CHAPTER 10 – Gases

Section 10.1 – Characteristics of Gases

(a) Name three characteristic properties of gases mentioned in this section.

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(b) These properties of gases can be explained by the fact that gas particles are relatively

______________________ from each other.

Section 10.2 – Pressure

(a) Write the definition of pressure as given by equation 10.1. ______________________

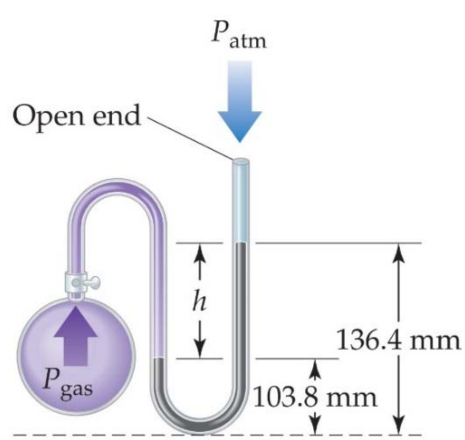
(b) ______________________________ can be defined as the force exerted by the atmosphere on a given surface area.

(c) In the 17th century, many scientists and philosophers believed that the atmosphere had no weight. Evangelista Torricelli was able to prove that air does in fact have weight. He invented the ____________________, which is illustrated in Figure 10.2.

(d) Although there are a wide variety of units that can be used for pressure, the AP Chemistry exam uses three units for pressure: atmospheres (atm), millimeters of mercury (mm Hg) and torr.

1 atm = __________ mm Hg = ___________ torr

(e) The diagram at right shows an open-end mercury manometer. Suppose that the atmospheric pressure in the laboratory is equal to 764.7 torr.



Is the gas pressure in the sealed flask higher than 764.7 torr or lower than 764.7 torr? ____________

How do you know? ________________________

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(f) Based on the information in part (e), determine the pressure of the gas in the sealed flask.

P_{gas} = ____________ torr = ____________ atm

(g) If you were to transport the open-end manometer shown in part (e) (Figure 10.3) to the top of a mountain, the value of *h* would (remain the same decrease increase) because

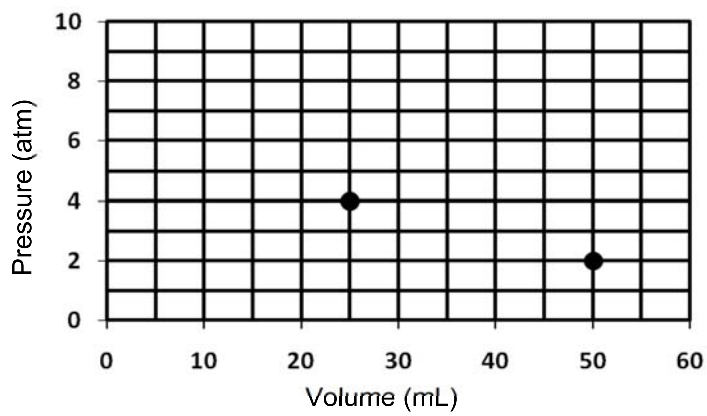
____

Section 10.3 – The Gas Laws

(a) Robert Boyle discovered that when the pressure on a fixed sample of gas is doubled (at constant temperature), the gas volume is ____

(b) Use the two points already on the graph to fill in the first two rows of the table below.
Use Boyle's Law to fill in the missing information, and plot the other two points on the graph.

Pressure (atm)	Volume (mL)
10.0	
	15.0



(c) Jacques Charles discovered the fact that the volume of a fixed sample of gas (at constant pressure) will _______________ when the temperature of the gas is increased.

(d) Using a ruler or a straight-edge, extrapolate the data in the graph below by drawing a line that passes through the data points and intersects the x-axis.

At what Celsius temperature would an ideal gas have a theoretical volume of zero?

