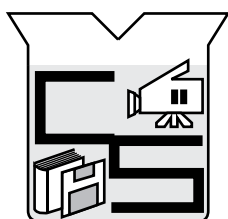
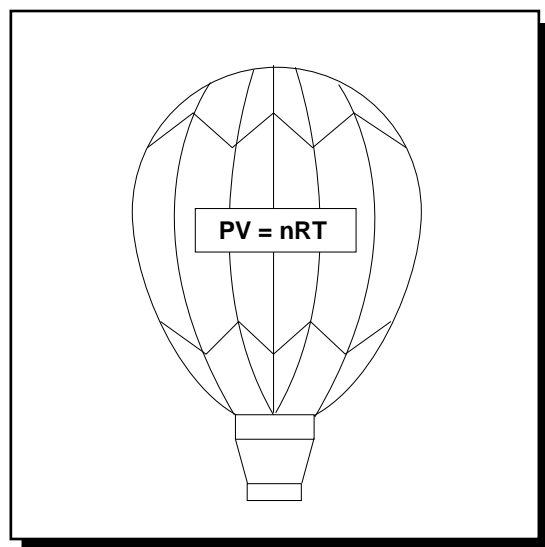


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*ChemSource Project Principal Investigator:
Mary Virginia Orna, OSU
Department of Chemistry
College of New Rochelle
New Rochelle, NY 10805
Phone: (914) 654-5302
FAX: (914) 654-5387*

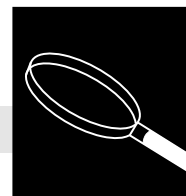


ChemSource

*Instructional Resources for Preservice and
Inservice Chemistry Teachers*

GASES

Topic Overview



CONTENT IN A NUTSHELL

Humans have been fascinated by gases since prehistoric times. Air was important enough to be considered one of the four elements by Aristotle—but only for about 200 years have scientists been able to distinguish one “air” from another and realize that the air we breathe is actually a mixture of gases.

Gases consist of submicroscopic particles that are relatively far apart from each other moving randomly with an average kinetic energy proportional to the temperature. (The word *gas* is derived from the Greek *chaos*.)

Gaseous particles have very little effect on each other; in fact an ideal gas particle is defined as having no volume and no attraction for other particles. We can show by our senses of smell or sight that gases spread out to fill a given volume.

Gases can be described in terms of four variables: pressure, volume, temperature, and amount (moles). Several sets of two variables—pressure and temperature (Gay Lussac’s Law), volume and temperature (Charles’ Law), volume and amount (Avogadro’s Law), and pressure and amount—are directly proportional, but pressure and volume are inversely proportional (Boyle’s Law). We can use simple logic to understand the relationships among these variables. Because gases are so much a part of our lives, many examples of these relationships are found in our environment.

The general gas law combines these variables into an equation with a single constant R , the gas law constant. The equation $PV = nRT$ may be used in calculations, assuming students have a background in algebra.

There are other relationships derivable from the gas law. Dalton’s law of partial pressures indicates that the sum of the pressures of two or more gases equals the total pressure. For many years the molar masses of gases and volatile liquids were determined experimentally using the gas laws. Gas densities can also be calculated using these relationships.

From the time of the 18th-century revolution in chemistry, scientists have used the volumes of gases in work involving stoichiometry of reactions. That two volumes of hydrogen react with one volume of oxygen to form two volumes of water vapor helped chemists understand that H_2 and O_2 were diatomic and that H_2O was a better formula for water than was HO . This knowledge was essential to development of modern atomic theory.

PLACE IN THE CURRICULUM

Many secondary school texts use gases as an introduction to the atomic theory since gases, like atoms, are generally invisible bodies with observable effects. Using gas laws in stoichiometric calculations makes a link between different topics in the study of chemistry.

Whether it is concern with pressure in an automobile tire, the expansion of popcorn on heating, the effect of a leak in a space suit, the understanding of absolute zero, or the solution to environmental problems such as the depletion of the ozone layer, the study of gases enriches the curriculum of high school students.

1. A gas sample is composed of submicroscopic particles moving with an average kinetic energy directly proportional to the temperature (in kelvins).
2. A gas sample may be described in terms of four variables—pressure, temperature, volume, and amount (moles).
3. The general gas law describes relationships among the variables (P , T , V , n) of an ideal gas. Relationships often described in textbooks as Boyle's law, Gay Lussac's law, Charles' law, and Avogadro's law can be regarded as subsets of the general gas law.
4. The behavior of gases can be explained in terms of the kinetic molecular theory.
5. Gas densities and molar masses can be determined experimentally through applications of the general gas law.
6. Gas particles, when compared with those of liquids and solids, are relatively far apart from one another.

CENTRAL CONCEPTS

1. Mass, volume, pressure, temperature, amount of substance
2. Energy
3. Molecules and atoms
4. Stoichiometry
5. Variables, constants
6. Density
7. Molar mass

RELATED CONCEPTS

1. Problem solving
2. Algebra
3. Measurement
4. Graphing
5. Temperature scales

RELATED SKILLS

After completing their study of gases, students should be able to:

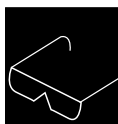
1. imagine gases moving in a given volume at a given temperature and describe the behavior of an ideal gas.
2. define and/or explain the meaning of pressure, temperature, volume, and amount of substance (moles).
3. predict the result of changes involving two variables (P , V , n or T) when others are held constant.
4. use the general gas law in calculations, either in one- or two-condition problems.
5. use the general gas law to calculate molar mass and gas density.
6. understand the importance of gases in the environment.
7. recognize and discuss environmental problems concerning atmospheric gases or global warming.

PERFORMANCE OBJECTIVES

Concept/Skills Development



LABORATORY ACTIVITY: STUDENT VERSION



Activity 1: Volume and Temperature Changes in a Gas

Introduction

Gases may be studied in relation to certain variables: pressure, volume, temperature, and amount (moles). In this laboratory activity we will develop a relationship between the volume and temperature of a gas sample and, by extrapolation, estimate the value of absolute zero, theoretically the lowest possible temperature obtainable.

Purpose

To determine the relationship between the volume and the temperature of a gas and to estimate the value of absolute zero.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Whenever a burner is used, care should be taken to avoid burns.
3. Do not leave thermometer in beaker between measurements.

Procedure

1. Completely fill an empty pipet with water.
2. Count the number of drops it takes to empty the pipet. (Repeat for better accuracy using a new pipet.)
3. Record the average number of drops. This number represents the volume of the pipet. It also represents the volume of gas at the higher temperature in this activity.
4. Fill a tray half full of water. Record room temperature. Adjust the temperature of the water until it equals that of the room.
5. Half fill the beaker with water. Heat until the water is about 20 °C above room temperature.
6. Holding a new pipet by the stem, immerse its bulb in the warm water in the beaker. Hold in the water for 1 min to allow the temperature of the air in the pipet become equal to that of the water.
7. Pinch the end of the pipet stem with pliers or a fingernail to seal it off. Record the temperature of the warm water in the beaker.
8. Fully submerge the pipet in the tray containing room-temperature water. Make sure that the stem is under the water and unseal. What happens in the stem of the pipet? Keep the pipet submerged until no further changes are noted (about 1 min).
9. Remove the pipet from the water and dry the outside. Count the number of drops of water that were drawn into the pipet. This number, subtracted from the original number, represents the volume of the air sample at the lower temperature.

10. Increase the temperature of the water bath by 10°C and repeat the procedure using a *new, dry* pipet.
11. Repeat Steps 8-10 until a temperature of about 90 °C is reached.
12. Thoroughly wash your hands before leaving the laboratory

	Temp. °C	Volume at higher Temp. (in drops) (Volume of empty pipet)	Volume at lower Temp. (in drops)
Room Temp.			
1			
2			
3			
4			

Data Analysis

Prepare a graph of your data, with temperature (in degrees Celsius) on the horizontal axis (x axis) and total air volume in the bulb (in drops) at that temperature on the vertical axis (y axis). The x axis should be numbered from $-350\text{ }^{\circ}\text{C}$ to $150\text{ }^{\circ}\text{C}$. You will have two data points—involving a high- and low-temperature volume—to include on each line. Draw a straight line connecting each set of two points. Extend each line so that it intersects the horizontal axis (x axis). You will thus obtain several estimates of absolute zero—one for each graph line. Decide on a “best” estimate for absolute zero, based on your overall graph results.

Implications and Applications

1. What is the relationship between the volume and the Celsius temperature as shown on your graph?
2. Absolute zero theoretically measures the limit of molecular motion. What is the volume of an ideal gas sample at absolute zero?
3. If we created a new temperature scale where the value of absolute zero were zero, what would be the value of $0\text{ }^{\circ}\text{C}$?
4. Superconductors function best at very low temperatures, below the boiling point of liquid nitrogen ($-195.8\text{ }^{\circ}\text{C}$). Select one of your graph lines. What would be the volume of that air sample at $-195.8\text{ }^{\circ}\text{C}$?



