AP Chemistry

**Unit 4 – The Gaseous State**

**Chapter 10 Problem Set p. 432-437: 4, 9, 17, 35, 43, 47, 52, 62, 75, 77, 84**

*This unit focuses on the gaseous state and the gas laws: Boyle's law, Charles's law, and Avogadro's law. All three empirical gas laws combine into a fundamental equation called the ideal gas law. Describe the properties of a mixture of gases and see how to apply Dalton's law of partial pressure to characterize the components of a gas mixture. Examine the postulates of kinetic molecular theory and learn how to apply them to explain the empirical gas laws and the ideal gas law in molecular terms.*

**Objectives:**

**5.1 Apply the Kinetic Theory of Matter to explain how the temperature and pressure of gases relate to the kinetic energy according to the Maxwell-Boltzman Distribution.**

**5.2 Use the three empirical gas laws to predict how gases respond to changing conditions.**

**5.3 Calculate gas properties with the Ideal Gas Law.**

**5.4 Evaluate the pressure of individual gas components with Dalton’s Law.**

**5.5 Describe the conditions under which gases deviate from ideal behavior.**

*Skills to Master:*

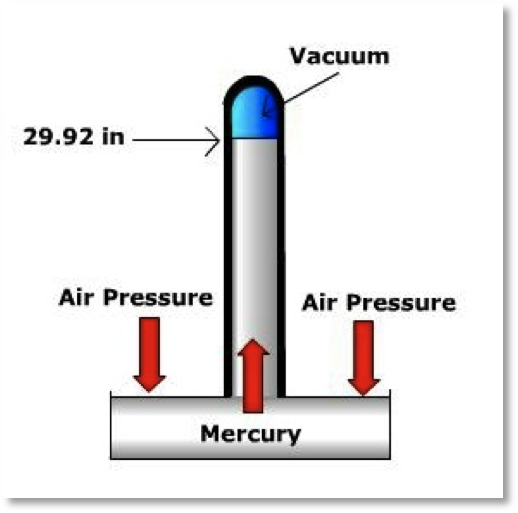
1. Distinguish a gas from a solid and a liquid.
2. Define the term pressure.
3. Explain gas pressure on the basis of the molecular nature of a gas.
4. Convert pressure from one unit to another.
5. Explain Boyle's law, Charles's law, and Avogadro's law in terms of the variables P, V, n, and T and the molecular behavior of gases. Apply these laws to predict the response of a gas to changing conditions.
6. Define standard conditions of temperature and pressure.
7. Manipulate variables in the ideal gas law to solve for unknown terms.
8. Apply the ideal gas law in numerical calculations involving gas density and molar mass.
9. Define the term "partial pressure." Write an expression for the total pressure of a mixture of gases in terms of the partial pressures of its components.
10. Calculate the partial pressure of a gas in a mixture, given the total pressure and the mole fraction of the gas.
11. Explain the molecular basis for the correction in the pressure of a gas collected over water.
12. Define the terms "effusion" and "diffusion."
13. Apply Graham's laws of diffusion and effusion to explain the behavior of gases in terms of their molar mass.
14. Summarize the characteristics of an ideal gas and explain the physical conditions under which a gas does not behave in an ideal manner. Evaluate the terms in the van der Waals equation for real gases for the correction to the ideal gas law.

**“Thankful for Chemistry” Project Due 11/08/2013**

1.      Research Three Substances: 1 Ionic Compound, 1 Molecular Compound, and 1 Element.  
2.      Make a PowerPoint Slide with RELEVANT images for each compound using the following template:  
  
I am thankful for Name of Compound  
Formula of Compound  
1.      Reason you are thankful for it  
2.      Reason you are thankful for it  
3.      Reason you are thankful for it.

**Podcast 5.1: Pressure**

Pressure:



Measured with a barometer.

Units of pressure

* 1 atmosphere = \_\_\_\_\_\_\_\_\_\_\_\_\_ mmHg
* 1 mm Hg = \_\_\_\_\_\_\_\_\_\_\_\_\_ torr
* 1 atm = \_\_\_\_\_\_\_\_\_\_\_ Pascals = \_\_\_\_\_\_\_\_\_\_ kPa

Example: What is 724 mm Hg in kPa?

* + in torr?
  + in atm?

**Podcast 5.2: The Gas Laws**

Factors needed to show the physical condition of a gas:

Boyle’s Law

* Pressure and volume are \_\_\_\_\_\_\_\_\_\_\_\_\_\_ related at constant \_\_\_\_\_\_\_\_\_\_\_\_\_\_.
* Sketch a graph to represent this kind of relationship

Equation:

Example 1: 20.5 L of nitrogen at 25ºC and 742 torr are compressed to 9.8 atm at constant T. What is the new volume?

Example 2: 30.6 mL of carbon dioxide at 740 torr is expanded at constant temperature to 750 mL. What is the final pressure in kPa?

Charles’s Law

* Volume of a gas varies \_\_\_\_\_\_\_\_\_\_\_ with the temperature at constant \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.
* Sketch a graph to demonstrate this type of relationship

Equation:

Example 3: What would the final volume be if 247 mL of gas at 22ºC is heated to 98ºC , if the pressure is held constant?

