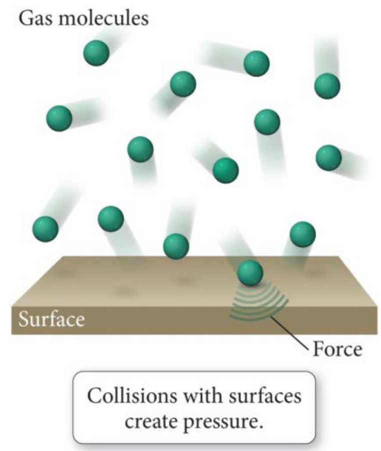
AP* Chemistry
GASES

The gaseous state of matter is the simplest and best-understood state of matter. You inhale approximately 8,500 L of air each day. This amounts to about 25 lbs of air. Breathing is a three-step process: inhaling, gas exchange with the circulatory system, and exhaling. Approximately 80% pressure differences created by your body allow you to breathe. Nearly the Earth's entire atmosphere is made up of only five gases: nitrogen, oxygen, water vapor, argon, and carbon dioxide.

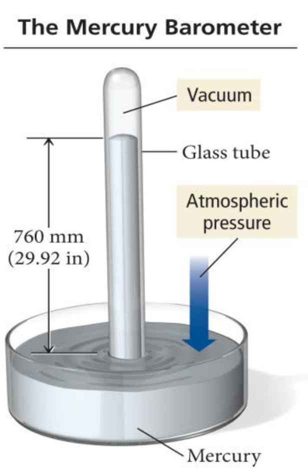


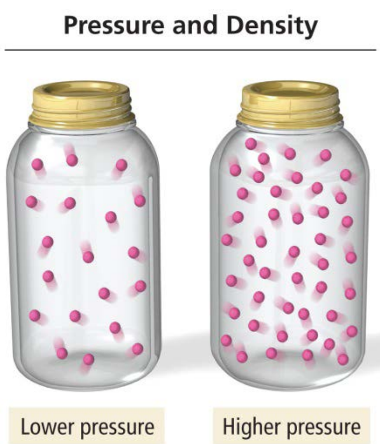
PROPERTIES OF GASES

Only FOUR quantities are needed to define the **state of a gas**:

1. the quantity of the gas, n (in moles)
2. the *temperature* of the gas, *T* (**in KELVINS)**
3. the *volume* of the gas, *V* (in liters)
4. the *pressure* of the gas, *P* (in atmospheres)

A gas uniformly fills any container, is easily compressed & mixes completely with any other gas.





GAS PRESSURE is a measure of the force that a gas exerts on its container. Force is the physical quantity that interferes with inertia. Gravity is the force responsible for weight.

Force = mass × acceleration; (Newton's 2nd Law)
The SI units follow:
N = kg × m/s^2
<u>Pressure</u>-- Force/ unit area; N/m^2; which is the definition of 1.0 Pascal

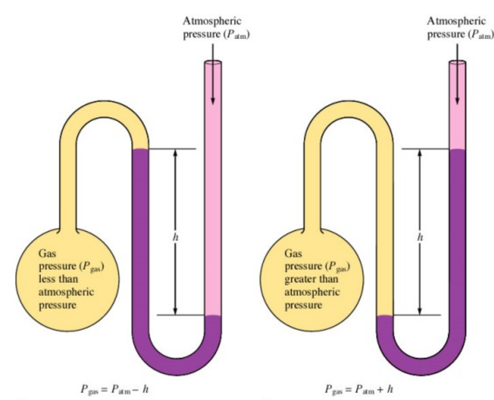
<u>Barometer</u>--invented by Evangelista Torricelli in 1643; uses the height of a column of mercury to measure gas pressure (especially atmospheric *P*)
1 mm of Hg = 1 torr

1.00 atm = 760.00 mm Hg = 760.00 torr = 29.92 in Hg = 14.7 psi = 101.325 kPa $\approx 10^5$ Pa

At sea level all of the above define STANDARD PRESSURE.
The SI unit of pressure is the Pascal (named after Blaise Pascal);
1 Pa = 1 N / m^2

The **manometer—**a device for measuring the pressure of a gas in a container. The pressure of the gas is given by *h* [the difference in mercury levels] in units of torr (equivalent to mm Hg).

a) Gas pressure = atmospheric pressure – *h*
b) Gas pressure = atmospheric pressure + *h*



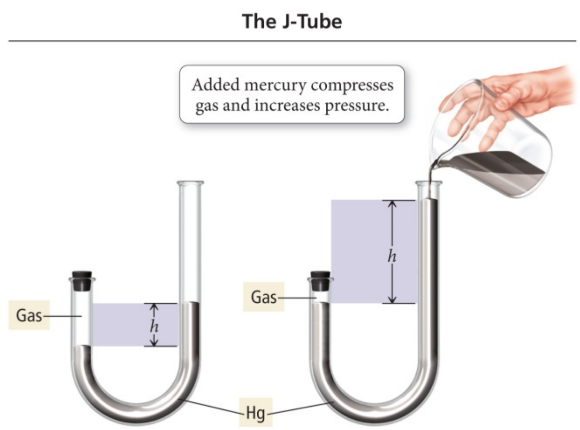
Exercise 1 Pressure Conversions

The pressure of a gas is measured as 49 torr. Represent this pressure in both atmospheres and Pascals.

Exercise 2 Pressure Comparisons

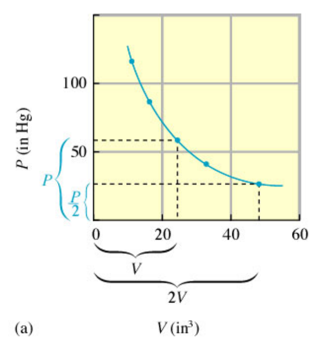
Rank the following pressures in decreasing order of magnitude (largest first, smallest last): 75 kPa, 300. torr, 0.60 atm and 350. mm Hg.

GAS LAWS: THE EXPERIMENTAL BASIS

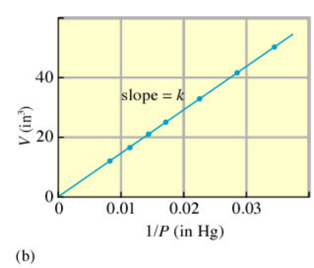


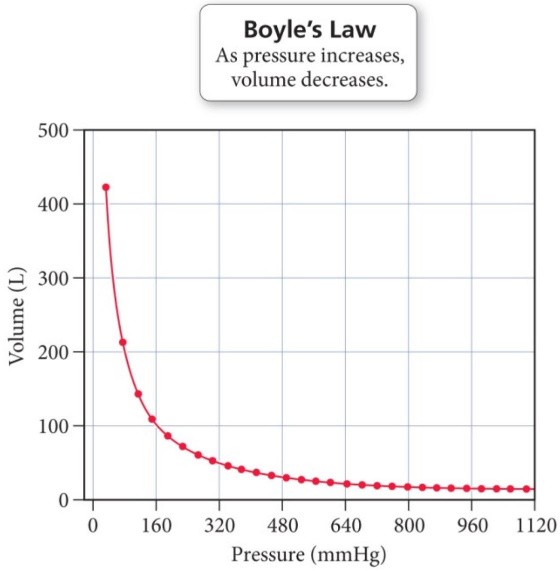
BOYLE'S LAW: "Father of Chemistry"--*the volume of a confined gas is inversely proportional to the pressure exerted on the gas.* ALL GASES BEHAVE IN THIS MANNER!

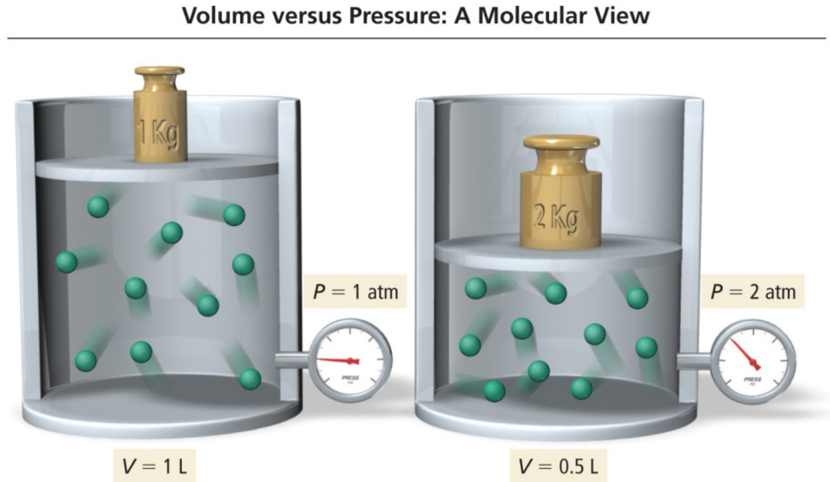
- Robert Boyle was an Irish chemist. He studied PV relationships using a J-tube set up in the multi-story entryway of his home. (Thus his was MUCH larger than the one shown right.)
  - o $P \propto 1/V$
  - o $\therefore$ pressure and volume are **inversely** proportional
  - o Volume ↑ Pressure ↓ **at constant temperature,** the converse is also true
  - o *for a given quantity of a gas at constant temperature, the product of pressure and volume is a constant*
    - $PV = k$
    - Therefore, $V = \frac{k}{P} = k\frac{1}{P}$
    - which is the equation for a straight line of the type $y = mx + b$, where m = slope, and b is the y-intercept
    - In this case, $y = V$, $x = 1/P$ and $b = 0$. Check out the plot on the right (b). The data Boyle originally collected is graphed on (a) above on the right.



  - o $P_1V_1 = P_2V_2$ is the easiest form of Boyle's law to **memorize**
  - o Boyle's Law has been tested for over three centuries. It holds true ***<u>only at low pressures</u>***.



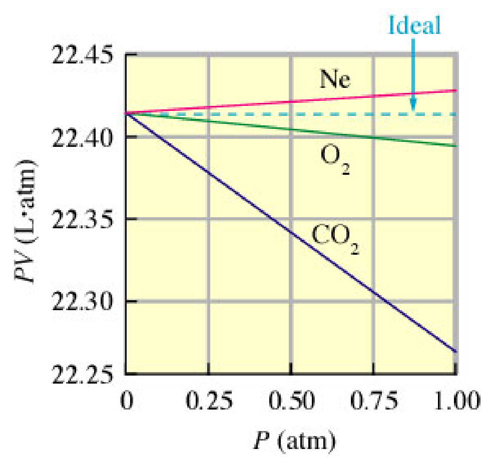




Exercise 3 **Boyle's Law I**

Sulfur dioxide (SO_2), a gas that plays a central role in the formation of acid rain, is found in the exhaust of automobiles and power plants. Consider a 1.53- L sample of gaseous SO_2 at a pressure of 5.6×10^3 Pa. If the pressure is changed to 1.5×10^4 Pa at a constant temperature, what will be the new volume of the gas?

Ideal Gases



At "normal" conditions such as standard temperature and pressure, most real gases behave qualitatively like an ideal gas. Many gases such as nitrogen, oxygen, hydrogen, noble gases, and some heavier gases like carbon dioxide can be treated like ideal gases within reasonable tolerances. Generally, a gas behaves more like an ideal gas at higher temperature and lower pressure, as the work which is against intermolecular forces becomes less significant compared with the particles' kinetic energy, and the size of the molecules becomes less significant compared to the empty space between them.

An **ideal** gas is expected to have a constant value of *PV*, as shown by the dotted line on the graph pictured right. CO_2 shows the largest change in *PV*, and this change is actually quite small: *PV* changes from about 22.39 L·atm at 0.25 atm to 22.26 L·atm at 1.00 atm. Thus Boyle's Law is a good approximation at these relatively low pressures.
So, why does CO_2 deviate from ideal behavior the most? It has more electrons, thus is more polarizable, thus has higher dispersion forces (a type of intermolecular force a.k.a. London dispersion forces or LDFs), therefore the molecules are more attracted to each other, so carbon dioxide gas deviates from ideal behavior more than oxygen or carbon dioxide does.

Exercise 4 **Boyle's Law II**

In a study to see how closely gaseous ammonia obeys Boyle's law, several volume measurements were made at various pressures, using 1.0 mol NH_3 gas at a temperature of 0ºC. Using the results listed below; calculate the Boyle's law constant for NH_3 at the various pressures. Calculate the deviation from ideal behavior in each case. Account for the trend apparent in the data.

