Gas Diary

Gases are the most abstract of the forms of matter that we will study because they are invisible, often odorless and tasteless, and because we live in the gases of the atmosphere and are nearly unaware of it because it's always all around us. We begin to miss it greatly when we are underwater for a bit too long, or if the air we are in suddenly fills with up with smoke or strange smells.

There are many gases that are atomic (exist as single atoms such as He, Ne, Ar, Kr, Xe, and Rn.) Many others are diatomic (paired, the HONClBrIF twins: H_2, O_2, N_2, Cl_2, Br_2, I_2, and F_2), and some gases are compounds (CO_2, SO_3, ozone O_3, methane CH_4, etc.).

All substances can be phase changed into gases, but for many this happens only at unusually high temperatures called their boiling point (iron will form a gas at 3023 K! Remember that water boils at 373 K at standard pressure. 3023 K is remarkably hot!

Gases have many uses, such as in neon signs. The term "neon sign" is misleading because all the colors of light that are in these sort of tubes are called neon, but neon is an element with specific properties. Neon lights only glow orange in color. Other "neon light" colors are in fact different gases with a silly, nonscientific name. This one at left has neon gas in the letter tubes (orange) and NOT NEON gas in the blue tubes. It might be water vapor, or mercury gas instead. Only neon gas glows orange.

Kinetic Molecular Theory of Gases

We covered this already, but it's worth repeating now because it's so important. It's the Kinetic Molecular Theory that allows us to think about, discuss, and understand gases. It tells us how gases normally act, why they are gases, and what's different about gases than the other 2 phases of matter (solids and liquids).

The kinetic molecular theory of gases states that gases

1. Are made up of small particles such as atoms or molecules.
   They are, this is an easy idea to grasp.
2. These particles will act as if they are small, hard spheres.
   They aren't really, they do have shapes, and are not spheres, but they act as if this is true.
3. They have no attraction for or any repulsion for any other gas particles.
   This is not true either, but the attraction and repulsion they have for one another is small, and unless crazy cold, no real effect on gases.
4. The particles move very fast, and only in straight lines.
   It's very geometric, no spiraling particles, no loops, etc.

5. All collisions are elastic which means that when the gas particles hit each other all of their energy is transferred, none is lost.
   This is not true, but the loss of energy is small, and the addition of energy all the time from the Sun, and the Earth more than makes up for it. Gases do stay gases usually.

6. Collisions result in pressures being exerted.
   The more collisions the higher the pressure. The stronger the collisions, the higher the gas pressure too.

7. Particles are separated by vast distances from each other relative to the size of the gas particles.
   Gases are mostly empty space, and particle size is insignificant. The particles do take up some space, but it's tiny. In theory, the particles act as if they take up no space at all, but that's silly.

Measuring Gases - Pressure

Gases are measured by 4 ways in our class, the gas pressure, the gas volume, the gas temperature, and the number of moles of gas. Each of these three is closely related to the others for gases, we'll see those relationships soon enough.

Pressure is measured in several units that you are aware of.

kPa	kilo-Pascals or kPa is primary. It's standard is 101.3 kPa
atm	Atmospheres is next. Normally, a "whole atmosphere" is pressing on you. That measure is 1.0 atm of gas pressure.
mm Hg	The original barometers which measure pressure measured with Hg. When the mercury in the tube was at 760 mm, that was normal
psi	The English system still used in tires, engines, and basketballs is pounds per square inch. Normal is 14.7 psi
These gas pressures count for air or for any gases. Gas pressure is produced when the particles of gas hit the surface of a container (or your face). The more collisions, the greater the pressure. The stronger the collisions, the greater the pressure.	

Measuring Gases - Volume

Volume is measured in liters of space, or milliliters of space (mL)
Converting from these units means knowing that 1 Liter = 1000 milliliters

Measuring Gases - Volume

Gases will be measured in Kelvin only, because when we use our formulas, a temperature of zero (as in 0°C) or a negative number (such as -4.5°C) will collapse the math.
Kelvin has only one zero, and when it's zero, it really is!