Example 4: At what temperature would 40.5 L of gas at 23.4ºC have a volume of 81.0 L at constant pressure?

Avogadro's Law

* Equal volumes of gases at the same temperature and pressure contain equal numbers of \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.
* 22.4 L = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ molecules
* At constant \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ and \_\_\_\_\_\_\_\_\_\_\_\_\_\_, the volume of gas is directly related to the number of moles.

Equation

Gay- Lussac Law

Pressure and temperature are \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ related at constant \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

Equation

Combined Gas Law

Equation

Example 5: A deodorant can has a volume of 175 mL and a pressure of 3.8 atm at 22ºC. What would the pressure be if the can was heated to 100.ºC?

Example 6: What volume of gas could the can release at 22ºC and 743 torr?

**Assignment #1 Gas Laws Example Problems**

1. During Hurricane Katrina, the atmospheric pressure (barometric pressure) within the eye dropped to 27.33 inches of mercury. Express this pressure in each of the following units:
   1. mmHg b. atm c. kPa d. torr
2. The volume of a balloon is 485 mL when filled with 0.022 moles of helium gas at a temperature of 20.0 oC. What is the pressure that the helium atoms are exerting on the sides of this balloon?
3. When 5.27 moles of nitrogen gas at 745 mmHg of pressure at a volume of 2.00 L is compressed to a new pressure of 300. kPa – what will be the new volume for the nitrogen?
4. Determine the temperature of the gas if a sample of oxygen gas with an original volume of 4.55 liters and at a temperature of -45 C has its volume reduced to 2.00L.

# Assignment #2 Gas Laws Practice Problems RALLY COACH

1. If the pressure of 50.0 liters of oxygen is 700 mmHg, what will the pressure be if the volume changes to 60.0 liters (temp. remains constant)?
2. If a gas occupies 25 mL at 800 mmHg and 0oC, what volume would it occupy at STP?
3. What would be the pressure of a 200 mL sample of hydrogen gas if a sample of hydrogen consisting of the same number of molecules has a volume of 300 mL when the pressure is 500 mmHg? Both samples are at the same constant temperature.
4. If 500 mL of nitrogen at STP is placed in a 1000 mL container, what would the new pressure be?
5. Calculate the volume of a gas if the original volume of 2.00 L at 700 mmHg is subjected to a pressure of 1000 mmHg. Assume constant temp.
6. If the pressure on 100 mL of a gas at 740 mmHg and 25oC is changed to 700 mmHg, what is the new volume?
7. If the temperature of 550 mL of oxygen changes from 25 oC to -25 oC, what is the volume of the gas? (assume constant pressure)
8. 5.00 L of nitrogen gas at 13 oC is transferred to a 10.00 L container. What is the new temperature of the gas?
9. 25.0 mL of nitrogen at 20 oC is cooled to standard temperature. What is its new volume if the pressure remains the same?
10. If 2.0 L of methane at 27 oC is heated to 127 oC at constant pressure, what would be the new gas volume?
11. 200.0 mL of butane at 200 oC is cooled to standard temperature. What is its new volume?
12. 1.00 L of hydrogen at STP is allowed to expand to 1.50 L. What is the temperature of this gas at standard pressure?
13. If a container of hydrogen at 15 mmHg and 23 oC is heated to 46 oC, what is the pressure of the gas?
14. Calculate the temperature of nitrogen gas if 2.0 L at STP is subjected to a pressure of 920 mmHg. Assume constant volume.
15. A container of oxygen at 22 oC and 740 mmHg is cooled to -22 oC. What is the new pressure of the gas?
16. A certain metal container of a gas will explode when the pressure reaches 5450 mmHg. If the container of a gas at standard temperature has a pressure of 750 mmHg, at what temperature would the container explode?
17. A gas has a volume of 50.0 mL when measured under a pressure of 740 mmHg at 25 oC. What will the volume be at STP?
18. If a 2.00 L sample of a gas at 700 mmHg and 27 oC is allowed to expand to 4.00 L at 54 oC, what will be the pressure under the new conditions?
19. A 25 mL sample of methane at 800 mmHg and 25 oC is subjected to a pressure of 600 mmHg at 100 oC. What is its new volume?
20. A cylinder contains 5.0 L of hydrogen gas at 7600 mmHg and 0 oC. What would be the volume of gas at standard pressure and a temperature of 27 oC?
21. What would be the temperature of a gas if 50.0 mL of it at STP were compressed under 1500 mmHg to 30.0 mL?
22. For a party, you need to fill 100 balloons with a capacity of 5.0 L each with helium. The barometer reading for the day is 740 mmHg and the temperature is 27 oC. Under what pressure would this gas be if it were bought in a 20.0 L cylinder at 20 oC?