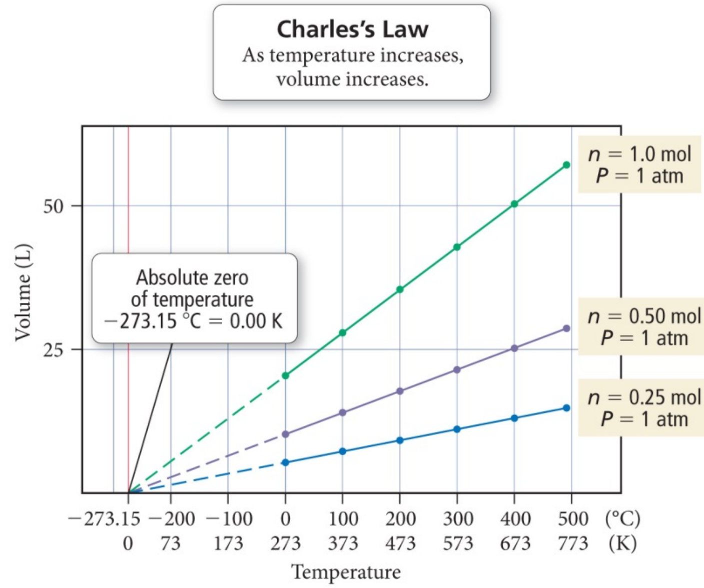
Experiment	*Pressure (atm)*	*Volume (L)*
1	0.1300	172.1
2	0.2500	89.28
3	0.3000	74.35
4	0.5000	44.49
5	0.7500	29.55
6	1.000	22.08

Exercise 5 **Boyle's Law III**

Next, PLOT the values of PV vs. P for the six experiments in **Exercise 4**. Extrapolate to determine what PV equals at a ***hypothetical*** 0.00 atm pressure. Compare it to the PV vs. P graph on page 3 of these lecture notes. What is the value of the y-intercept?

CHARLES' LAW: *If a given quantity of gas is held at a constant pressure, then its volume is directly proportional to the absolute temperature.* **<u>Must use KELVINS</u>** Why?

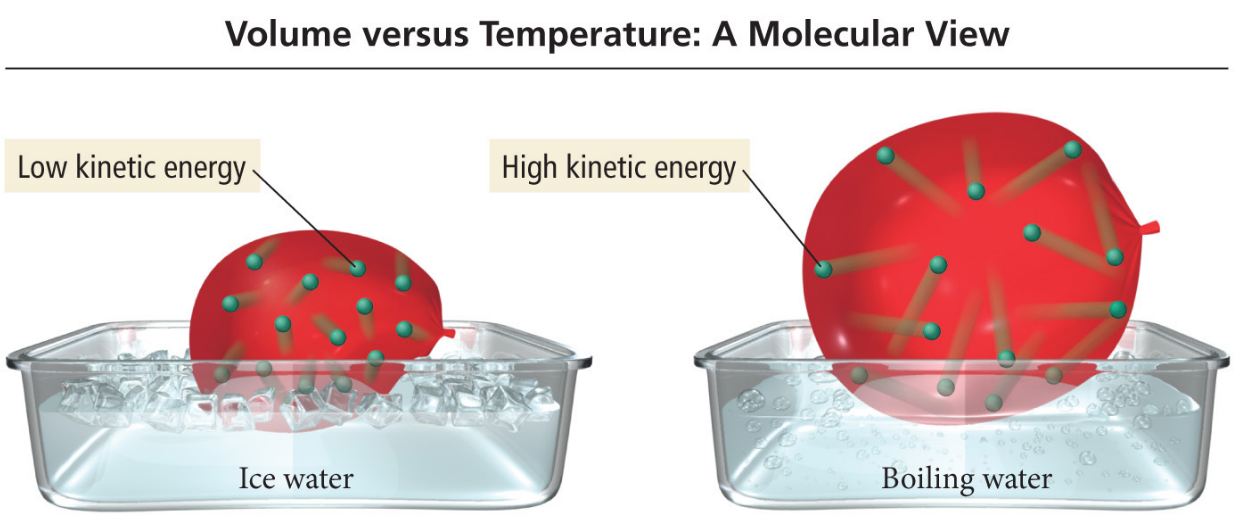
- Jacques Charles was a French physicist and the first person to fill a hot "air" balloon with hydrogen gas and made the first solo hot air balloon flight!
  - $V \propto T$ plot = straight line
  - $V_1T_2 = V_2T_1$
  - Temperature $\propto$ Volume **at constant pressure**
  - This figure shows the plots of V vs. T (Celsius) for several gases. The solid lines represent experimental measurements on gases. The dashed lines represent extrapolation of the data into regions where these gases would become liquids or solids. Note that the samples of the various gases contain different numbers of moles.
  - What is the *temperature* when the *volume* extrapolates to zero? Sound familiar?



Exercise 6 **Charles's Law**

A sample of gas at 15ºC and 1 atm has a volume of 2.58 L. What volume will this gas occupy at 38ºC and 1 atm?

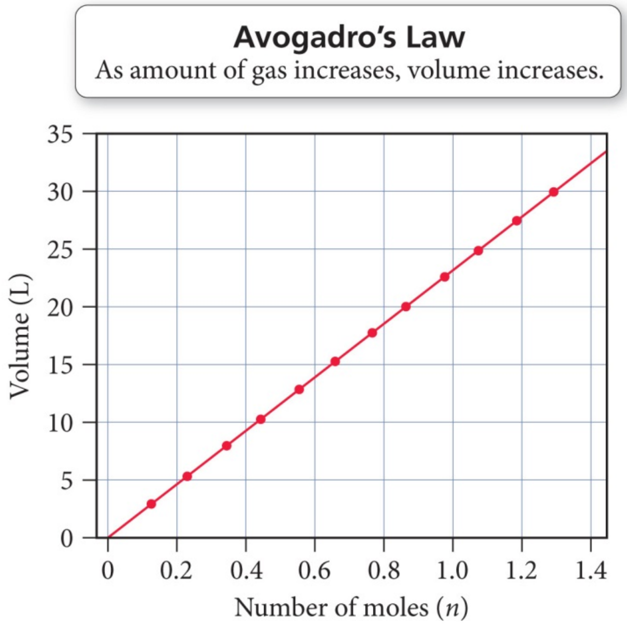
Each of the balloons below contains the same number of particles in the gas phase. This is yet another example of "heat 'em up and speed 'em up!". As the molecules warm, they gain kinetic energy, move faster, and thus collide with the walls of their container with more energy. (Umph! If you prefer) In this case, the "walls" of the container are made of rubber which can expand and contract.

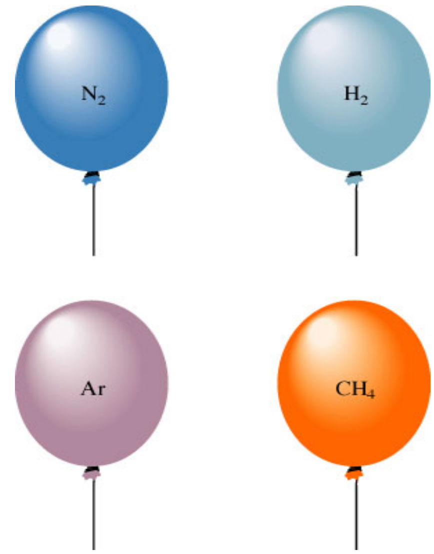


GAY-LUSSAC'S LAW of combining volumes: volumes of gases always combine with one another in the ratio of small whole numbers, as long as volumes are measured at the same T and P.

- $P_1T_2 = P_2T_1$

- Avogadro's hypothesis: *equal volumes of gases under the same conditions of temperature and pressure contain equal numbers of molecules.*





These balloons each hold 1.0 L of gas at 25°C and 1 atm. Each balloon contains 0.041 mol of gas, or 2.5×10^{22} gas molecules.

AVOGADRO'S LAW: *The volume of a gas, at a given temperature and pressure, is directly proportional to the quantity of gas.*

- $V \propto n$
- $n \propto$ Volume **at constant *T* & *P***

HERE'S AN EASY WAY TO MEMORIZE ALL OF THIS!
Start with the combined gas law:

$$P_1V_1T_2 = P_2V_2T_1$$

Memorize just this use a simple pattern to figure the rest out:

- Place the scientist names in alphabetical order.
- Boyle's Law uses the first 2 variables, Charles' Law the second 2 variables & Gay-Lussac's Law the remaining combination of variables.
  Whichever variable doesn't appear in the formula is being held CONSTANT!

Exercise 7 **Avogadro's Law**
Suppose we have a 12.2-L sample containing 0.50 mol oxygen gas (O_2) at a pressure of 1 atm and a temperature of 25ºC. If all this O_2 were converted to ozone (O_3) at the same temperature and pressure, what would be the volume of the ozone?

THE IDEAL GAS LAW

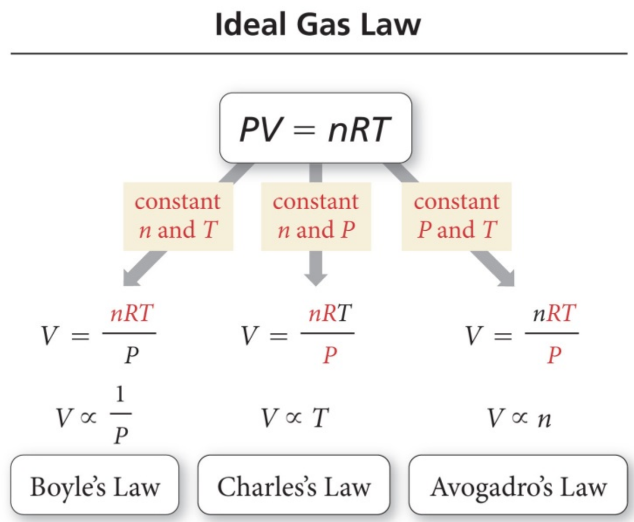
Four quantities describe the state of a gas: pressure, volume, temperature, and # of moles (quantity).

Combine all 3 laws:

$$V \propto \frac{nT}{P}$$

Replace the $\propto$ with a constant, R, and you get:

$$PV = nRT$$



The ideal gas law ***is an equation of state***.

R = 0.08206 L• atm/mol • K also expressed as 0.08206 L atm mol^{-1} K^{-1}

Useful only at low Pressures and high temperatures!

Guaranteed points on the AP Exam!

These next exercises can all be solved with the ideal gas law, BUT, you can use another if you like!

Exercise 8 Ideal Gas Law I

A sample of hydrogen gas (H_2) has a volume of 8.56 L at a temperature of 0ºC and a pressure of 1.5 atm. Calculate the moles of H_2 molecules present in this gas sample.

Exercise 9 Which Gas Law?

Suppose we have a sample of ammonia gas with a volume of 3.5 L at a pressure of 1.68 atm. The gas is compressed to a volume of 1.35 L at a constant temperature. Use the ideal gas law to calculate the final pressure.

Exercise 10 Which Gas Law?

A sample of methane gas that has a volume of 3.8 L at 5ºC is heated to 86ºC at constant pressure. Calculate its new volume.

Exercise 11 **Which Gas Law?**

A sample of diborane gas (B_2H_6), a substance that bursts into flame when exposed to air, has a pressure of 345 torr at a temperature of -15ºC and a volume of 3.48 L. If conditions are changed so that the temperature is 36ºC and the pressure is 468 torr, what will be the volume of the sample?

Exercise 12 **Ideal Gas Law II**

A sample containing 0.35 mol argon gas at a temperature of 13ºC and a pressure of 568 torr is heated to 56ºC and a pressure of 897 torr. Calculate the change in volume that occurs.

GAS STOICHIOMETRY

Use $PV = nRT$ to solve for the volume of one mole of gas at STP:

Look familiar? This is the **molar volume** of a gas at STP. Work stoichiometry problems using your favorite method, dimensional analysis, mole map, the table way…just work FAST! Use the ideal gas law to convert quantities that are NOT at STP.

Exercise 13 **Gas Stoichiometry I**

A sample of nitrogen gas has a volume of 1.75 L at STP. How many moles of N_2 are present?

Exercise 14 Gas Stoichiometry II

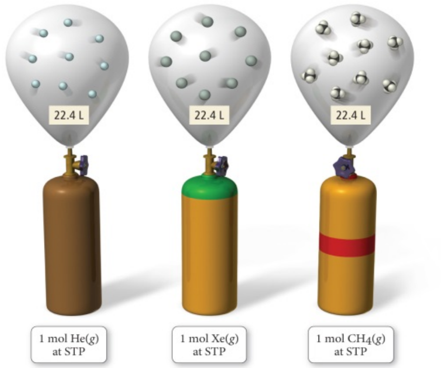
Quicklime (CaO) is produced by the thermal decomposition of calcium carbonate ($CaCO_3$). Calculate the volume of CO_2 at STP produced from the decomposition of 152 g $CaCO_3$ by the reaction

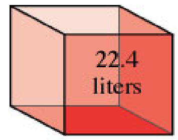
$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

Exercise 15 Gas Stoichiometry III

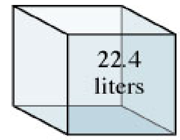
A sample of methane gas having a volume of 2.80 L at 25ºC and 1.65 atm was mixed with a sample of oxygen gas having a volume of 35.0 L at 31ºC and 1.25 atm. The mixture was then ignited to form carbon dioxide and water. Calculate the volume of CO_2 formed at a pressure of 2.50 atm and a temperature of 125ºC.

