the ideal-gas equation applies, calculate the molar mass of the gas.

Solve The gas volume equals the volume of water the flask can hold, calculated from the mass and density of the water. The mass of the water is the difference between the masses of the full and evacuated flask:

$$1067.9 \text{ g} - 134.567 \text{ g} = 933.3 \text{ g}$$

Rearranging the equation for density ($d = m/V$), we have $V = \frac{m}{d} = \frac{(933.3 \text{ g})}{(0.997 \text{ g/mL})} = 936 \text{ mL} = 0.936 \text{ L}$

Mass (gas) is the difference between the mass of the *flask + gas* and the mass of the evacuated flask:

$$137.456 \text{ g} - 134.567 \text{ g} = 2.889 \text{ g}$$

Knowing the mass of the gas (2.889 g) and its volume (0.936 L), we can calculate the density of the gas:

$$d = 2.889 \text{ g} / 0.936 \text{ L} = 3.09 \text{ g/L}$$

After converting pressure to atmospheres and temperature to kelvins, we can use Equation 10.11 to calculate the molar mass:

$$\begin{aligned} \mathcal{M} &= \frac{dRT}{P} \\ &= \frac{(3.09 \text{ g/L})(0.08206 \text{ L}\cdot\text{atm/mol}\cdot\text{K})(304 \text{ K})}{(0.967 \text{ atm})} \\ &= 79.7 \text{ g/mol} \end{aligned}$$

LECTURE QUIZ #19C

1. Given that a random sample of air was determined to be composed of N_2 , O_2 , and Ar with a mole fraction assay of 0.78, 0.21, and 0.010, respectively; What is the density of that air sample at standard pressure and temperature?

2. A 100.0 mL sample of air is analyzed and found to contain 0.835 g N_2 , 0.0640 g CO_2 and 0.197 g O_2 at 35 °C. What is the total pressure of the sample and the partial pressure of each component?

Practice Problems

A1. A gas is known to be either sulfur dioxide or sulfur trioxide. Its density at 98°C and 1.08 atm is 2.84 g/L. Which gas is it?



B1. Calculate the density of fluorine gas at 30.0 °C and 725 torr.

1.46 g/L

C1. What is the density of ethane gas at a pressure of 183.4 kPa and a temperature of 25.0 °C?

2.22 g/L

Kinetic Molecular Theory

- Matter is composed of tiny particles (atoms, molecules or ions) with definite and characteristic sizes that never change.
- The particles are in constant random motion, that is they possess kinetic energy. $E_k = \frac{1}{2} mv^2$
- The particles interact with each other through attractive and repulsive forces (electrostatic interactions), that is they possess potential energy. $U = mgh$
- The velocity of the particles increases as the temperature is increased therefore the average kinetic energy of all the particles in a system depends on the temperature.
- The particles in a system transfer energy from one to another during collisions yet no net energy is lost from the system. The energy of the system is conserved but the energy of the individual particles is continually changing.

Kinetic Molecular Theory of Gases – an explanation of the properties of an ideal gas in terms of the behavior of continuously moving molecules that are so small that they can be regarded as having no volume. This theory can be summed up with the following five postulates about the molecules of an ideal gas:

- 1. Gases are composed of molecules that are in continuous motion. The molecules of an ideal gas move in straight lines and change direction only when they collide with other molecules or with the walls of the container.**
- 2. The molecules of a gas are small compared to the distances between them; molecules of an ideal gas are considered to have no volume. Thus, the average distance between the molecules of a gas is large compared to the size of the molecules.**
- 3. The pressure of a gas in a container results from the bombardment of the walls of the container by the molecules of the gas.**
- 4. Molecules of an ideal gas are assumed to exert no forces other than collision forces on each other. Thus the collisions among molecules and between molecules and walls must be elastic; that is, the collisions involve no loss of energy due to friction.**

5. The average kinetic energy of the molecules is proportional to the Kelvin temperature of the gas and is the same for all gases at the same temperature.