**Podcast 5.3: The Ideal Gas Law**

Ideal Gas Law

Equation:

Standard Conditions: \_\_\_\_\_\_\_\_\_\_\_\_ atm, \_\_\_\_\_\_\_\_\_\_\_\_ oC , \_\_\_\_\_\_\_\_\_\_\_\_\_\_ moles

R = ideal gas constant

R = \_\_\_\_\_\_\_\_\_\_\_\_ L atm/ mol K

When to use PV=nRT

Values of R

R= 0.0821 L atm / mol K

R= 8.314 J / mol K

R = 1.987 cal / mol K

R=8.314 m3 Pa / mol K

R=62.36 L torr / mol K

Ideal Gases

* A hypothetical substance - the \_\_\_\_\_\_\_\_\_\_\_\_\_ gas whose pressure, volume, and temperature is completely described by the equation
* Think of it as a limit: Gases only approach ideal behavior at \_\_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_ (< 1 atm) and very high \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
* Even though gases are not ideal, use the laws anyway, unless told to do otherwise because they give good estimates.

Example 1: A 47.3 L container containing 1.62 mol of Helium is heated until the pressure reaches 1.85 atm. What is the temperature?

Example 2: Krypton gas in a 18.5 L cylinder exerts a pressure of 8.61 atm at 24.8ºC What is the mass of Kr?

Example 3: A sample of gas has a volume of 4.18 L at 29ºC and 732 torr. What would its volume be at 24.8ºC and 756 torr?

Density and Molar Mass

* The density of a gas can be determined:
* The higher the pressure and molar mass, the \_\_\_\_\_\_\_\_\_ dense the gas
* To determine Molar Mass of a gas:

Example 4: What is the density of ammonia at 23 oC and 735 torr?

Example 5: A compound has the empirical formula CHCl. A 256 mL flask at 100. oC and 750 torr contains 0.80 g of the gaseous compound. What is the molecular formula?

**Assignment #3 Ideal Gas Practice Problems RALLY COACH**

1. How many moles of gas does it take to occupy 120 liters at a pressure of 2.3 atmospheres and a temperature of 340K?
2. If I have a 50 liter container that holds 45 moles of a gas at a temperature of 2000 C, what is the pressure inside the container?
3. It is not safe to put aerosol canisters in a campfire, because the pressure inside the canisters gets very high and they can explode. If I have a 1.0 liter canister that holds 2 moles of gas, and the campfire temperature is 14000 C, what is the pressure inside the canister?
4. How many moles of gas in a 30 liter scuba canister if the temperature of the canister is 300K and the pressure is 200 atm?
5. I have a balloon that can hold 100 liters of air. If I blow up this balloon with 3 moles of oxygen gas at a pressure of 1 atm, what is the temperature of the balloon?
6. 6. Calculate the volume of 20.5g of NH3 at 0.658 atm and 25.00 C.
7. Calculate the volume of a 359g of CH3CH3 at 0.658 atm and 750 C.
8. A 2.00L container is place in a constant temperature bath and is filled with 3.05g of CH3OH. The pressure stabilizes at 800mmHg. What is the temperature of the water bath?
9. How many grams of O2 are needed to fill a 2.50L container at 104.7kPa at 250 C?
10. What is the pressure of SF6 when 3.75 mol of this gas are placed in a 150ml container at 500C

**Assignment #4 Ideal Gas Law**

1. Calculate the volume occupied by 2.5 moles of an ideal gas at STP.
2. The density of a gas was measured at 4.97 atm and 96.2 oC and found to be 0.873 g/L. Calculate the molar mass of this gas.
3. HCl (g) can be prepared by reacting NaCl with H2SO4. What mass solid NaCl is required to prepare enough HCl to fill a 340. mL cylinder to a pressure of 151 atm at 20.0 oC?
4. Ammonia, NH3, is generated by mixing hydrogen gas with nitrogen gas. What volume of ammonia can be generated if 30.5 liters of hydrogen at 143.0 oC and a pressure of 2.27 atm is mixed with excess nitrogen gas under the same conditions?

**Podcast 5.4: Applying the Gas Laws, Dalton’s Law and Vapor Pressure**

Gases and Stoichiometry

* Reactions happen in \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
* At Standard Temperature and Pressure (STP, 0 oC and 1 atm) 1 mole of gas occupies \_\_\_\_\_\_\_\_\_\_\_\_ L.
* If not at STP, use the \_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_ to calculate moles of reactant or volume of product.

Example 1: Mercury can be produced by the following reaction 

What volume of oxygen gas can be produced from 4.10 g of mercury (II) oxide

(a) at STP? (b) at 400. oC c) at 740 torr?

Example 2: Using the following reaction:



a) Calculate the mass of sodium hydrogen carbonate necessary to produce 2.87 L of carbon dioxide at 25 oC and 2.00 atm.

b) If 27 L of gas are produced at 26 oC and 745 torr when 2.6 L of HCl are added what is the concentration of HCl?

Example 3: Consider the following reaction



1. What volume of NO at 1.0 atm and 1000 oC can be produced from 10.0 L of NH3 and excess O2 at the same temperature and pressure?
2. What volume of O2 measured at STP will be consumed when 10.0 kg NH3 is reacted?

Example 4



1. What mass of H2O will be produced from 65.0 L of O2 and 75.0 L of NH3 both measured at STP?
2. What volume of NO would be produced?
3. What mass of NO is produced from 500. L of NH3 at 250.0 oC and 3.00 atm?

