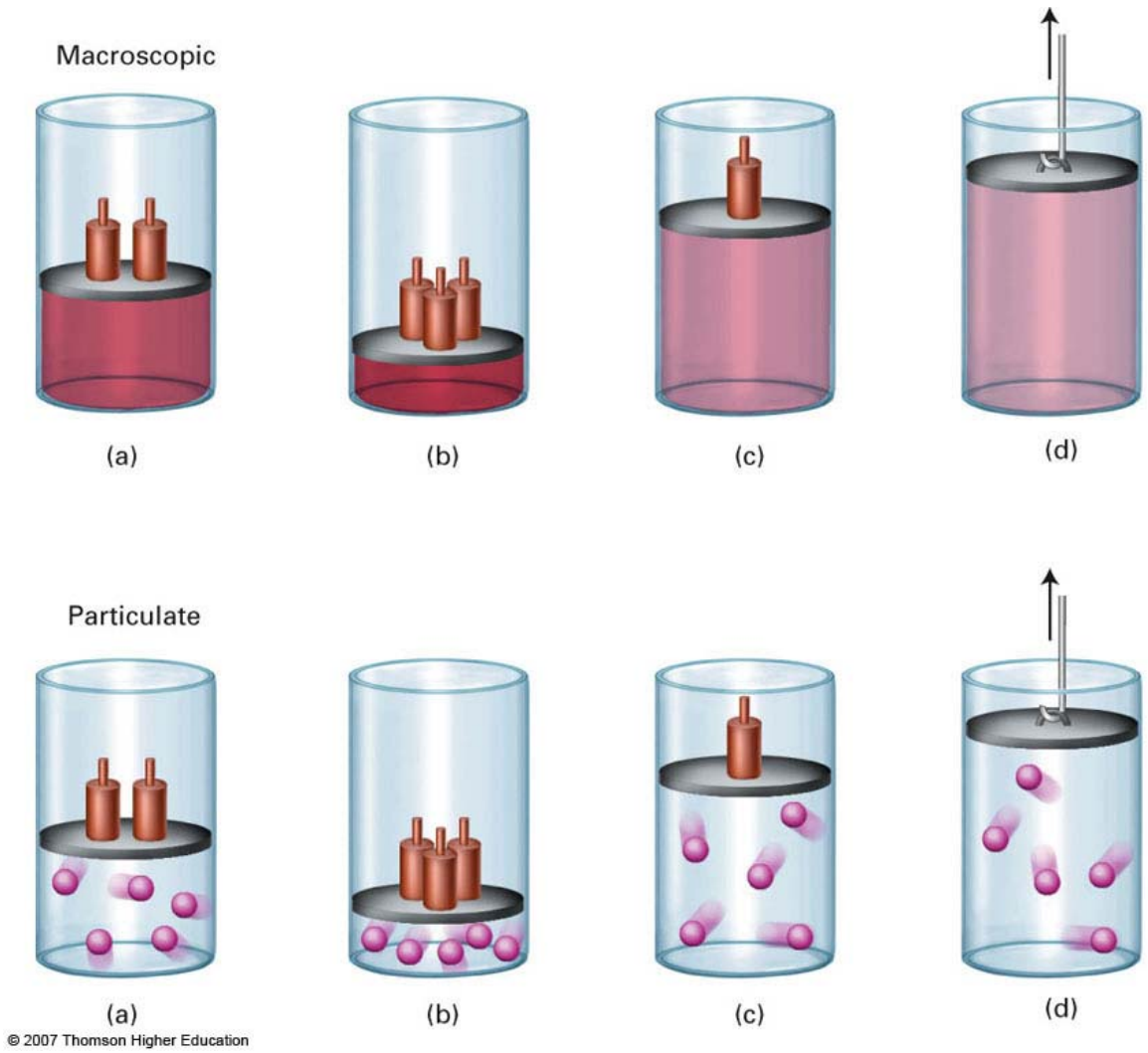
Chapter 6: Gases

Properties of Gases

1. No definitive shape or volume – expand to fill container
2. Easily compressed
3. Low density
4. Mix uniformly in any proportion
5. Exert pressure on container walls uniformly in all directions.
6. Diffuse relatively rapidly

Compression and expansion of gases:
The volume of the gas depends on the pressure exerted on the movable piston.



Comparing gases with liquids:

Density	
• $H_2O(l)$ = 1090 g/L	• Air = 1.3 g/L
Mixing	
• Not all liquids mix	• All gases mix
Dissolving	
• Solids dissolving in liquids have their limits	• Gases readily dissolve in each other
Pressure	
• Liquid pressure increases with depth in container	• Gas pressure is identical throughout closed container

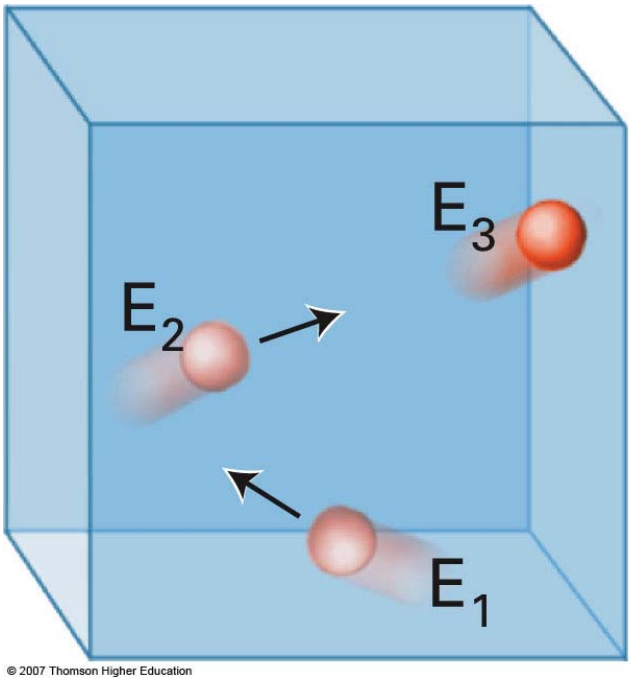
The Kinetic Molecular Theory of Gases and the Ideal Gas Model

The simple gas laws (to be discussed below) were developed from experimental observations. They are natural laws used to predict gas behavior. Specifically, they serve to explain observed macroscopic behavior in terms of particulate behavior (at the molecular level)

The Ideal Gas: Animation
To illustrate **the behavior of an ideal gas molecule,** click on the link below.
http://www.chm.davidson.edu/ChemistryApplets/KineticMolecularTheory/BasicConcepts.html (note: some aspects of the animation are currently non-functional)

1. A Single Molecule of a Gas
Consider the first box in the animation, containing a *single molecule*. Start the animation and observe the molecule, represented by the blue ball, bouncing and traveling back and forth across the box.
What do you observe about the behavior of the gas molecule? It:

- Travels at a constant speed
- Travels in a straight line
- Is in constant motion
- The collisions with the walls are perfectly elastic. Energy is neither gained nor lost from the collision. Because the walls do not move, the molecule's speed is unaffected by the collision.



$$E_1 = E_2 = E_3$$

The energy (E) of the gas molecule is the same at all times

This is the behavior of an ideal gas

You already know from experience that some of this behavior is true.
For example: A closed canister of propane gas retains its pressure indefinitely. A unopened can of soda pop or hair spray does not lose its gas pressure over time.

