Chapter 12 – The Behavior of Gases
Part 1 – Notes: Introduction to Pressure and Its Units

Objectives: To explain and define pressure and how it relates to and applies to the gaseous state.
Explain how the amount of gas and the volume of the container affect gas pressure.
Explain the effect of temperature changes on the pressure exerted by a gas sample.
Differentiate between force and pressure.
Identify, define, and explain: pressure, force, atmospheric pressure, and atmosphere (unit of pressure).
Text Reference: Section 12.2 – pages 330-332

Matter on earth occurs basically in three states or phases – solid, liquid, and gas.

Properties of a **GAS**:

What does a gas "look like" on a molecular level?

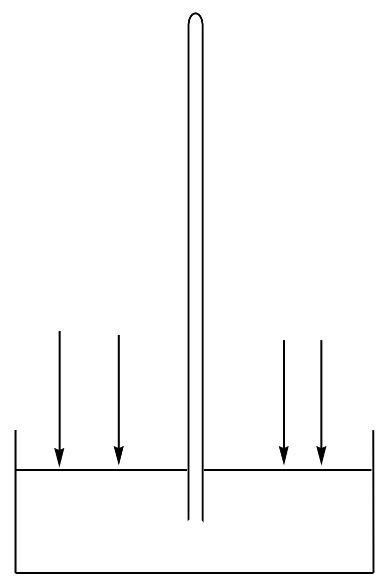
What may be said about the density of a gas?

All gases exert pressure. At any point, a gas exerts an equal **pressure** in all directions at any point within a gas. Gases at rest exert a pressure equal to the pressure exerted on them.

PRESSURE is defined as a force per unit area.

A **FORCE** is a push or a pull.

The difference between a force and pressure is important. You might not object if someone pushed against your shoulder with an open palm with a force of 10 pounds. However, if someone pushed the point of a knife against your shoulder with a force of 10 pounds, you would surely object. The same force exerted over the tiny area of the point of a knife would cause serious damage.

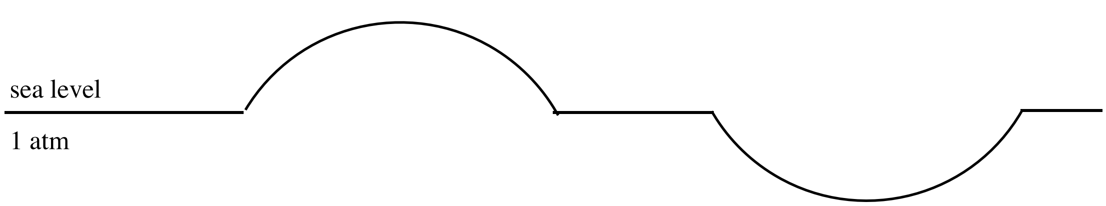


A **BAROMETER** is a simple piece of equipment used to measure pressure.

A simple barometer may be made by obtaining a long tube sealed at one end and filling it completely with mercury. The tube is then inverted and the open end is placed in a dish filled partially with mercury and the tube is held vertically. The mercury in the tube falls to a height, ***h***, determined by the pressure of the air on the surface of the mercury in the dish. There is essentially no pressure in the tube above the mercury.

There are two forces on the mercury in the tube – the force of gravity pulling down and the force due to the air pressure pushing up. When these two forces balance each other, the mercury stops falling. The greater the air pressure, the higher the mercury stands in the tube above the level of mercury in the dish.

HOW DOES ATMOSPHERIC PRESSURE VARY WITH ALTITUDE???



UNITS OF PRESSURE

A simple units of pressure is **MILLILMETERS OF MERCURY**, abbreviated *mm Hg*. The vertical height of the mercury in the column, corresponding to the *h* in the figure, is a measure of pressure.

TORR – a unit of pressure equal to the pressure needed to support 1 mm Hg is the torr – named for Evangelista Torricelli – an Italian physicist who discovered the principle of the barometer. NOTE: 1 torr = 1 mm Hg.

ATMOSPHERE (*atm*) or **STANDARD ATMOSPHERE** – at sea level at $0^{o}C$ on a "normal" day, the atmosphere can hold the mercury at a height of 760. mm; its pressure is 760. torr. The standard atmosphere is defined as 760. torr.

1 atm = 760. torr = 760. mm Hg Note there are 3 significant digits in 760. mm Hg.

Note the difference between 1 atmosphere of pressure and ***atmospheric pressure****. The first is a constant and the second varies widely from place to place and even varies over time at the same place. Atmospheric pressure is often referred to as* ***barometric pressure****.*

PASCAL – the SI unit of pressure is the pascal, abbreviated *Pa*. It is such a small unit that the ***kilopascal*** is most often used for atmospheric pressure under ordinary circumstances. 1 kilopascal = 1000 pascals. In order to convert from pascals to atmospheres and torr, you need to know the following equivalencies:

$$1 \text{ atm} = 101.3 \text{ kPa} = 1.013\text{x}10^{5} \text{ Pa}$$

POUNDS PER SQUARE INCH – abbreviated *psi* – used to describe how many pounds (force) are exerted over a square inch (an area). This is the most direct unit of pressure. The conversion is: 14.69 psi = 1 atm

CONVERSIONS BETWEEN UNITS OF PRESSURE:

Example 1: Convert 45.998 kPa to the equivalent pressure in atmospheres.

Example 2: Convert 3987 mm Hg to the equivalent pressure in atmospheres.

Example 3: Convert 6594 Pa to the equivalent pressure in torr.

Example 4: Convert 2.87 kPa to the equivalent pressure in mm Hg.

WHAT CAUSES A GAS TO HAVE PRESSURE?????

HOW DO YOU CHANGE THE PRESSURE OF A GAS?????

Chapter 12 – The Behavior of Gases
Part 1 – Assignment: Introduction to Pressure and Its Units

Solve the following pressure conversion calculations using dimensional analysis.