**LABORATORY
ACTIVITY:
TEACHER
NOTES**

***Activity 1: Volume and Temperature
Changes in a Gas***

Major Chemical Concept

The volume of a gas sample is directly proportional to its absolute (Kelvin) temperature. The volume of an ideal gas theoretically becomes zero at $-273\text{ }^{\circ}\text{C}$, or 0 K .

Level

This laboratory activity can be done by basic, general, and honors students.

Expected Student Background

Students should be aware of the variables (pressure, temperature, moles of substance, and volume) by which gas samples are described. They should know the relationship between gas temperature and its average kinetic energy. They should be reminded about construction of graphs and how to draw the best-fit line.

Time

If students work in pairs, the experimental work can be completed in 40 min. Additional time is needed for analysis of data.

Safety

No additional precautions.

Materials (For 24 students working in pairs)

- 12 Beakers, 800-mL, half filled with tap water (Allow to come to room temperature)
- 12 Trays or shallow plastic containers, containing enough water to cover pipet. (Allow to come to room temperature)
- 96 Plastic (Beral™) pipets, small, new. These must be dry. (Pipets may be obtained from many supply houses including Flinn Scientific Inc., Batavia, IL 60510.)
- 12 Thermometers ($^{\circ}\text{C}$)
- 12 Burners
- 12 Ring stands with iron rings
- 12 Wire gauze
- 12 Pinch pliers

NOTE: A hot plate may be used in place of the burner and ring stand.

Advance Preparation

Remind students of the variables in the gas laws. In this activity, temperature and volume are measured, while pressure remains constant. Four graph lines are drawn; each data pair involves a sample with different gas amount (moles).

Pre-Laboratory Discussion

Before students start the laboratory activity, demonstrate how to fill the pipet and how to count drops. To fill, squeeze as much air as possible from the pipet and insert stem in water. There will be some trapped air; it can be removed by holding the pipet vertically (bulb down), squeezing out the air so the stem is filled with water, then inserting the stem in water to fill completely. The bulb must be squeezed very gently to count each separate drop. A typical pipet contains between 100 and 175 drops.

Suggested pre-laboratory discussion question: What happens to the volume of an inflated balloon when you place it in hot water or very cold water?

Teacher-Student Interaction

Students will need help filling their pipets. The best way to seal a pipet is to pinch it with a fingernail. Students should count drops carefully and be sure to use a new pipet for each determination. Students will need help drawing graphs—scaling should include temperatures in the range +90 °C to -300 °C.

Some suggested questions to pose to students during the activity: What gas are we using? [*Air—a mixture of nitrogen and oxygen gases with traces of other gaseous substances.*] What is being held constant in this laboratory activity? [*The pressure.*] How could we determine the pressure? [*Barometer*] Why do you need to keep the pipet stem below the water? [*To keep air from getting in or out.*] Why does some water enter the pipet as it cools? [*When the air is heated the volume increases and some air leaves the pipet. As the pipet cools, the air contracts and draws some water into the stem.*] Do the four graph lines have the same slope? [*No*] Why do they all converge at approximately the same temperature at a “zero” volume? [*V is directly proportional to T (absolute temperature)—if V becomes zero, then T (absolute) also becomes zero, despite the amount of gas in the sample.*]

Anticipated Results

	Temp. °C	Volume at higher Temp. (in drops) (Volume of empty pipet)	Volume at lower Temp. (in drops)
Room Temp.	23	170	170
1	44	170	157
2	67	170	147
3	78	170	141
4	89	170	136

Answers to Implications and Applications

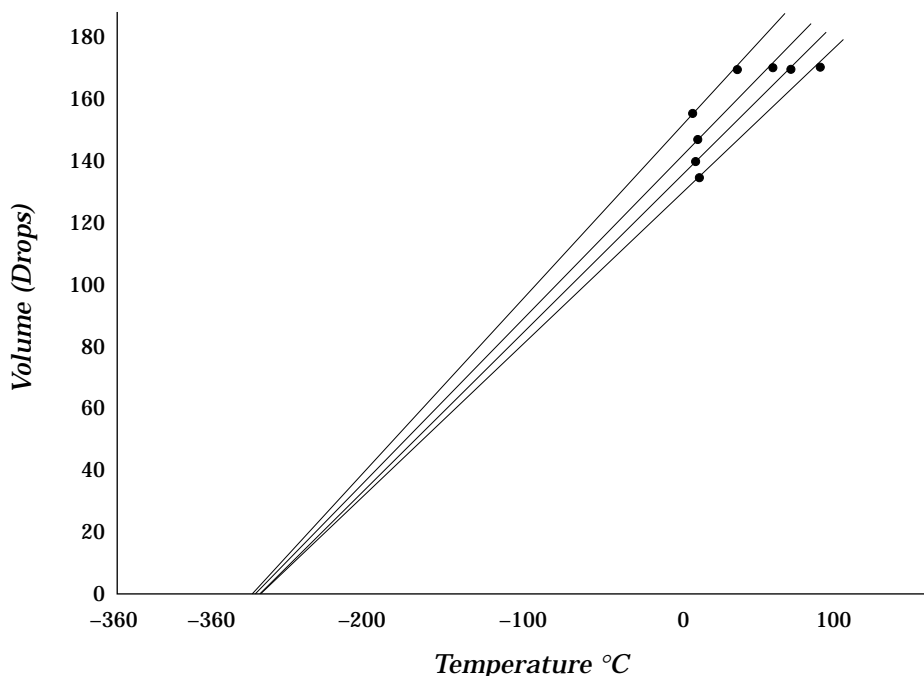


Figure 1. Temperature versus volume (drops) of sample data.



1. There is a direct relationship.
2. Volume is zero.
3. Guide students to draw new temperature scale along the side of graph starting at the intercept that should be about -300°C . Give them the accepted value of -273°C or 0 K . With the new scale, 0°C would equal 273 K .
4. Can be found from students' graph.

Post-Laboratory Activities

1. Show sample problems that relate gas temperature to volume.
2. Use graph to determine temperature at which air will have a certain volume and *vice versa*.
3. Demonstrate Charles' Law using a 30-mL plastic hypodermic syringe. Trap about 20 mL of air inside and close with cap. Place syringe in hot water; measure the change in volume. Trap 10 mL of air and repeat, showing that graph of T-V slopes differ.
4. Have students convert temperatures from kelvins to degrees Celsius, and vice versa.
5. Discuss the concept of absolute zero and note that although molecular motion will cease, motion within molecules will continue.

Extensions

1. Have students complete a related activity using a syringe rather than a pipet. If two series are done, one with 10 mL trapped air and one with 20 mL, two lines will be obtained with different slopes. With more able students the slopes can be determined; they will differ by a factor of two due to the difference in the moles of gas:

$$PV = nRT$$

$$V = (nR/P)T$$

$$V = (k) T$$

$$(nR/P) \text{ is a constant} = k$$

$$k = \text{slope of plotted line on a } V \text{ vs. } T \text{ graph}$$

Thus, if n doubles, k (slope) doubles

2. Allow students to experiment with inflated balloons in a dry ice/isopropyl alcohol mixture (*CAUTION: extremely cold*) and in warm water, comparing balloon sizes.
3. Have students construct graphs from other data and determine the value of absolute zero. Use milliliters rather than drops for the volume units.
4. Charles discovered the temperature-volume relation for gases due to personal interest in balloons. Discuss the use of balloons in warfare and for domestic purposes before airplanes were invented.

Assessing Laboratory Learning

1. Use the questions suggested (in *Teacher-Student Interaction* section) on a test.
2. Have students determine the value of absolute zero from a graph constructed from other data.

Activity 2: Analyzing the 'Pop' in Popcorn

Introduction

Corn is a common foodstuff, native to the Americas, which appears in many forms: corn on the cob, corn off the cob, creamed corn, and popcorn. Field corn, which is also fed to livestock, is hung on our doors at Halloween. Each variety of corn contains different amounts of water, sugars, and starches. Popcorn is a favorite as an evening snack. Popping popcorn involves heating the corn until the pressure inside the kernel is great enough to cause it to burst, turning the kernel inside out and releasing the trapped moisture.

Purpose

To determine the percent water in popcorn, to use the ideal gas law to determine the pressure inside the kernel when it pops, and to solve gas law problems.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Heat the flask evenly to prevent spattering the oil.
3. Heat with care. Oil is flammable.
4. Do not eat the popcorn.

Procedure

1. Record the mass of a weighing cup.
2. Add 16 kernels of one type of popcorn and reweigh.
3. Calculate the mass of the popcorn kernels.
4. Using the water displacement method, find the volume of the 16 kernels using a 10-mL graduated cylinder. Dry the kernels.
5. Add two medicine droppers of cooking oil and the 16 kernels to an empty, dry 125-mL Erlenmeyer flask.
6. Determine the mass of the flask, oil, and popcorn.
7. Assemble a ringstand with ring, placing a wire gauze on top of the ring.
8. Using a utility clamp, fasten the flask to the stand so it sits on the wire gauze. Cover the flask with a piece of copper gauze. Bend the sides of the copper gauze around the flask's mouth.
9. Light the burner.
10. Hold the burner, moving it back and forth slowly until the popcorn just begins to pop. Remove the heat when most of the kernels have popped. *Do not burn the popcorn.* If you do, you'll need to start over.
11. If you see water condensed on the upper part of the flask, heat that part very gently with the burner until it disappears.
12. Let the flask cool, then remove and determine the mass of the flask and contents. (Remove the copper gauze before weighing.)
13. Thoroughly wash your hands before leaving the laboratory.