Pressure would decrease and temperature would drop if molecules lost energy when colliding with the container walls. This does not happen, so energy must not be lost.

2. Multiple Molecules of a Gas
Now consider the second animation box in which *several* molecules exist. Start the animation and observe the blue molecule. Unlike the system above, this system has multiple molecules and thus collisions between molecules occur.

Observe how the speed (and direction) of the molecule changes as a result of a collision. Here, the speed of a given molecule is not constant. From one collision to the next the molecule speeds up, slows down, speeds up, *etc.*

Now, a wide range of velocities can be seen. However, *the total energy of the gas molecules is constant. The average energy is constant.*

Note that the ideal gas model is again followed:

- Molecules travel in straight lines, and obey Newton's laws
- Collisions are perfectly elastic (no energy is lost in collisions)
- Molecules do not interact with each other (except for collisions)
- Molecules have no volume
- The gas molecules are far apart (on average).

The Kinetic Molecular Theory of Gases

One theory to explain the observed behavior of gases is **the kinetic molecular theory (KMT) of gases**. The main assumptions (postulates) of the ideal gas model are:

Postulates of the Kinetic Molecular Theory of Gases

1. Gases consist of tiny particles (atoms or molecules).
2. These particles are so small, compared with the distances between them, that the volume (size) of the individual particles can be assumed to be negligible (zero).
3. The particles are in constant random motion, colliding with the walls of the container. These collisions with the walls cause the pressure exerted by the gas.
4. The particles are assumed not to attract or to repel each other.
5. The average kinetic energy of the gas particles is directly proportional to the Kelvin temperature of the gas.

Real gases do not conform exactly to all these postulates, particularly #'s 2 and 4. However, they do explain ideal gas behavior – and they also come very close to the behavior shown by real gases at typical conditions. Real gases only deviate significantly from the KMT model at very low temperatures or very high pressures, i.e. where gases are close to liquefying.

The implications of the KMT theory to be discussed below are:

- Temperature = proportional to the average speed of gas molecules
- Pressure = proportional to temperature. This implies that increased pressure is due to increased number of collisions with the walls of the container.
- Volume = proportional to temperature. This implies that increased pressure causes increase in volume.

Note:
The speed of various gases can be calculated using:
"Speed" = $(3RT/M)^{1/2}$ where M is the molar mass of the gas molecule, T is the Kelvin temperature, and R is the ideal gas constant (discussed shortly).

- Hotter gases move faster

The average speeds at 0 °C are:

- O_2 ~ 300 m/s
- N_2 ~ 400 m/s
- H_2 ~ 1500 m/s

What are these speeds in miles / hour (mph)?

The average distance traveled between collisions for small molecules can be calculated to be around 10^{-7} m (100 nm).

Note:
Two different gas molecules at the same temperature will have the same energy; therefore at the same temperature the smaller molecule will be traveling faster.

Kinetic Energy = $\frac{1}{2} mv^2$

- m = mass
- v = velocity

Note also: There has been no mention of what type of gas, or any other characteristics of the gas. As we will see, the gas laws are largely independent of the type of gas molecule.

The Gas Laws

There are four of them. The variables are:

1. Amount, n (moles)
2. Volume, V
3. Temperature, T
4. Pressure, P

Pressure

Atmospheric Pressure:

We live at the bottom of an ocean of gas (recall that the air we breathe is actually a ***solution*** of numerous gases). Just like the ocean, or in a swimming pool, where the pressure increases as you swim down to the bottom, we are subjected to a large amount of pressure due to the weight of the air. Usually we are entirely unaware of the weight of the air.

Demonstration of air pressure

1. Lab demo: air pressure: Why aren't we crushed?
2. Atmospheric pressure > 14 lbs / inch^2. (compare with a 10 lb weight)
3. Inverted column of water demo. Why doesn't the water run out?
4. How does a straw work? Demo

   Note: there is no such thing as the pulling force of suction. Gases can only push on things; they can never "pull" on things. When we "suck" on a straw, the liquid rises due to the pressure or "pushing" forces of the atmosphere. Suction, then, is really the result of the net difference between two gas pressures.

What causes air pressure?

Force vs. Pressure:

- Force is the total impact of one object (e.g. a gas) on another object.
- Pressure (P) is the ratio of force (F) to the surface area (A) over which it is applied to.

$$\mathbf{P = F/A}$$

To begin with, then, perhaps air could simply be viewed as having weight. Gravity pulls on our atmosphere just as it pulls on all matter on earth. The amount of "pull" or weight depends on the amount of material in question.

Example: Place a 10 lb weight on a 1 $inch^2$ area, e.g. the tips (first joints) of two fingers. The pressure that is exerted is a function of the weight per unit area. What would be the force exerted by a column of air that is 120 miles high and applied to a 1 $inch^2$ in area? The answer is: more than the 10 lb weight. If the air could be viewed as a column of air, the weight of a column of air might be viewed as 14.7 lb / $inch^2$.

However, this is an oversimplified analogy because remember: unlike solids and liquids, a gas is a substance which has neither a fixed volume nor a fixed shape. A gas will expand to fill up its container. Hence, pressure is not simply gravity pulling on a mass of gas molecules.

Question: When we walk into a room, why don't we find all the air molecules in a layer on the ground, or maybe bunched up to one side of the room?

The energy of any attractive forces that the molecules might have for each other, causing them to aggregate and fall down, or the simple force of gravity, are vastly exceeded by the kinetic energy of the gas molecules. As shown below, gas molecules are in rapid motion, constantly colliding with each other, changing energy, velocity, and direction with each collision.