Furthermore, the speed (or velocity) of these molecules can be related to temperature via the following, known as the root mean square speed (u_{rms})

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

where $R = 8.3145 \text{ J/K mol}$; $T = \text{temperature in Kelvin}$, and $M = \text{molar mass in kg/mol}$. The model of the kinetic molecular theory of gases is consistent with the ideal gas law and provides the aforementioned expression for the root mean square speed of molecules.

When combining root mean square speed with the expression for kinetic energy (which we know is $\frac{1}{2} mv^2$ PER MOLECULE), one can derive an equation for the kinetic energy of an ideal gas PER MOLE:

$$\text{KE (per mole)} = \frac{1}{2} mv^2 = \frac{1}{2} m \left(\sqrt{\frac{3RT}{M}} \right)^2 = \frac{3}{2} RT$$

Once again, we see that molar kinetic energy of a gas is proportional to the temperature.

How Fast Do Gas Molecules Move?

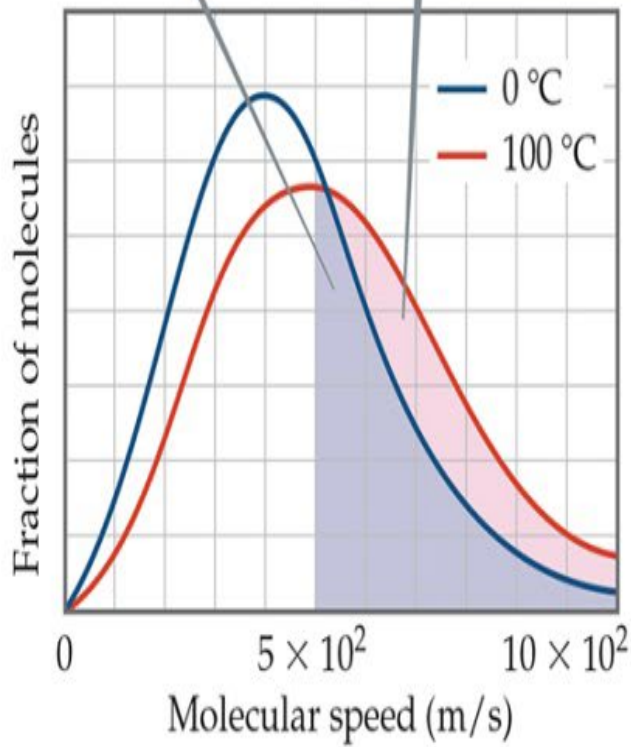
At 0 °C, fewer than half the molecules move at speeds greater than 500 m/s.

At 100 °C, more than half the molecules move at speeds greater than 500 m/s.

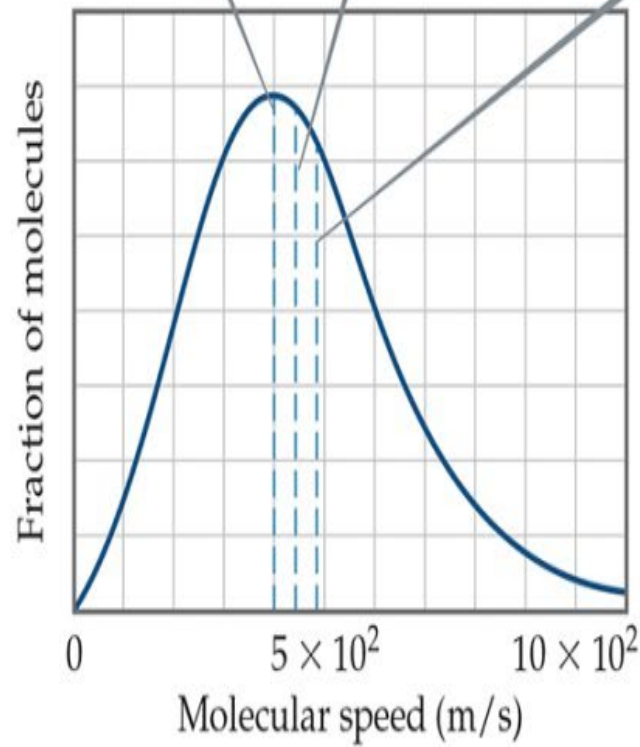
u_{mp} = the speed exhibited by the largest number of molecules

u_{av} = the average (mean) speed of all the molecules

u_{rms} = the speed of a molecule whose kinetic energy is equal to the average (mean) kinetic energy of all the molecules