1. 432 mm Hg = ??? atm

2. 0.682 atm = ??? Pa

3. 1.07925 atm = ??? psi

4. 1.21 atm = ??? mm Hg

5. 8367 Pa = ??? mm Hg

6. 518 torr = ??? kPa

7. Why is atmospheric pressure lower at the top of a mountain than it is at sea level?

8. What causes the pressure of a gas. How would increasing the molecule's speed affect the pressure of the gas?

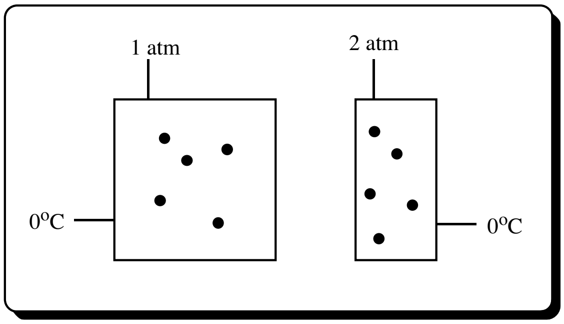
9. How is pressure different from weight?

10. ***On a separate sheet of paper answer the following:***
You are performing an experiment to determine the boiling point of an unknown substance. You perform six trials.
The results are: 115.67^{o}C, 115.24^{o}C, 114.88^{o}C, 116.13^{o}C, 115.85^{o}C, and 115.04^{o}C.
The actual boiling point of the substance is 122.56^{o}C?
    a. Are the measurements precise? Explain.
    b. Are the measurements accurate? Explain.
    c. Calculate the mean.
    d. Calculate the average absolute deviation.
    e. Calculate the relative deviation.
    f. Calculate the absolute error of the mean of the measurements.
    g. Calculate the percent error of the measurements.

Chapter 12 – The Behavior of Gases
Part 2A – Notes: The ABCs of Gases – Boyle's Law

Objectives: Investigate the relationship between the volume and the pressure of a given sample of a gas at a set temperature.
Apply gas laws to problems involving the temperature, volume, and pressure of a contained gas sample.
State Avogadro's hypothesis and use it to answer conceptual problems involving gases.
Identify, define, and explain: Kelvin, absolute temperature scale, direct relationship, and inverse relationship.

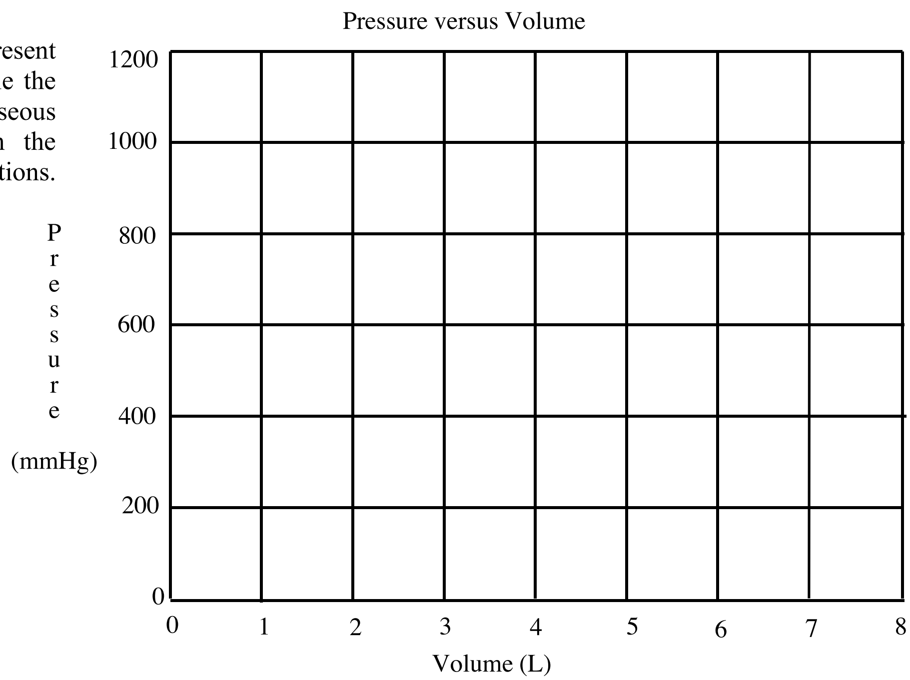
Text Reference: Section 12.2 (Part) – pages 330-332, Section 12.3 (Part) – pages 333-338, Section 12.5 (Part) – pages 347-349.



1. What observations may be made about the gas in the two containers?

2. The following sets of data represent pressure and volume readings taken while the amount and the temperature of a gaseous substance were held constant. Graph the following data points and answer the questions.

Pressure (mm Hg)	Volume (L)
1200	1.04
1000	1.25
800	1.563
600	2.083
400	3.125
300	4.167
200	6.25



3. From the graph and the illustration, state *Boyle's Law* in your own words.

4. According to the graph, what is the volume of the gas when the pressure is 900 mm Hg?

According to the graph, what would be the pressure of the gas when the volume is 2.500 L?

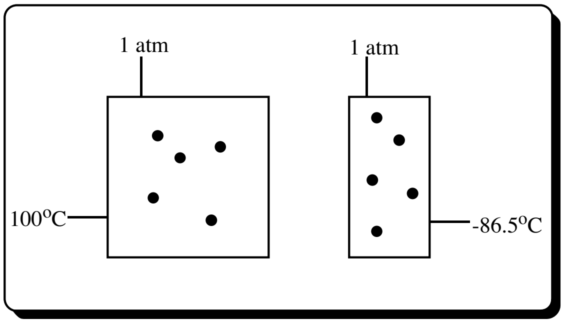
5. What type of relationship is indicated by the graph?

6. What type of equation would you expect for Boyle's Law?

7. What pressure is required to reduce the volume of a sample of air from 6.00 L to 4.50 L? The original pressure on the sample is 675 torr. Assume the temperature and quantity remain constant.