Gas pressure is a direct result of the **collision of gas molecules with the walls of its container.**

Gas Pressure: An Overview

Returning to our ideal gas model, what happens to the motion and speed of the blue ball when we add additional [red] gas molecules to the container?

http://www.chm.davidson.edu/ChemistryApplets/KineticMolecularTheory/BasicConcepts.html

The gas molecules quickly develop a wide range of speeds, therefore a wide range of kinetic energy displayed between the molecules.

The following Java-based model also explores the effect of temperature, pressure, and concentration on the motion of gas molecules.

Pedagogica Activities.jnlp

What is the effect of varying the temperature, volume, and number of gas molecules on the pressure of the gas?

Summary:
A gas is composed of a large number of particles in very rapid motion.

"Pressure" exerted by a gas is a result of collisions of gas molecules with the walls of its container.

- **The amount of pressure is directly related to the frequency with which the molecules collide with the walls of the container.**
- **More *frequent* collisions mean higher gas pressures:**

The frequency of collisions within a container is dependent on:

- **The speed of the gas molecules.** More *forceful* collisions also mean higher gas pressure.
- **Temperature.** Increasing the temperature increases the speed of the molecules and therefore the kinetic energy or force they have when they hit the wall of the container
- **The number (or concentration) of the gas molecules.**

The speed or energy of the molecules is also related to the *mass* of the gas molecule.

The following animation shows the motion of two different gas molecules (He and Kr) that have a large difference in mass (4 vs 84).

Note that initially, all are moving at the same speed. But, that quickly changes.

http://www.chm.davidson.edu/ChemistryApplets/KineticMolecularTheory/Pressure.html

In order for smaller particles to have the same amount of kinetic energy as larger particles, they must move faster.

Measuring Atmospheric Pressure: Torricelli's Barometer (Torricelli, 1608 – 1647)

We can fill a glass tube with water, invert it into a beaker of water, and the water does not run out of the tube. The pressure on the surface of the water in the beaker exceeds the force of gravity exerted on the column of water (the weight of the water). Indeed, the tube can be as high as 34 feet and still be completely full of water. In other words, the pressure of the atmosphere approximately equals the force exerted by the weight of 34 feet of water. If the initial glass tube is longer than 34 feet, for example 50 feet, and the tube is completely filled with water and inverted into the beaker of water, the level of water in the tube will drop from 50 feet to 34 feet.

Question:
If we did this experiment in Denver, would you observe something different?

***Problem*:**
If, in place of water, we substituted mercury, what is the height of the mercury that the pressure of the atmosphere would support?
Note:

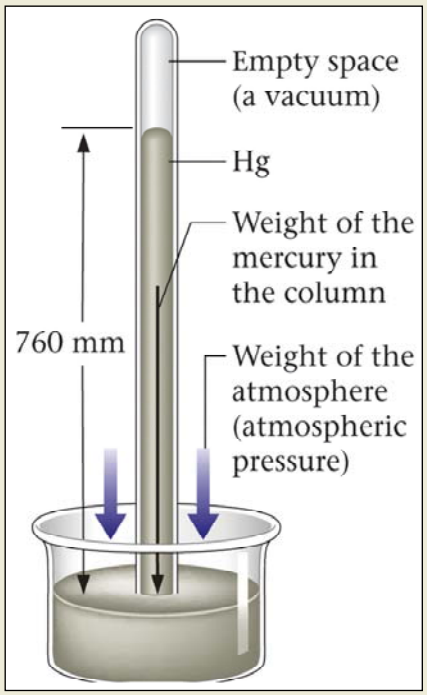
- Density of H_20 = 1.0 g / cm^3
- Density of Hg = 13.6 g / cm^3

What is the height of the column of Hg measured in:

- Cm?
- Mm?
- Inches?

Air pressure can hold up a column of Hg up to 760 mm at sea level. If the initial column of Hg is longer than 760 mm before it is inverted into the container of mercury, the mercury will drop until it reaches a height of 760 mm (at sea level). Above the Hg will be empty space, a vacuum.

Figure 12.2: When a glass tube is filled with mercury and inverted in a dish of mercury at sea level, the mercury flows out of the tube until a column approximately 760 mm high remains.



If the height of the column of Hg is dependent on atmospheric pressure, what will happen to the height of the column if:

- Atmospheric pressure decreases?
- Atmospheric pressure increases?

What situations can you think of where measuring atmospheric pressure might be important?

Note: There are numerous units of measure applied to atmospheric pressure, and they must all be learned because all are used. At sea level, units of atmospheric pressure are:

- 760 mm Hg
- 760 Torr (Torricelli)
- 14.69 psi (pounds / square inch), the English measure of pressure
- 101,325 Pascals (Pa), the SI unit for pressure

- 1 atm (standard atmosphere)

Problem: The height of mercury in a barometer was measured as 732 mm. What is this pressure in:

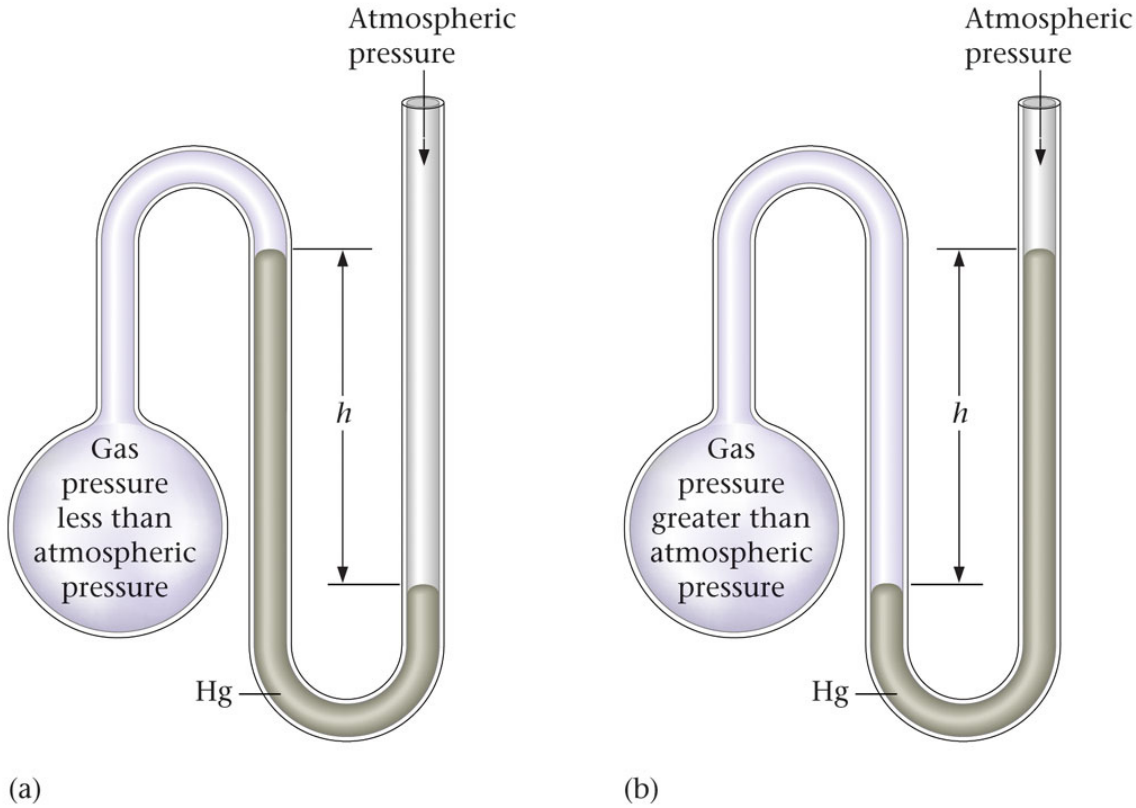
- Atm?
- Torr?
- Pa?

