# Gases, Chapter 5
# *Essentials*, Chang

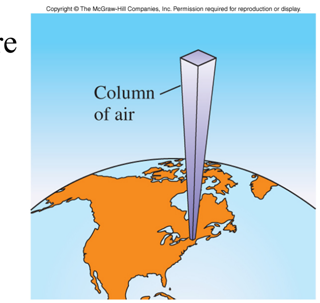

• Lets take a look at the oldest recorded chemical demonstration. It was performed in 440 B.C.E. by the Greek physician *Empedocles*.

• Lets define pressure: Pressure = force/area

http://hillcrestdiv5.weebly.com/uploads/1/3/4/8/13488001/710381125.jpg?423

http://www.nativegrace.com/v/vspfiles/assets/images/oil_drum.jpg

# Activity 5.01: Understanding Pressure




• Atmospheric pressure is very important. The atmosphere exerts a pressure in all directions. What evidence is there that the gasses in the atmosphere exert a pressure?

a. Lets take a standard oil drum and remove all of the air from inside the drum.

Predict what will happen as the air is being removed.

1. The oil drum will expand
2. Nothing will happen, its an oil drum
3. The oil drum will contact slightly
4. The oil drum will get crushed

removing air from
inside the drum →

b. Lets now take a standard lab marshmallow and remove the air outside of the marshmallow.

Predict what will happen as the air surrounding the marshmallow is removed.

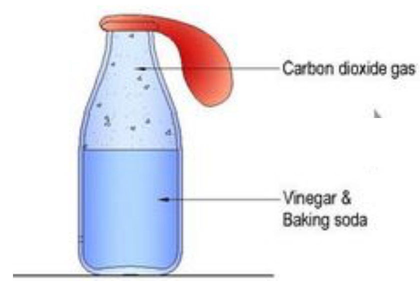
1. The marshmallow will expand
2. Nothing will happen, it's a marshmallow
3. The marshmallow will contact slightly
4. The marshmallow will get crushed

removing air
surrounding
the marshmallow →

• Chemical reactions can produce gases that exert a pressure. What do you think this pressure from gases produced in a chemical reaction is related to? Lets look a a chemical reaction where baking soda is added to vinegar to produce carbon dioxide.

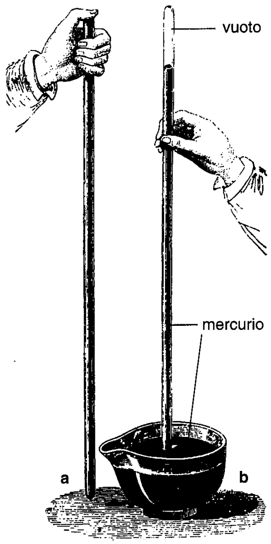
c. Predict what will happen if we add more baking soda to the vinegar

1. No change in the amount of carbon dioxide produced
2. More carbon dioxide is produced when more baking soda is added to the vinegar



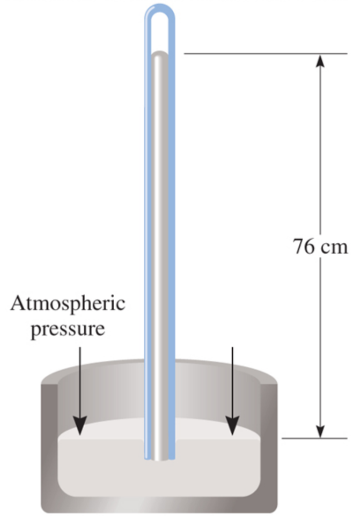
# Measuring the Pressure of Gases

- How can we measure the pressure of the atmosphere? How do we measure the pressure of any gas?




http://www.vacuum-guide.com/images/museum_torricelli02.gi






- What units do we use to describe pressure?

## More Experiments on Gas Properties and Volume

- We have looked at the relationships between pressure and volume, and we found that the two are inversely related to each other

- We have looked at the relationship between number of moles of gas and volume and found that the two are directly related to each other

- Consider a sample of a gas that is in an expandable container. What are some examples of expandable containers?

a. Predict what will happen to the container when we lower the temperature of the gas in the balloon.