LABORATORY ACTIVITY: STUDENT VERSION



Sample Data Table



Brand Name _____

Mass of corn + weighing cup (g)

Mass of weighing cup (g)

Volume of water (mL)

Volume, water and corn (mL)

Mass of flask, oil, and unpopped popcorn (g)

Mass of flask, oil and popped popcorn

Sample Calculations Table

Mass of unpopped kernels (g)

Mass of popped kernels (g)

Volume of popcorn (mL)

Volume of popcorn (L)

Mass of water lost (g)

Moles of water lost

*Percent water in kernels

**Pressure of steam inside kernel at time of pop

Help with Your Calculations

Please show all work here:

$$\text{*Percent of water in "popcorn"} = \frac{\text{Mass of water lost}}{\text{Mass of 16 kernels}} \times 100$$

** To estimate the pressure of the water vapor at the time of "pop," use $PV = nRT$

Here are some hints regarding this calculation:

Step 1. Rearrange the equation to solve for P .

Step 2. Use the "moles of water lost" from the data table.

Step 3. $R = 0.0821 \text{ L atm/(mol K)}$

Step 4. Assume that the popcorn pops at the boiling point of the cooking oil (225°C) and convert this temperature to kelvins.

Step 5. Pressure will be in atm.

Implications and Applications

1. What brand of popcorn did you use?
2. What was the percent water in your popcorn?
3. Name one way in which popcorn, corn on the cob, and field corn are different.
4. What is the pressure of the water (as a gas) inside the popcorn just as it pops?
5. What was the atmospheric pressure in the laboratory during this activity?
6. Compare the pressure required to “pop” corn (Question 4) with atmospheric pressure (Question 5).
7. What assumptions have you made that might cause errors in your calculation of pressure?



**LABORATORY
ACTIVITY:
TEACHER
NOTES**

Activity 2: Investigating the 'Pop' in Popcorn

Major Chemical Concepts

This activity involves using the gas laws to determine the pressure inside popcorn kernels when they pop. Students will apply general gas law equations and discuss differences among various types of corn.

Level

This laboratory activity is written for a general or honors level student. It involves solving the gas law equation and requires algebra skills.

Expected Student Background

Prior to conducting this laboratory activity, students should have extensive practice with Boyle's law, Charles' law, combined gas law, ideal gas equation, the kelvin temperature scale, converting mass of water to moles of water, percent composition by mass, and using the water displacement method to determine the volume of irregular solids.

Time

A 55-min laboratory period is ample time to complete the activity. It will help to place needed hardware at laboratory stations and the oil and popcorn at a central location.

Safety

1. Vegetable oil is flammable. Students should heat the flask with care.
2. Students must fasten the ring onto the ring stand securely. The flask must also be securely fastened to the ringstand with a utility clamp.
3. Warn students not to eat the popcorn. Prepare other popcorn for them to enjoy outside the laboratory area.

Materials (For 24 students working in pairs)

12 Erlenmeyer flasks, 125-mL
12 Graduated cylinders, 10-mL
12 Copper gauze, or aluminum window screens, 1 piece, 5-cm x 5-cm
Pan balance
12 Ringstands
12 Iron rings
12 Wire gauze, 5" x 5"
12 Burners
12 Utility clamps
12 Packs of matches or flint lighter
12 Weighing cups, plastic

At a Central Location

Cooking oil, 3-4 containers
3-4 Medicine droppers
Popcorn, 3-4 brands
Barometer (or call radio/television station for barometric pressure)

Pre-Laboratory Discussion

1. Provide samples of corn on the cob, corn off the cob, field corn, and popcorn for students to inspect. What are some differences among these types of corn?

2. Discuss the process of popping popcorn.
 - a. When you heat popcorn, what happens to the water inside the popcorn kernels?
 - b. Why don't some popcorn kernels pop?
3. What is the pressure of 0.010 mol of water vapor at 200 °C if it occupies a volume of 0.0020 L?
4. Calculate the pressure of the 0.010 mol of water vapor at 150 °C if its volume is held constant at 0.0020 L.
5. What is the relationship between pressure and temperature?

Teacher-Student Interaction

While circulating through the room, ask students questions such as: Do you expect the mass of popcorn to increase or decrease after it has popped? Why? *[The mass of the popcorn should decrease after it is popped because popping releases water vapor into the atmosphere.]* Do you expect the pressure inside popcorn kernels just before it pops to be greater than or less than atmospheric pressure? *[The pressure inside the kernels must be greater than atmospheric pressure to rupture the kernel "shell."]* What is the purpose of the cooking oil? *[The cooking oil surrounds and softens the kernel shell and provides a medium through which heat can be transferred from the glass to the popcorn.]*

Advance Preparation

1. Place all hardware at each laboratory station.
2. Place cooking oil, medicine droppers, and popcorn at a central location.
3. Decide whether all students should pop the same brand of popcorn or compare different brands of popcorn.
4. Prepare popcorn using a hot-air popper while students are completing the activity. Invite them to eat the popcorn *you* prepared, in a separate area, after they have cleaned all materials, washed their hands, and are working on the calculations.

Calculations

Mass of corn + weighing cup (g)	2.17 g
Mass of weighing cup (g)	0.17 g
Mass of popcorn (g)	2.00 g
Volume of popcorn (mL)	1.4 mL
Volume of popcorn (L)	0.0014 L
Mass of flask, oil, and "unpopped" popcorn (g)	80.7g
*Mass of flask, oil and "popped" popcorn (g)	80.5g
** Percent water in kernels	10%
*** Pressure of water inside kernel at time of pop (atm)	300 atm

*From this value (and the preceding value) students can find the mass of water lost in popping.

$$\text{**Percent water in kernel} = \frac{\text{Mass of water lost}}{\text{Mass of 16 kernels}} \times 100 = \frac{0.2 \text{ g}}{2.0 \text{ g}} \times 100 = \mathbf{10\%}$$

***Pressure inside kernel at time of pop. Use $PV = nRT$; solve for P .



$$n = 0.2 \text{ g} \times \frac{1 \text{ mol H}_2\text{O}}{18 \text{ g}} = 0.01 \text{ mol H}_2\text{O}$$

$$P = \frac{nRT}{V} = \frac{(0.01 \text{ mol}) (0.0821 \text{ L} \cdot \text{atm} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}) (273 + 225 \text{ K})}{0.0014 \text{ L}}$$

$$P = 292 \text{ atm, or } 3 \times 10^2 \text{ atm}$$

Answers to Questions

Pre-Laboratory Discussion Questions

- Different types of corn contain varying amounts of water. Corn on the cob and corn fresh off the cob contain the largest percent water. Popcorn and field corn contain less water. Fresh corn on cob will not work as popcorn because there is no hard kernel to allow pressure buildup.
 - When the popcorn is heated, water in the kernels becomes hotter and vaporizes.
 - Unpopped kernels possibly were too dry (insufficient moisture) and failed to build up sufficient pressure to pop, or had porous shells allowing gas to escape.

3.

$$P = \frac{nRT}{V} = \frac{(0.010 \text{ mol}) (0.0821 \text{ L} \cdot \text{atm} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}) (273 + 200 \text{ K})}{0.0020 \text{ L}}$$

$$P = 1.9 \times 10^2 \text{ atm}$$

4.

$$\text{Use } \frac{P_1}{T_1} = \frac{P_2}{T_2} \quad \frac{(1.9 \times 10^2 \text{ atm})}{(273 + 200 \text{ K})} = \frac{x}{(273 + 150 \text{ K})}$$

$$x = P_2 = 1.7 \times 10^2 \text{ atm}$$

Answers to Implications and Applications

- Answers will vary depending on the brand used.
- Answers will range from 2% to 10% water.
- Corn on the cob contains a much larger percentage of water. Field corn and popcorn contain much less water.
- The pressure inside the popcorn will vary greatly. Answers should, however, range from 100 to 350 atm.
- Read barometer and convert to atmospheres.
- It is much larger.
- Throughout the laboratory activity, the volume of popcorn kernels was assumed equal to the volume of displaced water. The pulp of the popcorn occupies much of the space inside the kernel; the volume recorded is larger than the actual volume occupied by the water. The kernel temperature at the time of popping was not directly measured. The assumption of 225 °C used in calculations is probably not highly accurate. Thus, for all these reasons, calculated results contain some degree of error.

Bell Jar Demonstrations

- Using a vacuum pump, evacuate a bell jar containing a small quantity of shaving cream or a marshmallow. Expansion rules!
- Obtain a small vial (without cap). Put a bit of shaving cream in vial mouth to “plug” the opening. Place vial in the bell jar. Evacuate the bell jar and watch the shaving cream expand.

Diffusion and Convection

- Pictures in the mind of gas filling a room.