Chapter 12 – The Behavior of Gases
Part 2B – Notes: The ABCs of Gases – Charles' Law

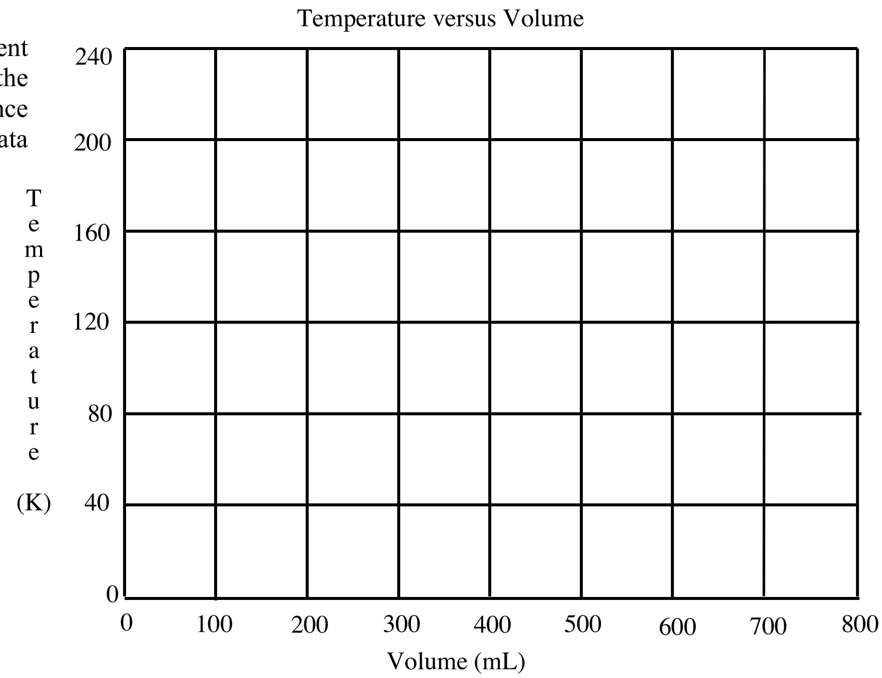
Objective: Investigate the effect of a temperature change on the volume of a given sample of gas at a constant pressure.



1. What observations may be made about the gas in the two containers?

2. The following sets of data represent temperature and volume readings taken while the amount and the pressure of a gaseous substance were held constant. Graph the following data points and answer the questions.

Vol (mL)	Temp (K)
0	0
50	15
150	45
250	75
350	105
450	135
550	165
650	195



3. From the graph and the illustration, state *Charles' Law* in your own words.

4. According to the graph, what is the volume of the gas when the temperature is 150K?

According to the graph, what would be the temperature of the gas when the volume is 500 mL?

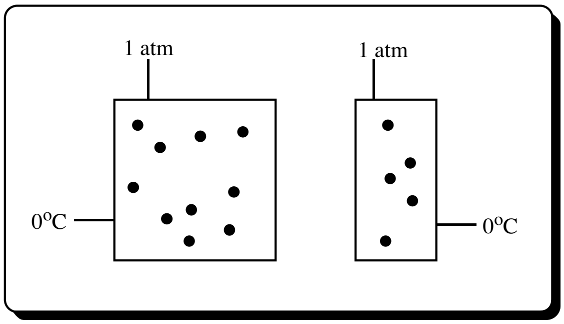
5. What type of relationship is indicated by the graph?

6. What type of equation would you expect for Charles' Law?

7. Initially a gas is at 45°CK and has a volume of 740 mL. What would the temperature be if the volume is increased to 975 mL? Assume the pressure and quantity remain constant.

Chapter 12 – The Behavior of Gases
Part 2C – Notes: The ABCs of Gases – Avogadro's Hypothesis

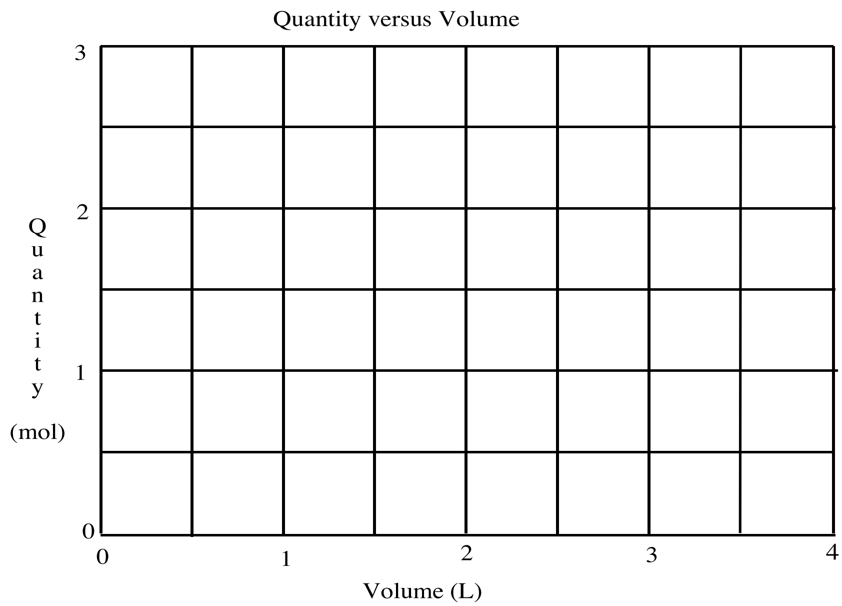
Objective: Investigate the relationship between the quantity of a gas and the given volume of the sample at set T and P.



1. What observations may be made about the gas in the two containers?

2. The following sets of data represent quantity and volume readings taken while the temperature and the pressure of a gaseous substance were held constant. Graph the following data points and answer the questions.

Quantity (mol)	Volume (L)
0	0
0.5	0.625
1.0	1.25
1.5	1.875
2.0	2.50
2.5	3.125
3.0	3.75



3. From the graph and the illustration, state *Avogadro's Relationship* in your own words.

4. According to the graph, what is the volume of the gas when the quantity present is 1.75 moles?

   According to the graph, what is the quantity present when the volume of the gas is 2.00 L?