Dalton’s Law

* The total pressure in a container is the \_\_\_\_\_\_\_\_\_ of the pressure each gas would exert if it were alone in the container.
* The total pressure is the sum of the \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Partial Pressure: The pressure exerted by a particular component in a \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ of gases

* PTotal = P1 + P2 + P3 + P4 + P5 ...
* For each gas, P = nRT/V

Dalton's Law

In the same container R, T and V are the same, therefore.

Mole Fraction: the dimensionless number that expresses the \_\_\_\_\_\_\_\_\_\_\_\_ \_\_\_\_ \_\_\_\_\_\_\_\_\_\_ of a component substance to the total moles in a mixture.

symbol is Greek letter chi 

Example 5: The partial pressure of nitrogen in air is 592 torr. Air pressure is 752 torr, what is the mole fraction of nitrogen?

Example 6: What is the partial pressure of nitrogen if the container holding the air is compressed to 5.25 atm?

Vapor Pressure

* Collecting Gases over Water: commonly used so that pressure inside and outside can be equalized.
* Ptotal = Pgas + PH2O
* If equalization does not occur, then the pressure difference must be accounted for as well

Example 7: N2O can be produced by the following reaction 

What volume of N2O collected over water at a total pressure of 94 kPa and 22 oC can be produced from 2.6 g of NH4NO3? (the vapor pressure of water at 22 oC is 21 torr)

**Assignment #5 Dalton’s Law of Partial Pressures Practice Problems**

1. Mixtures of helium and oxygen can be used in scuba diving tanks to help prevent “the bends.” For a particular dive, 33 liters of helium at 20.0 oC and at 1.0 atm and 14 liters of oxygen at 20.0 oC and at 1.0 atm were pumped into a tank with a volume of 4.5 liters. Calculate the partial pressure of each gas and the total pressure in the tank at 20.0 oC.
2. The partial pressure of oxygen gas was observed to be 160. torr in the air with a total atmospheric pressure of 760. torr. Calculate the mole fraction of O2 present.
3. The mole fraction of nitrogen in the air is 0.7808. Calculate the partial pressure of N2 in the air when the atmospheric pressure is 760 torr.



1. A sample of solid potassium chlorate was heated in a test tube as shown above and decomposed by the following reaction:

2 KClO3 (s) 2 KCl (s) + 3 O2 (g)

The oxygen produced was collected by displacement of water at 22 oC at a total pressure of 754 torr. The volume of the gas collected was 0.650 liters, and the vapor pressure of water at 22 oC is 21.0 torr. Calculate the partial pressure of O2 in the gas collected **AND** the mass of KClO3 in the sample that was decomposed.

**Podcast 5.5: Kinetic Molecular Theory**

* Theories tell **why** the things happen.
* Explains why ideal gases behave the way they do.

1. The particles are in \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.
2. Collisions of particles cause \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.
3. Particles exert \_\_\_\_\_\_\_\_\_\_ attractive or repulsive forces on one another.
4. The average kinetic energy of the particles is proportional to the \_\_\_\_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

To describe temperature, use the formula for kinetic energy, KE = 1/2 mv2 and KE per mole = 3/2 RT

Temperature – \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_





Example 1: Calculate the average kinetic energy of the SO2 molecules in a sample of SO2 gas at (a) 295 K and at (b) 733 K.

Combine these two equations

(KE)avg = n(1/2 mυ 2 )

(KE)avg = 3/2 RT



Example 2: Calculate the root mean square velocity of carbon dioxide at 25ºC.

Example 3: Calculate the root mean square velocity of hydrogen at 25ºC.

Effusion and Diffusion

* \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_: the escape of gas through a tiny hole into evacuated space
* \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_: the spread of a substance throughout a given volume or throughout another substance

Graham’s Law: the rate of effusion is \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ proportional to the square root of the mass of its particles.

The less massive the gas, the higher the speed

Equation:

**Kinetic Molecular Theory of Gases Example Problems**

1. Calculate the average kinetic energy of a chlorine molecule in a sample of chlorine gas, Cl2 , at (a) 295 K and at (b) 733 K.
2. Calculate the root mean square velocity for the atoms in a sample of argon gas at 35.0 0C.
3. Calculate the root mean square velocity for the molecules SO2 molecules in a sample of SO2 gas at (a) 295 K and at (b) 733 K

**Assignment #6 Heat: Kinetic Molecular Theory**

**Learning Goals:** Students will be able to describe heat in terms of the motions of molecules and atoms. They will describe how the particle mass and temperature affects the model.They will compare the size and speed of gas particles to everyday objects, and identify the differences and similarities between solid, liquid, and gas particle motion.