THE DENSITY OF GASES

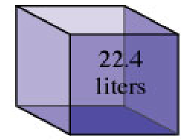




32 g of O_2 (1 mole)



2 g of H_2 (1 mole)



17 g of NH_3 (1 mole)

$$d = \frac{m}{V} = \frac{P(MM)}{RT} \text{ \{for ONE mole of gas\}} = \frac{MM}{22.4\ L} \quad \text{AND} \quad \text{Molar Mass} = MM = \frac{dRT}{P}$$

"Molecular Mass kitty cat"—all good cats put dirt [*dRT*] over their pee [*P*]. Corny? Yep! Crude and socially unacceptable? You bet! But … you'll thank me later when you're flying through such gas law problems.

Just remember that densities of gases are reported in g/L NOT g/mL.

What is the approximate molar mass of air expressed in g/L?

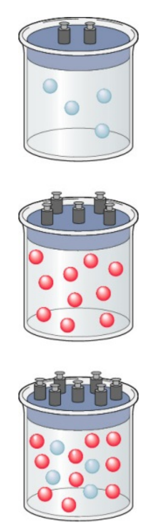
List 3 gases that float in air:

List 3 gases that sink in air:

Exercise 16 **Gas Density/Molar Mass**

The density of a gas was measured at 1.50 atm and 27ºC and found to be 1.95 g/L. Calculate the molar mass of the gas.

GAS MIXTURES AND PARTIAL PRESSURES



The pressure of a mixture of gases is the sum of the pressures of the different components of the mixture:

$$P_{total} = P_1 + P_2 + \ldots P_n$$

John Dalton's Law of Partial Pressures also uses the concept of mole fraction, χ

$$\chi_A = \frac{\text{moles of A}}{\text{moles A + moles B + moles C + . . .}}$$

so now, $P_A = \chi_A P_{total}$

The partial pressure of each gas in a mixture of gases in a container depends on the number of moles of that gas. The total pressure is the SUM of the partial pressures and depends on the total moles of gas particles present, no matter what they are!

Exercise 17 **Dalton's Law I**

Mixtures of helium and oxygen are used in scuba diving tanks to help prevent "the bends." For a particular dive, 46 L He at 25ºC and 1.0 atm and 12 L O_2 at 25ºC and 1.0 atm were pumped into a tank with a volume of 5.0 L. Calculate the partial pressure of each gas and the total pressure in the tank at 25ºC.

Exercise 18 **Dalton's Law II**

The partial pressure of oxygen was observed to be 156 torr in air with a total atmospheric pressure of 743 torr. Calculate the mole fraction of O_2 present.

Exercise 19 **Dalton's Law III**

The mole fraction of nitrogen in the air is 0.7808. Calculate the partial pressure of N_2 in air when the atmospheric pressure is 760. torr.

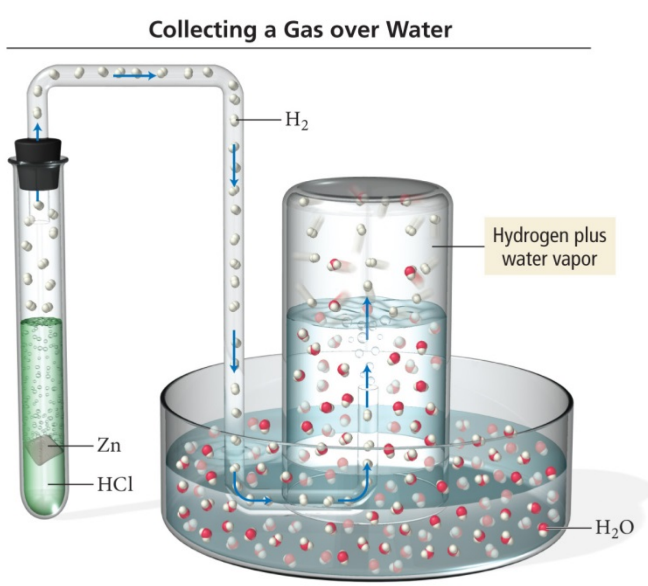
WATER DISPLACEMENT

It is common to collect a gas in the laboratory by water displacement. The confounding factor is that some of the pressure in the collection vessel is due to water vapor collected as the gas was passing through!

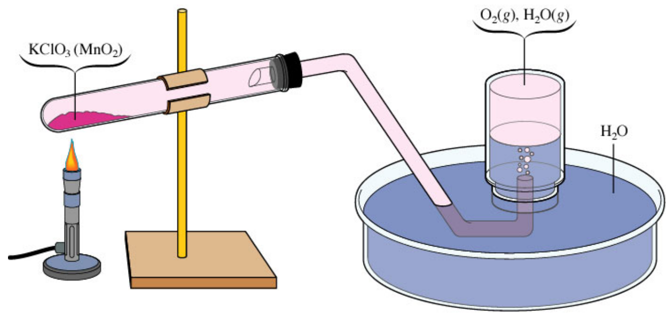
You must correct for this in order to report the *P* of "dry" gas.

How? You simply look up the partial pressure due to water vapor at a given ***temperature*** and subtract that value from the total pressure.

The experiment pictured is a classic! You may have done it with Mg rather than zinc and used a eudiometer or inverted and sealed buret to measure the volume of the gas collected "over water".



Exercise 20 **Gas Collection over Water**



A sample of solid potassium chlorate ($KClO_3$) was heated in a test tube (see the figure above) and decomposed by the following reaction:

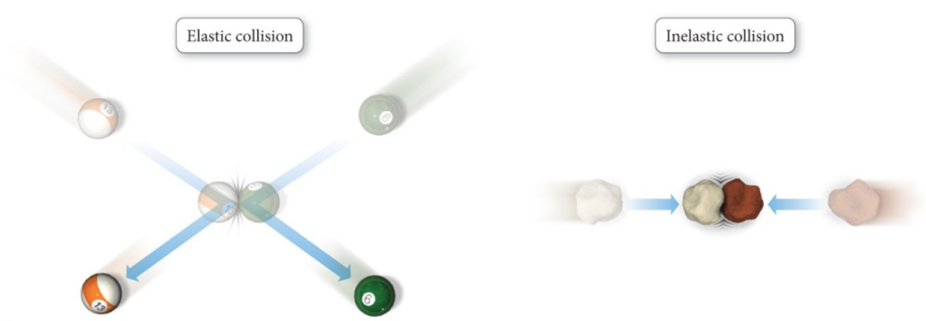
$$2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)$$

The oxygen produced was collected by displacement of water at 22ºC at a total pressure of 754 torr. The volume of the gas collected was 0.650 L, and the vapor pressure of water at 22ºC is 21 torr. Calculate the partial pressure of O_2 in the gas collected and the mass of $KClO_3$ in the sample that was decomposed.

KINETIC MOLECULAR THEORY OF GASES

Assumptions of the MODEL:

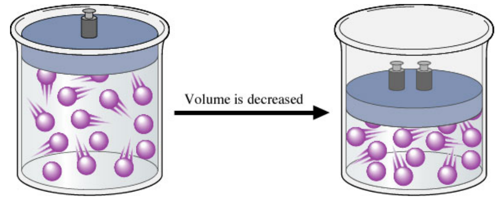
1. All particles are in constant, random, motion.
2. All collisions between particles are perfectly elastic.
3. The volume of the particles in a gas is negligible
4. The average kinetic energy of the molecules is its Kelvin temperature.



This theory neglects any intermolecular forces as well. And it is important to note that gases expand to fill their container, solids/liquids do not. And that gases are compressible; solids/liquids are not appreciably compressible.

This helps explain Boyle's Law:

If the volume is decreased that means that the gas particles will hit the wall more often, thus increasing pressure

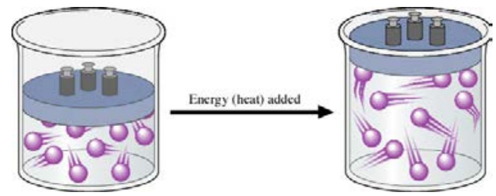


$$P = (nRT)\frac{1}{V}$$

Constant

And also helps explain Charles' Law

When a gas is heated, the speeds of its particles increase and thus hit the walls more often and with more force. The only way to keep the *P* constant is to increase the volume of the container.

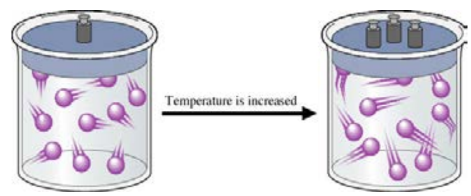


$$V = \left(\frac{nR}{P}\right)T$$

Constant

Yep, you guessed it! It also helps explain Gay-Lussac's Law

When the temperature of a gas increases, the speeds of its particles increase, the particles are hitting the wall with greater force and greater frequency. Since the volume remains the same this would result in increased gas pressure.

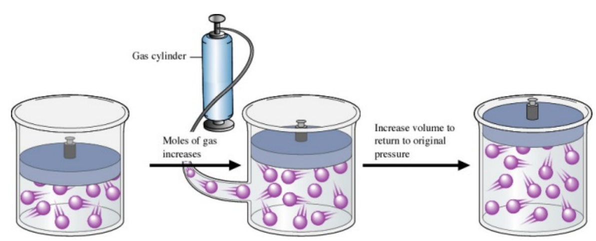


$$P = \left(\frac{nR}{V}\right)T$$

Constant

And it even helps explain Avogadro's Law

An increase in the number of particles at the same temperature would cause the pressure to increase if the volume were held constant. The only way to keep constant *P* is to vary the *V*.



$$V = \left(\frac{RT}{P}\right)n$$

Constant

What about Dalton's Law? The *P* exerted by a *mixture* of gases is the SUM of the partial pressures since gas particles are acting *independent* of each other and the volumes of the individual particles DO NOT matter.

DISTRIBUTION OF MOLECULAR SPEEDS

Plot # of gas molecules having various speeds vs. the speed and you get a curve. Changing the temperature affects the *shape* of the curve NOT the area beneath it. Change the # of molecules and all bets are off!

Maxwell's equation:

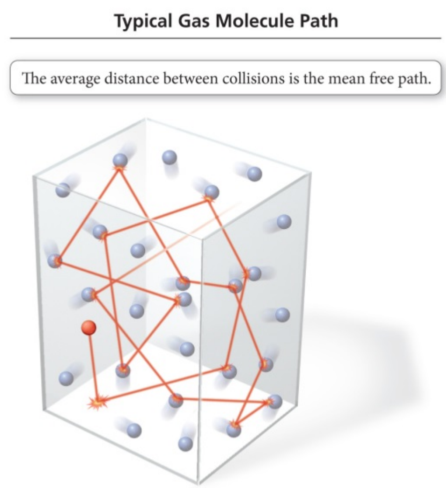
$$\sqrt{\overline{u^2}} = u_{rms} = \sqrt{\frac{3RT}{MM}}$$

Use the "energy *R*" or 8.314510 J/K• mol for this equation since kinetic energy is involved.

Exercise 21 **Root Mean Square Velocity**

Calculate the root mean square velocity for the atoms in a sample of helium gas at 25ºC.

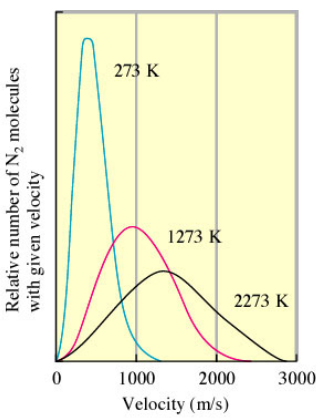
If we could monitor the path of a single molecule it would be very erratic.



<u>Mean free path</u>—the average distance a particle travels between collisions. It's on the order of a tenth of a micrometer. WAAAAY SMALL!

Examine the effect of temperature on the numbers of molecules with a given velocity as it relates to temperature.

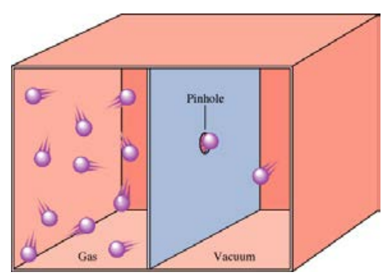
HEAT 'EM UP, SPEED 'EM UP!!



Drop a vertical line from the peak of each of the three bell shaped curves—that point on the *x*-axis represents the AVERAGE velocity of the sample at that temperature. Note how the bells are "squashed" as the temperature increases. You may see graphs like this on the AP exam where you have to identify the highest temperature based on the shape of the graph!

GRAHAM'S LAW OF DIFFUSION AND EFFUSION

Effusion is closely related to diffusion. **<u>Diffusion</u>** is the term used to describe the mixing of gases. The ***rate*** of diffusion is the ***rate*** of the mixing.



<u>Effusion</u> is the term used to describe the passage of a gas through a tiny orifice into an evacuated chamber as shown on the right.
The rate of effusion measures the speed at which the gas is transferred into the chamber.