5. What type of relationship is indicated by the graph?

6. What type of equation would you expect for Avogadro's Relationship?

7. If the volume of 3.79 moles of a gas is 8.95 L, then what is the volume of 4.62 moles of the same gas? Assume that pressure and temperature are held constant?

Chapter 12 – The Behavior of Gases
Part 2 – Assignment: The ABCs of Gases – Avogadro. Boyle, & Charles

Solve the following problems and answer the questions. Show all set-ups, units, formulas, etc. BE NEAT!!!

1. A sample of gas in a 25.0-L container exerts a pressure of 3.20 atm. Calculate the pressure exerted by the gas if the volume is changed to 45.0-L at a constant temperature.

2. A sample of gas in a 21.5-L container at 45ºC is cooled at constant pressure to a temperature of –37ºC at constant pressure. Determine the volume of the cooled gas.

3. A 32.8 g sample of hydrogen gas occupies a volume of 21.6 L at a particular temperature and pressure. What volume does 12.3 g of hydrogen gas occupy at the same pressure and temperature?

4. For Charles' Law to hold true, why must the pressure and amount of gas remain the same?

5. One of the most obvious properties of gaseous materials is the pressure they exert on their surroundings. In particular, the pressure exerted by the atmospheric gases is important. How does the pressure of the atmosphere arise, and how is this pressure commonly measured?

6. What temperature scale is defined with its lowest point as the absolute zero of temperature? What is absolute zero in Celsius degrees?

7. You are holding two balloons of the same volume. One contains helium and one contains hydrogen. Complete each of the following statements with "different" or "the same" and justify your answer.

   a. The pressure of the gas in the two balloons are ____________________.

   b. The temperature of the gas n the two balloons are ____________________.

   c. The number of moles of gas in the two balloons are ____________________.

   d. The densities of the gas in the two balloons are ____________________.

8. A balloon whose volume is 5.00 L contains 0.275 mole nitrogen. What mass, in grams, must be present to have a volume of 11.0 L at the same temperature and pressure conditions?

9. A gas has a volume of 426 mL at 118ºC. What is the temperature of the gas if the volume increases to 636 mL?

10. The volume of a gas is changed from 475 mL to 195 mL. The initial pressure was 775 mmHg, if the quantity and temperature remain the same, what is the final pressure?

Chapter 12 – The Behavior of Gases
Part 3 – Notes: Combined Gas Law and Standard Conditions

Objectives: Explore the relationship between pressure, volume, and temperature of a sample of a gas.
Calculate the moles, masses, and volumes of various samples of gas at STP.
Apply gas laws to problems involving the temperature, volume, and pressure, of a contained gas.
Identify, define, and explain: standard condition, standard temperature and pressure, and combined gas law.

Text Reference: Section 12.3 (Part) – pages 339-340.

Boyle's and Charles' Laws may be combined into a single law that relates the pressure, volume, and temperature of a single sample of gas. Using this **Combined Gas Law** you may explore what happens to one of these variables when the other two are varied.

The **Combined Gas Law** may be expressed as:

While both the temperature and pressure of a sample may be changed to alter the volume, *the quantity must remain constant.* The units for the Combined Gas Law are the same used for Boyle's and Charles' Laws, individually.

Example 1: Calculate the volume of a sample originally occupying 505 mL at 750. torr and 25^{o}C after its temperature and pressure are changed to $50.^{o}$C and 1.21 atm.

Example 2: Calculate the final pressure of a gas that is expanded from 1.50 L at 45^{o}C and 850. torr to 1.73 atm at 68^{o}C.

<u>Standard Conditions – Standard Temperature and Pressure</u>

A temperature of 0^{o}C and a pressure of exactly 1 atm constitute a set of *standard conditions* for a gas called **STP** – or **Standard Temperature and Pressure**. When STP is involved in a problem, substitute 1 atm and 0^{o}C for the appropriate variables and solve as usual. (Note: Only part of STP may be required. If they ask for a gas at standard pressure only, just use 1 atm.)

Example 3: A 4.00 L sample of gas originally at standard temperature and pressure is changed to 2.95 L and 2020 mm Hg. Calculate the final temperature.

Chapter 12 – The Behavior of Gases
Part 3 – Assignment: Combined Gas Law and Standard Conditions

Solve the following problems and answer the questions. Show all set-ups, units, formulas, etc. BE NEAT!!!

1. A sample of gas at 38°C occupies a volume of 2.97 L and exerts a pressure of 3.14 atm. The gas is heated to 118°C and the volume is decreased to 1.04 L. Determine the new pressure exerted by the gas.

2. What would happen to the pressure of a sample of chlorine gas if the volume and the Celsius temperature were both tripled?

3. What would happen to the volume of a sample of nitrogen if the pressure is doubled and the absolute temperature is tripled?

4. A gas at 118°C and 745 mm Hg has a volume of 782 mL. What would be the volume of the gas if the temperature were increased to 135°C and the pressure changes to 842 mm Hg?

5. A gas has a volume of 417 at STP. What will be the pressure of the gas if the volume is changed to 683 mL and the temperature becomes 212°C?

6. A gas has a volume of 1635 mL at 792 Torr and –32°C. What is the volume of the gas at STP?

7. If the temperature is held constant, what will happen to the pressure exerted by a gas if the volume of the container is increased?

   A. Pressure will increase.
   B. Pressure will decrease.
   C. Pressure will remain the same.
   D. Pressure will increase or decrease depending on other conditions.

Chapter 12 – The Behavior of Gases
Part 4 – Notes: The Ideal Gas Law and Standard Molar Volume

Objectives: Identify and give properties of an ideal gas.
Use the ideal gas law to perform calculations reflecting the state of a gas.
Calculate the molar volume of a gas at standard conditions and use standard molar volume in calculations.
Calculate the amount of gas at any specified conditions of pressure, volume, and temperature.
Differentiate between a real gas and an ideal gas.
Identify, define, and explain: ideal gas, real gas, equation of state, standard molar volume, ideal gas constant.