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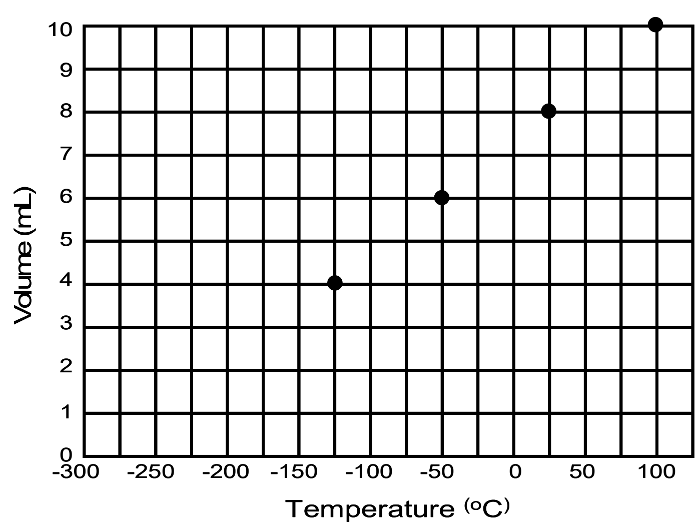


(e) Absolute zero (0 K) is equal to ______________ ^oC

If a 1-liter sample of hydrogen gas at 25^{o}C is heated to 75^{o}C (at constant pressure), the volume of gas should increase to 3 liters.

(f) Do you agree or disagree with the statement above? If you agree, justify your answer by performing a calculation using Charles' law. If you disagree, calculate what the new volume of hydrogen gas should be at 75^{o}C.

(g) Joseph Louis Gay-Lussac observed the law of combining volumes in 1808. Based on this law (and our knowledge of stoichiometry), we would predict the following results (at constant pressure and temperature).

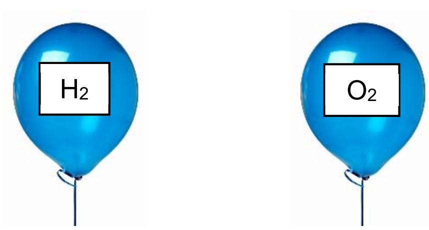
- 2 liters of H_2 gas should react with 1 liter of O_2 gas

  to produce ____ liters of water vapor.

- 1 liter of N_2 gas should react with 3 liters of H_2 gas

  to produce ____ liters of ammonia (NH_3) gas.

(h) Amedeo Avogadro interpreted Gay-Lussac's observation by proposing that equal volumes of gases at the same ____________________ and __________________ contain the same ____. This is known as Avogadro's hypothesis.



(i) Each balloon shown above has the same volume. The gases are at a pressure of 1.0 atm and a temperature of 0^{o}C. The balloon on the left contains 2.016 grams of $H_2(g)$ and the balloon on the right contains $O_2(g)$.

How many molecules of H_2 are present in the left balloon? ________________________

How many grams of O_2 are in the right balloon? ____________________

What is the volume of each balloon (in units of liters)? _______________

Each gas sample has the **same average kinetic energy** because they are both at the **same temperature**.

Are the molecules in each balloon travelling at the **same average speed**?
If you agree, explain why. If you disagree, explain why not.

Consider each of the following situations and decide whether or not certain variables will increase, decrease, or remain the same. Assume that the amount of gas is constant.

(j) A sample of gas is heated in a rigid container of constant volume.

The pressure of the gas should ____________________

The density of the gas should _____________________

The average distance between the gas particles should _________________

The average speed of the gas particles should ________________________

(k) A sample of gas is in a cylinder fitted with a movable piston. The gas volume is reduced from 1.0 L to 0.50 L at constant temperature.

The pressure of the gas should ____________________

The density of the gas should _____________________

The average distance between the gas particles should _________________

The average speed of the gas particles should ________________________

(l) A sample of gas is in a rigid container of constant volume. Additional gas is injected into the container at constant temperature.

The pressure of the gas should ____________________

The density of the gas should _____________________

The average distance between the gas particles should _________________

The average speed of the gas particles should ________________________

Section 10.4 – The Ideal Gas Equation

(a) The ideal gas law is written as ___________________

(b) If we use a value of 0.08206 for the gas constant R, we must make sure that pressure is in units of ___________, volume is in units of _________, and temperature is in units of ___________.

(c) If we use a value of 62.36 for the gas constant, then the only units that will change (relative to the information in part (b) will be the units for _____________, which become _____________.

(d) Use the ideal gas law to calculate the pressure inside a helium tank that has a volume of 17.8 liters and contains 30.0 grams of helium at a temperature of 22°C.

(e) Use the combined gas law (Equation 10.8 on p. 395) to solve the following problem.

An inflated balloon has a volume of 30. L at sea level (1.0 atm). During its vertical ascent, the temperature changes from 25°C to –19°C, and the volume of the balloon increases to 55 L. What is the atmospheric pressure when the balloon has reached this higher altitude?