The rates of effusion of two gases are inversely proportional to the square roots of their molar masses at the same temperature and pressure.

$$\frac{\text{Rate of effusion of gas 1}}{\text{Rate of effusion of gas 2}} = \sqrt{\frac{MM_2}{MM_1}}$$

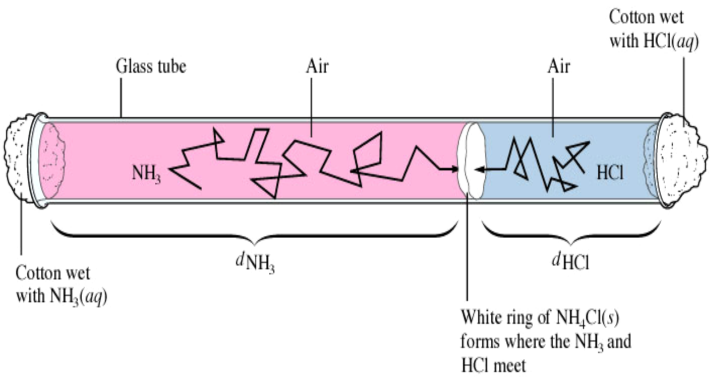
REMEMBER ***rate*** is a change in a quantity over time, NOT just the time!

Exercise 22 **Effusion Rates**
Calculate the ratio of the effusion rates of hydrogen gas (H_2) and uranium hexafluoride (UF_6), a gas used in the enrichment process to produce fuel for nuclear reactors.

Exercise 23
A pure sample of methane is found to effuse through a porous barrier in 1.50 minutes. Under the same conditions, an equal number of molecules of an unknown gas effuses through the barrier in 4.73 minutes. What is the molar mass of the unknown gas?

Diffusion

This is a classic! https://www.youtube.com/watch?v=GRcZNCA9DxE (animation) Effusion of a Gas:
http://www.youtube.com/watch?v=0uBK7VxT00E
https://www.youtube.com/watch?v=o7C4lo5n0zU (actual, but not great)



$$\frac{\text{Distance traveled by } NH_3}{\text{Distance traveled by HCl}} = \frac{u_{rms} \text{ for } NH_3}{u_{rms} \text{ for HCl}} = \sqrt{\frac{MM_{HCl}}{MM_{NH_3}}} = \sqrt{\frac{36.5}{17}} = 1.5$$

The observed ratio is LESS than a 1.5 distance ratio—why?

This diffusion is slow despite considering the molecular velocities are 450 and 660 meters per second—which one is which?

This tube contains air and all those collisions slow the process down in the real world. Speaking of real world….

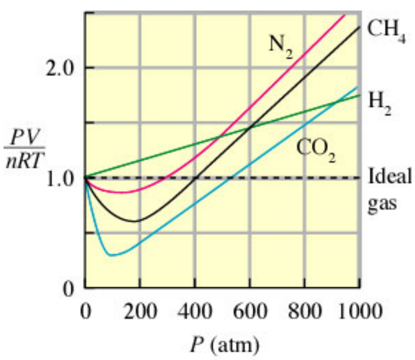
REAL, thus *NON*IDEAL GASES
Most gases behave ideally until you reach high pressure and low temperature. (By the way, either of these can cause a gas to liquefy, go figure!)

van der Waals Equation--corrects for negligible volume of molecules and accounts for inelastic collisions leading to intermolecular forces (his real claim to fame).

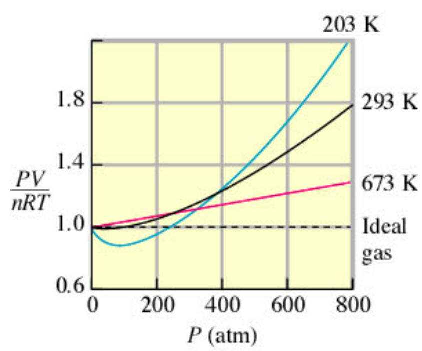
$$[P + a(\frac{n}{V})^2][V - bn] = nRT$$

The coefficients *a* and *b* are van der Waals constants; no need to work problems, it's the concepts that are important! Notice pressure is increased (intermolecular forces lower real pressure, you're correcting for this) and volume is decreased (corrects the container to a smaller "free" volume).

These graphs are classics and make great multiple choice questions on the AP exam.



When PV/nRT = 1.0, the gas is ideal
All of these are at 200K.
Note the P's where the curves
cross the dashed line [ideality].



This graph is just for nitrogen gas.
Note that although nonideal behavior
is evident at each temperature, the
deviations are smaller at the higher Ts.

Don't underestimate the power of understanding these graphs. We love to ask question comparing the behavior of ideal and real gases. It's not likely you'll be asked an entire free-response gas problem on the real exam in May. Gas Laws are tested extensively in the multiple choice since it's easy to write questions involving them! You will most likely see $PV = nRT$ as one part of a problem in the free-response, just not a whole problem!

GO FORTH AND RACK UP THOSE MULTIPLE CHOICE POINTS!!

And, just for fun:

Collapsing Can: http://www.graspr.com/videos/The-Collapsing-Can-1

Shall we try a bigger can? http://www.youtube.com/watch?v=Uy-SN5j1ogk&NR=1

How about the biggest can we can find? http://www.youtube.com/watch?v=E_hci9vrvfw

Boiling water with ice: http://www.youtube.com/watch?v=zzVtbvVS2lQ

Cooling gases with liquid nitrogen MIT: http://www.youtube.com/watch?v=ZvrJgGhnmJo

Getting a boiled egg into a bottle: http://www.youtube.com/watch?v=xZdfcRiDs8I&NR=1&feature=fvwp

Gravity has nothing to do with it! http://www.youtube.com/watch?v=BofIBaYk5e0&feature=related

Getting egg OUT of bottle! http://www.youtube.com/watch?v=x--4l-SL77Y&feature=related

Peeps in a vacuum: http://www.youtube.com/watch?v=lfNJJEdKgLU&NR=1

and another: http://www.youtube.com/watch?v=ciPr4Tg9k78&feature=related

Ideal Gas Law: http://www.youtube.com/watch?v=Mytvt0wlZK8&feature=related

Diffusion of ammonia and HCl (animation): http://www.youtube.com/watch?v=L41KhBPBymA&feature=related

Diffusion of ammonia and HCl (real deal): http://www.youtube.com/watch?v=WAJAslkwolk&feature=related

Putting it all together:
http://www.mhhe.com/physsci/chemistry/animations/chang_7e_esp/gam2s2_6.swf

Why *R* as in *PV* = *nRT*?
This is the kind of stuff that drives me crazy as well! I went in search of the answer some time ago, so I'll share what I found:

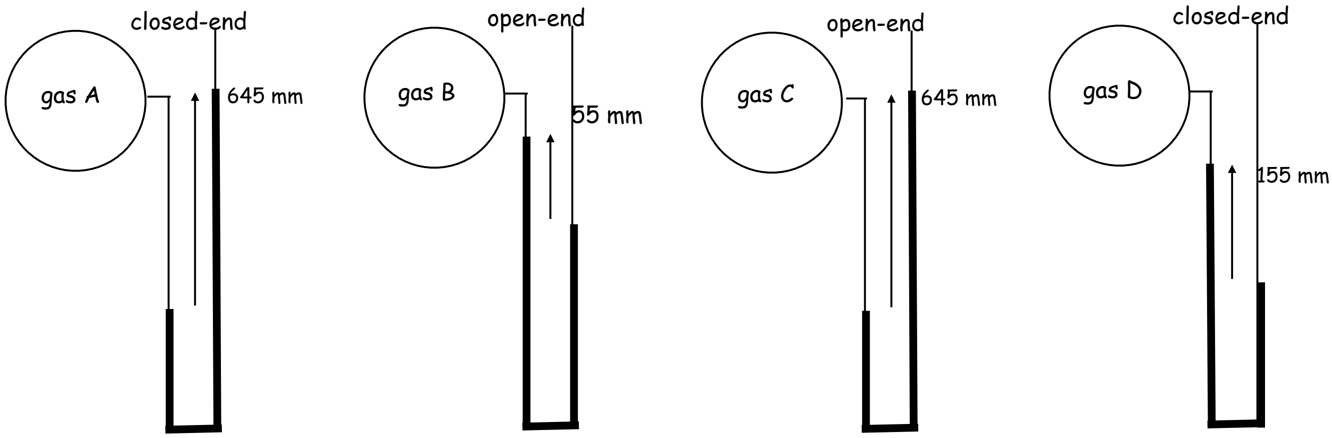
As usual, the answer relates to a history lesson. It was Clausius in the mid 1800’s that refined the conversion factor for converting degrees Celsius to Kelvins (from adding 267 to adding 273 to the Celsius temperature. He did so using the careful experimental data of another French scientist, Regnault.

Clausius also noted that Regnault’s data indicated that the farther the temperature and pressure conditions were from the condensation point of the gas, the more correctly the Ideal Gas Law applies. So, there is speculation that the “constant” was assigned the letter “R” to honor Regnault’s work.

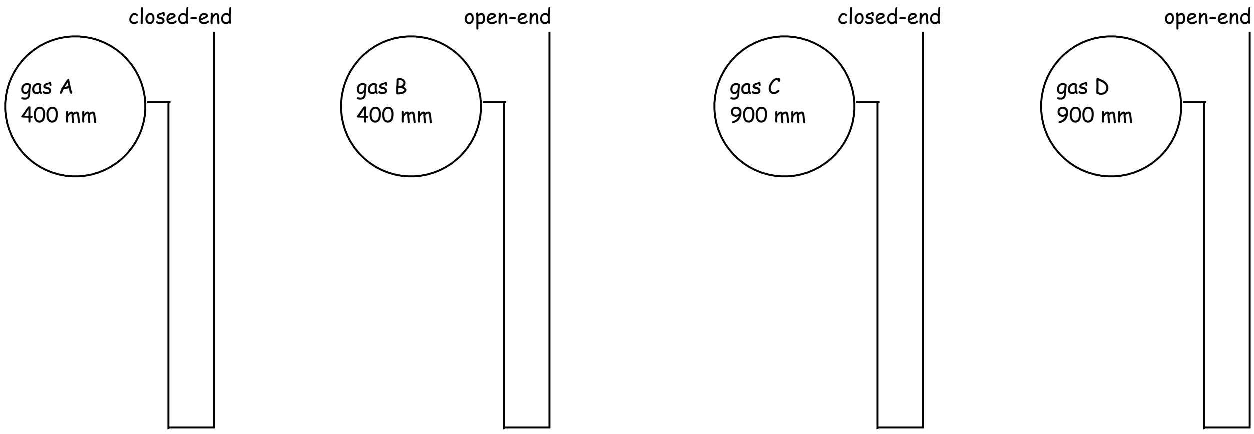
In the spirit of giving credit where credit is due, my source was a Journal of Chem. Ed article written by William B. Jensen Department of Chemistry, University of Cincinnati, Cincinnati, OH 45221-0172

P 13.1 A (pg 1 of 2) Manometers and Barometers

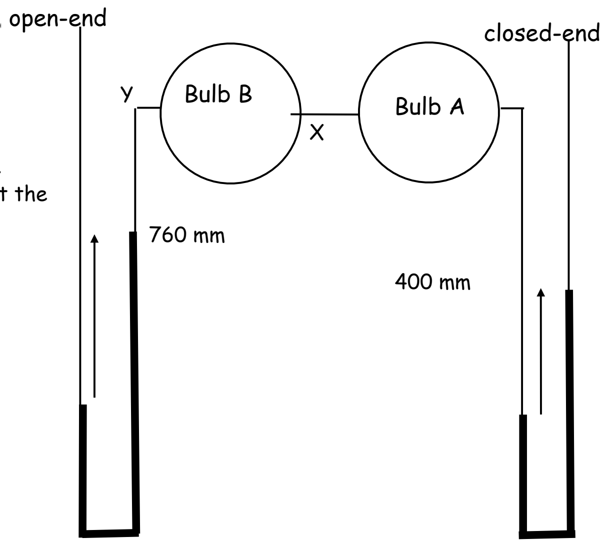
1. Determine the gas pressure inside each bulb. Assume the atmospheric pressure is 755 mm.



2. Draw the position and relative heights of the columns of mercury for each bulb. Assume the atmospheric pressure is 760 mm.



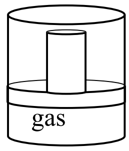
3. At the start all tubes were completely empty, Valves at X and Y are
   and Bulb B is empty, and gas was put into Bulb A.
   a. What is the pressure in B?
   b. What is the pressure in Bulb A?
   c. What is the air pressure in the room?
   d. When Valve X is opened, determine the pressure in Bulbs A & B.
   e. When Valve Y is opened, draw and label the appropriate levels at the open-end



P 13.3 (pg 1 of 3) Combined Gas Law ("Before & After" Problems)

1. Imagine you have a sealed 20.0 L balloon filled with helium gas at 750 mmHg in the house at 25ºC.
   a. If you brought it outside on a winter day and its temperature dropped to -10ºC. Determine the volume of the balloon. Assume that the pressure in the balloon will stay constant.
   b. Then if you bring the same balloon into a sauna at 35ºC, what would the volume of the balloon become?