Text Reference: Section 12.4 (Part) – pages 341-346

You have considered three laws that describe the behavior of gases as revealed by experimental observation: *Boyle's Law, Charles' Law,* and *Avogadro's Law*. These laws describe the new conditions in which a gas exists after its initial conditions are changed. These relationships show how the volume of a gas may depend on pressure, temperature, and the number of moles of a gas present and may be combined as follows:

$$V = R\,(T\,n / P)$$

where R = the combined proportionality constant called the **universal gas constant**.

The above formula may be rearranged into the most familiar form of the ***IDEAL GAS LAW***: $P\,V = n\,R\,T$

P = __________________________ with a unit of __________________________ or __________________________.

V = __________________________ with a unit of __________________________.

n = __________________________ with a unit of __________________________.

T = __________________________ with a unit of __________________________.

The *universal gas constant*, ***R***, is dependent upon the unit of pressure used.

If the unit of pressure used in the problem is atmospheres, then ***R*** = __________________________.

If the unit of pressure used in the problem is kilopascals, then ***R*** = __________________________.

The Ideal Gas Law is an ***equation of state*** for a gas, where it describes the state of the gas as its conditions at a given time. A particular ***state*** of a gas is described by its pressure, volume, temperature, and quantity. Knowledge of any three of these properties is enough to completely define the state of a gas, since the fourth property may be determined from the Ideal Gas Law.

The laws of Boyle, Charles, and Avogadro are exactly true only if the gas is IDEAL. Since you do not live in an ideal world, you know the laws are not going to work perfectly. However, you can closely approximate an ideal gas and therefore have a close approximation of Boyle, Charles, and Avogadro, as well as the Ideal Gas Law.

PROPERTIES OF AN IDEAL GAS

Although real gases are not ideal, a *real gas may approach ideal conditions when there is. . .*

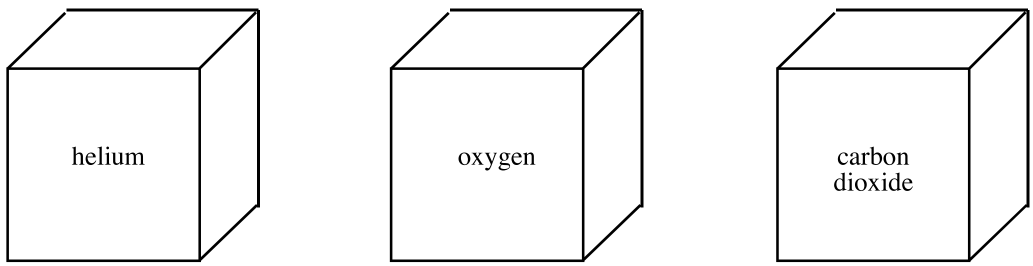
Example 1: A sample of hydrogen gas has a volume of 8.56 L at a temperature of $0^{o}C$ and a pressure of 1.00 atm. Calculate the mass, in grams, of this sample.

Example 2: A 5.00 L flask contains 0.650 grams oxygen at 22ºC. What is its pressure, in kilopascals? In atm?

MOLAR VOLUME – the volume of one mole of a gas at specified temperature and pressure conditions

STANDARD MOLAR VOLUME – the volume of one mole of gas at STP

One mole of ANY gas will occupy the same volume as one mole of ANY OTHER gas at STP



molar mass = 4.00 g
density = 0.178 g/mL
22.4 L/mol at STP

molar mass = 32.00 g
density = 1.42 g/mL
22.4 L/mol at STP

molar mass = 44.01 g
density = 1.96 g/mL
22.4 L/mol at STP

In other words, 1 mol of ANY GAS at STP will occupy a volume of 22.4 L!!!

Example 3: What is the volume of 78.08 g of sulfur trioxide at STP?

Example 4: What is the density of nitrogen dioxide at STP?

Chapter 12 – The Behavior of Gases
Part 4 – Assignment: The Ideal Gas Law and Standard Molar Volume

Solve the following problems and answer the questions. Show all set-ups, units, formulas, etc. BE NEAT!!!

1. What mass of oxygen gas exerts a pressure of 475 mm Hg in a volume of 1.25 L at a temperature of –22ºC?

2. What volume will 1.87 mol of an ideal gas occupy at 397 K and 134.6 kPa of pressure?

3. Calculate the number of moles of an ideal gas if 4.50 L of the gas is at 785 torr and 23.7ºC?

4. 5.74 g of carbon dioxide are confined in a container at STP. What is the volume of the container?

5. If 28.0 g of a gas occupy 22.4 L at STP, what would the gas be?
   A. CO B. CO_2 C. C_2H_2 D. C_2H_6

6. How many moles of oxygen are contained in a sample that occupies 55.0 L at STP?
   A. 1.61 moles B. 2.01 moles C. 2.46 moles D. 1230 moles

7. The mass of 11.2 L of a gas is 20.0 g at STP. What is the mass of 6.02×10^{23} molecules of this gas?
   A. 10.0 g B. 20.0 g C. 40.0 g D. 80.0 g

8. At STP, 5.6 L of CH_4 contains the same number of molecules as:
   A. 1.4 L oxygen B. 2.8 L ammonia C. 5.6 L hydrogen D. 11.2 L neon

9. How many molecules of carbon dioxide are present in a 2850 mL sample at STP?

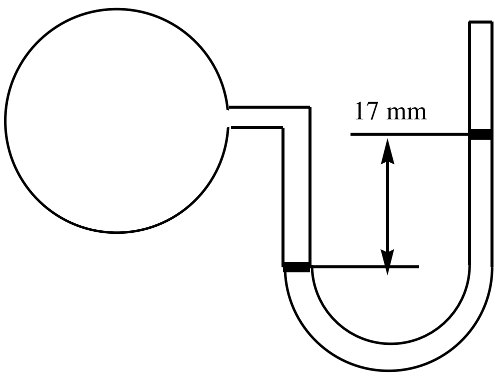
Chapter 12 – The Behavior of Gases
Part 5 – Notes: Dalton's Law of Partial Pressure and Manometers

Objectives: Explain why and how a manometer works and how is measures pressure.
Determine the pressure of a gas sample from a manometer.
Explain the difference between an open and a closed manometer.
Calculate the partial pressures of gaseous components in a mixture.
Identify, define, and explain: manometer, open manometer, closed manometer, and partial pressure.