1. The balloon will expand
2. Nothing will happen, it's a balloon
3. The balloon will shrink

- We can now develop a relationship between pressure, volume, number of moles of gas and temperature from each of three experiments. In each of these experiments we noted how the **volume of the container** changed

# The Ideal Gas Law: What is it? Why can we use it?

- We can combine all of these individual “gas laws” into a single mathematical relationship if we add a constant.

- Why is this equation called the **ideal gas law**?

- What is the **ideal gas constant R**?

a. Determine the volume of 1 mole of He gas in a balloon at a pressure of 1 atm and at a temperature of -192 $^{o}$C (this is the temperature of liquid nitrogen)

This volume will be:

1. Negative
2. Zero
3. Positive

## Activity 5.02 What about when a gas changes? (General Gas Equation)

- We can use the ideal gas law to calculate the properties of any gas that undergoes some kind of change.

a. Write the gas law expression for a gas with the following **initial (i)** values:

Volume = $V_i$
Temperature = $T_i$
Number of moles = $n_i$
Pressure = $P_i$
Gas constant = R

b. Now write the expression for the gas with the following **final (f)** values:

Volume = $V_f$
Temperature = $T_f$
Number of moles = $n_f$
Pressure = $P_f$
Gas constant = R

c. Set these two equations equal to each other and write down the resulting expression:

d. Helium balloons are used to carry climate monitoring instruments into our upper atmosphere. If a helium balloon on the ground has as initial conditions of V =1 x $10^6$ liter, T= 300K, and P= 1 atm. What is the final volume balloon in the upper atmosphere if its final conditions are T= 240 K and P = 0.1 atm?

http://www.esrl.noaa.gov/gmd/ozwv/ozsondes/images/spo_ozsonde.jpg

# The Relationship between the Density of a Gas and its Molar Mass

a. We have expressed concentration in units of molarity, can we rearrange the gas law to find an expression for the concentration of a gas?

• How can we use this information to find the relationship between the density of a gas and molar mass?

# Activity 5.03: Chemical Reactions Involving Gases

- We have already seen that adding vinegar ($CH_3COOH$) to baking soda ($NaHCO_3$) produces carbon dioxide ($CO_2$). Below is the complete balanced and net ionic equation for this reaction. NOTE: more than just $CO_2$ is produced in this reaction

$$CH_3COOH_{(aq)} + NaHCO_{3\,(aq)} \rightarrow CO_{2\,(g)} + H_2O_{(l)} + NaCH_3COO_{(aq)}$$

$$CH_3COOH_{(aq)} + HCO_3^-{}_{(aq)} \rightarrow CO_{2\,(g)} + H_2O_{(l)} + CH_3COO^-{}_{(aq)}$$

a. Using the net ionic equation for this reaction, calculate the volume of $CO_2$ that will be produced at 23 $^oC$ and 720 torr, if 10.0 grams of sodium bicarbonate ($NaHCO_3$) reacts with 250 ml of vinegar ($CH_3CCOH$) .
Household vinegar is acetic acid at a concentration of 0.837 M NaHCO$_3$ has molar mass of 84 g/mol.

$$CH_3COOH_{(aq)} + HCO_3^-{}_{(aq)} \rightarrow CO_{2\,(g)} + H_2O_{(l)} + CH_3COO^-{}_{(aq)}$$

# Activity 5.04: Understanding Mixtures of Gases

- Our atmosphere is not a made up of one single gas but a mixture of gases. Let's discover the properties of mixtures of gases.

- Consider the three gases each in its own separate 1 L container at the pressures indicated and at a temperature of 298K. The gases are then mixed into the same 1 L container while the temperature is held constant at 298K.