Problem: The pressure of a gas is measured as 2.79×10^5 Pa. What is this pressure in:

- Atm?
- Torr?
- Psi?

Measuring the Pressure of a Gas: The Manometer

Pressure of gas in a container can be measured with a device called a manometer.



The difference in the height of the two columns is directly related to the gas pressure in the container.

a) Gas pressure is atmospheric pressure (mm Hg) – h (mm Hg)
b) Gas pressure is atmospheric pressure (mm Hg) + h (mm Hg)

http://www.chm.davidson.edu/ChemistryApplets/GasLaws/Pressure.html
http://www.chm.davidson.edu/ChemistryApplets/KineticMolecularTheory/Pressure.html

4.5 Boyle's Law: Pressure and Volume

Demo: squeeze a plastic bottle with and without the cap on.

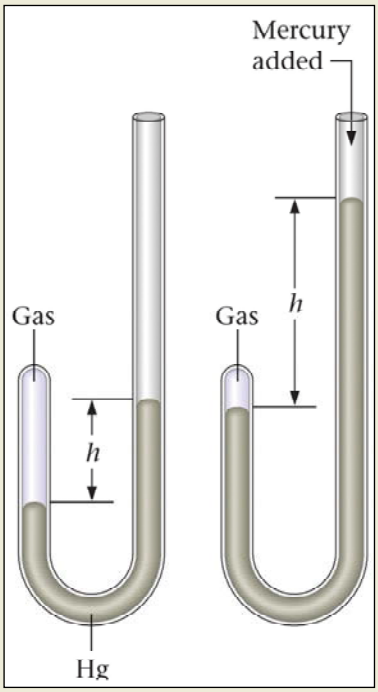
Observation?

Robert Boyle (1627 – 1691) was the first to systematically study gases. He used Hg in a glass tube shaped like a "J" to study the effect of the pressure

exerted by a column of Hg on the volume of a trapped quantity of air above the Hg.

The "J" tube apparatus is shown below:

Figure 12.4: A J-tube similar to the one used by Boyle.



He varied the amount of Hg (the height of Hg in the open end of the J tube) and recorded the effect on the volume of the gas in the closed end of the tube. He found the following data:

Table 12.1 A Sample of Boyle's Observations (moles of gas and temperature both constant)

Experiment	Pressure (in Hg)	Volume (in.^{3})	*Pressure × Volume (in Hg) × (in.^{3})* Actual	Rounded*
1	29.1	48.0	1396.8	1.40×10^3
2	35.3	40.0	1412.0	1.41×10^3
3	44.2	32.0	1414.4	1.41×10^3
4	58.2	24.0	1396.8	1.40×10^3
5	70.7	20.0	1414.0	1.41×10^3
6	87.2	16.0	1395.2	1.40×10^3
7	117.5	12.0	1410.0	1.41×10^3

*Three significant figures are allowed in the product because both of the numbers that are multiplied together have three significant figures.

What is interesting about this data? What relationship can you discern?

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Pressure $\propto$ **1/Volume**
Pressure = k/Volume where k is a proportionality constant
P = k/V

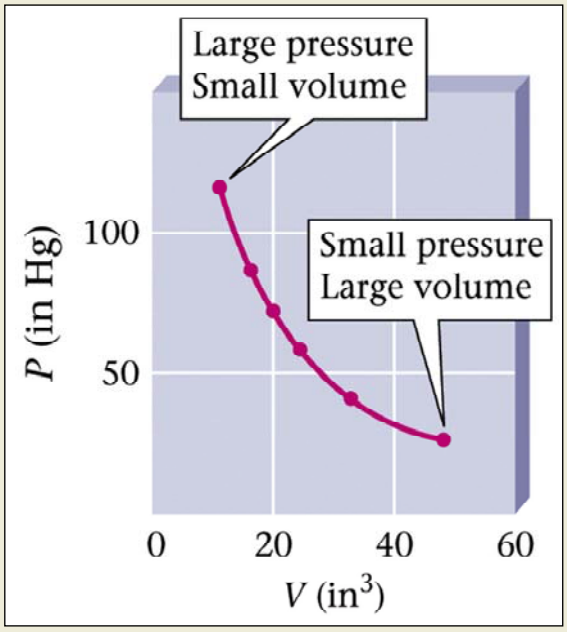
Inspection of the data reveals that the *mathematical product* obtained by multiplying the pressure P (in inches of Hg) times the volume of the gas V (inches^3) is always the same number, i.e. **the product of P x V is a constant.** That is, for every incremental increase in P, there is an equal and opposite decrease in V, so the product of multiplying P x V is a number which is the same for each value of P and V.

P x V = k (a constant)

This relationship is know as **Boyle's Law**

A plot of this pressure and volume data shows a near linear (inverse) relationship between P and V:

Figure 12.5: A plot of *P* versus *V* from Boyle's data in Table 12.1.



Volume is inversely related to pressure.