Section 10.5 – Further Applications of the Ideal Gas Equation

(a) The density (d) of a gas is normally measured in units of grams per liter. $d = \frac{mass}{V}$

Show how the density equation can be converted into this one: $d = \frac{MP}{RT}$

where *M* = molar mass

(b) The density of dry air at STP is 1.29 g/L. What is the density of dry air at 100°C (assuming that pressure remains at 1.00 atm)?

(c) Does your answer to part (b) seem logical, based on what you know about hot air balloons? Explain.

The following free response question was included on the 2009 AP Chemistry Exam.

A student was assigned the task of determining the molar mass of an unknown gas. The student measured the mass of a sealed 843 mL rigid flask that contained dry air. The student then flushed the flask with the unknown gas, resealed it, and measured the mass again. Both the air and the unknown gas were at 23.0°C and 750. torr. The data for the experiment are shown in the table below.

Volume of sealed flask	843 mL
Mass of sealed flask and dry air	157.70 g
Mass of sealed flask and unknown gas	158.08 g

(a) Calculate the mass, in grams, of the dry air that was in the sealed flask. (The density of dry air is 1.18 g L^{-1} at 23.0°C and 750. torr.)

(b) Calculate the mass, in grams, of the sealed flask itself (i.e., if it had no air in it.)

(c) Calculate the mass, in grams, of the unknown gas that was added to the sealed flask.

(d) Using the information above, calculate the value of the molar mass of the unknown gas.

(e) After the experiment was completed, the instructor informed the student that the unknown gas was carbon dioxide (44.0 g/mol). Calculate the percent error in the value of the molar mass calculated in part (d).

(f) For each of the following two possible occurrences, indicate whether it by itself could have been responsible for the error in the student's experimental result. You need not include any calculations with your answer. For each of the possible occurrences, justify your answer.

<u>Occurrence 1:</u> The flask was incompletely flushed with $CO_2(g)$, resulting in some dry air remaining in the flask.

<u>Occurrence 2:</u> The temperature of the air was 23.0°C, but the temperature of the $CO_2(g)$ was lower than the reported 23.0°C.

(g) Describe the steps of a laboratory method that the student could use to verify that the volume of the rigid flask is 843 mL at 23.0°C. You need not include any calculations in your answer.

One application of the ideal gas law is a situation involving stoichiometry calculations where gases are consumed or formed. In an automobile air bag, the chemical reaction that takes place to fill the bag is the decomposition of sodium azide (NaN_3) to produce sodium metal and nitrogen gas.

Calculate the volume of nitrogen gas that is produced at 25°C and 1.1 atm from the complete decomposition of 75 grams of sodium azide.

<u>Section 10.6 – Gas Mixtures and Partial Pressures</u>

(a) Dalton's Law of partial pressures states that the total pressure of a mixture of gases is equal to

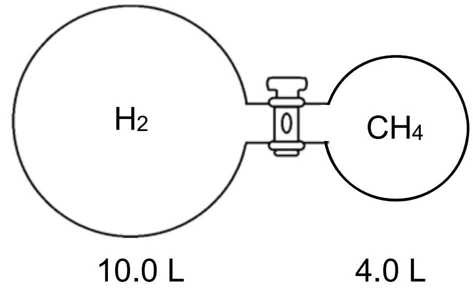
____

(b) Write the expression for the mole fraction of a gas (*X*) as defined in Equation 10.15 on page 400.

(c) How can you use the partial pressure of gas 1 in a gas mixture (P_1) and the total pressure of the gas mixture (P_{total}) to calculate the mole fraction of gas 1?

(d) A mixture contains equal masses of oxygen gas and helium gas. The total pressure of the gas mixture is 2.70 atm. What is the partial pressure of each gas?

(e) In the diagram at right, two flasks are connected by a valve. The flask on the left contains 10.0 L of H_2 gas at a pressure of 3.0 atm. The flask on the right contains 4.0 L of CH_4 gas at a pressure of 2.0 atm. Calculate the total pressure of the system after the valve is opened. Assume that the temperature remains constant at 25°C and that the volume in between the flasks is negligible.



(f) Calculate the mole fraction of each gas in the mixture created in part (e).

The mass of a butane lighter was recorded as 22.24 g. It was used to collect butane gas in a tube by the method of water displacement. The temperature of the water bath was 22°C. The barometric pressure in the laboratory was 755.0 mm Hg. After the experiment, the mass of the butane lighter was recorded as 22.03 g.

(g) Once this experiment was complete, the height of the collection tube was adjusted so that the water levels inside and outside the tube were the same. Explain why this was done.