**Procedures:** Go to <http://www.colorado.edu/physics/phet>.

1. Under “Chemistry,” open *Gas Properties* and then use the pump to put a little gas into the box.
   1. What happens to the clump of particles in the box?
   2. Use the heat control box to increase the temperature in the box of gas. What happens to the particles when you add heat? When you remove it? Try to make the particles stand still. Why do you find it difficult?
   3. Reset the simulation. Now pump in some blue (heavier) particles, and add some red (lighter) particles by clicking on the “light species” button. Describe the similarities and differences that you see between heavy and light particles.
   4. Write a description for your model of a heated gas based on your observations. Include diagrams to help with your description.
2. Summarize the behavior of gases as observed in this simulation.
3. Compare the rate of effusion of a light gas to that of a heavy gas using the following experiment. Set up the experiment by:
4. Selecting “Temperature” under “Constant Parameter”.
5. Keep gravity at zero.
6. Move the ruler close to the top of the gas chamber.
7. Select 100 heavy species particles and 100 light species particles and allow the particles to mix thoroughly.
8. Slide the lid on the top of the chamber to create a 0.2 nm opening and quickly toggle “Start” on the timer at the bottom of the screen.
9. Count the number of each type of particle in the box every 30 ps for 150 ps.
10. How fast do you think the air particles in this room are moving compared to a car going 50 miles per hour? Predictyour answer before testing it with the simulation: “I think an air molecule travels \_\_\_ as fast as a car.” To test your answer, adjust the molecules to a typical room temperature (293 K), click on the “species information” box, and assume the heavier species is the air molecule. How close were you? You will need to convert miles per hour to m/s (or vice versa):

50 miles 1 hr 1 min 1609 m m

------ x ----- x ------- x --------- = \_\_\_\_\_ -----

hr 60 min 60 sec 1 mile s

3. Now open a new simulation, *Microwaves*, which is found under the “Light & Radiation” section. You will study how a microwave heats your coffee. Click on the “coffee” tab at the top of the screen: you are now studying a liquid. Compare and contrast the behavior of the particles in the two simulations. You should find a few differences.

4. How do you think particles in a solid behave when heated? Come up with a theory of particle motion within a heated solid. Draw diagrams to help your explanation.

**Assignment #7** 

**Podcast 5.6: Deviations from Ideal Gas Behavior**

* Ideal gases are assumed to occupy \_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_\_ and have NO \_\_\_\_\_\_\_\_\_\_\_\_ for one another.
* Real gases have \_\_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ and DO attract one another
* Engineers and chemists that work with gases at \_\_\_\_\_\_\_\_\_\_\_ pressures cannot use the ideal gas law because the gases are not “ideal” at these conditions and must correct for \_\_\_\_\_\_\_\_\_\_\_\_\_\_

Van der Waals Equation

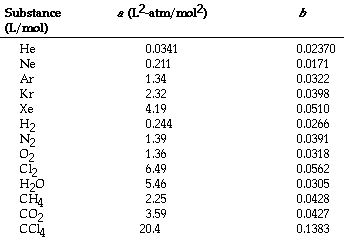
Correction Correction for

for volume molecular

of molecules attraction

Where a and b are Van der Waals constants which differ for each type of gas

Van der Waals Constants for Gaseous Molecules



Rearranged Van der Waals Equation

* a and b generally increase with an increase in mass of the molecule and with an increase in complexity of its structure (polarity)
* Larger more massive molecules have larger volumes and greater intermolecular attractive forces

Example: Calculate the pressure exerted by 0.5000 mol Cl2 in a 1.000 L container at 25.0ºC using the ideal gas law and Van der Waal’s equation

* a = 6.49 atm L2 /mol2
* b = 0.0562 L/mol

**Assignment #8 Unit 5 Review – Gases**

1. Calculate the density of manganese(III) sulfide vapor at 3025 torr and at a temperature of 1275 0C?
2. Calculate the root mean square velocity for the xenon atoms in a sample of xenon gas at –144.68 0C.
3. How many times faster would phosphorus penta-oxide gas diffuse than sulfur hexa-bromide gas?
4. Gas “X” diffuses one-sixth as fast as gas “Y”. Gas “Y” has a molecular mass = 18.572 g mol-1. Calculate the molar mass of gas “X” and identify the gas as either, CO2, Pb3(PO4)4 , Sn, Os2(SO4)3 , or O2
5. Propane gas, C3H8 , is collected over water at 30.0 0C. The atmospheric pressure on that day was recorded at 0.986 atm. of pressure. Calculate the volume of propane gas that must be collected to obtain 7.55 grams of propane gas? (At 30.0 0C the vapor pressure of water is 31.824 torr.)
6. A sample of inert gases mixed together with a pressure of 2580 torr contains 25.0 % nitrogen gas and 75.0 % radon gas by mass. What are the partial pressures of the individual gases?
7. Consider the flasks in the diagram to the below. In flask 1, 5.00 liters of N2 at 720 torr. In flask 2, 1.50 liters of Ne at 1.350 atm. In flask 3, 4.00 liters of argon gas at a pressure of 433 mm Hg. What is the final partial pressure of each gas after the valve between the three flasks is opened? And then calculate the total pressure after the valve is open. (Assume the final volume is 10.50 liters)