What this means is:

- if we change the pressure, we can predict the new volume
- if we change the volume, we can predict the new pressure

For example:
The initial pressure and volume are P_1 and V_1
The pressure is changes to a new pressure P_2, the new volume will be V_2

Since $P_1 \times V_1 = k$ (a constant)
And $P_2 \times V_2 = k$ (a constant)

Then $P_1 \times V_1 = P_2 \times V_2$ (Boyle's law rewritten)

One way to visualize this is to consider the piston in an automobile engine. We understand that as we push on the piston to compress the gas above it,

we have to exert a larger and larger force, or pressure, as the volume above the piston grows smaller and smaller, e.g.

http://www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/flashfiles/gaslaw/boyles_law.html

Start the simulation shown below. The pressure-volume data is plotted on the graphs at the lower right.

http://www.chm.davidson.edu/ChemistryApplets/KineticMolecularTheory/PV.html

Slowly drag the top barrier to change the volume of the system. The new pressure-volume data will be automatically plotted.

Questions:
1. Does the system obey Boyle's Law?
2. What property of the gas causes the change in pressure when the volume is changed?
Bear in mind that the temperature of the system is held constant. Consequently the average kinetic energy of the particles does not change when the volume changes. The mass of each particle is also held constant. Thus the average speed of the particles and the average force exerted on a wall during a collision are independent of the volume. Carefully examine the behavior of the particles at a small volume and at a large volume. What properties of the gas do change with a change in volume?

http://www.chm.davidson.edu/ChemistryApplets/GasLaws/BoylesLaw.html

Note: *Two assumptions about the experimental conditions must be made for this relationship to be true. What are they? (What other factors can affect pressure or volume?)*

Answer: Temperature must remain constant and the number of moles of gas molecules must remain constant. If the temperature changes, then the frequency and force of impacts changes, thus changing the pressure. If the number of molecules changes, then the frequency of impacts changes, thus changing the pressure.

Problem: Calculating a *new volume* using Boyle's law:

1.5 L of the refrigerant, Freon (CCl_2F_2), at a pressure of 56 torr and a specified temperature has the pressure changed to 150 torr at the same temperature. What will be the new volume of this gas?

A 23.1 L container of He gas has a pressure of 3.54 atm. The gas is transferred to a new container and the pressure in this container is 1.87 atm. What is the volume of the new container?

Examples: Calculating a *new pressure* using Boyle's law:

A gasoline engine cylinder has an initial cylinder volume of 0.725 L. When the piston compresses the gas, which is initially at 1.0 atm pressure, to its final volume of 0.075 L prior to ignition, what is the final pressure in the cylinder? (assume constant temperature)

A 45.1 L container of Ne has a pressure of 7.43 atm. The gas is transferred to a new container which has a volume of 18.4 L. What is the pressure of the Ne in this new container? (assume constant temperature)

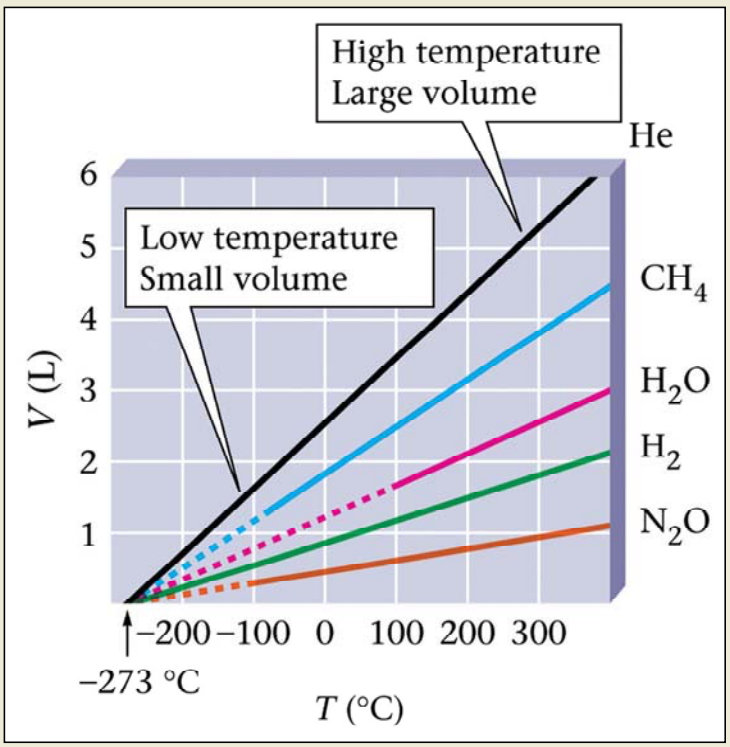
4.4 Volume and Temperature: Charles' Law

Jacques Charles (1746 – 1823) was the first person to make a solo balloon flight using a balloon filled with H_2. He was a French mathematician who studied the relation of volume of a gas to its temperature. This relationship between T and V can be demonstrated: Click on the link below: http://www.chm.davidson.edu/ChemistryApplets/KineticMolecularTheory/PT.html

Start the simulation shown. Slowly drag the liquid in the thermometer to change the temperature of the system. (This is admittedly a rather unlikely way to change the temperature of the system.) The new pressure-temperature data will be automatically plotted on the graphs at the lower right.

A plot of the relationship of temperature (T) to volume (V) for several gases is shown in Fig. 12.7 below.

Figure 12.7: Plots of V (L) versus *T* (°C) for several gases.

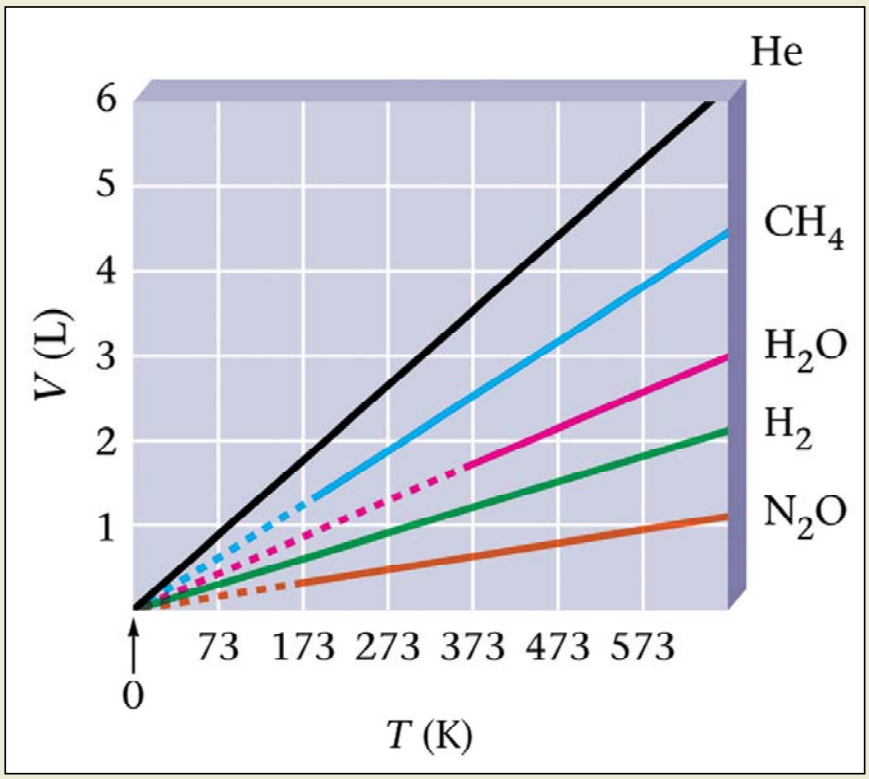


Two things to note:

1. There is a linear relationship between V and T for **all** gases

2. Charles discovered that for each Celsius degree that the temperature of the gas was lowered, the volume of the gas will decrease by 1/273rd of the volume that the gas occupies at O °C. This implies that if a gas was cooled to – 273 °C, it would have a volume of zero. When V versus T values for several gases are plotted, the plot lines seem to converge as the volume is reduced. Although all gases will eventually turn to liquid at some point, when the gas volume is *extrapolated* to zero volume, the plot lines converge at a single temperature, -273 °C (actually, -273.15 °C). Because you cannot have a volume of less than zero, the implication is that -273.15 °C is the lowest possible temperature. This temperature is now known as **absolute zero**.

Figure 12.8: Plots of *V* versus *T* as in Figure 12.7, except that here the Kelvin scale is used for temperature.



http://www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/flashfiles/gaslaw/charles_law.html

This direct proportionality between volume of a gas, and temperature (in Kelvin) is known as Charles's law: $\mathbf{V} \propto \mathbf{T}$

$V = bT$ (Charles law)

b = a proportionality constant
T = temperature in Kelvin (!)

Volume is directly related to temperature.

Note that this can be rewritten as: **$V/T = b$**

What this means is:

- if we change the temperature, we can predict the new volume
- if we change the volume, we can predict the new temperature

For example:
The initial pressure and temperature are P_1 and T_1
The pressure is changes to a new pressure P_2, the new temperature will be T_2

Since $V_1 / T_1 = b$ (a constant)
And $V_2 / T_2 = b$ (a constant)

Then: $V_1/T_1 = V_2/T_2$ (Charles's law rewritten)

Calculating Volume Using Charles's Law

A 2.0 L sample of air at 298 K is cooled to 278 K. Pressure is kept constant at 1 atm. What is the new volume?

$V_1/T_1 = V_2/T_2$

$2.0\ L / 298\ K = V_2 / 278\ K$

Answer: $V_2 = 1.9\ L$

Problem: A sample of H_2 gas is collected at 34 °C and heated to 68 °C where the volume is determined to be 26.0 L. Assuming constant pressure, what was the initial volume at 34 °C?

14.2 Avogadro's Law: Moles and Volume

When the number of moles of a gas is doubled (at constant temperature and pressure), the volume of the gas doubles. Thus, there is a direct relationship.

$\mathbf{V \propto n}$
V = a • n (Avogadro's Law)

Where:
n = number of moles
a = a proportionality constant

This can be rewritten as: $V/n = a$ (a constant)

Since $V_1 / n_1 = a$
And $V_2 / n_2 = a$

Then $\mathbf{V_1/n_1 = V_2/n_2}$ **(Avogadro's Law rewritten)**

Example calculation:

A 12.2 L volume contains 0.50 mol of O_2 at 1 atm pressure and 25 °C. All the O_2 is converted to O_3, also at 1 atm pressure and 25 °C. What is the new volume of gas?

$O_2 \rightarrow O_3$

The balanced equation is:
$3O_2 \rightarrow 2O_3$

Therefore,
0.50 mol O_2 ($2O_3$ / $3O_2$) = 0.33 mol O_3

From Avogadro's Law: $V_1/n_1 = V_2/n_2$

12.2 L / 0.50 mol O_2 = V_2 / 0.333 mol O_3

V_2 = 8.2 L

Problem: If 2.55 mol of helium gas occupies a volume of 59.5 L at a given temperature and pressure, what volume does 7.83 mol of helium gas occupy at the same temperature and pressure?

Molar Volume of a Gas
It was determined that when n = 6.02214×10^{23} molecules of gas (a mole of molecules), at standard temperature and pressure (**STP**: 0 ˚C and 1 atmosphere pressure), the volume that that gas will occupy is 22.4 L. As point of reference, 22.4 L is approximately the volume of three basketballs. *This relationship was found to be the case for any gas.*

In other words,

- **1 mole of any gas (at STP) = 22.4 L of gas**
- **This is the molar volume of any gas (at STP).**

Note: The molar volume of a gas, 22.4 L, is independent of the specific gas. Because we know that a mole is simply a "count" of something, then the volume occupied by a gas must be dependent only on the ***number of molecules.*** This is indeed expressed by Avogadro's Law: **V = an** .

Interestingly, this independence from the specific type or identity of the gas molecule suggests two things:

1. That the volume of a gas molecule must be so small as to be unimportant
2. That the forces among the gas molecules must not be very important.

Example 1:
What is the mass of 1.00 L of C_3H_6 (cyclopropane gas, an anesthetic) measured at STP?

Solution: use dimensional analysis
1.00 L of gas x (1mole of any gas / 22.4 L) x (42.08 g of C_3H_6 / mole of C_3H_6) = 1.88g of C_3H_6

Example 2:
Your backyard gas barbecue releases 30.0 L of C_3H_8 (propane) at STP. What ***mass*** of propane was released?

Solution:
30.0 L x

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Example 3:
A 128 g piece of dry ice [CO_2(s)] evaporates into CO_2 gas at STP. How many liters of CO_2 (g) are formed at STP?

Solution: use dimensional analysis
128 g CO_2 x (1 mol CO_2 / 44.02 g CO_2) x 22.4 L CO_2 / 1 mole C
O_2

14.3 The Ideal Gas Law and the Combined Gas Law

The effects of individual changes in the three different variables, on the volume of a gas, are described by the three laws:

Boyle's Law: **V = k / P** (volume is inversely proportional to P)

Charles's Law: **V = b T** (volume is directly proportional to T)

Avogadro's Law: **V = a n** (volume is directly proportional to n)

If volume is proportional to three quantities, it is logical that volume is proportional to the *product* of these three quantities. Thus, these three simple gas laws can be combined into a *single equation* which can be applied to all gas law problems.