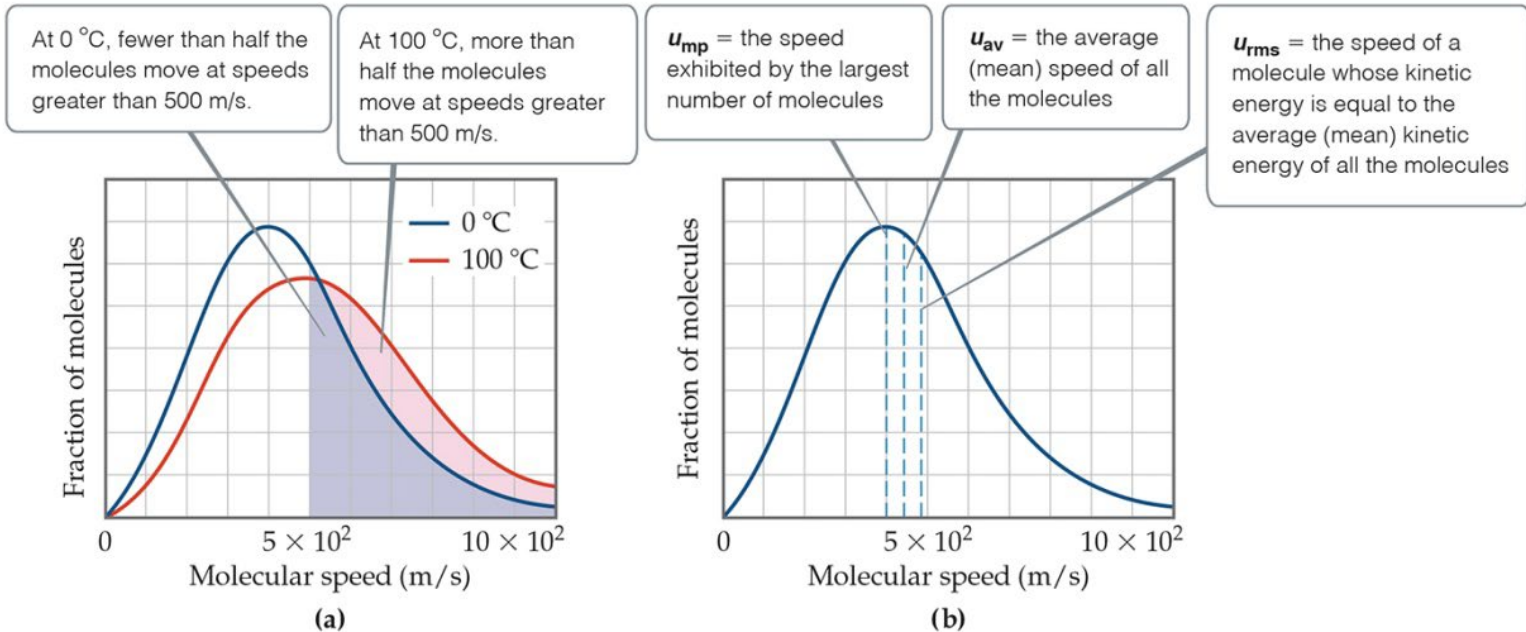


(a)



(b)