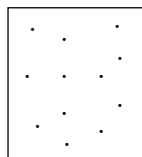
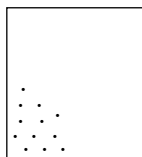
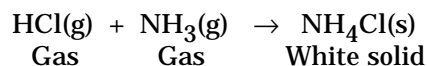


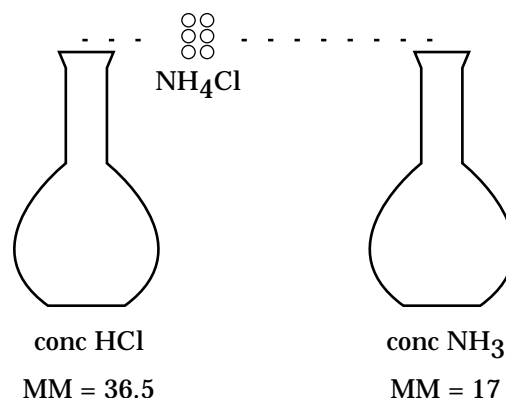
Figure 3. Gas filling a room.

- Pictures and discussion of formation of ammonium chloride, NH_4Cl :



The solid is formed closer to the HCl source because the less dense gas (NH_3) diffuses faster.

Figure 4. Formation of NH_4Cl from HCl and NH_3 .



GROUP AND DISCUSSION ACTIVITIES

Pressure = Force/Area

- Difference between spike heel and large heel** Because spike heel has smaller area, it will exert more pressure on the floor than large heel, assuming the force (body weight) is the same.
- Atmospheric pressure** (effect on body, changes on mountain). Compare to changes of state. Atmospheric pressure decreases as you move away from the earth because there are fewer gas molecules per unit volume. At high altitudes, we may have trouble breathing because of a lower O_2 concentration and because of the reduced pressure. The reduced pressure also affects the change from liquid to gas; the boiling point is lowered because the atmospheric pressure is lower. (Boiling point is the temperature at which the vapor pressure equals atmospheric pressure). Cake mixes will have different directions for higher altitudes. (Show cake mix box.) It also takes longer for boiled foods to cook at high altitudes due to the lowered boiling point.
- Design and use of a barometer** (construction of barometer involves safety hazards). Convert barometric pressure in inches of mercury to cm Hg, to mmHg.

$$1 \text{ inch} = 2.54 \text{ cm.}$$

- Measure blood pressure**
- Question** How tall would a barometer column under one atmosphere pressure need to be if the column were filled with water, not mercury?
[$13.53 \times 760 \text{ mmHg} = \text{the ratio of density of Hg to that of water}$]

Gas Densities

Balloons, dirigibles, and blimps float in air because they contain gases that are less dense (lighter) than ordinary air. Helium, hydrogen, or hot air may be used. The densities of these gases are all less than the surrounding air, since the mass of gas in a given volume is less. When unconfined air is heated, its molecules move faster (on average) and the volume expands. Hydrogen and helium have lower molar masses than most gases composing air (such as oxygen, nitrogen, water vapor, carbon dioxide, etc.). Dirigibles were used often in World Wars I and II for observation and transportation. Germany used hydrogen gas because it had no access to helium, which is found in large quantities in the U.S. Hydrogen is very explosive; the German airship *Hindenberg* exploded near Lakehurst, New Jersey in 1937, with great loss of life.

Mixtures of Gases and the Atmosphere

The atmosphere is a mixture of 80% nitrogen and 20% oxygen with traces of argon and other gases. Total atmospheric pressure is the sum of pressures of each individual gas. Even though oxygen and nitrogen have different densities, they mix completely in the atmosphere because of the ability of a gas to fill any space it occupies.

Analogies and Metaphors

1. Use plastic ruler. Hang sign indicating temperature (T) in middle. Show that if T is constant, see-saw effect takes place and P is inversely proportional to V . If V is held constant, increasing P , increases T , etc.

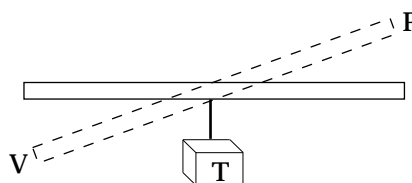


Figure 5. Seesaw analogy.

2. The analogy of a school dance can be used to illustrate many gas relationships. If you increase the size of the room (volume), the pressure of dancers decreases.

Pictures in the Mind

The study of gases provides a good example of relationships among variables and the identification of variables and constants. It also provides excellent opportunities to create “pictures in the mind.”

General Questions

1. Predict what will happen as you change certain conditions of gases, holding other conditions constant (see *Tips for the Teacher*).
2. How can you determine the molar mass of a gaseous substance? [If you measure the mass of a sample of gaseous substance at a known temperature, pressure, and volume, it is possible to determine the molar mass using the general gas law, $PV = nRT$. Using the known temperature, pressure, and volume, the amount (moles) of gas can be calculated. The measured mass divided by the amount of gas provides the molar mass. See *Tips for the Teacher* for further explanation.]
3. What is an ideal gas? [One in which the molecules occupy no volume and in which there are no attractions among molecules. Of course this is an ideal, not a real, situation, but real gases at room temperature and standard atmospheric pressure (1 atm) can behave nearly ideally since their molecules are so far apart. When gases are under high pressure and low temperature, their molecules come closer together and no longer act ideally.]
4. Why is it necessary to change the temperature scale to kelvins in applications of gas laws? [So that “zero” on the temperature scale represents the lowest possible temperature. With such a temperature scale, ideal gas volume is directly proportional to temperature.]



TIPS Language of Chemistry FOR THE TEACHER

absolute zero temperature at which molecular motion ceases; $0\text{ K} = 273.15\text{ }^\circ\text{C}$.

barometer device used to measure atmospheric pressure.

Boyle's law states that the volume of a fixed amount of gas at a constant temperature is inversely proportional to its pressure.

Charles' law states that the volume of a fixed amount of gas at a constant pressure is directly proportional to its absolute (kelvin) temperature.

gas form or state of matter in which a material assumes the shape of its container and expands to fill the container (thus having neither definite shape nor volume).

gas constant R numerical constant appearing in ideal gas equation ($PV = nRT$).

ideal gas one whose behavior can be predicted by the ideal gas equation.

ideal gas equation relates the pressure, volume, temperature, and amount of gas through the expression $PV = nRT$.

manometer device used to measure the pressure of a gas, usually by comparing the gas pressure to barometric pressure.

non-ideal gas departs from the behavior predicted by the ideal gas equation. Its behavior can only be predicted by other equations of state.

partial pressure pressure exerted by an individual gas in a mixture, independently of other gases.

Common Student Misconceptions

Misconceptions such as those listed are best dispelled with demonstrations followed by class discussion.

1. "There cannot be any pressure around us, since we cannot feel it."

a. In an airplane you can feel the pressure in your ears, particularly during take offs and landings. When opening a soda can you can hear the change in pressure as the gas escapes.

b. Demonstrate the breaking of a ruler covered with one sheet of newspaper by placing a cheap yardstick on a table with about half of it extending past the edge. Carefully spread a full page of newspaper over the portion of the yardstick on the table. (Press paper against the yardstick to expel air, smoothing the paper flat against the tabletop.) Then, with a karate-like blow, chop off the portion of the yardstick that extends beyond the table. Students can calculate amount of pressure on newspaper.

2. "Gases do not weigh anything." (See *Demonstrations* section of module.)

a. Wrap a 250-mL Erlenmeyer flask with strips of masking tape to prevent injury in case of an implosion. Fit the flask with a one-hole stopper, a 5-cm piece of glass tubing, and a 5-cm piece of rubber tubing with a clamp. Measure the mass of the assembly when full of air. Evacuate moderately with a vacuum or hand pump. Fasten clamp tightly and quickly weigh again to show the loss of mass. Demonstrate that the assembly returns to its original mass when the clamp is opened.

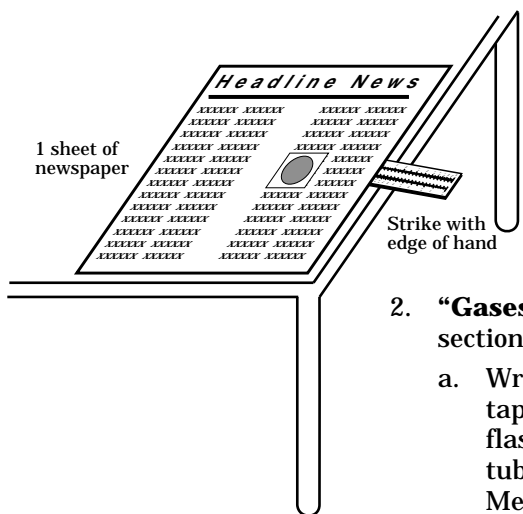


Figure 6. Air pressure demonstration set up.

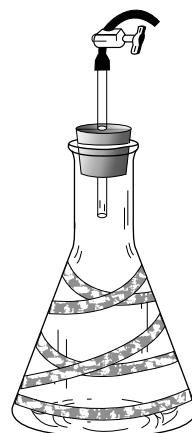


Figure 7. Wrapped Erlenmeyer flask.