2. If you have a sealed syringe set up like the picture on the right and you have 5 ml of gas that weighs 0.0034 g inside at 45ºC and 1 atm pressure. When you draw the syringe back to a volume of 20 ml and cool the gas to 30ºC,
   a. What will be the mass of the gas?
   b. What will be the pressure of the gas inside the syringe?



3. If you had a closed 10 L container of gas at 2 atm pressure and 20ºC, and you reduced the volume to 4 L what would the temperature of the gas need to be, to be sure that the pressure increased to 5 atm?

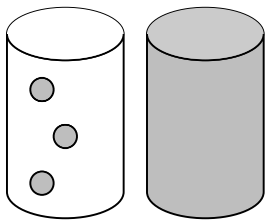
4. If you had the same 10 L container of gas at 2 atm pressure and 20ºC, and you reduced the volume to only 8 L what would the temperature of the gas need to be, to be sure that the pressure increased to 5 atm?

5. A sample of air in a sealed rigid container is at 30ºC to what temp in celsius do you have to heat it to double the pressure?

6. If you had 3 moles of gas in a rigid container at 20ºC and 5 atm pressure and you wanted to double the pressure while maintaining a constant temperature, how many moles of gas should you squirt in?

7. If you have 18.0 g of gas in a 5.0 L glass flask at 10ºC and 3 atm pressure what mass of gas do you need to release so that the pressure will drop to 1 atm

8. The two sealed containers shown at the right are at the same temp and pressure, how many molecules must be in the gray container?



9. The same containers as shown to the right, but this time the pressure in the right container is twice that of the left container, how many molecules must be in the gray container?

P 13.4 (pg 1 of 2) Ideal Gas Law & Density

Please note that both mm Hg and atm are used as units of pressure. It is a simple conversion: 760 mm = 1 atm

1. Given 3.43 g of gas in a 2.00 L container at 25.0°C and a pressure of 1140 mm Hg:
   a. Determine the number of moles of gas in the container.
   b. Recalling that molar mass (molecular weight) is nothing more than a quotient of grams per mole (mass/moles), determine the molar mass of this gas.
   c. What might be the identity of this gas?

2. A 3.0 L flask at 30.0°C contains 0.250 mole of Cl_2 gas.
   a. What is the pressure in the flask?
   b. What is the mass of the gas in the flask?
   c. What is the density of the chlorine gas in this flask?

3. A 500.0 ml flask contained O_2 gas at 25.0°C at a pressure of 4.5 atm.
   a. What is the number of moles in the flask?
   b. What is the mass of the gas in the flask?
   c. What is the density of the oxygen in the flask

4. A 5.0 L flask of carbon dioxide gas at a pressure of 4.54 atm had a mass of 36 g?
   a. How many moles of gas are in this flask?
   b. What is the temperature, in Kelvin and °C, of the gas in this flask?

5. How large of a metal gas canister would you need to contain 20.0 moles of compressed gas at a pressure of 22 atm and at room temperature, 25.0°C?

6. The density of SO_2 gas n a container at room temperature, 25.0°C is 2.51 g/L.
   a. Determine the number of moles of 1.00 L of this gas.
   b. Determine the pressure in this flask.

7. Determine the density of O_2 at STP.

Unit A: Free Response Questions

The following free responses are taken DIRECTLY from released AP exams, the rubrics can be found at the end of this packet. Try these and see how many points you can get on your own

1971

$$2\ HCOONa + H_2SO_4 \rightarrow 2\ CO + 2\ H_2O + Na_2SO_4$$

A 0.964 gram sample of a mixture of sodium formate and sodium chloride is analyzed by adding sulfuric acid. The equation for the reaction for sodium formate with sulfuric acid is shown above. The carbon monoxide formed measures 242 milliliters when collected over water at 752 torr and 22.0°C. Calculate the percentage of sodium formate in the original mixture.

1971

At 20°C the vapor pressure of benzene is 75 torr, and the vapor pressure of toluene is 22 torr. Solutions in both parts of this question are to be considered ideal.

(a) A solution is prepared from 1.0 mole of biphenyl, a nonvolatile solute, and 49.0 moles of benzene. Calculate the vapor pressure of the solution at 20°C.

(b) A second solution is prepared from 3.0 moles of toluene and 1.0 mole of benzene. Determine the vapor pressure of this solution and the mole fraction of benzene in the vapor.

1972

A 5.00 gram sample of a dry mixture of potassium hydroxide, potassium carbonate, and potassium chloride is reacted with 0.100 liter of 2.0 molar HCl solution.

(a) A 249 milliliter sample of dry CO_2 gas, measured at 22°C and 740 torr, is obtained from the reaction. What is the percentage of potassium carbonate in the mixture?

(b) The excess HCl is found by titration to be chemically equivalent to 86.6 milliliters of 1.50 molar NaOH. Calculate the percentages of potassium hydroxide and of potassium chloride in the original mixture.

1973

A 6.19 gram sample of PCl_5 is placed in an evacuated 2.00 liter flask and is completely vaporized at 252°C.

(a) Calculate the pressure in the flask if no chemical reaction were to occur.

(b) Actually at 252°C the PCl_5 is partially dissociated according to the following equation:

$$PCl_5(g) \rightarrow PCl_3(g) + Cl_2(g)$$

The observed pressure is found to be 1.00 atmosphere. In view of this observation, calculate the partial pressure of PCl_5 and PCl_3 in the flask at 252°C.

1976

When the molecular weight of a volatile liquid is calculated from the weight, volume, temperature, and pressure of a sample of that liquid when vaporized, the assumption is usually made that the gas behaves ideally. In fact at a temperature not far above the boiling point of the liquid, the gas is not ideal. Explain how this would affect the results of the molecular weight determination.

1982

(a) From the standpoint of the kinetic-molecular theory, discuss briefly the properties of gas molecules that cause deviations from ideal behavior.

(b) At 25°C and 1 atmosphere pressure, which of the following gases shows the greatest deviation from ideal behavior? Give two reasons for your choice.

CH_4 SO_2 O_2 H_2

(c) Real gases approach ideality at low pressure, high temperature, or both. Explain these observations.

1984

The van der Waals equation of state for one mole of a real gas is as follows:

$$(P + {}^{a}/_{V^2})(V - b) = RT$$

For any given gas, the values of the constants *a* and *b* can be determined experimentally. Indicate which physical properties of a molecule determine the magnitudes of the constants a and b. Which of the two molecules, H_2 or H_2S, has the higher value for *a* and which has the higher value for *b*? Explain.

One of the van der Waals constants can be correlated with the boiling point of a substance. Specify which constant and how it is related to the boiling point.

1986

Three volatile compounds X, Y, and Z each contain element Q. The percent by weight of element Q in each compound was determined. Some of the data obtained are given below.

Compound	Percent by Weight of Element Q	Molecular Weight
X	64.8%	?
Y	73.0%	104.
Z	59.3%	64.0

(a) The vapor density of compound X at 27 degrees Celsius and 750. mm Hg was determined to be 3.53 grams per liter. Calculate the molecular weight of compound X.

(b) Determine the mass of element Q contained in 1.00 mole of each of the three compounds.

(c) Calculate the most probable value of the atomic weight of element Q.

(d) Compound Z contains carbon, hydrogen, and element Q. When 1.00 gram of compound Z is oxidized and all of the carbon and hydrogen are converted to oxides, 1.37 grams of CO_2 and 0.281 gram of water are produced. Determine the most probable molecular formula.

1990

A mixture of $H_2(g)$, $O_2(g)$, and 2 millilitres of $H_2O(l)$ is present in a 0.500 litre rigid container at 25°C. The number of moles of H_2 and the number of moles of O_2 are equal. The total pressure is 1,146 millimetres mercury. (The equilibrium vapor pressure of pure water at 25°C is 24 millimetres mercury.)
The mixture is sparked, and H_2 and O_2 react until one reactant is completely consumed.

(a) Identify the reactant remaining and calculate the number of moles of the reactant remaining.

(b) Calculate the total pressure in the container at the conclusion of the reaction if the final temperature is 90°C. (The equilibrium vapor pressure of water at 90°C is 526 millimetres mercury.)

(c) Calculate the number of moles of water present as vapor in the container at 90°C.

1993

Observations about real gases can be explained at the molecular level according to the kinetic molecular theory of gases and ideas about intermolecular forces. Explain how each of the following observations can be interpreted according to these concepts, including how the observation supports the correctness of these theories.

(a) When a gas-filled balloon is cooled, it shrinks in volume; this occurs no matter what gas is originally placed in the balloon.

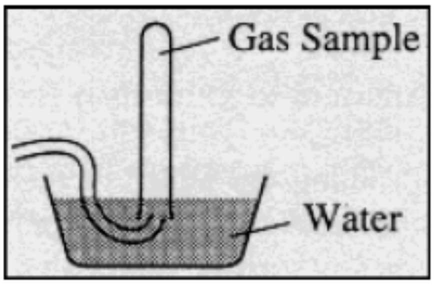
(b) When the balloon described in (a) is cooled further, the volume does not become zero; rather, the gas becomes a liquid or solid.

(c) When NH_3 gas is introduced at one end of a long tube while HCl gas is introduced simultaneously at the other end, a ring of white ammonium chloride is observed to form in the tube after a few minutes. This ring is closer to the HCl end of the tube than the NH_3 end.

(d) A flag waves in the wind.

1994

A student collected a sample of hydrogen gas by the displacement of water as shown by the diagram below The relevant data are given in the following table.



GAS SAMPLE DATA	
Volume of sample	90.0 mL
Temperature	25°C
Atmospheric Pressure	745 mm Hg
Equilibrium Vapor Pressure of H_2O (25°C)	23.8 mm Hg

(a) Calculate the number of moles of hydrogen gas collected.

(b) Calculate the number of molecules of water vapor in the sample of gas.

(c) Calculate the ratio of the average speed of the hydrogen molecules to the average speed of the water vapor molecules in the sample.

(d) Which of the two gases, H_2 or H_2O, deviates more from ideal behavior? Explain your answer.

1995

Propane, C_3H_8, is a hydrocarbon that is commonly used as fuel for cooking.

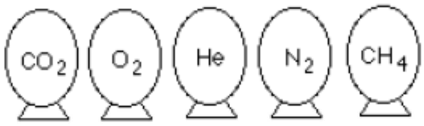
(a) Write a balanced equation for the complete combustion of propane gas, which yields $CO_2(g)$ and $H_2O(l)$.

(b) Calculate the volume of air at 30°C and 1.00 atmosphere that is needed to burn completely 10.0 grams of propane. Assume that air is 21.0 percent O_2 by volume.

(c) The heat of combustion of propane is −2,220.1 kJ/mol. Calculate the heat of formation, ΔH_f°, of propane given that ΔH_f° of $H_2O(l)$ = −285.3 kJ/mol and ΔH_f° of $CO_2(g)$ = −393.5 kJ/mol.

(d) Assuming that all of the heat evolved in burning 30.0 grams of propane is transferred to 8.00 kilograms of water (specific heat = 4.18 J/g·K), calculate the increase in temperature of water.

1996



Represented above are five identical balloons, each filled to the same volume at 25°C and 1.0 atmosphere pressure with the pure gas indicated.

(a) Which balloon contains the greatest mass of gas? Explain.

(b) Compare the average kinetic energies of the gas molecules in the balloons. Explain.

(c) Which balloon contains the gas that would be expected to deviate most from the behavior of an ideal gas? Explain.

(d) Twelve hours after being filled, all the balloons have decreased in size. Predict which balloon will be the smallest. Explain your reasoning.

2002B

A rigid 8.20 L flask contains a mixture of 2.50 moles of H_2, 0.500 mole of O_2, and sufficient Ar so that the partial pressure of Ar in the flask is 2.00 atm. The temperature is 127°C.

(a) Calculate the total pressure in the flask.

(b) Calculate the mole fraction of H_2 in the flask.

(c) Calculate the density (in g L^{-1}) of the mixture in the flask.

The mixture in the flask is ignited by a spark, and the reaction represented below occurs until one of the reactants is entirely consumed.

$$2\ H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

(d) Give the mole fraction of all species present in the flask at the end of the reaction.