Text Reference: Section 12.5 (Part) – Page 350-351

A ***manometer*** is used for measuring the pressure of a gas enclosed in the vessel.
There are two types of manometers: open and closed.

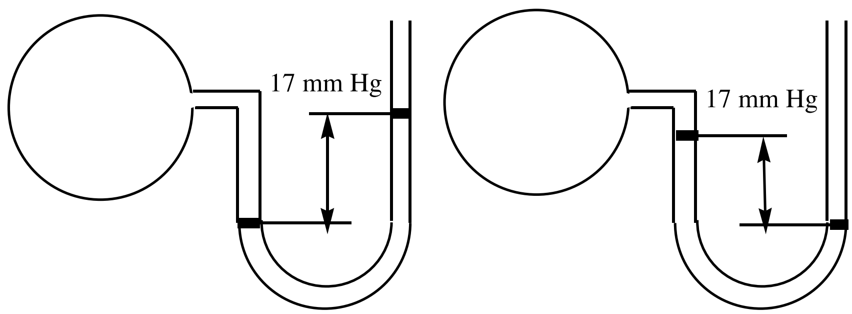
CLOSED MANOMETERS



Since a closed manometer cannot be affected by the atmospheric pressure, the pressure of a gas in the manometer vessel is simply the difference in the heights of the mercury in the U-tube. In the example to the left, the pressure of the gas is _______________.

OPEN MANOMETERS

Since, in an open manometer, the armospheric pressure can exert a pressure down on one of the arms of the U-tube, the pressure of the gas in the vessel equals the atmospheric pressure *plus* ***or*** *minus* the difference in the mercury heights. In the examples on the right, assume the atmospheric pressure is 735 mm Hg.



In the figure on the left side of the diagram, the pressure of the gas in the vessel is ____________________.

In the figure on the right side of the diagram, the pressure of the gas in the vessel is ____________________.

Examples: In each of the following, the atmospheric pressure is 750. mm Hg. What is the pressure of the gas in each case?

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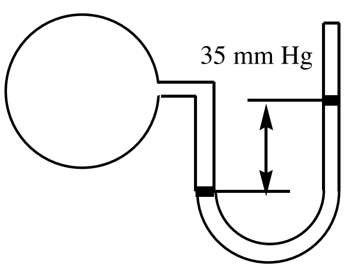

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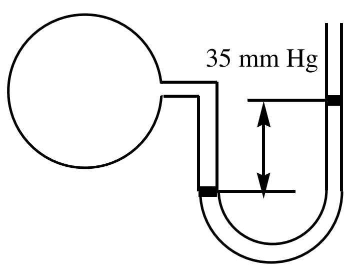

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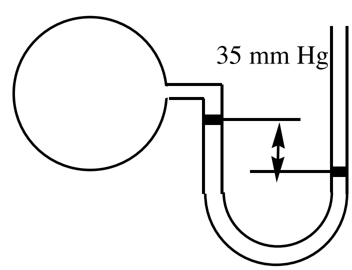

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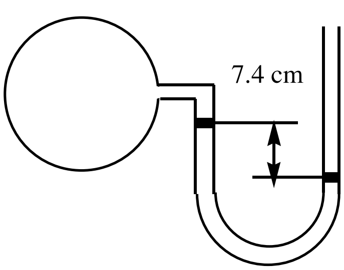

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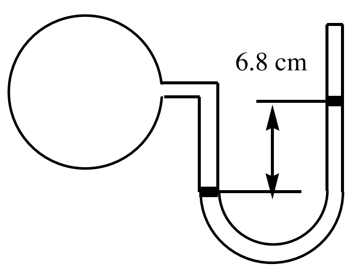

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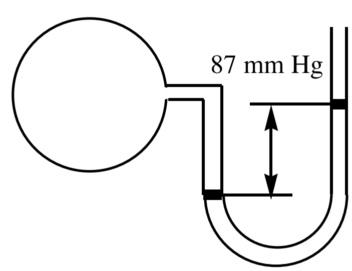

DALTON'S LAW OF PARTIAL PRESSURE

The total pressure* (P$_T$) *of the mixture of gases is the sum of the partial pressures of the components in that mixture.

In other words, the total pressure in a container is the SUM of the pressure of each individual component gases present in the container at constant temperature and pressure.

So $P_T = P_1 + P_2 + P_3 + \ldots$

P_T = total pressure

$P_1, P_2, P_3, \ldots$ = partial pressure of the various components of the mixture

Also Partial pressure of a gas = (mole fraction) (total pressure)

And Mole fraction = moles of component gas / total moles of all gases in the mixture

Example 1: A gaseous mixture contains 1.01 g hydrogen and 32.00 g oxygen in a volume of 22.4 L at a constant temperature. Determine the partial pressure of each gas in the mixture if the total pressure is 1.64 atm.

Example 2: A gaseous mixture contains 5.75 g of hydrogen gas, 22.5 g neon gas, and 18.0 g helium. The partial pressure of the neon is 0.685 atm. What is the total pressure of the gaseous mixture?

Example 3: H_2gas collected over water at 20.°C and at 750. mm Hg has a volume of 24.2 mL. What would be the volume at STP of the "dry" gas? The vapor pressure of water at 20.°C is 17 mm Hg.

Chapter 12 – The Behavior of Gases
Part 5 – Assignment: Dalton's Law of Partial Pressure and Manometers

If the atmospheric pressure is 775 mm Hg, what is the pressure of the gas in the manometer vessel in the following diagrams?

