Chapter 12: Gas Diary

Gases are the most abstract of the forms of matter that we will study because they are invisible and often odorless and tasteless. The gases of the atmosphere are always all around us, though we remain unaware of their presence. We recognize the importance of gases when we are underwater for a bit too long, or if the air we're in suddenly fills with smoke or strange smells.

There are many gases that are atomic (exist as single molecules such as He, Ne, Ar, Kr, Xe, and Rn. Many others are diatomic ("7up"), and some are compound or polyatomic (carbon dioxide, sulfur trioxide, ozone, propane, butane, etc.).

Most elements can change into gaseous phase, but usually that happens only at unusually high temperatures. Iron for instance will in fact be 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. Technically the neon signs are only the orange ones, this one at left has other noble gases or compound gases other than neon inside the tubes, which gives us the many colors of "neon" lights.

The lamp at RIGHT is actually a NEON light.

Measuring Gases

There are several ways to measure gases, including PRESSURE, VOLUME, TEMPERATURE, and MOLES. With these measurements we can determine much about a gas, and make changes to up to three of these measures and predict the final measure. The relationships between Pressure and Volume, Pressure and Temperature, and Pressure and Volume are very important in our studies. We need to know the relationships between these variables, how to GRAPH them, and how to recognize their graphs .

Mathematically we will use the **Combined gas law**, found on Table on the back of your reference tables. We can use starting or INITIAL CONDITIONS to predict what the final conditions will be when changes are made to pressure, volume or temperature of a gas.

$$\frac{P_1\,V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Pressure and Volume

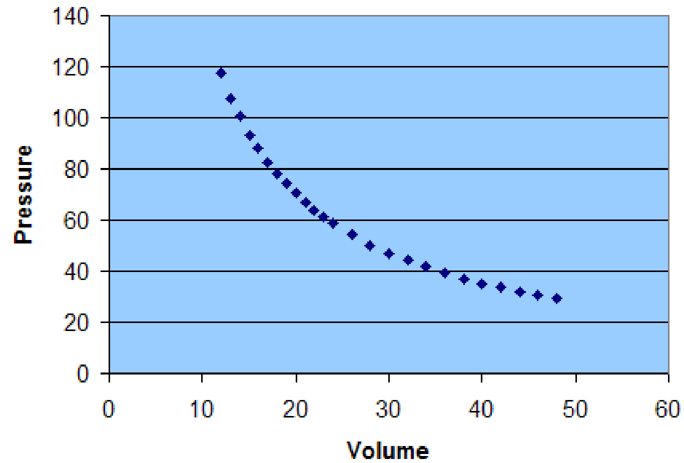
If we were to take the pressure and the volume of ANY GAS at ANY CONDITION, and multiply them together, we would get a particular gas constant for that particular gas sample.

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 and the VOLUME multiplied together will give the SAME ANSWER
(the gas constant for this sample of gas).
Since this is true for every point where pressure and volume exist, not only is
PV = Constant,
but, we can also say:

$$P_1V_1 = P_2V_2$$



This is known as **Boyle's Law**

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

Boyle's Law states the volume of a definite quantity of dry gas is inversely proportional to the pressure, provided the temperature remains constant.

Mathematically Boyle's law can be expressed as $P_1V_1 = P_2V_2$

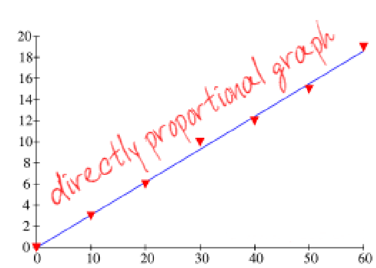
- V_1 is the original volume
- V_2 is the new volume
- P_1 is original pressure
- P_2 is the new pressure

Example: Suppose you have a gas with 45.0 ml of volume and has a pressure of 760.mm. If the pressure is increased to 800mm and the temperature remains constant then according to Boyle's Law the new volume is 42.8 ml.

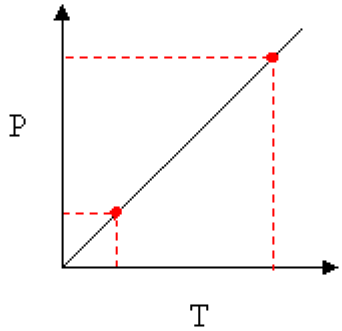
(760mm)(45.0ml) = (800mm)(V_2)
V_2=42.8ml

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



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 (or per minute, hour, etc.) 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.