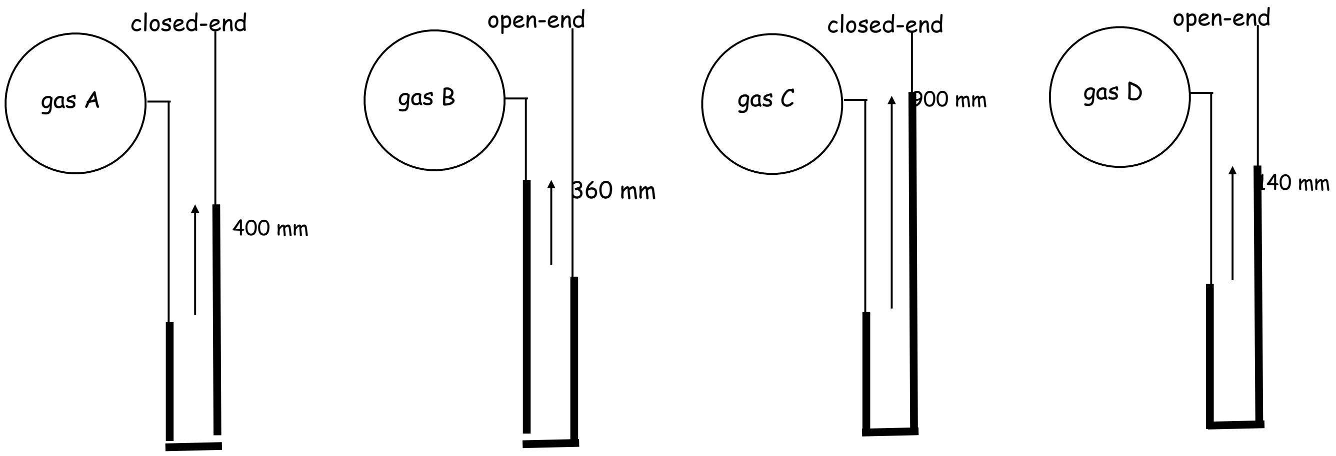
Answer Keys

Don’t look back here till you’ve tied the worksheets on your own

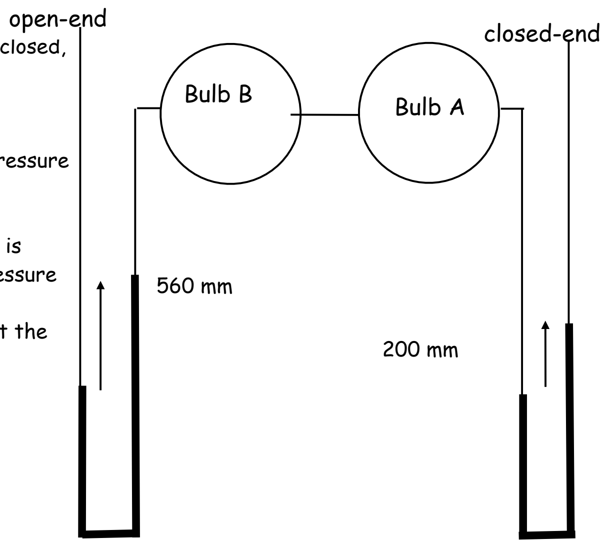
ANSWERS

1. Determine the gas pressure inside each bulb. Assume the atmospheric pressure is 755 mm.
   a. The gas pressure inside is 645 mm Hg. You can read the pressure of a closed-end manometer directly.
   b. The gas pressure inside is 700 mm Hg. You can tell this because the inside gas is losing the push-pull and is 55 mmHg less pressure than the outside air pressure which is stated in the directions to be 755 mmHg.
   c. The gas pressure inside is 1400 mm Hg. You can tell this because the gas inside is pushing harder than the gas outside. It is pushing harder by 645 mmHg. Thus 755 + 645 would be the total pressure inside.
   d. This closed-end manometer must be broken. Since there is "nothing", in the closed-end, there would be no push thus the lowest it could go would be flat with the other side.

2. Draw the position and relative heights of the columns of mercury for each bulb. Assume the atmospheric pressure is 760 mm.



4. At the start all tubes were completely empty, Valves at X and Y are closed, and Bulb B is empty, and gas was put into Bulb A.
   a. Since B is empty, the pressure would be 0 mm Hg
   b. The pressure in Bulb A is 400 mmHg
   c. Since the tubes were empty, you can assume that the outside pressure is 76o mmHg because the outside air is pushing the mercury up towards valve Y to a height of 760 mm
   d. Assuming that Bulb A and B are the same size, the when valve x is opened, the volume will be twice as big which will reduce the pressure to half. Thus both bulbs will have a pressure of 200 mmHg
   e. When Valve Y is opened, draw and label the appropriate levels at the open-end, and the new level at the closed-end.



For all of these problems you should use the equation from NS 13.2: $\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$

If any of the variables stays constant, it can cancel out of the equation.

All of these problems MUST be done in the Kelvin temperature scale. Refer to NS 10.2 to learn how to convert Celsius to Kelvin.

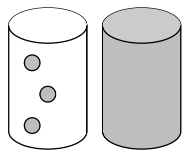
1. Since a balloon is flexible, you can assume that it's internal pressure stays more or less constant, and since the balloon is sealed, the number of moles stays constant. So the equation reduces to $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ and solve for $V_2 = \frac{V_1T_2}{T_1}$
   a. $V_2 = \frac{(20L)(263K)}{(298K)}$ Solve $V_2 = 17.7$ L
   b. $V_2 = \frac{(20L)(308K)}{(298K)}$ Solve $V_2 = 20.7$ L

2. Assume the syringe is sealed
   a. Thus the mass will stay the same = 0.0034 g.
   b. Since the mass stays constant throughout the problem, the moles will remain constant, so the equation reduces to $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ and solve for $P_2 = \frac{P_1V_1T_2}{V_2T_1}$
   $P_2 = \frac{(1atm)(5ml)(303K)}{(20ml)(318K)}$ $P_2 = 0.24$ atm

3. Again the container is sealed, so the n remains constant throughout the problem. You might notice that the volume is reduced by a factor of 2.5 while the pressure in creased by the same factor. This means that these factor changes will cancel each out, meaning that the temp will need to stay the same, 20ºC
   Alternatively, you can use the equation which reduces to $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ then solve for $T_2 = \frac{P_2V_2T_1}{P_1V_1}$
   so $T_2 = \frac{(4L)(5atm)(293K)}{(10L)(2atm)}$ solve and $T_2 = 293$ K which equals 20ºC, the same starting temp.

4. This time the volume is not reduced as much as the pressure is increased. So you will find it easiest to use the equation.
   which reduces to $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ then solve for $T_2 = \frac{P_2V_2T_1}{P_1V_1}$
   so $T_2 = \frac{(8L)(5atm)(293K)}{(10L)(2atm)}$ then solve and $T_2 = 586$ K which = 313ºC (586 K – 273K)

5. The temperature must be doubled. But that is a doubling of the Kelvin temp. So 30ºC = 303K which would double to 606 K which is 333ºC.
   Alternatively you can use the equation. The container is rigid, so the volume will stay constant, and it is sealed so the number of moles is also constant so the equation reduces to $\frac{P_1}{T_1} = \frac{P_2}{T_2}$ and then you can solve for $T_2 = \frac{P_2T_1}{P_1}$ since you are not given a pressure, you can just pick any value, and then double it. I'll pick a value of 1 and double it to 2.
   so $T_2 = \frac{(2L)(303K)}{(1L)}$ then solve and $T_2 = 606$ K which is 333ºC



6. To double the pressure while holding the volume and temperature constant, you would need to double the number of moles of gas.
   Alternatively you can use the equation. The volume is constant because the container is rigid, and the temperature is held constant.
   so the equation reduces to $\frac{P_1}{n_1} = \frac{P_2}{n_2}$ and then you can solve for $n_2 = \frac{P_2n_1}{P_1}$ $n_2 = \frac{(10atm)(3mol)}{5atm}$ so $n_2 = 6$ moles

7. It is important to realize that mass of gas is directly proportional to its number of moles. So even though we don't know the number of moles, we can assume that the mass will change by the same proportion as moles would. In this problem, the volume is constant because it is a glass container, and it says the temp is held at 10ºC. Thus if the pressure is reduced by one third, the amount of moles must be reduced by the same one third. Thus the new mass must be 6 g (one third of 18g) meaning that 12 g must be released.

8. We can assume the volumes are the same, so it the pressure and temp are the same, the number of moles, therefore number of molecules must be the same = 3

9. If the pressure is doubled and the temp and volume are constant, it can only be caused by more gas molecules. They must be doubled. Thus there must be 6 molecules in the gray container.

 Ideal Gas Law Problems

ANSWERS

1. Again apply the ideal gas law solving for n. Be sure your temp is in Kelvin and select the R that matches the P units. Remember that molar mass is mass/moles.
   a. n = (1140 mmHg * 2 L)/(62.4 mmHgL/mole K * 298 K) = 0.123 mole
   b. 3.43 g/0.123 mole = 27.9 g/mole
   c. It's likely to be diatomic nitrogen, N_2 with MM = 28 g/mole

2. Apply the ideal gas law: PV = nRT solving for P. Depending which R value you use, tells you which label to put on your resulting pressure value. Don't forget that your temp must be in Kelivin.
   a. P = (0.25 mole * 62.4 mmHg L/mole K * 303 K)/3 L = 1577 mmHg then round off to 1600 mmHg
      OR P = = (0.25 mole * 0.0821 mmHg L/mole K * 303 K)/3 L = 2.1 atm
   b. 0.25 mole * 71 g/1 mole = 17.8 g
   c. Remember that D = M/V so 17.8 g / 3 L = 5.9 g/L

3. Apply the ideal gas law: PV = nRT solving for n. Here you can use the pressure given in atm, but you must choose the 0.0821 atm L/mole K gas constant. Or you can change 4.5 atm to 3420 mm Hg and use the 62.4 mmHg L/mole K gas constant. You must change 500 ml to 0.5 L because both gas constants have units in L, and so the volume must be in liters so it can cancel out. And of course, you must change the temperature to Kelvin.
   a. n = (4.5 atm * 0.5 L)/(0.0821 atm L/mole K * 298 K) = 0.092 mole
   b. 0.092 mole * 32 g/1mole = 2.9 g
   c. Remember that D = M/V so 2.9 g/0.5 L = 5.9 g/L

4. First change mass to moles using the molar mass of CO_2 the apply the ideal gas law, PV = nRT solving for T. Remember that the answer will come out in Kelvin, and you must change to report it in Celsius.
   a. 366 g * 1 mole/44 g = 0.82 mole
   b. T = 4.54 atm * 5 L /(0.82 mole * 0.0821 atm L/mole K) = 337 K which converts to 64°C

5. Fill the values into the ideal gas law, PV = nRT then solve for V. Be sure you convert your temp to Kelvin.
   V = (20 mole * 0.0821 atmL/moleK * 298 K)/22 atm = 22.2 L

6. The density is 2.52 g/ 1 L. Convert 2.52 g to moles using molar mass, and use 1 L in PV = nRT, solve for P.
   a. 2.52 g * 1 mole/64 g = 0.0394 mole
   b. P = (0.0394 moles * 62.4 mmHg L/mole K * 298 K)/1 L = 732 mm Hg which is also equal to 0.9663 atm

7. Take the simple route by dividing the molar mass, 32 g/mole by the molar volume 22.4 L/mole to get 1.42 g/L

1971

$P_{CO} = P_{atm} - P_{H_2O} = (752 - 19.8)\text{ torr} = 732.2\text{ torr}$

$$n = \frac{PV}{RT} = \frac{(732.2\text{torr})(0.242\text{ L})}{(62.4\frac{\text{L\_torr}}{\text{mol\_K}})(295.15\text{K})} = 9.62\times10^{-3}\text{ mol}$$

$$9.62\times10^{-3}\text{mol} \times \frac{2\text{ mol HCOONa}}{2\text{ mol CO}} \times \frac{68.0\text{g}}{1\text{ mol}} = 0.654\text{ g}$$

0.654 g/0.964 g × 100 = 67.9%

1971

(a) $P_{C6H6} = \chi P°_{C6H6} = ({}^{49}/_{50})(75\text{ torr}) = 73.5\text{ torr}$

(b) $P_T = \chi P°_{tol.} + \chi P°_{benz.}$

$= ({}^{3}/_{4})(22\text{ torr}) + ({}^{1}/_{4})(75\text{ torr}) = 35.3\text{ torr.}$

$$\chi_{benz.} = \frac{({}^1\!/_4)(75\text{torr})}{35.25\text{torr}} = 0.532$$

1972

(a)

$$n = \frac{PV}{RT} = \frac{(\frac{740}{760}\text{ atm})(0.249\text{L})}{(0.08205\frac{\text{L atm}}{\text{mol\_K}})(295\text{K})} = 0.0100\text{mol } CO_2$$

$K_2CO_3 + 2\ HCl \rightarrow 2\ KCl + CO_2 + H_2O$

$$\frac{0.0100\text{mol } CO_2}{} \times \frac{1\text{ mol } K_2CO_3}{1\text{ mol } CO_2} \times \frac{138.2\text{g } K_2CO_3}{1\text{ mol } K_2CO_3}$$

$= 1.38\text{ g } K_2CO_3$

$$\frac{1.38\text{ g } K_2CO_3}{5.00\text{ g mix}} \times 100\% = 27.7\%\ K_2CO_3$$

(b) $KOH + HCl \rightarrow K^+ + Cl^- + H_2O$

$$\frac{0.100\text{L HCl}}{} \times \frac{2.0\text{ mol}}{1\text{L}} = 0.200\text{ mol HCl}$$

2(0.0100 mol) = 0.0200 mol HCl reacted with K_2CO_3

1 mol NaOH = 1 mol HCl

$$\frac{0.0866\text{L NaOH}}{} \times \frac{1.5\text{ mol}}{1\text{L}} = 0.130\text{ mol HCl excess}$$

mol HCl reacted = (0.200 - 0.0200 - 0.130)mol = 0.050 mol

$$\frac{0.050\text{mol HCl}}{} \times \frac{1\text{mol KOH}}{1\text{mol HCl}} \times \frac{56.1\text{g KOH}}{1\text{mol KOH}} = 2.81\text{g}$$

$$\frac{2.81\text{ g KOH}}{5.00\text{ g mix}} \times 100\% = 56.2\%\text{ KOH}$$

KCl = (100 - 27.7 - 56.2)% = 16.1% KCl

1973

(a) 6.19 g PCl_5 / 208.22 g/mol = 0.0297 mol PCl_5

$$P = \frac{nRT}{V} = \frac{(0.0297\text{ mol})(0.08205\frac{\text{L atm}}{\text{mol\_K}})(525.15\text{K})}{2.00\text{L}}$$

= 0.640 atm = 487 mm Hg

(b) $P_{PCl3} = P_{Cl2} = X$; $P_{PCl5} = (0.640 - X)$ mm Hg

$P_T = 1.00\text{ atm} = (0.640 - X) + X + X$

$X = 0.360\text{ atm} = P_{PCl3} = P_{Cl2}$

$P_{PCl5} = (0.640 - 0.360)\text{ atm} = 0.290\text{ atm} = 220\text{ mm}$

1976

Useful realtionship is: M= (gRT)/(PV). Significant intermolecular attraction exists at temperatures not far above boiling point.
Therefore, the compressibility of the gas is greater and the value of PV is smaller than predicted.
This would lead to a higher value for the molecular weight than the true value.