- b. Set up a balloon balance and add air to one balloon and CO₂ or methane to the other. The balance will tip to the side containing the higher-density gas sample.

3. “Hand me an empty glass.”

This demonstration shows that the air will be displaced by water flowing into the flask; thus, the flask was not really “empty”. This demonstrates also how gases can be collected “over water,” provided the gas is not very soluble in water. (Hydrogen and oxygen can be collected this way; chlorine, ammonia and hydrogen chloride cannot.)

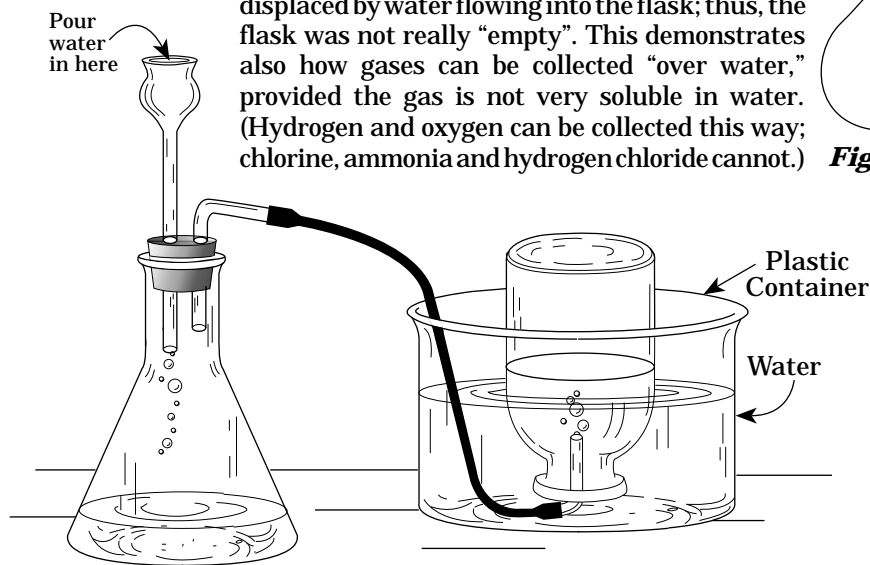


Figure 9. Apparatus to demonstrate water displacement by air.

4. “When a gas expands the particles get larger.”

The molecules of a gas sample stay the same size, but distances between the molecules get larger when the sample expands.

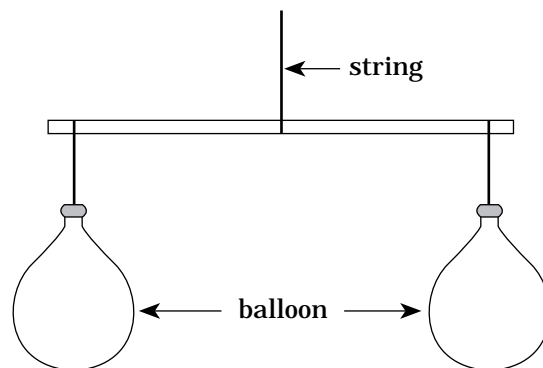


Figure 8. Balloon balance set up.

Problem Solving

Although many texts approach the gas laws as separate conditions, with separate equations such as $P_1 V_1 = P_2 V_2$ for students to memorize and use; these concepts can also be taught in terms of only one equation, $PV = nRT$.

Because some students may not have an adequate background in algebra, presenting the gas laws logically will help them understand these concepts. For instance, if the two variables are pressure and volume (the text may call this Boyle’s law) we know that they vary inversely. As one increases, the other decreases. Therefore we can determine logically the direction of the change. It is useful to start by setting up a data table:

Example: What will the final pressure of a gas sample be if it is initially at 2.0 atm and the volume is changed from 42 L to 124 L at constant temperature?

Remember that T must always be in kelvins (or other absolute temperature scale). Students should predict whether the pressure will increase or decrease (if volume increases, pressure will decrease) and set up the relationship accordingly.

$$2.0 \text{ atm} (42 \text{ L} / 124 \text{ L}) = ? \text{ atm}$$

The fraction must be less than one to realize a decrease in pressure.

	Condition I (initial)	Condition II (final)
P	2.0 atm	? atm
V	42 L	124 L
n	Constant	Constant
T	Constant	Constant



If the two variables are pressure and temperature (Gay-Lussac's law), they are directly proportional. As one increases the other increases.

If the two variables are volume and temperature (Charles' law), they are also directly proportional.

If there are more than two or more variables, each set or pair should be treated individually, either in two steps (using two tables) or in a combined equation.

Example: What is the volume of a gas sample at 0.95 atm and 25°C, if its volume is 26 L at 1.2 atm and 14°C?

First, use the relationship between P and V . Since the pressure *decreases*, the volume will *increase*.

	Condition I (initial)	Condition II (final)
P	1.2 atm	0.95 atm
V	26 L	? L
T	14 °C + 273 = 287 K	25 °C + 273 = 298 K
n	Constant	Constant

$$26 \text{ L} (1.2 \text{ atm}/0.95 \text{ atm}) = 33 \text{ L}$$

If the temperature *increases*, the volume will also *increase*. Use the new volume found above:

$$33 \text{ L} (298 \text{ K}/287 \text{ K}) = 34 \text{ L}$$

If the table shows that only one set of conditions is involved, then $PV=nRT$ should be directly used.

Example: How many moles of a gas will occupy 2.7 L at 36 °C and a pressure of 0.95 atm?

Note that all units used must match those in the constant (R).

	Condition I (initial)	Condition II (final)
P	0.95 atm	-
V	2.7 L	-
n	?	-
T	36 °C + 273 = 309 K	-

$$PV = nRT \qquad n = PV/RT$$

$$n = \frac{(0.95 \text{ atm})(2.7 \text{ L})}{(0.0821 \text{ L} \cdot \text{atm} \cdot \text{mol}^{-1} \cdot \text{K}^{-1})(309 \text{ K})} = 0.10 \text{ mol}$$

Note that all units used must match those in the constant (R).

	Condition I (original problem)	Condition II (STP)
P	0.95 atm	1 atm
V	2.7 L	22.4 L
n	? mol	1 mol
T	36 °C + 273 = 309 K	273 K

If students lack the algebra background to complete such one-condition problems, suggest that they set up such problems using the molar volume of gas at standard temperature and pressure. This will allow them to treat it the same way as previous problems.

HISTORY: ON THE HUMAN SIDE

1. Robert Boyle (1627-1691) was the son of one of the richest men in England, who had made his fortune acquiring land in Ireland, some of it from Sir Walter Raleigh. The wealth of his father allowed Boyle the leisure to become a scientist; his talent at experimentation resulted in many new insights into nature. He was a chemist as well as a physicist and approached the study of science in a manner more modern than most of his contemporaries.

2. Before development of airplanes, balloons were a common type of airship. The French were pioneers in ballooning, beginning with the historic ascent by the Montgolfier brothers in 1783 in a hot air balloon. In the same year Charles introduced hydrogen gas that caused the balloon to rise without a burner. During the wars of the French Revolution, balloons were used as observation posts, although they were difficult to steer and were not used as transport. Gay-Lussac's ascents in 1804 were the first scientific expeditions using balloons. Readings of temperature, pressure, humidity, and of magnetic field were taken as well as samples of air, which were later analyzed. Gay-Lussac's ascent broke a record for altitude—7 km above sea level—which was not surpassed for nearly 50 years.
3. Boyle discovered the relationship between the pressure and the volume of a gas in 1662, but Charles and Gay-Lussac did not state the relationship between temperature and volume or temperature and pressure until about 1800. Why? Because accurate thermometers were not available.
4. Individual gases were not discovered until late in the 18th century. Oxygen was first recognized by C.W. Scheele in Sweden, but his work was not published until after that of Joseph Priestley, an Englishman, who was forced to leave his native country in 1794 because of his political beliefs. Priestley also described carbon monoxide, sulfur dioxide, and ammonia. Henry Cavendish, the eccentric grandson of a Duke, isolated hydrogen and nitrogen.
5. About 50 years ago, the scourge of polio appeared each summer, with an annual average of 21,000 cases of crippled Americans. An effective vaccine was developed in 1955 and polio virtually disappeared. Iron lungs were invented to help victims who had their diaphragms paralyzed by the disease. Many polio victims lived for years in these artificial respiration machines, which used pressure to draw air in and out of the lungs.

1. Student responses on an exam:
 - a. Hydrogen is often found in snow and rain water.
 - b. Hydrogen is prepared by freezing air.
 - c. Hydrogen is produced by formation of sugar.
 - d. Two zincs plus two hydrochloric acids plus two chlorides will produce hydrogen.
 - e. Oxygen has a strong odor; it extinguishes a flame.
 - f. Oxygen can be prepared by hydrochloric acid on hydrogen chloride.
 - g. Test for carbon dioxide by burning a flint.
 - h. Test for carbon monoxide: Canary plus CO equals no canary plus CO.