Measuring Gases - Moles

Just what is says, how many moles of gas are present.

The connection between pressure, volume, temperature and the number of moles of a gas comes in a special formula (that you don't have to learn, but it's so cool) called the ideal gas law: PV=nRT

P for pressure, V for volume, n for number of moles, R for the universal gas constant (R = 0.0821 L x atm / (mol x K), and T for temperature in Kelvin.

We'll just imagine how amazing it is to use such an amazing formula, and we'll learn an easier one (with only PV and T).

The Combined Gas Law

States that the pressure, volume and temperature of a gas at initial conditions is equal to the final conditions. It's on your reference table too.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

This shows that the original pressure and volume (any units are allowed here, as long as they are the same units on right right side of the math) multiplied together, then divided by the original temperature (in Kelvin) equals the final conditions for any one sample of gas.

The relationship between **PRESSURE and VOLUME**

If we were to take the pressure and the volume of ANY GAS at ANY CONDITION, and multiply them together, we would get a particular gas constant for that particular gas sample. This relationship is:

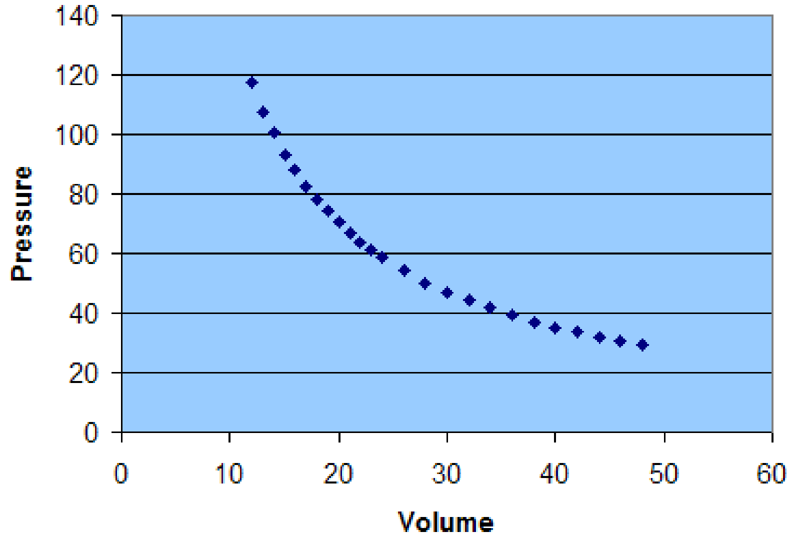
PV = Gas Constant

Any gas constant is the CONSTANT JUST FOR THAT SAMPLE.

But if you change the pressure, you can calculate the new volume since PV = that constant.

If you were to change to volume, you could calculate the new pressure as well.

At any point, the PRESSURE + the VOLUME multiplied together will give the SAME ANSWER.



Since this is true for every point where pressure and volume exist, not only is PV = a constant, any PV point will equal the same constant, so we can also say:

$$P_1V_1 = P_2V_2$$

Pressure and volume are inversely proportional. That means as one goes up, the other must go down.

We showed this relationship with in class demos, such as: The Bell Jar with the Balloon put inside the jar and as the air pressure is decreased outside the balloon, the volume of the balloon increases.

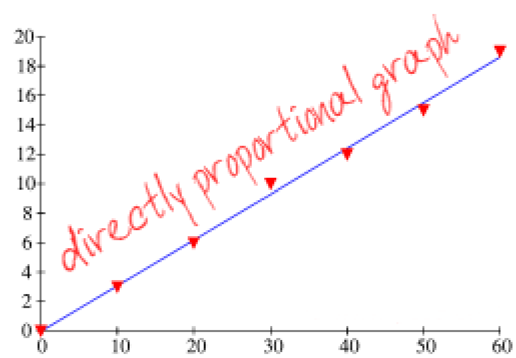
The relationship between **PRESSURE and TEMPERATURE**

When it comes to pressure and temperature, this relationship is directly proportional. As one goes up, so does the other, and the reverse too, as pressure decreases, so does temperature.