1982

a) 2 points

Real molecules exhibit finite volumes, thus excluding some volume from compression.

Real molecules exhibit attractive forces, thus leading to fewer collisions with the walls and a lower pressure.

---

b) 3 points

SO_2 is the least ideal gas.

It has the largest size or volume.

It has the stongest attractive forces (van der Waals forces or dipole-dipole interactions).

---

c) 3 points

High temperature results in high kinetic energies.

This energy overcomes the attractive forces.

Low pressure increases the distance between molecules. (So molecules comprise a small part of volume or attractive forces are small)

1984

Scoring Standards for Question #7 - - - - - - - - - - - - - - (8 points total)

Constant *a* is related to the attractive forces that exist between real molecules.	1 point
Constant *b* is related to the fact that real molecules occupy space or volume.	1 point
H_2S has a larger *a* value	1 point*
because H_2 is a polar molecule and therefore has stronger intermolecular forces.	2 points
H_2S has a larger *b* value because of its additional atom.	1 point**
The constant *a* correlates with the boiling point since it is related to the intermolecular forces which must be overcome in the process of boiling.	1 point* 1 point

* 1 point granted in the absence of an explanation only if the constant is correctly identified somewhere in the discussion.

** Explanation not required for point provided constant is correctly identified somewhere in the discussion.

Question 3. 9 points. Average score 5.18

1986

Three volatile compounds X, Y, and Z each contain element Q. The percent by weight of element Q in each compound was determined. Some of the data obtained are given below.

Compound	Percent by Weight of Element Q	Molecular Weight
X	64.8%	?
Y	73.0%	104.
Z	59.3%	64.0

(a) The vapor density of compound X at 27°C and 750. mm Hg was determined to be 3.53 grams per liter. Calculate the molecular weight of compound X.

(b) Determine the mass of element Q. contained in 1.00 mole of each of the three compounds.

(c) Calculate the most probable value of the atomic weight of element Q.

(d) Compound Z contains carbon, hydrogen, and element Q. When 1.00 gram of compound Z is oxidized and all of the carbon and hydrogen are converted to oxides, 1.37 grams of CO_2 and 0.281 gram of water are produced. Determine the most probable molecular formula of compound Z.

Part a: PV = (grams/m wt) × RT

m wt = (3.53 grams/liter) (0.0821 liter-atm/mole K) × (300 K) (1/[750/760]) = 88.1 grams/mole (3 points)

OR

(3.53 grams/liter) (760/750) (300/273) × (22.4 liters/mole) = 88.1 grams/mole

OR other equivalent solutions

Part b: gram Q/mole X = 0.648 × 88.1 = 57.1

gram Q/mole Y = 0.730 × 104 = 75.9

gram Q/mole Z × 0.593 × 64.0 = 38.0

One correct (1 point)
All correct (1 additional point)

Part c: Masses in (b) must be integral multiples of atomic weight. Largest common denominator is 19. (1 point)

Note: Credit given for incorrect at. wt. if consistent with values in (b).

Part d: 1.37 grams CO_2 (1 mole/44.0 grams CO_2) = 0.0311 mole CO_2 ≡ 0.0311 mole C

0.281 gram H_2O (1 mole/18.0 grams H_2O) = 0.0156 mole H_2O = 0.0312 mole H (2 points)

1.00 gram Z (1 mole/64 grams) = 0.0156 mole Z

Each mole Z contains 2 moles of CH, or 26 grams, which leaves (64 − 26) = 38 grams, corresponding to 2 moles of Element Q. Mol. formula is $C_2H_2Q_2$. (1 point)

Note: Other equivalent solutions received credit.

1990

number of moles of H_2 and the number of moles of O_2 are equal. The total pressure is 1,146 millimeters mercury. (The equilibrium vapor pressure of pure water at 25° C is 24 millimeters mercury.)

The mixture is sparked, and H_2 and O_2 react until one reactant is completely consumed.

(a) Identify the reactant remaining and calculate the number of moles of the reactant remaining.

(b) Calculate the total pressure in the container at the conclusion of the reaction if the final temperature is 90° C. (The equilibrium vapor pressure of water at 90° C is 526 millimeters mercury.)

(c) Calculate the number of moles of water present as vapor in the container at 90° C.

) $2H_2 + O_2 \longrightarrow 2H_2O$

Moles H_2 = moles O_2 initially but 2 moles of H_2 react for every mole of O_2. ∴ O_2 is left (1)

$P_{Tot} = P_{H_2+O_2} + P_{H_2O}$

$1146 = P_{H_2+O_2} + 24$ (1)

$P_{H_2+O_2} = 1122$ mm Hg

$1122\text{ mm}/4 = P_{O_2}$ left (½ of init. O_2 which is ½ total)

$P_{O_2} = 280.5$ mm (1)

$P_1V_1/T_1 = P_2V_2/T_2$ OR $PV = nRT$ (1)

$\frac{(280.5)(.5\ell)}{298°} = \frac{(760\text{ mm})V_2}{273°}$ OR $\frac{(280.5/760)\text{atm}\,(.5\ell)}{.0821\frac{\ell\,\text{atm}}{\text{mol deg}}\ 298\text{deg}} = n$ (1)

$V_2 = 0.169\ell$

$n = 0.169\ell/22.4$

$= 7.55 \times 10^{-3}$ mol OR $n = 7.55 \times 10^{-3}$ mol (1)

) $\frac{P_{O_2}(90°)}{363} = \frac{280.5}{298}$ OR $P = \frac{(7.55\times10^{-3}\text{ mol})(.0821)(363)}{0.5}$

$= .45$ atm

$P_{O_2}(90°) = 342$ mmHg OR $P = 342$ (1)

$P_{tot} = P_{O_2} + P_{H_2O}$

$= 342 + 526$

$= 868$ mm Hg (1)

c) $V_{H_2O}(STP) = \frac{526\text{ mm Hg}\,(.5\ell)\,273}{760\ (363)}$ OR $n = \frac{(526/760)\text{atm}\,(0.50\ell)}{\left(0.0821\frac{\ell\text{-atm}}{\text{mol deg}}\right)(363\text{deg})}$

$V_{H_2O} = 0.260\ell$ (1)

$n_{H_2O} = 0.260\ell/22.4\ell\,\text{mol}^{-1}$ OR $n = 0.0116$ mol

$= .0116$ mol

#9 1993

STANDARDS

(a) Reducing the temperature of a gas reduces the average kinetic energy (or velocity) of the gas molecules. This would reduce the number (or frequency) of collisions of gas molecules with the surface of the balloon (OR decrease the momentum change that occurs when the gas molecules strike the balloon surface.) In order to maintain a constant pressure vs the external pressure, the volume must decrease. 1 pt. 1 pt

[If other 2 pts not awarded]

(b) The molecules of the gas do have volume 1 pt

When they are cooled sufficiently, the forces of attraction that exist between them cause them to liquefy or solidify. 2 pts

(c) The molecules of a gas are in constant motion so the HCl and NH_3 diffuse along the tube. Where they meet, NH_4Cl is formed. Since HCl has a higher molar mass, its velocity (avg) is lower. Therefore it doesn't diffuse as fast as the NH_3. 1 pt 1 pt

(d) The wind is moving molecules of air that are going mostly in one direction. Upon encountering a flag, they transfer some of their energy (momentum) to it and cause it to move (flap!). 1 pt. 1 pt.

1994

CHEMISTRY

STANDARDS

Question 3

(a) $n = \frac{PV}{RT} = \frac{(721)(0.090)}{(62.4)(298)} = 3.49 \times 10^{-3}$ mol H_2

25°C ⟶ 298 K (1 pt.)

745 - 24 = 721 mm Hg (1 pt.)

calculation of moles of H_2 (1 pt.)

(b) $\frac{(23.8)(0.090)}{(62.4)(298)} = 1.15 \times 10^{-4}$ mol H_2O (1 pt.)

$(1.15 \times 10^{-4})(6.03 \times 10^{23}) = 6.92 \times 10^{19}$ molecules H_2O (1 pt.)

(c) The average kinetic energies are equal, so

$(\tfrac{1}{2}mv^2)_{H_2O} = (\tfrac{1}{2}mv^2)_{H_2}$

$\frac{v_{H_2}}{v_{H_2O}} = \sqrt{\frac{MM_{H_2O}}{MM_{H_2}}} = \sqrt{\frac{18}{2}} = 3$ (1 pt.) for formula (1 pt.) for calculation

Note: credit also given for correct use of $v_{rms} = \sqrt{\frac{3RT}{M}}$.

(d) H_2O deviates more from ideal behavior. (1 pt.)

Explanation:

EITHER

i) The volume of the H_2O molecule is larger than that of the H_2 molecule

OR,

ii) The intermolecular forces among H_2O molecules are stronger than those among H_2 molecules (1 pt.)

1995

QUESTION 2

(9 pts.)

a) $C_3H_8 + 5\,O_2 \rightarrow 3\,CO_2 + 4\,H_2O$ *(1 pt.)*

Notes: ignore phases (even when wrong)
multiples are OK
if balanced wrong, parts b and c should be consistent

b) $10.0\text{ g } C_3H_8 \times \frac{1\text{ mol } C_3H_8}{44.1\text{ g } C_3H_8} = 0.227\text{ mol } C_3H_8$ *(1 pt.)*

$0.227\text{ mol } C_3H_8 \times \frac{5\text{ mol } O_2}{1\text{ mol } C_3H_8} = 1.13\text{ mol } O_2$ *(1 pt.)*

$V = \frac{(1.13\text{ mol } O_2)(0.0821\text{ L·atm·mol}^{-1}\text{·K}^{-1})(303\text{ K})}{1.00\text{ atm}} = 28.1\text{ L } O_2$ *(1 pt.)*

$28.1\text{ L } O_2 \times \frac{100\text{ L air}}{21.0\text{ L } O_2} = 134\text{ L air}$ *(1 pt.)*

Note: answer must be consistent with part a

c) $\Delta H^0_{rxn} = \sum \Delta H^0_f(products) - \sum \Delta H^0_f(reactants)$

$-2{,}220.1\text{ kJ} = [4(-285.3\text{ kJ}) + 3(-393.5\text{ kJ})] - [5(0\text{ kJ}) + \Delta H^0_f(C_3H_8)]$ *(1 pt.)*

$-2{,}220.1\text{ kJ} = -1{,}141.2\text{ kJ} - 1{,}180.5\text{ kJ} - \Delta H^0_f(C_3H_8)$

$-2{,}220.1\text{ kJ} = -2{,}321.7\text{ kJ} - \Delta H^0_f(C_3H_8)$

$-101.6\text{ kJ} = \Delta H^0_f(C_3H_8)$ *(1 pt.)*

Notes: answer should be consistent with part a
1 point deducted if negative sign missing from answer
1 point deducted if −2,220.1 kJ substituted for $\Delta H^0_f(C_3H_8)$
no points earned if coefficients are inconsistent and not set equal to ΔH^0

d) $30.0\text{ g } C_3H_8 \times \frac{1\text{ mol } C_3H_8}{44.1\text{ g } C_3H_8} \times \frac{2{,}220.1\text{ kJ}}{1\text{ mol } C_3H_8} = 1.51 \times 10^3\text{ kJ}$ *(1 pt.)*

$1.51 \times 10^3\text{ kJ} = 1.51 \times 10^6\text{ J} = (8{,}000\text{ g})(4.18\text{ J·g}^{-1}\text{·K}^{-1})(\Delta T)$

$45.1\text{ K (or °C)} = \Delta T$ *(1 pt.)*

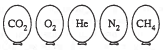
Notes: must correctly substitute into $q = mc\Delta T$ for 1 point
1 point earned if q value wrong but ΔT consistent

1996

QUESTION 5

(8 points)

5.