Gay-Lussac's Law states that the pressure of a sample of gas at constant volume, is directly proportional to its temperature in Kelvin.

Example: Find the final pressure of gas at 150 K, if the pressure of gas is 210 kPa at 120 K.
P1= 210 kPa T1= 120 K
P2= ? T2= 150 K

Formula:

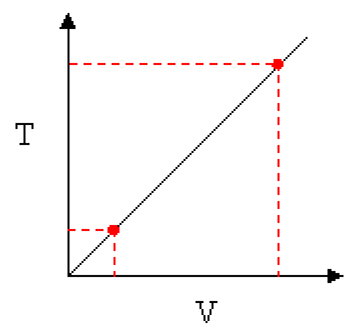
P2= P1xT2/T1
= (210 kPa) (150 K) / 120 K
P2= 262.5 kPa

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:

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



This is **Charles's Law**

Important: Charles's Law only works when the pressure is constant.

Note: Charles's Law is fairly accurate but gases tend to deviate from it at very high and low pressures.

Example
A sample of gas at 15ºC and 1 atm has a volume of 2.50 L. What volume will this gas occupy at 30ºC and 1 atm? You must first convert the temperatures to Kelvin.
T_1= 15C +273=288K and T_2 =30C +273=303K

Then plug and chug:

$$\frac{2.50\ \text{L}}{288\text{K}} = \frac{V_2}{303\ \text{K}} \qquad V_2 = 2.63\ \text{L}$$

The Combined Gas Law

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

The **Combined gas law** is a combination of Boyle's Law and Charles's Law; hence its name the combined gas law. In the combined gas law, the volume of gas is directly proportional to the absolute temperature and inversely proportional to the pressure.
This can be written as PV / T = constant. Since for a given amount of gas there is a constant then we can write $P_1V_1 / T_1 = P_2V_2 / T_2$.

- P_1 is the initial pressure
- V_1 is the initial volume
- T_1 is the initial temperature (in Kelvin)
- P_2 is the final pressure
- V_2 is the final volume
- T_2 is the final temperature (in Kelvin)

This equation is useful if you have the current volume, temperature, and pressure of a gas, and if you have two of the three final values of the gas.

For example if you have 4.0 liters of gas at STP, and you want to know the volume of the gas at 2.0 atm of pressure and 30° C, the equation can be setup as follows:

$(1.0)(4.0) / 273 = (2.0)(V_2) / 303$

$(V_2)(2)(273) = (1)(4)(303)$

$V_2 = 2.2$

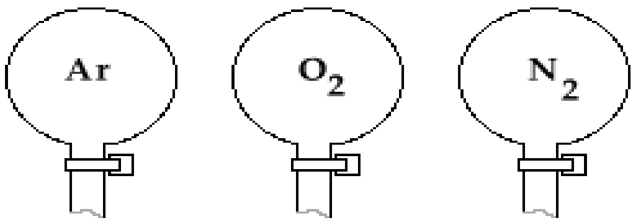
Therefore the new volume is 2.2 liters.

Avogadro's Hypothesis

Amedo Avogadro's ideas 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.

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!

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}

Kinetic Molecular Theory and Gases

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 and the gas will remain a gas, never changing phase

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.

Ideal Gas Law

The ideal gas law is a combination of all the gas laws. The ideal gas law can be expressed as PV = nRT.

- P is the pressure in atm
- V is the volume in liters
- n is the number of moles
- R is a constant
- T is the temperature in Kelvin
-

The constant R is calculated from a theroretical gas called the ideal gas. The most commonly used form of R is .0821 L * atm / (K * mol). This R will allow the units to cancel so the equation will work out.

To find the volume of 2.00 moles gas that is at 1.00 atm of pressure and 235 Kelvin, use the ideal gas law equation.

(1.00 atm)(V) = (2.00 mol)(.0821 L * atm / (K * mol))(235 kelvin)
V = (38.587 L * atm) / (1.00 atm)

Which Units to Use?

STP means standard temperature and pressure. Those are in Table A on the reference table. 101.3 kPa and 273 Kelvin.

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. A negative number will also create math havoc for you. SO, always use Kelvin, even if the problem uses centigrade.

Volume 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 volume. They don't matter, as long as you use the same one in the whole problem.

When doing gas law 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!

Additional resources:

http://www.kentchemistry.com/moviesfiles/Units/GasLawsMovies.htm

http://www.brightstorm.com/science/chemistry/kinetic-molecular-theory/combined-gas-law/