CHEM 13 NEWS, February 1983, p. 14

HUMOR: ON THE FUN SIDE



3. TWO GAS PUNS:

- Albert the German astronomer goes out tonight to watch the stars. "It must be cloudy," he exclaims, "for I Xenon."
- Said Benny the burglar, "It's time that I made another Radon the bank."

CHEM 13 NEWS, April 1980, p. 7

4. BOYLE'S LAW: TWO ANECDOTES

- When stewardesses on commercial planes complained that their skirts fitted on the ground, but not in the air, an airline spokesman blamed it all on Sir Robert Boyle's Law. This law says that if the quantity and temperature of a gas remain constant, its volume will vary inversely with pressure.

The application of the law to the skirts is simple: Air pressure decreases as an airplane ascends, and thus the pressure on the gas in a stewardess' stomach lessens, which means the volume of the gas increases. In other words, her tummy bulges. Stewardesses now wear adjustable skirts.

- During the construction of a tunnel under a river, a party of politicians went down to celebrate the meeting of the two shafts. They drank champagne and were disappointed that it was under depth pressure—the carbon-dioxide bubbles remained in solution. When the town fathers arrived at the surface, the wine popped in their stomachs, distended their vests and all but frothed from their ears. One dignitary had to be rushed back into the depths to undergo champagne recompression.

CHEM 13 NEWS, March 1973, p. 572

5. Word Search (see Appendix for master copy)

A B E W L A E Z S A D M T T C B U
D A E M Y Q V R E T E M O N A M W
P R E S S U R E Z B L U B Z H Z E
Z O E D A N I E U P S W N A Q T T
W M W S Z Q G H X L B A A S D T L
P E L O M E N O Q E L Q G P A U I
X T I Q I P G W R P T J E N U Y W
E E X D O Q C Z I E M U Q L A M E
Y R K V L T Y V A G G I L O Y T K
K K N A C H A R L E S S O Q B O M
G U E P G S W C N J U F T V S A B
W D P V B U P D K L N U L W I B T
I K H I C N C C Q Q O Q N V A E A

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

- Gas whose behavior can be predicted by $PV = nRT$.
- Device used to measure the pressure of a gas.
- State of matter with no definite shape or volume.
- Gas law that states: $V = k/P$.

5. Device used to measure atmospheric pressure.
6. 0 K or -273.15 degrees Celsius (2 words)
7. Gas law that states: $V = kT$
8. This quantity of gas at STP occupies 22.4 L. (2 words)
9. Gaseous mixture of nitrogen, oxygen, and trace gases found in the atmosphere.
10. Force per unit area exerted by a gas in a closed container.

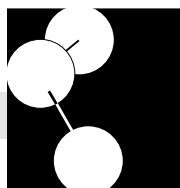
Answers: 1. IDEAL 2. MANOMETER 3. GAS 4. BOYLE 5. BAROMETER
6. ABSOLUTE ZERO 7. CHARLES 8. ONE MOLE 9. AIR 10. PRESSURE

6. See cartoons at end of the module.

1. CHEM Study film and video on gases, *Gases and How They Combine*. Available from Ward's Natural Science Establishment, P.O. Box 92912, Rochester, NY; (800) 962-2660.
2. The Cable Weather Channel periodically runs films on environmental issues that may be taped for class.
3. Software published by *JCE: Software*, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).
Ideal Gas, by Richard Hiatt, Vol. IV, No. 1, for the Apple II computer.
4. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 263-2837 (voice) or (608) 262-0381 (FAX).
 - a. For the Apple II computer: AP 401, AP 402, AP 404
 - c. For IBM PCs and PC-compatibles: PC 2601, PC 2602, PC 2603
5. Videodiscs published by *JCE: Software*, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

"Experimenting with Gas Pressure I," "Experimenting with Gas Pressure II," "Effect of Temperature on Pressure," and "Experimenting with Gas Pressure III," four chapters on *The World of Chemistry: Selected Demonstrations and Animations*: Disc I (double sided, 60 min.), Special Issue 3.

MEDIA



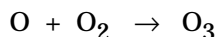
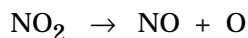
Links/Connections

WITHIN CHEMISTRY

1. Use of gases in stoichiometry (see *Extensions*).
2. Effect of pressure on changes of state; change of boiling point with change in atmospheric pressure.
3. Determination of molar masses of unknown gaseous substances by researchers.

TO THE CONTEMPORARY WORLD

1. **Acid Rain** The gases NO_2 and SO_2 contribute to acid rain. These gases dissolve in water droplets or aerosol particles, causing the pH to decrease. This moisture precipitates as acid rain or acid snow. NO_2 produces nitric acid (HNO_3) and nitrous acid (HNO_2); SO_2 produces sulfurous acid (H_2SO_3) and, after conversion to SO_3 , sulfuric acid (H_2SO_4). Rain is naturally acidic (pH = 5.6) from the CO_2 gas that dissolves in water to form carbonic acid (H_2CO_3). Acid rain has pH values below 5.6.
2. **Ozone Depletion by Chlorofluorocarbons (CFCs)** CFCs are halogenated hydrocarbons that find use as refrigeration fluids, degreasing solvents for electronic parts, fire extinguishers (CFCs are unreactive) and, at one time, aerosol propellants for hair sprays, deodorants, and food. Their use as propellants is now banned in the U.S. due to evidence that they help destroy the ozone layer. In the stratosphere where ozone is the most abundant, ultraviolet light breaks down CFCs forming a very reactive chlorine atom (Cl). The chlorine atom and an ozone molecule combine producing an oxygen molecule and another reactive atom. A chain reaction occurs consuming many more ozone molecules. The stratospheric ozone layer filters out harmful ultraviolet rays, preventing skin cancer.
3. **Photochemical Smog** Originally smog meant a combination of smoke and fog. Today it generally refers to the polluted air found over cities as a result of industrial emissions and automotive exhaust. The major chemical substances involved are hydrocarbons (unburned fuel), nitrogen oxides (which form at the high temperatures and pressures of internal combustion engines) and ozone. Ozone production is a light activated reaction and an example of a process contributing to photochemical smog. NO_2 decomposes in light to nitrous oxide and free oxygen. The latter combines with molecular oxygen to make ozone.

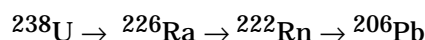


The ozone then reacts with various hydrocarbons to produce aldehydes, peroxy nitrates, acrolein, and other irritants. These induce health problems in some humans and produce a phytotoxic effect in plants. Even some carcinogens are found in smog. Air pollution problems are aggravated by local geographic conditions in cities such as Denver and Los Angeles. Consequently, gaseous emissions are restricted in these places.

4. **The Greenhouse Effect** Carbon dioxide is transparent to visible and ultraviolet light but absorbs infrared radiation strongly. As a consequence, the atmosphere of the earth is warmer than it would be without carbon dioxide. Incoming higher energy visible light is absorbed at the surface by pigmented objects and organisms. Some of the absorbed energy is emitted as lower energy infrared radiation that is then trapped in the atmosphere. Water vapor also contributes (and, indeed, is the major contributor) to the effect. This warming process is referred to as the Greenhouse Effect.

The CO₂ composition of the atmosphere has increased with the combination of fossil fuel use and decreasing forested land acreage. There is much concern that this effect might, or already has begun to, lead to global warming that could have disastrous changes in the size of the polar ice caps and distribution of arable land. Scientists are carefully monitoring this situation, and many individuals and governments are acting to reduce CO₂ emissions.

5. **Radon** A member of the noble gas family, radon-222 is a product of the uranium decay series and is radioactive. It is a direct result of the decay of radium-226.

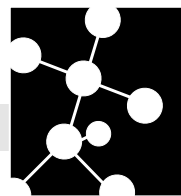


Radon-222 is found in soil and rocks. When radon-222 decays it produces alpha particles and several short-lived radioisotopes—all of which are alpha emitters. Since radon is a gas, it can be present in the air one breathes. If it decays while in the lungs (half life of radon-222 is about four days) atoms of Po and Pb (both solid elements) and alpha particles are formed. Although the Po and Pb cause problems, alpha radiation damages lung tissue because of close contact with the tissue. This can result in a higher than normal risk of lung cancer.

6. **Commonly Encountered Gaseous Air Pollutants**

Gaseous Pollutant	Comment
SO ₂ , sulfur dioxide	Related to acid rain
O ₃ , ozone	Pollutant in lower atmosphere, beneficial in stratosphere
CO, carbon monoxide	From incomplete combustion of gasoline in automobile
NO ₂ , nitrogen dioxide	From combination of N ₂ and O ₂ in air, in automobiles; related to acid rain problem
CO ₂ , carbon dioxide	From burning oil, gasoline, coal, and natural gas
CFCs, chlorofluorocarbons	Believed responsible for depletion of ozone layer
Rn-222, radon	Alpha-emitting gas found in rocks and soil

Extensions

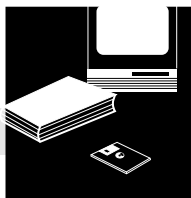


The stoichiometry of chemical reactions involving gases can be interpreted through the relationship between amount and volume. Because the volume of a gas is directly proportional to the amount of gas (in moles), volumes of gases can be used experimentally in place of amounts in stoichiometric problems, if temperature and pressure remain constant.