How Fast Do Gas Molecules Move? 2

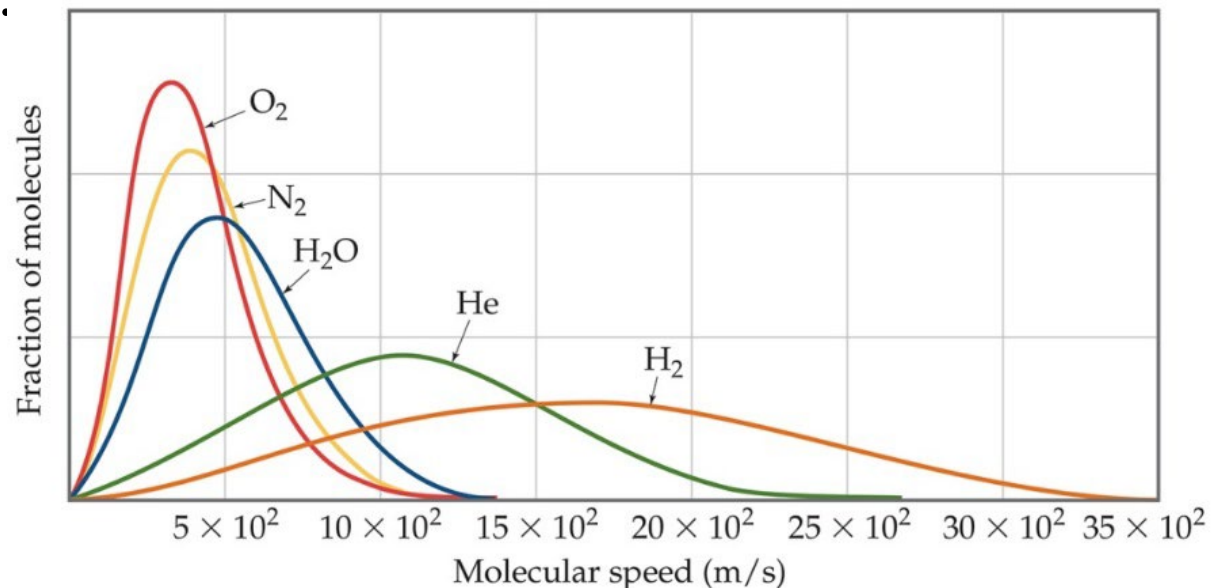


- Temperature is related to the **average** kinetic.
- Individual molecules can have different speeds.
- The figure shows three different speeds:
 - u_{mp} is the most probable.
 - u_{av} is the average speed of the molecules.
 - u_{rms} , the root-mean-square speed, is the one associated with their average kinetic energy.

u_{rms} and Molecular Mass

- At any given temperature, the average kinetic energy of molecules is the same.
- So, $\frac{1}{2}m(u_{rms})^2$ is the same for two gases at the same temperature.
- If a gas has a low mass, its speed will be greater than for a heavier molecule.

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



Calculate the rms speed of the molecules in a sample of N₂ gas at 25 °C

Solve We must convert each quantity in our equation to SI units. We will also use R in units of J/mol-K (Table 10.2) to make the units cancel correctly.

$$T = 25 + 273 = 298 \text{ K}$$

$$\mathcal{M} = 28.0 \text{ g/mol} = 28.0 \times 10^{-3} \text{ kg/mol}$$

$$R = 8.314 \text{ J/mol-K} = 8.314 \text{ kg-m}^2/\text{s}^2\text{-mol-K (Since } 1 \text{ J} = 1 \text{ kg-m}^2/\text{s}^2\text{)}$$

$$u_{\text{rms}} = \sqrt{\frac{3RT}{\mathcal{M}}}$$
$$= \sqrt{\frac{3(8.314 \text{ kg-m}^2/\text{s}^2\text{-mol-K})(298 \text{ K})}{28.0 \times 10^{-3} \text{ kg/mol}}}$$

$$\mathbf{5.15 \times 10^2 \text{ m/s}}$$

TABLE 10.2 Numerical Values of the Gas Constant R in Various Units

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

*SI unit

LECTURE

EXAMPLE on rms

Fill in the blanks for the following statement:

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

The rms speed of the molecules in a sample of H_2 gas at 300 K will be _____ times larger than the rms speed of O_2 molecules at the same temperature, and the ratio $u_{\text{rms}}(\text{H}_2)/u_{\text{rms}}(\text{O}_2)$ _____ with increasing temperature.