**a) How can we determine the total number of moles of gas after mixing?**

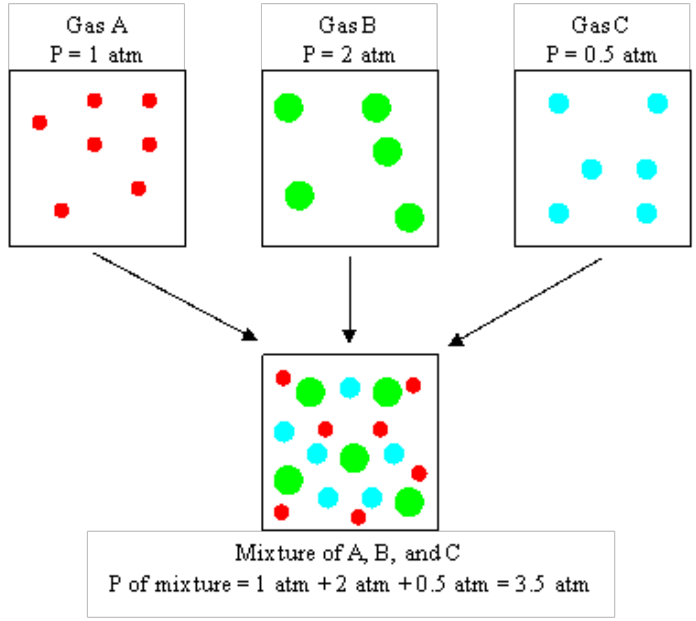
*True* or *False*: Total number of moles after mixing can be found by adding the initial number of moles of each gas together.

**b) How can we determine the total volume of gases after mixing?**

*True* or *False*: Total volume after mixing can be found by adding the initial volume of each gas together.

**c) How can we determine the total pressure of the gasses after mixing:**

*True* or *False*: Total pressure after mixing can be found by adding the initial pressure of each gas together.




http://www.chemcool.com/regents/physicalbehaviorofmatter/aim9.h17.gif

- How can we relate the *total* **n**, **V** and **P** for a mixture of gasses in one equation? .

# Mixture of Gases: Mole Fraction

- What equation can we derive that relates the **partial pressure** of each gas in a mixture to the **total pressure** of all the gases?

# Activity 5.05: Water Vapor

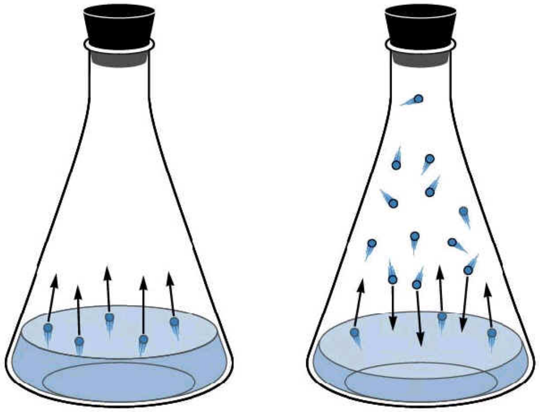
a. If you left a small cup of water outside and then returned a few hours later, would you have the same volume of water in the cup?

- All liquids will evaporate producing a gas which we call a ***vapor*** because the liquid and gas exist together. Liquid water evaporates, producing water as a gas; this gas is referred to as ***water vapor*** and it will exert a pressure on the walls of the container.

b. If you put a tight lid on a cup of water, would it evaporate?

The ***vapor pressure of water*** is dependent upon temperature.

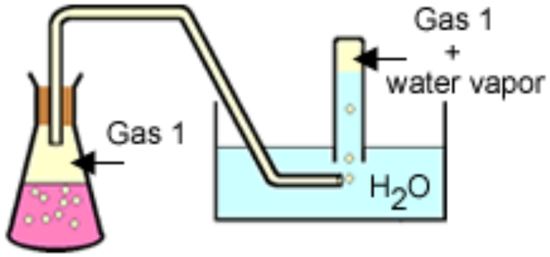
| Temperature (C) | Vapor Pressure (torr) |
|---|---|
| 0 | 4.58 |
| 10 | 9.21 |
| 20 | 17.54 |
| 25 | 23.76 |
| 30 | 31.82 |
| 50 | 92.5 |
| 80 | 355.1 |
| 100 | 760.0 |



http://www.unit5.org/christjs/Vapor_2.jpg

# Activity 5.06: Mixture of Gases: Collecting Gases Over Water

- A simple technique to isolate a gas is to collect the gas over water. This technique allows us to determine the volume of gas collected but what is its limitation?




- Calcium carbide ($CaC_2$) decomposes in water to yield acetylene gas ($C_2H_2$).
  For every mole of calcium carbide that decomposes 1 mole of Acetylene gas is produced.

$$CaC_2 + 2\,H_2O \rightarrow C_2H_2 + Ca(OH)_2$$

- When a sample of calcium carbide is decomposed in water, 500 ml of acetylene gas collected over water at a pressure of 1 atm and a temperature of 30°C.

a. Determine the partial pressure of the acetylene gas collected over water.