Example: For the reaction: $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g})$

2 volumes of $\text{H}_2(\text{g})$ react with 1 volume of $\text{O}_2(\text{g})$ to form 2 volumes of $\text{H}_2\text{O}(\text{g})$. By measuring the volume of water vapor formed, one can determine the volume of H_2 that reacted.

To determine the number of moles or grams of H_2 involved, one would use the general gas law in the form $n = PV/(RT)$.



References

Module developed by Jennifer Hubert, Jane Miller, and Marie Sherman, the Missouri team.

Crosland, M. (1978). *Gay-Lussac scientist and bourgeois*. Cambridge, MA: Cambridge University Press.

Good biography.

Davenport, D. A. (1983). How the right Professor Charles went up in the wrong kind of balloon. *ChemMatters*, 1 (4), 14-16.

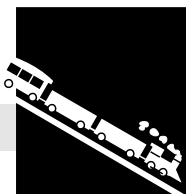
Kuslan, Louis, and Stone, A. (1970). *Robert Boyle*. Englewood Cliffs, NJ: Prentice-Hall, Inc.

Russo, T. (1985, 1986). *Microchemistry for high school general chemistry*.

A pioneer in microscale laboratory methods. Excellent reference.

Sibley, L. K. (1984, October). Popcorn. *ChemMatters*, 2(3), 10-13.

Appendix



- **Additional Demonstrations**

1. Place reagent bottles of concentrated NH_3 and concentrated HCl near each other. Remove both stoppers at the same time and note where the “white smoke” (NH_4Cl) forms. *CAUTION: Handle both reagents with care and avoid breathing vapors.*
2. Environmental
 - a. Permanently stretch a rubber band indoors and one outdoors. See how long each takes to deteriorate in the presence of ozone. (The destructive property of ozone was originally discovered due to its action on the rubber coverings of telephone lines.)
 - b. Reaction of SO_2 on leaves, stockings, cloth.
 - c. To illustrate temperature inversion, rest a 1-L graduated cylinder bottom on an ice bath. Heat cylinder top with lamp. Light a cigarette; drop it into the cylinder.

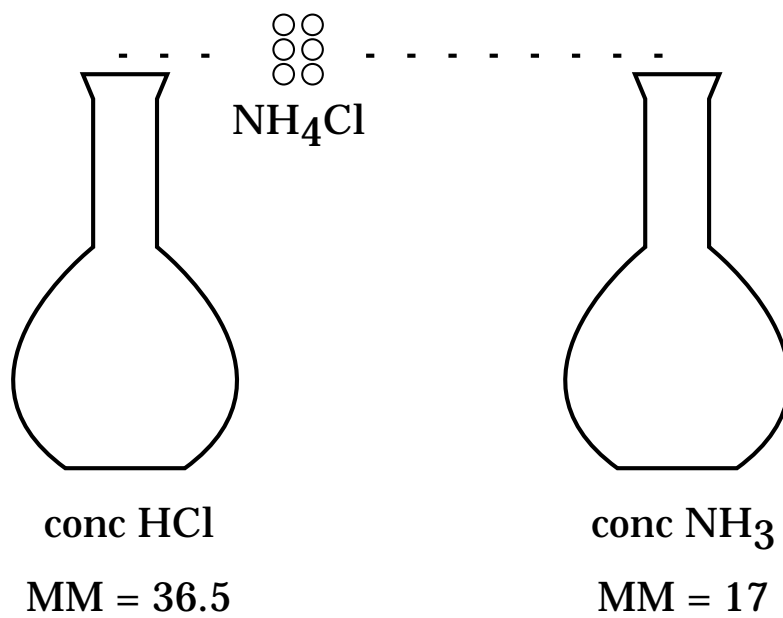
- **Transparency Masters**

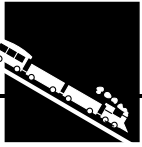
1. Formation of Gaseous Ammonium Chloride
2. Effect of Changing Pressure of a Gas on its Volume at Constant Temperature
3. Word Search

- **Humor**

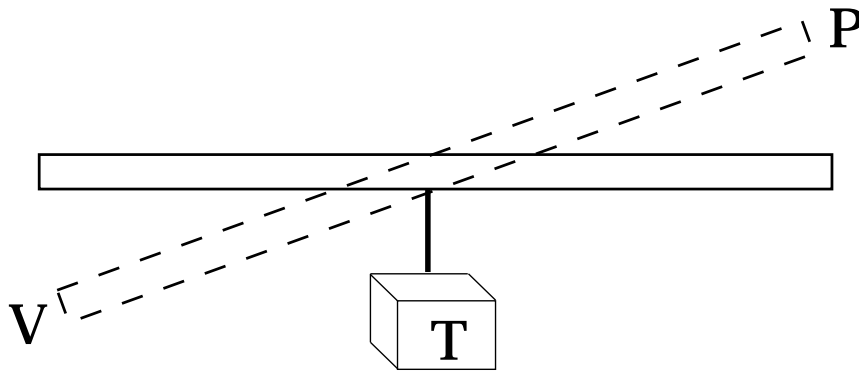
1. What is this gas?
2. Cartoons

Formation of Gaseous Ammonium Chloride





**Effect of Changing Pressure of a Gas on its
Volume at Constant Temperature**

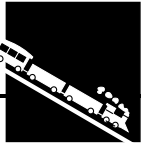


Word Search

A B E W L A E Z S A D M T T C B U
D A E M Y O V R E T E M O N A M W
P R E S S U R E Z B L U B Z H Z E
Z O E D A N I E U P S W N A Q T T
W M W S Z Q G H Z L B A A S D T L
P E L O M E N O Q E L Q G P A U I
X T I Q I P G W R P T J E N U Y W
E E X D O Q C Z I E M U Q L A M E
Y R K V L T Y V A G G I L O Y T K
K K N A C H A R L E S S O O B O M
G U E P G S W C N J U F T V S A B
W D P V B U P D K L N U L W I B T
I K H I C N C C Q Q O Q N V A E A

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Find these words in the block of letters:

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8. This quantity of gas at STP occupies 22.4 L. (2 words)
9. Gaseous mixture of nitrogen, oxygen, and trace gases found in the atmosphere.
10. Force per unit area exerted by a gas in a closed container.



What is this gas?

This gas is very toxic and an extreme fire hazard. It is fatal in concentrations of 0.000001 ppm. Organisms exposed to these gas concentrations die within a few minutes. Symptoms resemble those of cyanide poisoning (blue face, *etc.*). In higher concentration, *e.g.*, about 20%, the toxic effect is delayed, and it takes about 2.5 billion inhalations before death takes place. The reason for the delay is the different mechanism of the toxic effect of this gas in 20% concentration. It apparently contributes to a complex process called aging, of which very little is known, except that it is always fatal.

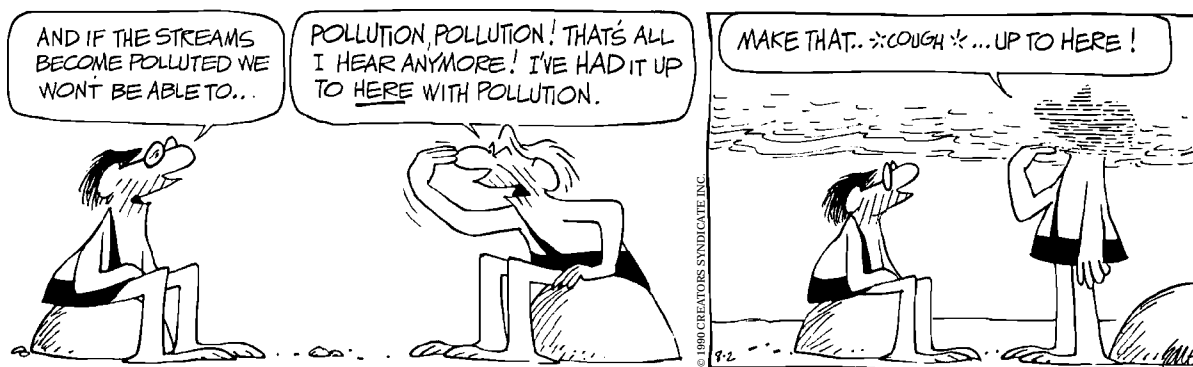
However, the main disadvantage of this 20% gas concentration is the fact that it is habit forming. The first inhalation (occurring at birth) is sufficient to make this gas addiction permanent. After that, any considerable decrease in the daily doses of this gas results in death.

Concentrations higher than 20% decrease the above-mentioned delay. High concentrations of this gas provokes a condition known as retrolental fibroplasia in prematurely born babies placed in incubators. This condition results in blindness. Lung irritation has been reported on experimental animals exposed to high concentrations of this gas for several days.

This gas is an extreme fire hazard. All the fires that were reported in continental U.S. for the period of the past 25 years were attributed to the presence of this gas in the atmosphere surrounding the fire areas.