Directly Proportional ALWAYS is a STRAIGHT LINE GRAPH

The formula showing this relationship is:

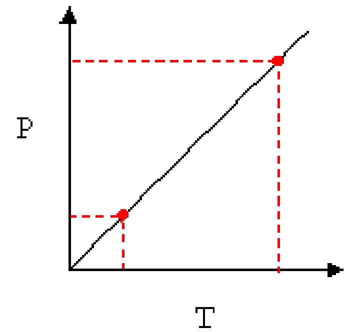
$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$



Pressure and Temperature continued.

As temperature rises, so does pressure, because the increase in temperature is an increase in Kinetic Energy of the gas particles. Since they are now moving faster, they have more collisions PER SECOND, and the collisions they have are stronger because they are moving faster. This results in a greater pressure of gas if the volume is held constant.

We saw this in class demonstrations when the hardboiled egg was placed upon the top of a flask that was filled with steam. When the steam condenses as the flask is cooled rapidly in cold water, the volume of the gas decreases dramatically as the steam turns to liquid water. Since the volume of the gas is almost GONE, the air pressure holding the egg out of the flask is also decreased dramatically, the egg is easily pushed into the flask.



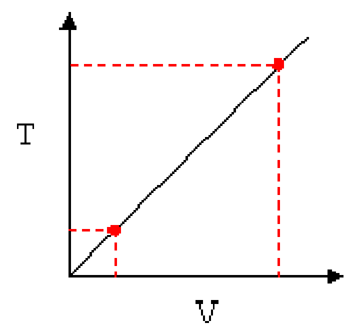
We also put balloons into ice water and hot water, and the balloons shrunk when cold and expanded when hot. This is clearly showing the temperature affects the pressure (and therefore the size of the balloons (balloon volume cannot be held constant with a rubber balloon).

The relationship between **VOLUME and TEMPERATURE**

When it comes to volume and temperature, this relationship is directly proportional. As one goes up, so does the other, and the reverse too, as volume decreases, so does temperature.

Directly Proportional ALWAYS is a
STRAIGHT LINE GRAPH
The formula showing this relationship is:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$



We saw this relationship in class demos with balloons in hot water and in ice water. In the hot water, the balloons kinetic energy and temperature rose, as did the volume. In the ice water, or cold bath, the temperature and kinetic energy decreased, as did the volume.

The Combined Gas Law

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Because of the three different relationships between pressure, volume and temperature of gases, we can combine the three into one formula. The left side, the "1" side, is the initial conditions of a gas. The "2" side is the changed conditions of a gas. In a gas problem you will always be given the initial conditions, and 2 of 3 conditions of the final changed ones. Deciphering the questions is less tricky if you always start with the formula, then fill in the parts you do know, and then solving for the missing part.

Gas law problems can seem cryptic (like those Thermo chemistry problems). STP means standard temperature and pressure. Those are in Table A on the reference table.

Pressure can be done in kPa, or mm of Hg (standard is 760 mm Hg), or even atmospheres (standard is 1.0 atm.). As long as you use the SAME UNITS on both sides of the equal sign for pressure, it does NOT MATTER which unit you do use.

For **Temperature**, standard is 273 Kelvin. You MUST USE KELVIN, for this reason:
If the temperature ever is zero, then it can only mean absolute zero.
Standard temperature in centigrade is 0°C, and since it comes up a lot, it should be clear that putting in zero centigrade will make many of the formulas crash mathematically.

ALSO: temperatures under 0°C are not uncommon (right now if feels like −15°C with the wind chill factor!). A negative number will also create math havoc for you. SO, always use Kelvin, even if the problem uses centigrade (to try to throw you off the right path!)