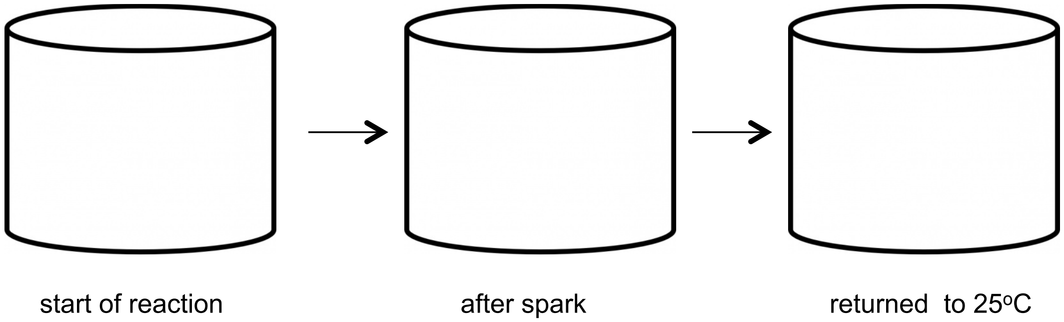
b. Using this information, determine the number of moles of calcium carbide that decomposed.

# Activity 5.07: Reactions Involving Mixtures of Gases

- A 1.00-liter steel vessel contains 1.0 atm $H_2$ and 1.0 atm $O_2$ at 25°C. A spark is passed through the vessel and water is produced. Assume that the reaction proceeds to completion. The vessel and its contents are then returned to 25°C.

a) Calculate the number of moles of water in the vessel at 25°C.

$$H_2 \quad + \quad \tfrac{1}{2}\, O_2 \quad \rightarrow \quad H_2O$$




b) Determine the mass of liquid water at 25°C.

# Activity 5.08: Using Rice Tables Tabulated in Units of Pressure Reactions of Gasses

- Rice tables for gases can be tabulated in number of moles and in units of pressure (like atm).
  In order to use units of pressure for the RICE table:
  **(1) the initial and final volume of the container must not change**
  **(2) the initial and final temperature must be the same**

- Consider the reaction below between gas A and gas B to form gas C and gas D. This reaction happens in a sealed steel container.

a. If Gas A is at an initial pressure of 4 atm and gas B is at an initial pressure of 2 atm and these gasses react **completely,** determine the pressure of each gas at the end of the reaction.

$$A + 3B \rightarrow C + D$$

- Consider the reaction below where gas E **partially decomposes** into gas F and gas G. This reaction happens in a sealed steel container.
  NOTE: when **some** gas reacts or there is a **partial** decomposition, **you will need to assign a variable for the amount of gas that has reacted**.

b. If Gas E has an initial pressure of 5 atm and the total pressure of all of the gases at the end of the reaction is 6 atm, determine the final pressure of every gas in the vessel.

$$E \rightarrow 2F + G$$

## Kinetic Energy, Temperature and the Speeds of Gases

Look at the gas simulation again. Can you see how the average kinetic energy of a gas is related to its temperature?

- Let's write an expression for the *average kinetic energy (KE)* of one single gaseous molecule or atom. NOTE: the units of energy are the **Joule** which we talked about on the first day of class and that **1 J = kg m²/s²**

$$KE = \tfrac{1}{2} m\overline{u}^2$$

m = mass of single atom or molecule
u = average speed of atom or molecule

- There are two ways of expressing the average kinetic energy of 1 mole of gas particles:
  (1) one is derived using our expression for average kinetic energy above
  (2) the other is derived from statistical theoretical approach which we are only going to show the results of here

  **Both expressions are equivalent** even though they look different.

Let's look at the gas simulator again. What happens when I add a heavier gas "molecule"? Let's develop the relationship between the *speed* of the "molecules" and the *mass* of the "molecules" by setting the two equations above equal to each other.