V = RnT / P (where R is the *combined* proportionality constants)

This is usually rewritten as:

The ideal gas law: PV = nRT

R is known as the universal gas constant

The Universal Gas Constant
Its value will depend on the units chosen for the calculation.

Problem:
If: n = 1 mole
V = 22.4 L (molar volume of any gas at STP)
P = 1 atm
T = 273.15K (0 °C)

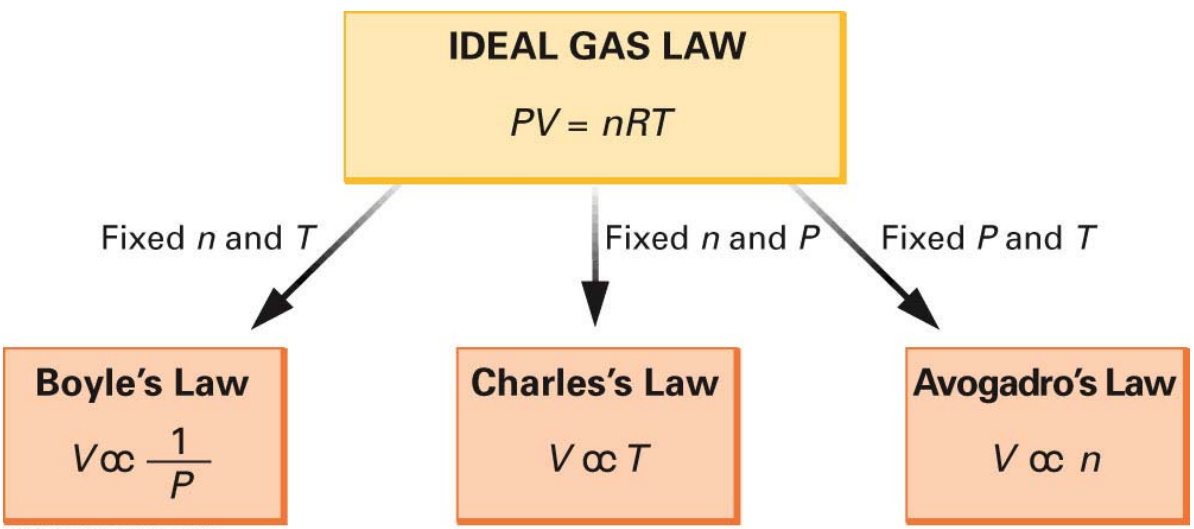
Solve for the universal gas constant, R, by substituting into the ideal gas law.

R = 0.08206 L • atm / mole • K

The ideal gas law and the universal gas constant must be memorized.
Note: If different units are chosen (e.g. SI units: cubic meters - m^3 – for volume; pascal – Pa – for pressure), a gas constant with a different number and units will be obtained.

Note: Ideal gas behavior should be assumed for problems presented here, and that means pressures near 1 atm. and temperatures near 0 °C.

A wide range of gas problems can be solved by rearranging the ideal gas law to solve for the unknown.



Problem:
Consider: What is the volume occupied 1.0 mol of He at 1 atm pressure and 0°C?

Example 1: 8.56 L of hydrogen gas at 0 °C and 1.5 atm pressure contain how many moles (n) of gas?

Rearranging to solve for n:
n = PV / RT

Substitute the known's and solve for the unknown.

n = (1.5 atm x 8.56 L) / [(0.08206 L • atm / mol • K) x 273.15 K]
n = 0.57 mol H_2

Example 2: 0.250 mol of carbon dioxide gas at 25 °C and 371 torr pressure will occupy what volume?

Note: Pressure in Torr must first be converted to the same pressure units as R (atm).

1.00 atm / 760 torr = P / 371 torr

Rearranging: P = 1.00 atm x (371 torr / 760 torr)
P = 0.488 atm.

Rearranging the ideal gas equation:

$V = nRT / P$

Solve for V.

$V = 0.250 \text{ mol} \times [(0.08206 \text{ L} \cdot \text{atm} / \text{mol} \cdot \text{K}) \times 298 \text{ K}] / 0.488 \text{ atm}$

$V = 12.5 \text{ L}$

Problem: A sample of neon gas has a volume of 3.45 L at 25 °C and a pressure of 565 torr. Calculate the number of moles of Ne present.

Answer: 0.105 mol

Problem: A 0.250 mol sample of argon gas has a volume of 9.00 L at a pressure of 875 mm Hg. What is the temperature of the gas (in °C)?

Answer: 232 °C

Using the Ideal Gas Law Under *Changing* Conditions:

1. A Change in Two Variables

Example: A 0.240 mol sample of ammonia gas occupies a volume of 3.5 L at 25 °C. If this gas is compressed to1.35 L at 25 °C, what is the new pressure?

Procedure for Gas Law Problems:

- Identify which of the four measures are variable and which are constant.
- Rearrange the ideal gas equation to put all variables on one side, and all constants on the other side.
- Combine the two equations
- Rearrange the equation to solve for one unknown variable

In this example, the two variables are P and V; n and T remain constant, and R is always a constant. Therefore,

Starting conditions: $P_1V_1 = nRT$

Final conditions: $P_2V_2 = nRT$

These two equations can be combined:

$$P_1V_1 = P_2V_2$$

Rearrange to solve for P_2: $P_2 = P_1V_1 / V_2$

$$P_2 = (1.68 \text{ atm})(3.5 \text{ L} / 1.35 \text{ L})$$
$$P_2 = 4.4 \text{ atm}$$

Problem: 5.60 L of He gas at 2.45 atm pressure and 23 °C is cooled to 15 °C and a pressure of 8.75 atm. What is the new volume?

Answer: 1.53 L

Problem: 3.80 L of He at 28 °C and 3.15 atm pressure is heated to 43°C and a new volume of 9.50 L. What is the new pressure?

Answer: 1.32 atm

2. A Change in Three variables: the Combined Gas Law

If pressure, temperature, and volume all change for a fixed amount of a gas, the ideal gas equation can be rewritten:

$P_1V_1/T_1 = nR$

$P_2V_2/T_2 = nR$

These can be combined into:

The Combined Gas Equation: $\mathbf{P_1V_1/T_1 = P_2V_2/T_2}$

If three variables change and two of the changes are known, the equation can readily be rearranged to solve to the unknown variable.

Example: 3.48 L of a gas with a pressure of 0.454 atm at -15 °C is heated to 36 °C and a pressure of 0.616 atm. What is the new volume?