1. A mixture of 25.36 grams of nitrogen dioxide gas and 76.33 grams of sulfur trioxide are placed in a 43.50 liter container at 152.0 0C. Calculate the partial pressure of each gas and the total pressure.
2. A compressed gas cylinder contains 45.22 liters of neon gas at a pressure of 4732.0 torr at a temperature of 36.4 0C. What is the new volume if the pressure is reduced to 3000.0 torr and at a new temperature of 13.56 0C.
3. A 8.00 liter sample of butane gas, C4H10, reacts with a 24.0 liter sample of oxygen gas at 40.2 0C and at a pressure of 1.460 atm. Calculate the volume of the carbon dioxide gas formed at a new temperature of 103 0C and at a pressure of 3.10 atm.
4. The density of a gas was measured at 3.33 atm and 16.45 0C and found to be 4.97 g/L. Calculate the molar mass of this gas.
5. Hydrogen gas is produced from a chemical reaction and collected over water at 25 0C and 1.07 atm total pressure. What total volume of gas must be collected to obtain 3.80 grams of hydrogen gas? (At 25 0C the vapor pressure of water is 23.8 torr.)
6. Calculate the root mean square velocity for the molecules in a sample of radon gas at 220 K.
7. How many times faster would nitrogen gas, N2, diffuse than sulfur tri-oxide gas, SO3 ?
8. Calculate the average kinetic energy of the molecules in a sample of radon gas at 220 K.

**Gas Law Labs**

**PART I – Determine the Gas Constant, R**

When Boyle’s Law (relating pressure to volume) and Charles’ Law (relating temperature to volume) are combined, the resulting equation PV=nRT contains a constant of proportionality designated by R. This equation is for ideal gases, but most real gases under ordinary conditions conform quite well to ideality. The units of R depend on the units of the other quantities in the equation, but one useful R value is 0.082056 liter atm/mole Kelvin. To use this value the volume must be in liters, the pressure in atmospheres and the temperature in Kelvin.

In this experiment you will use the reaction:

Mg + 2HCl → MgCl2 + H2 (g)

and its stoichiometry to determine the quantity (number of moles) of hydrogen. This value along with measurements of the volume, pressure and temperature allows R to be calculated and compared with the accepted value. The hydrogen gas will be collected over water, so the pressure of the gas must be adjusted to discount the pressure due to the water vapor, i.e.,

Phydrogen = Ptotal - Pwater

If the levels of water in your buret and the beaker cannot be equalized the weight of the water column pulls downward on the gas trapped in the buret, reducing the gas pressure by the “hydrostatic” pressure of the water column. So the Phydrostatic must be subtracted, i.e.

Phydrogen = Ptotal –Pwater -Phydrostatic

**Procedures:**

1. Accurately weigh approximately 0.015 grams of magnesium ribbon.
2. Make sure the stopcock in your buret is closed, then add approximately 8 mL concentrated HCl to the buret. Wash down any acid on the walls of the buret with a wash bottle and slowly fill the buret completely with water.
3. Crumple your magnesium ribbon and wrap it in a small piece of copper wire. Place the cage in the top of the buret. Make certain the buret is filled to its brim with water.
4. Hold your finger over the buret and quickly invert it into a beaker of water. Clamp the buret in the beaker of water with the buret resting on the bottom of the beaker. **Wash your hands in running water.**
5. Allow the reaction to reach equilibrium (when there is no further production of hydrogen).
6. Measure the temperature of the water, record the gas volume in the buret, and note the differences in mm between the water levels in the beaker and the buret.
7. Remove the buret and measure the uncalibrated space in the bottom. Add this to the recorded volume.
8. Repeat the experiment.
9. From your recorded data determine the value of the gas constant in units of liter atm/mole K for each experiment trial. Average your class’s determinations (two for each student).

**Data**

Mg

|  |  |  |
| --- | --- | --- |
|  | **Trial One** | **Trial Two** |
| Temperature of Water Bath (K) |  |  |
| Hydrostatic Difference (mm) |  |  |
| Atmospheric Pressure (Pa) |  |  |
| Water Pressure (Pa) |  |  |
| Volume of Gas (L) |  |  |

**Data Analysis:**

1. Use stoichiometry to convert the mass of magnesium to moles of Hydrogen.
2. Convert your measured temperature of the water bath to Kelvin.
3. Calculate the pressure of the hydrogen gas:
   1. Phydrostatic = Hydrostatic difference (mm)÷ 13.6 (density of Hg)

\*\* note the units from hydrostatic pressure are in mmHg. You will need to convert these to Pa)

* 1. PH2O read from table below.
  2. Patm= use current weather data (up-to-date info from noaa)

Patm = PH2 + P H2O + Phydrostatic

1. Convert Volume of H2 gas to liters.
2. Calculate the value of R using the ideal gas law.
3. Calculate your error based on the published value for the universal gas constant. PAY CLOSE ATTENTION TO UNITS!

**PART II – Determine the Molecular Mass of Butane**

**Purpose:** To experimentally determine the molecular mass of butane.

**Materials:** 125 mL Erlenmeyer flask, wax pencil, trough, thermometer, graduated cylinder, butane lighter, and index card.

**Discussion:** According to Avogadro, the molar volume of any gas at STP is 22.4 liters. In this experiment the mass of a specific volume of butane gas will be determined. **Butane has the chemical formula C4H10.** This information can be used to determine the experimental mass of one mole of butane or the molecular mass of the gas.