PROOF:

At 600K same ratio! A

$$\frac{1.93 \times 10^3 \text{ m/s}}{4.84 \times 10^2 \text{ m/s}}$$

(a) four, will not change

(b) four, will increase

(c) sixteen, will not change

(d) sixteen, will decrease

(e) Not enough information is given to answer this question.

$$\text{H}_2 \text{ } u_{\text{rms}} = 2735$$

$$\text{O}_2 \text{ } u_{\text{rms}} = 683.8$$

$$= 3.99 \text{ ratio}$$

$$u_{\text{rms}} = \sqrt{(3) (8.314 \text{ J/mol K}) (300 \text{ K}) / 2 \text{ g/1000 mol}} = 1.93 \times 10^3 \text{ m/s}$$

$$u_{\text{rms}} = \sqrt{(3) (8.314 \text{ J/mol K}) (300 \text{ K}) / 32 \text{ g/1000 mol}} = 4.84 \times 10^2 \text{ m/s}$$

Lecture example on RMS

What is the rms speed of an atom in a sample of H_2S gas at both 15.0 & 55.0 $^{\circ}\text{C}$?

Practice EXAMPLE

What is the rms speed of an atom in a sample of CO₂ gas at both 25 & 45.0 ° C?

$$u_{\text{rms}} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{(3) (8.314 \text{ J/mol K}) (25+273)}{44 \text{ g/mol (1kg/1000g)}}}$$

$$\mathbf{u_{\text{rms}} = 411 \text{ m/s}}$$

$$u_{\text{rms}} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{(3) (8.314 \text{ J/mol K}) (45+273)}{44 \text{ g/mol (1kg/1000g)}}}$$

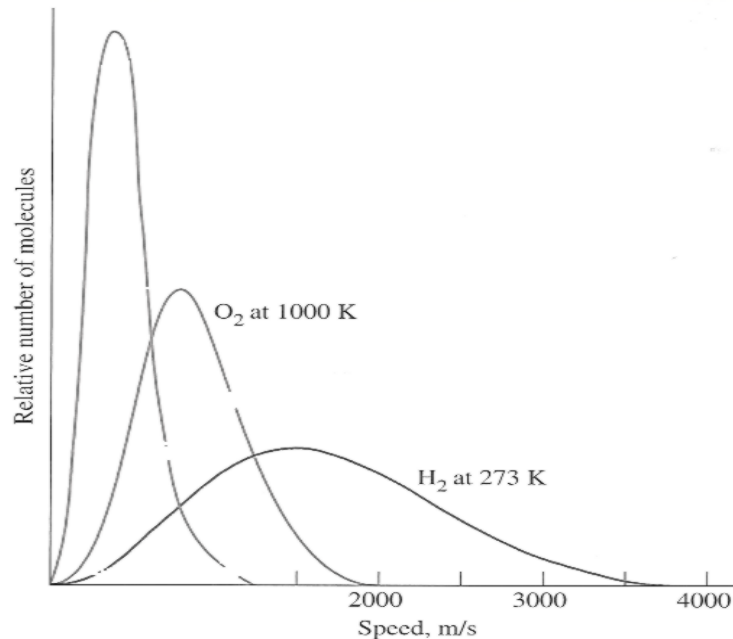
$$\mathbf{u_{\text{rms}} = 425 \text{ m/s}}$$

Note: the higher the temperature, the faster it moves!

Maxwell Distribution of Speeds – As useful as the root mean square equation is for most gases, it only represents an *average* speed. Individual molecules undergo several billion changes of speed and direction each second. The formula for calculating the fraction of gas molecules having a given speed at any instant was first derived by James Maxwell. You should be familiar with the conceptual implications of this equation:

1. The molecules of all gases have a wide range of speeds. As the temperature increases, the root mean square speed and the range of speeds both increase. The range of speeds is described by the Maxwell distribution (equation).

2. Heavy molecules (such as CO_2) travel with speeds close to their average values. The greater the molar mass, the lower the average speed and the narrower the spread of speeds. Light molecules (such as H_2) not only have higher average speeds, but also a wider range of speeds.



For example, some molecules of gases with low molar masses have such high speeds that they can escape from the gravitational pull of small planets and go off into space. As a consequence, hydrogen molecules and helium atoms, which are both very light, are rare in the Earth's atmosphere.