# Activity 5.09: Gas Speeds Experiment

- Listen to different gases passing through the instrument and producing a tone. The *higher* pitch the tone, the *faster* the molecules are moving. Describe the tone (highest to lowest) and predict which gases must be moving the fastest.
  Rank (1,2,3) the relative speeds of the materials below from fastest to slowest.

a) Air

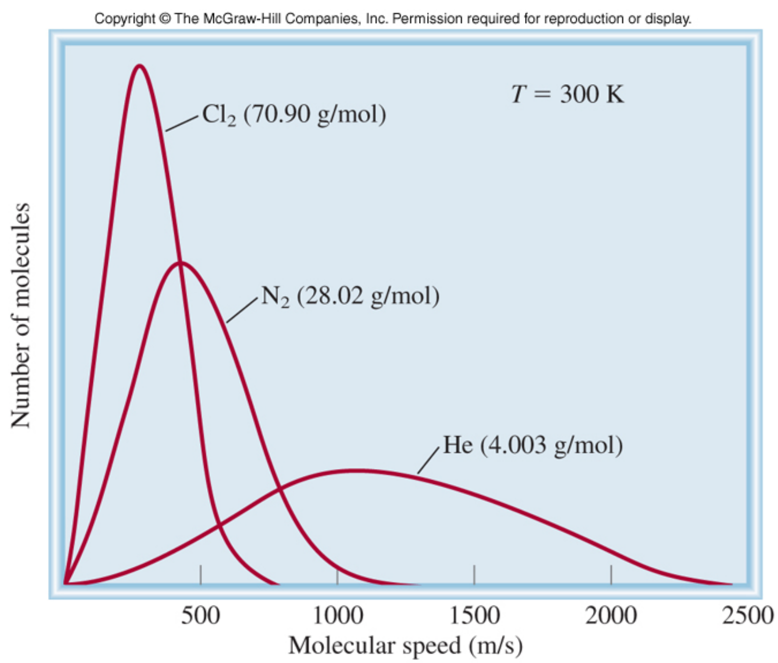
b) He

c) $SF_6$

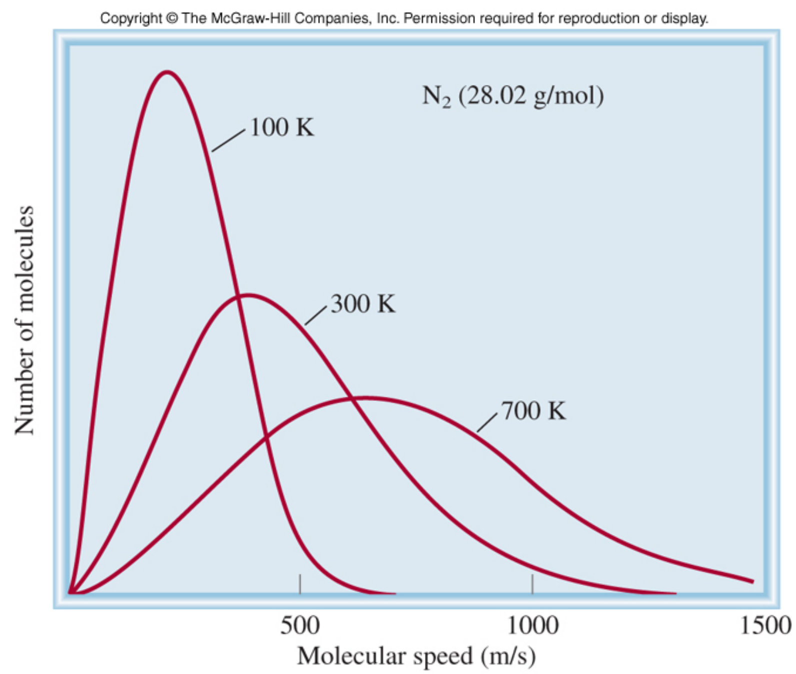
# Plotting the Speeds of Gases

- The *Maxwell-Boltzman Distribution* shows the distribution of gas molecules at a given temperature.

- Here are several distributions that show the *effect of varying the mass* (i.e. changing the element) of the gas molecules and the effect of *varying the temperature* of the gas for a single element. How do both of these graphs relate to what you have seen using our gas simulator and the equations we have developed?



(b)



(a)

NOTE: 1 m/s = 2.2 miles per hour