**Procedure:**

1. Fill the trough ¾ full of water.
2. Dip the lighter in the water and then wipe it dry with a paper towel.
3. Measure and record the mass of the lighter.
4. Fill the Erlenmeyer flask completely with water.
5. Cover the top of the flask with the index card and invert the flask so that it is upside down in the trough of water.
6. Remove the index card.
7. Hold the butane lighter under the water so that it is directly under the falsk opening.
8. Press the release lever and collect approximately 100 + mL of gas.
9. Remove the lighter and dry it with a paper towel. Set aside.
10. Carefully raise and lower the flask until the water level inside the flask is equal to the water level outside the flask. this is to equalize the pressure so that the pressure inside is equal to the atmospheric pressure.
11. Mark the water level with a wax pencil.
12. Remove the flask and fill it with water up to the recorded mark.
13. Pour the water in the flask into a graduated cylinder and measure and record the volume of water. This is equal to the VOLUME OF GAS COLLECTED.
14. Measure and record the temperature of the water in the trough.
15. Measure and record the atmospheric pressure.

**Data**

|  |  |
| --- | --- |
| Mass of lighter before the experiment (g) |  |
| Mass of lighter after the experiment (g) |  |
| Mass of butane (g) |  |
| Volume of butane (L) |  |
| Temperature of butane (K) |  |
| Atmospheric Pressure (Pa) |  |

The gas in the flask is not all from the butane lighter. There is also WATER VAPOR. Since you are only interested in the pressure due to butane, you must subtract the pressure due to the water vapor. The table below will give you the water vapor pressure based on temperature. Use the equation below to calculate the pressure of DRY BUTANE.

Pbutane = Patmosphere - Pwater

|  |  |  |  |
| --- | --- | --- | --- |
| Temperature (oc) | Pressure, (Pa) | Temperature (oC) | Pressure (Pa) |
| 15 | 1705.6 | 23 | 2810.4 |
| 16 | 1818.5 | 24 | 2985.0 |
| 17 | 1938.0 | 25 | 3169.0 |
| 18 | 2064.4 | 26 | 3362.9 |
| 19 | 2197.8 | 27 | 3567.0 |
| 20 | 2338.8 | 28 | 3781.8 |
| 21 | 2487.7 | 29 | 4007.8 |
| 22 | 2633.7 | 30 | 4245.5 |

Pressure of Dry Butane \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Gas Law Lab Report Rubric

|  |  |  |
| --- | --- | --- |
| **Table of Contents** | **Points Earned** | **Points Possible** |
| Includes the title, page numbers, and date of experiment |  | 2 |
| **Title** |  |  |
| Capitalized appropriately, relates to the experiments, NOT underlined, but rather just centered at the top of the lab report |  | 1 |
| **Problem Statement** |  |  |
| * Independent variable, dependent variables, constants (at least 3), and the control are stated |  | 5 |
| * Purpose and Problem are testable and clearly stated |  | 2 |
| **Hypothesis** |  |  |
| States what you are doing, what you predict will happen, and why you think that will happen. If…Then…Because |  | 3 |
| **Materials** |  |  |
| A list of all materials used in the experiment |  | 1 |
| **Procedure** |  |  |
| Write a complete, **DETAILED** procedure. |  | 3 |
| **Data** |  |  |
| Organized table that shows the data you have collected during the experiment |  | 3 |
| **Prelab Questions** |  |  |
| 1. For a party, you need to fill 100 balloons with a capacity of 5.0 L each with helium. The barometer reading for the day is 740 mmHg and the temperature is 27 oC. Under what pressure would this gas be if it were bought in a 20.0 L cylinder at 20 oC? |  | 3 |
| 2. Define STP |  | 1 |
| 3. Calculate the density of manganese(III) sulfide vapor at 3025 torr and at a temperature of 1275 °C? |  | 3 |
| 4. Propane gas, C3H8 , is collected over water at 30.0 °C. The atmospheric pressure on that day was recorded at 0.986 atm. of pressure. Calculate the volume of propane gas that must be collected to obtain 7.55 grams of propane gas? (At 30.0 °C the vapor pressure of water is 31.824 torr.) |  | 5 |
| 5. A sample of inert gases mixed together with a pressure of 2580 torr contains 25.0 % nitrogen gas and 75.0 % radon gas by mass. What are the partial pressures of the individual gases? |  | 3 |
| Questions: |  |  |
| 1.   Using the Ideal Gas Law and measured values, calculate the number of moles of butane released into the flask. |  | 3 |
| 2.   The molecular mass is defined as grams/mole. Calculate the molecular mass of butane (using the mass in your data table). |  | 3 |
| 3.   Determine the actual mass of butane using the formula mentioned in the discussion. |  | 3 |
| 4.   Find the average molecular mass for the data collected by your class. |  | 3 |
| 5.   Evaluate your percent error. |  | 3 |
| 6. Determine whether outliers are present in the class data set. If necessary, calculate a new average and a new percent error after discarding outliers. |  | 5 |
| Calculations |  |  |
| Part One: Determine Gas Law Constant |  | 5 |
| Part Two: Determine the Molecular Mass of Butane |  | 7 |
| Part Three: Apply Statistical Analysis to Evaluate Reliability and Relevance of Data |  | 6 |
| **Conclusion** |  |  |
| Written in paragraph form (minimum of 3 paragraphs) |  | 1 |
| Support or refute your hypothesis. Give reasons why. USE YOUR DATA! |  | 5 |
| Discuss any EXPERIMENTAL error you may have had in the experiment. |  | 4 |
| Discuss how to change the design to fix the errors. What further questions or investigations does this lead to? |  | 4 |
| **Discussion/Reflection** |  |  |
| Discuss what you learned from this experiment and how it relates to what we are learning in class and applications in the real world (your world). Examples: Medicine, Pharmacy, Industry, Technology, Mining |  | 4 |
| Explain how the theoretical concepts we are learning in class directly apply to the lab experience. What is the ideal gas law? How did this lab verify it? How did Dalton's law of partial pressures allow you to calculate the pressure of individual gases (such as hydrogen or butane). |  | 3 |
| PRELAB BONUS! |  | up to 5 |
| **Total Points** |  | 94 |
| **Grade** |  |  |