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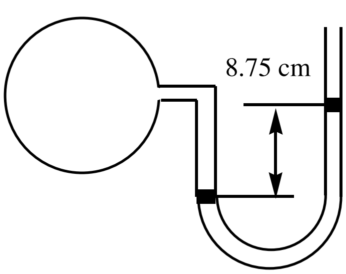

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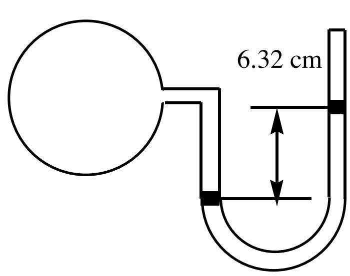
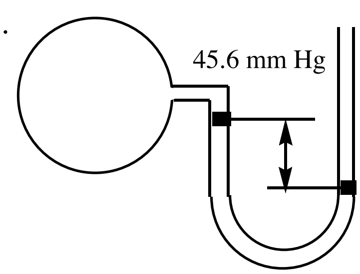


3.


Draw the following manometers registering the listed pressures. Atmospheric pressure for these problems is 737 mm Hg.

4. closed; 225 mm Hg

5. open; 560 mm Hg

6. open; 880 mm Hg

7. A gaseous mixture contains 20.0 g neon and 14.5 g helium in a 15.0 L container at constant temperature. The partial pressure of the helium gas is 3.25 atm. What is the total pressure of the system?

8. Calculate the number of moles in a 4.00 L sample of a gas at 300. K containing oxygen at 0.650 atm and nitrogen at 0.220 atm. Also calculate the number of moles of oxygen and the number of moles of nitrogen present.

9. 14.0 g nitrogen and 14.0 g hydrogen are placed in a container and the total pressure is 3.75 atm. (a) What is the partial pressure of hydrogen? (b) What is the partial pressure of the nitrogen?

10. A gas collected over water at 20.^{o}C and at 780. mm Hg has a volume of 42.65 mL. What is the volume of the "dry"gas at STP? The vapor pressure of water at 20.^{o}C is 17 mm Hg.

11. A sample containing 0.80 mol oxygen gas is collected over water at 30.0^{o}C. The total pressure is 1.10 atm and the water vapor pressure at 30.0^{o}C is 31.8 torr. Determine the volume of the oxygen gas.

12. A student is preparing for a laboratory experiment in which three gases will be mixed together. There is a 2-L sample of neon gas at a pressure of 2 atm, a 2-L sample of carbon dioxide gas at a pressure of 3 atm, and a 2-L sample of nitrogen gas at a pressure of 4 atm. All three samples are at room temperature. In the experiment the three gases will be transferred to the same rigid 2-L container.

(A) What pressure should the student expect the nitrogen gas to exert in the final mixture?

a. 2 atm b. 3 atm c. 4 atm d. 9 atm

(B) What total pressure should the student expect the gases to exert in the final mixture?

a. 4 atm b. 5 atm c. 9 atm d. 12 atm

Chapter 12 – The Behavior of Gases
Part 6 – Notes: Graham's Law of Diffusion

Objectives: Identify, define, explain, and differentiate between effusion and diffusion.
Qualitatively discuss the rates of diffusion of two gases at the same temperature.
Calculate relative rates of diffusion/effusion.
Text Reference: Section 12.5 (Part) – Pages 352-353.

"UNDER IDENTICAL CONDITIONS, THE RATES AT WHICH GASES DIFFUSE ARE INVERSELY PROPORTIONAL TO THE SQUARE ROOT OF THEIR MOLAR MASSES (OR DENSITIES)."

You may use this to determine information about the speeds at which gases diffuse, relative to one another and also information about the molar mass of an unknown gas, or even its density.

Example 1: *You have equal amounts of helium and carbon dioxide at the same temperature and pressure conditions. Which gas will diffuse faster? How much faster?*

Example 2: *An unknown gas diffuses 1.50 times faster than oxygen. What is the molar mass of the unknown gas?*

Example 3: *What is the molar mass of a gas if oxygen diffuses 1.41 times faster than it does?*

Chapter 12 – The Behavior of Gases
Part 6 – Assignment: Graham's Law of Diffusion Plus Some Extras

Solve the following problems showing all set-ups, formulas, units, etc.

1. What is the molar mass of a gas that diffuses 3.00 times faster than SO_2?

2. What is the molar mass of a gas that diffuses exactly one-half as fast as carbon dioxide?

3. You have equal amounts of two gases at the same temperature and pressure: neon and fluorine. Which diffuses faster? How much faster?

4. An unknown gas diffuses 2.75 times faster than sulfur trioxide. What is the molar mass of the unknown gas?

5. Given oxygen and helium at the same T and P conditions, which one diffuses more quickly? How much faster?

6. What is the molar mass of a gas if oxygen diffuses 1.41 times faster than it does?

7. H_2 collected over water at 20.^{o}C and 750. mm Hg has a volume of 24.2 mL. What would be the volume at STP of the "dry" gas? (The vapor pressure of water at 20^{o}C is 17 mm Hg. Remember Dalton's Law.)

8. 14.0 g nitrogen and 14.0 g hydrogen are placed in a container and the total pressure of the mixture is 3.75 atm. (A) What is the partial pressure of hydrogen? (B) What is the partial pressure of the nitrogen?

9. 5.6 L of oxygen (at STP) are placed in a container along with 1.0 g of hydrogen. The partial pressure of the hydrogen is 1.0 atm. What is the total pressure of the mixture?

10. A mixture of 14.0 g nitrogen and 34.0 g NH_3 is contained in a 2500. mL vessel at 27^{o}C. (A) Calculate the total pressure. (B) Calculate the mole fraction of NH_3. (C) Calculate the partial pressure of the nitrogen.

Chapter 12 – The Behavior of Gases
Part 7 – Notes: Kinetic Molecular Theory

Objectives: Differentiate between a law and a theory.
State the postulates of and explain the kinetic molecular theory.
Describe the properties of gas particles, including the volume, pressure, temperature, and compressibility.
Explain how the kinetic energy of gas particles relate to the absolute temperature.

Text Reference: Section 12.1 – Pages 326-328.

What is the difference between a ***LAW*** and a ***THEORY***?

KINETIC MOLECULAR THEORY – KMT:

Five Postulates of the *Kinetic Molecular Theory:*

1.