Represented above are five identical balloons, each filled to the same volume at 25°C and 1.0 atmosphere pressure with the pure gases indicated.

(a) Which balloon contains the greatest mass of gas? Explain.

(b) Compare the average kinetic energies of the gas molecules in the balloons. Explain.

(c) Which balloon contains the gas that would be expected to deviate most from the behavior of an ideal gas? Explain.

(d) Twelve hours after being filled, all the balloons have decreased in size. Predict which balloon will be the smallest. Explain your reasoning.

Scoring Guide

Question 5

(a) CO_2 — *1 point*

because all contain same number of molecules (moles), and CO_2 molecules are the heaviest — *1 point*

(Note: total of 1 point earned if CO_2 not chosen but same number of molecules (moles) is specified)

(b) All are equal — *1 point*

because same temperature ⇒ same average kinetic energy — *1 point*
(Note: just restatement of "same conditions, etc." does not earn second point)

(c) CO_2 — *1 point*

it has the most electrons, hence is the most polarizable
it has the strongest intermolecular (London) forces } *either one* — *1 point*

(Note: also allowable are "polar bonds", "inelastic collisions"; claiming larger size or larger molecular volume does not earn second point)

(d) He — *1 point*

greatest movement through the balloon wall
smallest size
greatest molecular speed
most rapid effusion (Graham's law) } *any one* — *1 point*

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Question 2

10 points

2. A rigid 8.20 L flask contains a mixture of 2.50 moles of H_2, 0.500 mole of O_2, and sufficient Ar so that the partial pressure of Ar in the flask is 2.00 atm. The temperature is 127°C.

(a) Calculate the total pressure in the flask.

$P_{H_2} = \left(\frac{n_{H_2}RT}{V}\right) = \left(\frac{(2.50\text{ mol})(0.0821\frac{\text{L·atm}}{\text{mol·K}})(400K)}{8.20\text{ L}}\right) = 10.0\text{ atm}$	1 point earned for the partial pressure of H_2
$P_{O_2} = \left(\frac{n_{O_2}RT}{V}\right) = \left(\frac{(0.500\text{ mol})(0.0821\frac{\text{L·atm}}{\text{mol·K}})(400\text{ K})}{8.20\text{ L}}\right) = 2.00\text{ atm}$ $P_{Ar} = 2.0$ atm	1 point earned for the partial pressure of O_2
$P_T = P_{H_2} + P_{O_2} + P_{Ar} = 10.0\text{ atm} + 2.0\text{ atm} + 2.0\text{ atm} = 14.0\text{ atm}$	1 point earned for the total pressure

(b) Calculate the mole fraction of H_2 in the flask.

$\text{Mol fraction}_{H_2} = \left(\frac{\text{mol}_{H_2}}{\text{mol}_{H_2} + \text{mol}_{O_2} + \text{mol}_{Ar}}\right)$ $\text{mol}_{H_2} = 2.50$ mol $\text{mol}_{O_2} = 0.500$ mol	
$\text{mol}_{Ar} = \left(\frac{PV}{RT}\right) = \left(\frac{(2.00\text{ atm})(8.20\text{ L})}{(0.0821\frac{\text{L·atm}}{\text{mol·K}})(400\text{ K})}\right) = 0.500\text{ mol Ar}$	1 point earned for mol Ar
$\text{mol}_{H_2} + \text{mol}_{O_2} + \text{mol}_{Ar} = 2.50\text{ mol} + 0.500\text{ mol} + 0.500\text{ mol} = 3.50\text{ mol total}$ $\text{Mol fraction}_{H_2} = \left(\frac{\text{mol}_{H_2}}{\text{mol}_{H_2} + \text{mol}_{O_2} + \text{mol}_{Ar}}\right) = \left(\frac{2.50\text{ mol}}{3.50\text{ mol}}\right) = 0.714$	1 point earned for mol fraction of H_2

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Question 2 (cont'd.)

(c) Calculate the density (in g L^{-1}) of the mixture in the flask

$2.50\text{ mol H}_2\left(\frac{2.016\text{ g H}_2}{1\text{ mol H}_2}\right) = 5.04\text{ g H}_2$ $0.500\text{ mol O}_2\left(\frac{32.0\text{ g O}_2}{1\text{ mol O}_2}\right) = 16.0\text{ g O}_2$ $0.500\text{ mol Ar}\left(\frac{40.0\text{ g Ar}}{1\text{ mol Ar}}\right) = 20.0\text{ g Ar}$	1 point earned for mass of all species
total mass = 5.04 g + 16.0 g + 20.0 g = 41.0 g $\text{density} = \left(\frac{\text{total mass}}{\text{volume}}\right) = \left(\frac{41.0\text{ g}}{8.20\text{ L}}\right) = 5.00\text{ g L}^{-1}$	1 point earned for density

The mixture in the flask is ignited by a spark, and the reaction represented below occurs until one of the reactants is entirely consumed.

$$2\,H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

(d) Give the mole fraction of all species present in the flask at the end of the reaction.

$2\,H_2(g) + O_2(g) \rightarrow 2\,H_2O(g)$ I 2.50 0.500 0 C − 1.00 − 0.500 2(+0.500) E 1.50 0 1.00	1 point earned for 1.00 mol H_2O
total moles after reaction = $\text{mol}_{H_2} + \text{mol}_{H_2O} + \text{mol}_{Ar}$ = 1.50 mol + 1.00 mol + 0.500 mol = 3.00 mol total	1 point earned for total moles
$\text{mol fraction}_{H_2} = \left(\frac{1.50\text{ mol H}_2}{3.00\text{ mol}}\right) = 0.500$ $\text{mol fraction}_{O_2} = \left(\frac{0\text{ mol O}_2}{3.00\text{ mol}}\right) = 0$ (not necessary) $\text{mol fraction}_{Ar} = \left(\frac{0.500\text{ mol Ar}}{3.00\text{ mol}}\right) = 0.167$ $\text{mol fraction}_{H_2O} = \left(\frac{1.00\text{ mol H}_2\text{O}}{3.00\text{ mol}}\right) = 0.333$	1 point earned for any two mol fractions, excluding O_2

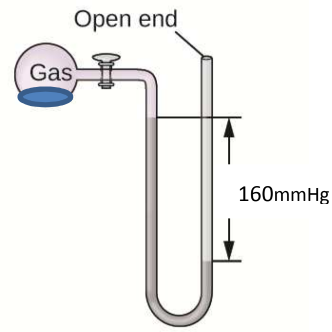
Unit A: Gases

Homework Problems

Directions: The AP exam does NOT allow you to use a calculator; therefore you should abstain from using one to complete this homework. As we move through the unit complete each topic and bubble in your answer on the bubble sheet provided at the end. Bring Coach L any section you complete and he will tell you which questions you missed. On the day of the unit test the bubble sheets will be collected and graded for correctness and a homework grade.

Pressure & the KMT

1. The system above is at equilibrium at 28°C. At this temperature the vapor pressure of water is 28mmHg. The partial pressure of $O_2(g)$ in the system is



a. 160 mm Hg
b. 600 mm Hg
c. 760 mm Hg
d. 920 mmHg

2. Equal numbers of moles of F_2, and Cl_2 are placed in a glass vessel at room temperature. If the vessel has a pinhole-sized leak, which of the following will be true?

a. The partial pressures of both gases will be the same
b. The partial pressure F_2 will be greater than the partial pressure of Cl_2
c. The partial pressure of Cl_2 will be greater than the partial pressure of F_2
d. The volume of the contain will decrease

3. A rigid metal tank contains nitrogen gas. Which of the following applies to the gas in the tank when additional nitrogen is added at a constant temperature?

a. The pressure of the gas increases
b. The volume of the gas increases
c. The average speed of the gas molecules increase
d. The total number of gas molecules remains the same

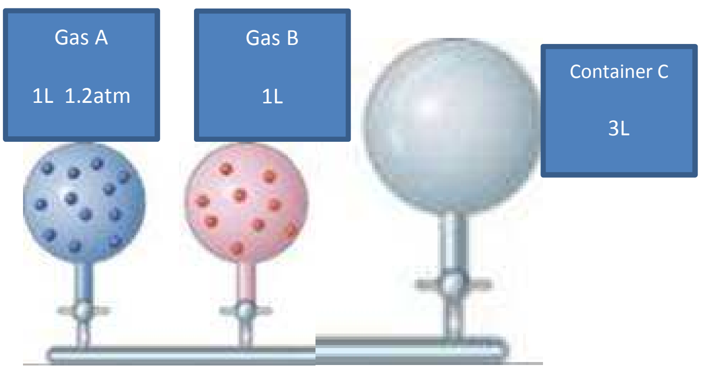
Use the table below to answer the next two questions:

	Gas A	Gas B	Gas C	D
Pressure (atm)	4.0	6.0	2.0	Gas A, B, and C are the same
Temperature (K)	290	290	290	
Molar Mass (g/mol)	44	44	44	

4. Which gas has highest kinetic energy (KE_{avg})?

5. Which gas contains the largest number of particles?

Gas Laws



1. Each one liter container contains a different noble gas at 298K. The pressure of gas A is 2.0atm. The ratio of gas A to B is 2 to 12 to 9. A student then combines all the gases into container C which has a volume of 3.0L. What is the total pressure of container C?

a. 0.7 atm
b. 1.1 atm
c. 2.1 atm
d. 7.0 atm

2. **When a sample of Ne is cooled to half its absolute temperature, which of the following is doubled?**
   a. The density of the gas
   b. The pressure of the gas
   c. The average velocity of the gas atoms
   d. The potential energy of the atoms

3. **A sample of gas occupies 3L at 127°C and has a pressure of 1.25 atm, what pressure would the gas exert at 27°C if the volume was held constant?**
   a. 0.156 atm
   b. 0.313 atm
   c. 0.938 atm
   d. 1.25 atm

4. **A 1.00 L sample of Nitrogen Gas at 27°C and 300mmHg is heated until it occupies a volume of 2.5 L. If the pressure remains unchanged, the final temperature of the gas is**
   a. 120°C
   b. 477°C
   c. 750°C
   d. 950°C

5. **A flask contains 2.5 mol H_2O, 5.0 mol N_2, and 2.5 mol Ar. The total pressure of the gases in the flask is 500 mmHg. What is the partial pressure of the N_2 in the flask?**
   a. 125 mmHg
   b. 250 mmHg
   c. 500 mmHg
   d. 1000 mmHg

The Ideal Gas Law

1. **The pressure, in torr, exerted by 2.1 mol of an ideal gas placed in a 5.00 L container at 15°C is given by which of the following expressions?**
   a. $\frac{(2.1)(62.4)(288)}{5.00} torr$
   b. $\frac{(2.1)(15)}{(62.4)5.00} torr$
   c. $\frac{5.00}{(2.1)(62.4)(288)} torr$
   d. $\frac{(2.1)(0.08206)(288)}{5.00} torr$

2. **Under which of the following conditions of temperature and pressure would 1.0mol of CH_4 exhibit the most ideal properties?**

	Temperature (K)	Pressure (atm)
A	200	0.01
B	200	200
C	1600	0.01
D	1600	200

3. **Which of the following gases deviate the LEAST from ideal behavior?**
   a. SO_2
   b. CH_4
   c. N_2
   d. H_2

4. **A sample of 1.2 grams of an ideal gas at 27°C and 1.0 atm has a volume of 0.5L. Which of the following expressions correctly provides the molar mass of the gas?**
   a. $\frac{[(1.0)(0.5)]}{[(1.2)R(300)]}$
   b. $\frac{[(1.2)R(27)]}{[(1.0)(0.5)]}$
   c. $\frac{[(1.0)R(27)]}{[(1.2)(0.5)]}$
   d. $\frac{[(1.2)R(300)]}{[(1.0)(0.5)]}$

5. **When the actual gas volume is LESS than the volume predicted by the ideal gas law, the explanation lies in the fact that the ideal gas law does NOT include a factor for:**

   a. Molecular attraction
   b. Molecular volume
   c. Molecular mass
   d. Molecular velocity

AP Chemistry Unit A

Name: ______________________________

Pressure

1. (A) (B) (C) (D) (E)
2. (A) (B) (C) (D) (E)
3. (A) (B) (C) (D) (E)
4. (A) (B) (C) (D) (E)
5. (A) (B) (C) (D) (E)

Gas Laws

1. (A) (B) (C) (D) (E)
2. (A) (B) (C) (D) (E)
3. (A) (B) (C) (D) (E)
4. (A) (B) (C) (D) (E)
5. (A) (B) (C) (D) (E)

The Ideal Gas

1. (A) (B) (C) (D) (E)
2. (A) (B) (C) (D) (E)
3. (A) (B) (C) (D) (E)
4. (A) (B) (C) (D) (E)
5. (A) (B) (C) (D) (E)