(h) Use Appendix B on p. 1058 to determine the vapor pressure of water under the conditions

of this experiment: ____________________

(i) Use equation 10.18 on p. 402 to calculate the partial pressure of the butane gas that was collected in this experiment.

(j) The volume of butane gas collected was 100.0 mL. Calculate the amount of butane gas (in units of moles) that was collected in this experiment. Assume that the temperature of the butane is the same as the temperature of the water bath.

(k) Calculate the experimental value for the molar mass of butane.

(l) Calculate the percent error for this experiment. (butane = C_4H_{10})

<u>Section 10.7 – The Kinetic-Molecular Theory of Gases</u>

(a) Summarize the five statements of the KMT.

(i) ____

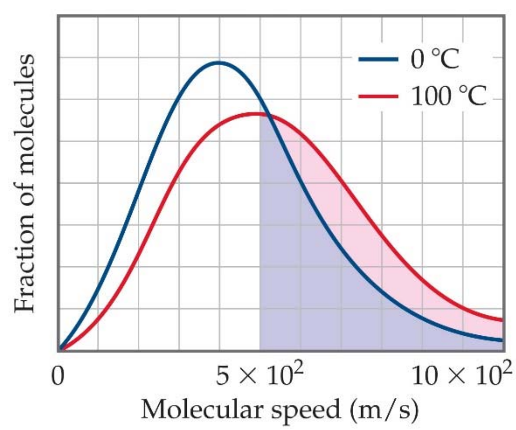
(ii) ____

(iii) ____

(iv) ____

(v) ____

(b) The diagram at right can be found on the bottom of p. 403. It shows the distribution of molecular speeds for nitrogen gas at two different temperatures.



(a)

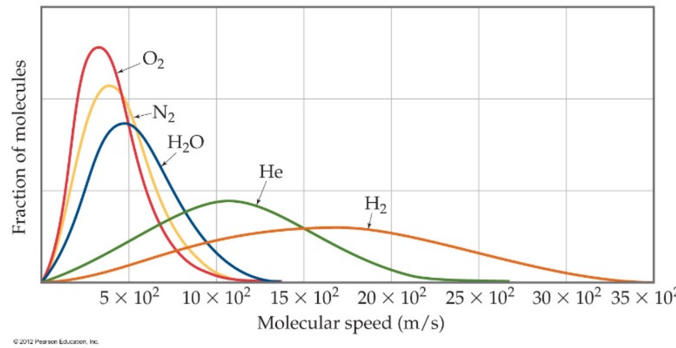
True or False: At a given temperature, each molecule in a gas sample is moving at the same speed.

True or False: As the temperature of a gas sample is increased, the average speed of the gas molecules increases.

True or False: The distribution curve tends to broaden as the temperature is increased, indicating that the range of molecular speeds increases at a higher temperature.

(c) The average kinetic energy per molecule of a gas is equal to ½ mv^2 (where m = mass and v = velocity). This information helps us to compare the average speeds of different gases at a given temperature. How should the following gases be arranged in order of increasing average speed at 298 K: $N_2(g)$, Ne(g), $CO_2(g)$

(d) What does the diagram below tell us about the relationship between molecular speed and molar mass?



Section 10.8 – Molecular Effusion and Diffusion

Even though a quantitative calculation involving Graham's law is not likely to appear on the AP Chemistry exam, it is still a good idea for you to be able to use this equation (10.24 on p. 407).

A certain gas has a rate of effusion that is twice as fast as that of methane, CH_4. Calculate the molar mass of this gas.

Section 10.9 – Real Gases: Deviations from Ideal Behavior

(a) Gases tend to experience the largest deviations from ideal behavior under conditions of

_____________ pressures and __________ temperatures.

(b) The particles of an ideal gas are assumed to occupy no space and have no attractive forces.

However the particles of a ___________ gas have finite volumes and experience some

attractive forces.

(c) Explain why under conditions of high pressure, the volume of a gas tends to be slightly greater than the value predicted by the ideal gas law equation.

(d) Explain why under conditions of low temperature, the pressure of a gas tends to be slightly smaller than the value predicted by the ideal gas law equation.

(e) In general, gases tend to experience larger deviations from ideal behavior as the strength of their intermolecular attractive forces ____________________. For this reason, which of the following molecules is more likely to show a deviation from ideal gas behavior: $H_2(g)$ or $H_2O(g)$? Justify your answer.