Volumes can also be in ANY UNIT, as long as the same units are on both sides of the equal sign! Liters, mL, deciliters, cubic centimeters, etc. are all fine units of pressure. They don't matter, as long as you use the same one in the whole problem.

When doing these problems, always use all of your units, and be sure that they cancel each other out. If the units end up strangely, you probably did something UPSIDE down. Check your set up and formula! Use your reference table, never guess!

AVOGADRO'S HYPOTHESIS

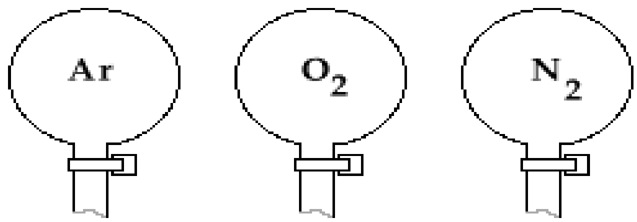
Our main man in chemistry is Amedo Avogadro, and he was a smarty. The ideas he had for gases can be summed up as his hypothesis. It certainly would be a law since it always works, but no one can ever count to his number 6.02×10^{23} number of particles. But if they could, this would become Avogadro's Law of Gases!

Avogadro said:

"Equal volumes of different gases, at the same temperature and pressure will have the same number of particles."

Avogadro's Hypothesis continued.

Look at this diagram, there are three containers of 22.4 Liters in volume, with 3 different gases in them. Each is at STP. Remember that 22.4 Liters of any gas at STP is ONE MOLE of gas. Therefore, each container has ONE MOLE, or 6.03 x 10 particles of gas, the same in each container.



Type of gas	argon	oxygen	nitrogen
volume in L	22.4	22.4	22.4
Pressure in kPa	101.3	101.3	101.3
Temp in Kelvin	273 K	273 K	273 K
# of moles	1.0 mole	1.0 mole	1.0 mole
# of particles	6.02×10^{23}	6.02×10^{23}	6.02×10^{23}

What is really cool is that as easy as it is to see with exactly ONE MOLE of gas, it holds true at any identical conditions of T, P, V.

The "opposite" is true too. Of these four conditions, PRESSURE, VOLUME, TEMPERATURE, and NUMBER OF PARTICLES, if you have any three of them with different gases, the fourth is true as well, for instance:

If you have equal numbers of particles of different gases, and they are at the same pressure and volume, they will also be at the same temperature.

Or………. If you have equal numbers of particles of different gases, and they are also at the same volume and temperature, they will then have the same volume as well.

Pretty amazing!

Vocabulary words to know:

Volume: space taken up by a substance, measured often in LITERS or MILLILITERS
Pressure: force of a gas upon the walls of the container that holds it, often in kPa, or atm.
Temperature: a measure of the kinetic energy of a substance
Directly proportional: when 2 variables react the same way, as one goes up, so does the other (temp and volume are directly proportional, so are pressure and temperature)
Inversely proportional: when 2 variables react in opposite ways, as one goes up, the other goes down. (pressure and volume are inversely proportional)

Ideal gases: are the idea of gases, not real gases. Real gases are real, they exist on the periodic table and in our air. Ideal gases are what we use to explain what gases are. They are a “model” of gases, not actual gases.

Kinetic Molecular Theory and Gases (revisited)

KMT for an ideal gas states that all gas particles…

1. are in random, constant, straight line motion
2. are separated by great distances from each other relative to their sizes
3. the volume of gas particles is considered to be negligible
4. have no attractive forces between them (attraction or repulsion)
5. have collisions that will transfer energy between particles, but the total energy of the gas system will remain constant
6. can be compressed indefinitely, the gas will remain a gas, always

Ideal gases are models, or conceptualizations of gases, they are not real.
A real gas is most like an ideal gas when.

it is at high temperature
it is at low pressure
and when comparing different gases, it has smaller particles.

Helium is the most ideal of the real gases.

Carbon dioxide is “more ideal” than octane, when both are at the same temperature and pressure, because the particles are smaller.