2.

3.

4.

5.

The meaning of temperature:

The relationship between temperature and pressure:

The relationship between temperature and volume:

The relationship between volume and pressure:

Chapter 12 – The Behavior of Gases

Part 7 – Assignment: Kinetic Molecular Theory

Solve the following problems and answer the questions. Show all set-ups, units, formulas, etc. BE NEAT!!!

1. Using the postulates of the kinetic molecular theory, explain how these postulates account for the following properties of a gas:
   a. the pressure of a gas
   b. pressure increases with an increase in temperature
   c. a gas fills its entire container
   d. volume of a gas increases with an increase in temperature

2. You are holding two balloons, an orange balloon and a blue balloon. The orange balloon is filled with neon gas and the blue balloon is filled with argon gas. Which of the following best represents the mass ratio of Ne:Ar in the balloons?

   A. 1:1 B. 1:2 C. 2:1

3. You have two containers at the same temperature and pressure. Container #1 has 1 mole oxygen gas and Container #2 has 1 mole hydrogen gas. Indicate how these containers compare in . . .
   a. volume.
   b. number of molecules in each container.
   c. density of sample.
   d. average kinetic energy.
   e. number of collisions per given area of container per second.

4. How does Dalton's law help us realize that for an ideal gas sample, the volume of an individual molecule is insignificant compared with the bulk volume of the sample?

Chapter 12 – The Behavior of Gases
Part 8 – Notes: Gas Laws and Stoichiometry

Objectives and Text Reference: See Chapter 9.

DETERMINING DENSITY AND MOLAR MASS

Density is mass divided by volume. You know the volume of 1 mole of a gas when it is at STP. You are also able to calculate the mass of 1 mole of a substance (the molar mass). So you are easily able to calculate the density of a gas when it is at STP.

Example 1: *The density of a gas at STP is 2.42 g/L. What is the molar mass of the compound?*

You may also work with non-ideal situations.

Example 2: *16.5 g CO_2 is in a container at 323K and 1.45 atm. What is the density of carbon dioxide at these conditions?*

You may also work with other issues regarding moles, mass, and gas laws. You know *molar mass* is the mass of one mole. If you know the mass in a certain number of moles, you may calculate the molar mass of the substance.

Example 3: *A sample of gas has a volume of 3.25 L at 1.35 atm and 318K. The mass of the sample is 4.32 g. What is the molar mass of this substance?*

GAS LAWS WITH BASIC STOICHIOMETRY

Sometimes you need to pull in topics from previous chapters. You can easily use stoichiometry and mole ratios with your gas law problems.

Example 4: *Calculate the volume of carbon dioxide produced and collected at 325K and 1.65 atm, when 24.0 g of carbon is combusted with oxygen.* (Remember all the steps you need to go through.)

Example 5: *How many grams of sodium chloride may be produced by a reaction of 121 mL chlorine gas at STP with excess sodium?*

Example 6: *What volume of chlorine gas at 24^{o}C and 0.872 atm would be required to react with 2.51 g silver to produce silver chloride?*

VOLUME-VOLUME PROBLEMS AT CONSTANT TEMPERATURE AND PRESSURE

If you need to find the volume of a gas ***(NOT AT STP)*** but with ***CONSTANT TEMPERATURE AND PRESSURE***, *use the coefficients from a balanced chemical equation* to make a proportion so the units cancel, and determine the answer using unit analysis. You know the coefficients from the balanced chemical equation allow you to go from substance A to substance B through a mole ratio. Now you have recently learned that, according to Avogadro's principle, at the same T and P, equal volumes of gases contain the same number of molecules, in other words the same number of moles. **So the coefficients in a balanced chemical equation give you not only a mole ratio, but they also allow you to have a volume ratio, provided the T and P are constant**.

Example 7: *Ammonia (NH_3) reacts with oxygen gas to form nitrogen monoxide and water. (A) Write a balanced chemical equation. (B) How many milliliters of oxygen are required to react with 0.270L of ammonia? Assume that all gases are at the same temperature and pressure.*

Example 8: *Methane gas (CH_4) combusts with oxygen. What volume of oxygen at 125^oC and 105.5 kPa is required to combust 6.84 L of methane at the same temperature and pressure?*

Chapter 12 – The Behavior of Gases
Part 8 – Assignment: Gas Laws and Stoichiometry

Solve the following problems using unit analysis where appropriate. Show all units, set-ups, equations, etc.

1. At a deep-sea station 200. m below the surface of the Pacific Ocean, workers live in a highly pressurized environment. How many liters of gas, at STP, must be taken from the surface and compressed to fill the underwater environment with $7.00x10^7$ L of a gas at 20.0 atm? Assume a constant temperature.

2. How many liters of hydrogen gas may be produced at 27^oC and 104 kPa, if 20.00 g sodium metal are reacted with water?

3. How many liters of hydrogen gas may be produced at $150.^oC$ and 0.765 atm, if 17.50 g lithium are reacted with water?

4. Magnesium will burn in the presence of oxygen to form magnesium oxide. What mass of magnesium will react with 500.0 mL of oxygen at STP?

5. What volume of ammonia gas (NH_3) at 85°C and 2.85 atm can be produced by the reaction of 4.50 g hydrogen gas with excess nitrogen gas?

6. When 3.00 moles of hydrogen gas react with 2.00 moles of nitrogen gas, ammonia gas (NH_3) is formed.
   a. Write a balanced chemical equation.

   b. Calculate the volume of 3.00 moles of hydrogen at STP. Calculate the volume of 2.00 moles nitrogen at STP.

   c. Find the limiting reactant.

   d. Calculate the liters of gas produced when it is collected at 1.87 atm and 125°C.

   e. Calculate how many grams of excess are left over after the reaction is completed.

   f. Calculate the volume of excess left over after the reaction is completed at 1.87 atm and 125°C.

7. 15.0 L of methane gas (CH_4) is at 85.0 kPa and 175°C. What volume of oxygen is required for the combustion of all the methane if the oxygen is at the same temperature and pressure conditions as the methane?