**Determination of the Molar Mass of a Volatile Gas Lab**

In this experiment you will determine the molar mass of butane, the gas that is used as a fuel in disposable lighters and fuel canisters for camping stoves and lanterns. Because this gas is insoluble in water, it can be collected by displacement of water. The partial pressure of the butane collected is calculated using Dalton’s Law. The combination of Charles’ and Boyle’s laws is used to correct for pressure and temperature differences.

**Purpose:** To measure the volume of gas collected by indirect methods and to understand the importance of correcting an experimental gas volume to the STP volume.

To determine the molar mass of butane by using experimental data and gas laws.

**Pre-lab Question:**

A 200. mL sample of gas is collected over water at a temperature of 20 oC and a total pressure of 750. mm Hg. Determine the volume of the dry gas at STP.

**Materials:**

Erlenmeyer flask

Glass plate

thermometer

butane candle lighter

graduated cylinder

**Safety alerts: goggles, flammability – DO NOT perform this experiment if anyone is using a burner or flame in the room!**

**Procedure:**

1. Measure the *mass of the butane candle lighter*.
2. Fill the flask completely full of water. Measure and record the *volume of water* in the flask by pouring the water into a graduated cylinder. Then refill the flask with water and cover with the glass plate.
3. Fill the trough with water. Carefully invert the flask in the trough without getting any air bubbles in the flask.
4. Attach the plastic tubing to a butane candle lighter and hold the free end beneath the mouth of the inverted flask. Press the trigger on the lighter, holding onto the tube and allowing the butane to flow through to the flask.Make sure the bubbles go into the flask. Release enough gas to half-fill the flask.
5. Remove the tubing from the lighter. Determine the *final mass of the butane candle lighter*.
6. Raise or lower the flask until the water levels inside and outside the flask are equal. *Note what happens to the volume of the gas as you move the position of the flask.*
7. Put the glass plate over the top of the inverted flask. Turn it upright and take it to the fume hood to release the gas.
8. Measure the *volume of water remaining* in the flask.
9. Measure the *temperature of the water* in the flask.
10. Record the *barometric pressure* according to the NOAA website.
11. Look up the *vapor pressure of water* at the appropriate temperature.

**Data Table**

|  |  |
| --- | --- |
| Initial mass of butane lighter |  |
| Final mass of butane can |  |
| Volume of flask |  |
| Volume of water remaining |  |
| in flask after gas is released |  |
| Temperature of water |  |
| Barometric pressure |  |
| Vapor pressure of water |  |

**Calculations: show work for the following calculations, and put answers in the table.**

1. Calculate the volume of the gas collected *in liters*.

2. Determine the pressure of the *dry* gas. Remember that the barometric pressure is the total pressure of the gas + water vapor.

3. Calculate the temperature of the water bath in Kelvin.

4. Calculate the mass of the gas released into the flask by subtracting data collected.

5. Calculate the volume of the dry gas at STP (as in the pre-lab question).

6. Calculate the density of butane in grams per liter

7. Calculate the molar mass of butane in grams per mole, using the following conversion factor: 22.4 L = 1mole.

8. Look up the accepted value for the molar mass of butane in the Handbook of Chemistry and Physics and calculate the percent error.

**Questions:**

1. How would your results be affected if you had bubbles in your flask in step 2?

2. Why is it necessary to raise or lower the cylinder until the water levels inside and outside are equal? How would your results be affected if the water level inside the flask was higher than outside?

3. Why is it necessary to find the pressure of the dry gas in calculation #2? How would your results be affected if you did not make this correction to the pressure?

4. Why is it necessary to change the volume of the butane gas to STP in calculation #5?

5. Propane (C3H8) is another member of the hydrocarbon series. If you wanted to buy 1000 grams of compressed gas to take on a camping trip, would you get more moles of propane or butane?

6. You could not use this procedure to find the molar mass of oxygen or carbon dioxide. Why not?