This gas is especially dangerous because it is odorless, colorless, and tasteless, so its presence cannot be readily detected until it is too late.

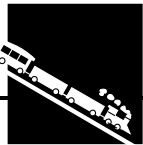
The gas described here is oxygen!



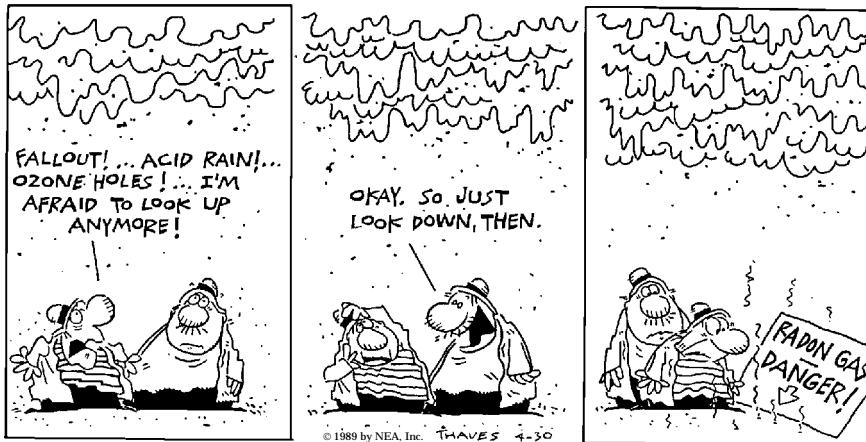
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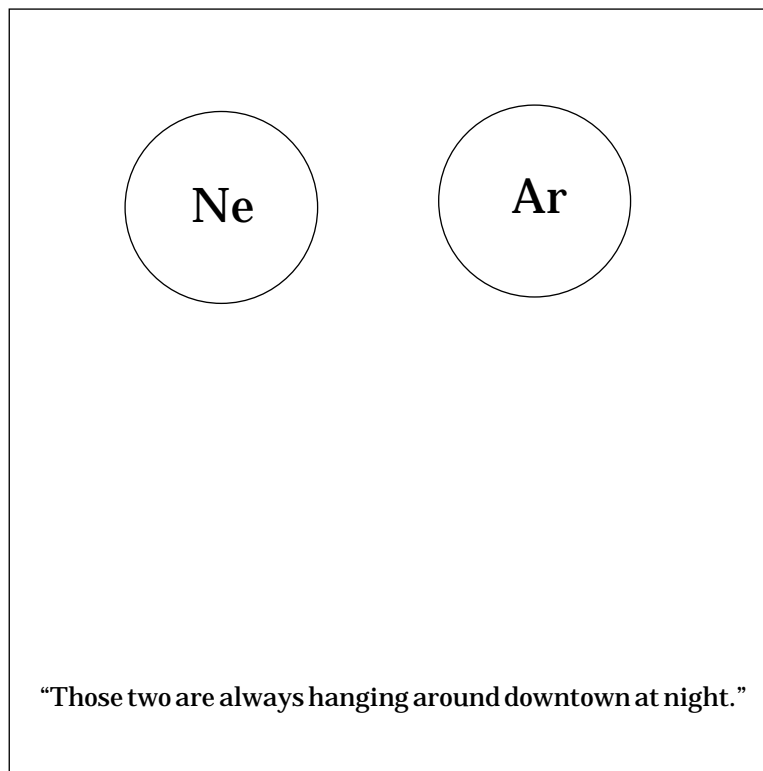
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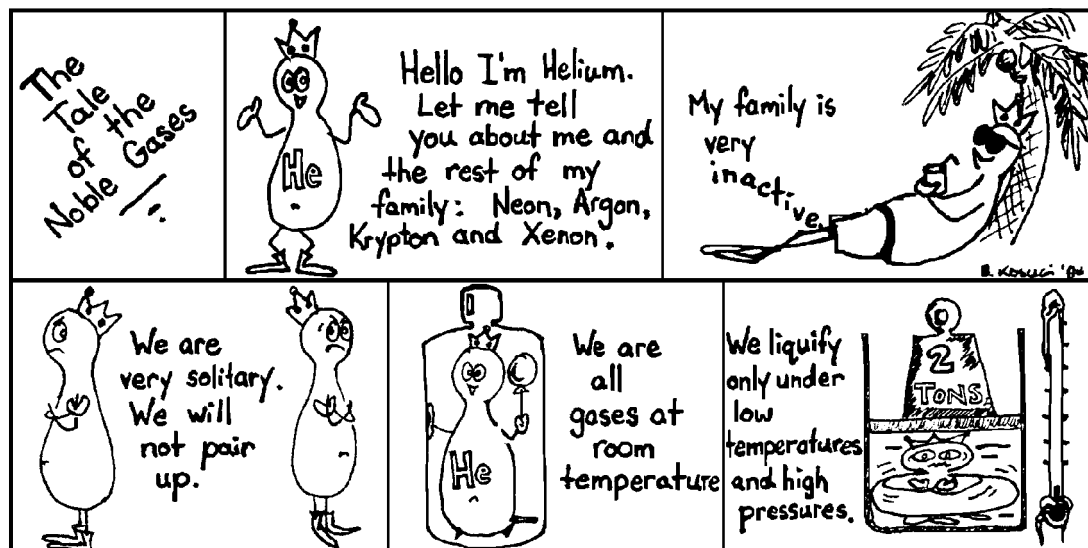
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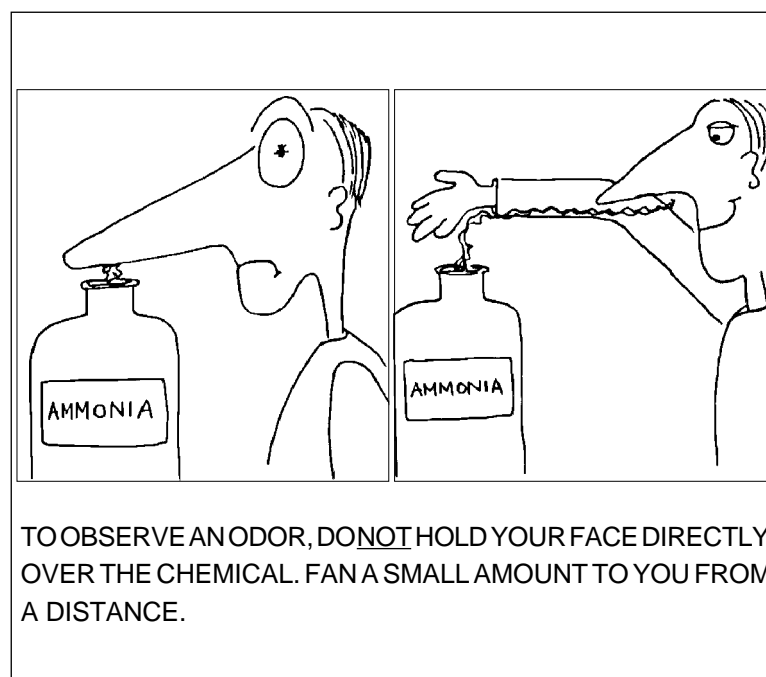
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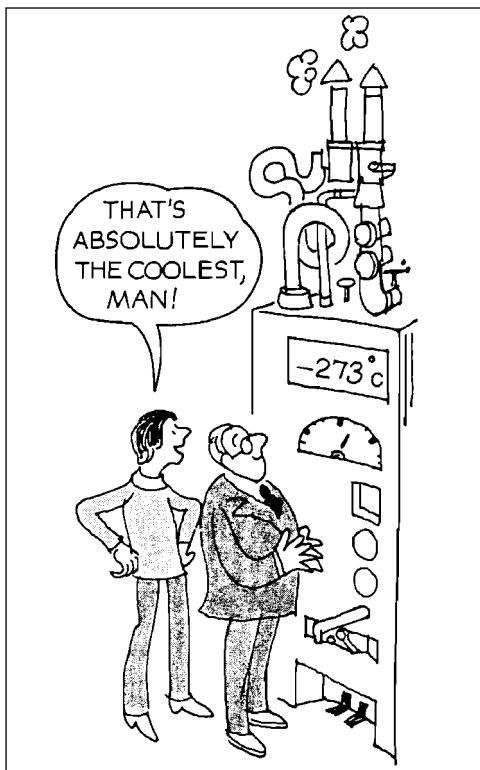
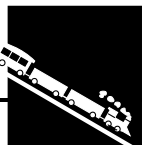
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WHAT'S THIS GAS?



NITROGEN

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**Percent water in kernel = $\frac{\text{mass of water}}{\text{total mass}} \times 100 = \frac{0.2}{2} \times 100 = \mathbf{10\%}$

***Pressure inside kernel at time of pop. Use $PV = nRT$; solve for P .

$$n = 0.2 \text{ g} \times \frac{1 \text{ mol}}{18 \text{ g}} = 0.011 \text{ mol H}_2\text{O}$$

$$P = \frac{nRT}{V}$$

$$P = 292 \text{ atm, or } 3 \times 10^2 \text{ atm}$$