$P_1V_1/T_1 = nR$

Because n = constant, and R = constant, we can use the combined gas equation:

$P_1V_1/T_1 = P_2V_2/T_2$

Rearranging to solve for V2,

$V_2 = P_1V_1T_2 / T_1 P_2$

Converting temperatures to Kelvin and substituting,

$V_2 = 3.07$ L

Problem: 3.65 L of hydrogen gas at 63 °C and 4.55 atm pressure is cooled to – 35 °C and 2.75 atm. What must the new volume be?

Answer: 4.28 L

Problem: 9.55 L of oxygen gas at 27 °C and 2.97 atm pressure is heated to 125 °C and a new pressure of 8.25 atm. What is the new volume?

Answer: 4.56 L

Stoichiometry of Gases

The ideal gas equation allows us to calculate the number of moles of a gas, making stoichiometric gas equation calculations possible.

Example:

$2\ KClO_3 \rightarrow 2KCl + 3O_2$

What volume of O_2 gas is produced at 1.00 atm and 25 °C by the decomposition of 10.5 g of $KClO_3$?

Strategy for solving:

1. Calculate the theoretical yield (number of moles) of O_2 gas that could be made.
2. Use the ideal gas equation to convert moles to volume (L).

Answer: 3.13 L

Problem:

3.45 of Ar at STP contains how many moles? How much does it weigh (mass of Ar)?

Answer: 6.15 g.

Problem:

The decomposition of 152 g of $CaCO_3$ will produce what volume of CO_2 at STP?

$CaCO_3 \rightarrow CaO + CO_2$

Answer: 34.1 L

Avogadro's Law: Mole – Volume Relationships

The automobile air bag saves thousands of lives every year. The inflation of the air bag is based on the following chemical reaction:

$Na_3N(s) + Fe_2O_3(s) \rightarrow Na_2O(s) + Fe(s) + \mathbf{N_2(g)}$

Two solids, sodium nitride, and iron (III) oxide, react to form solid sodium oxide, metallic iron, and nitrogen ***gas***.

How would you classify this reaction? _____

How would you balance this equation?

The balanced equation for this reaction is:

$6Na_3N(s) + Fe_2O_3(s) \rightarrow 3Na_2O(s) + 2Fe(s) + 9N_2(g)$

- You start with two solids whose volume is essentially negligible. So, the airbag is empty at the start
- You end up with two more solids, but also with a gas, nitrogen (N_2).

Question:
If you need 22.4 L of $N_2(g)$ to fully inflate the air bag, how many grams of $Fe_2O_3(s)$ and $NaN_3(s)$ will you need? The solution is simple when you know Avogadro's Law.

Dalton's Law of Partial Pressures

Note that every gas problem we have worked with so far has not required any knowledge of what gas we were working with. Indeed, the simple gas laws were all based on the behavior of air, a mixture of gaes. That's because

gas pressure is a direct result of the **number or frequency of collisions** of gas molecules with the walls of the container they are in. At a given temperature, the kinetic energy of the gas molecules is independent of the mass of the gas molecule, so pressure is only dependent on the force of the molecule collisions per unit area.

According to the Ideal Gas model, molecules in the gas phase are separated by vast spaces, and they are in rapid motion. Because of this, two properties of gases are:

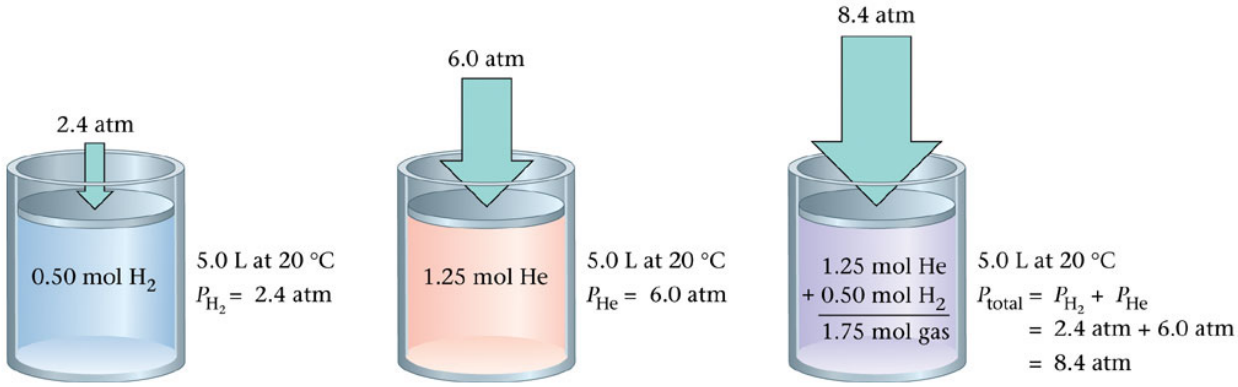
- Gas molecules rapidly fill the entire volume of the container that they are introduced into
- When a second gas is added, its molecules are distributed throughout the vast open space between the gas molecules to form a uniform mixture of gases

Dalton's observations of the behavior of mixtures of gases are summed up in….

Dalton's Law of Partial Pressures:
The total pressure (P) exerted by a mixture of gases is the sum of the partial pressures (p) exerted by each of the gases in the mixture. The partial pressure of one gas in a mixture is the pressure it would exert if it alone occupied the same volume at the same temperature.

$$\mathbf{P = p_1 + p_2 + p_3 + \ldots}$$

In the example shown in the figure below, the individual pressures of known amounts of two gasses are measured, and compared to the same amounts of these two gases mixed together in the same volume and at the same temperature.



Using PV = nRT, at constant volume and temperature, the pressures of the individual gases are calculated as shown. When the gases are combined in the same volume and at the same temperature, the new calculated value exactly equals the sum of the individual gas calculations.

Example: For SCUBA diving, 12 L O_2 at 1.0 atm and 25°C , and 46 L He at 1.0 atm and 25 °C are pumped into a 5.0 L tank. What is the partial pressure exerted by each gas, and what is the total pressure inside the tank at 25°C?

The strategy for solving the problem is to

1. First determine the number of moles of each gas using:

n = PV / RT

2. Then, determine the pressure that would be exerted by each gas when compressed into the smaller volume of the gas cylinder, using:

P = nRT/V

Problem:
A 2.0L flask contains a mixture of nitrogen gas and oxygen gas at 25 °C. The pressure in the flask is 0.91 atm. The amount of nitrogen gas in the

mixture is known: 0.050 mol. Calculate the pressure due to the oxygen (partial pressure of oxygen), and the number of moles of oxygen.

Answers: 0.024 mol oxygen; 0.29 atm

The Law of Combining Volumes