Gas BASICS

Gases are the most interesting phase because mostly they are invisible and you have to use your measuring and wits to grasp how perfect they are. You already know that gases follow the Kinetic Molecular Theory, which is covered again just below. Some of the theory is perfectly true but sometimes the theory "fibs" to help you grasp how gases are gases, and how they stay gases.

Nearly all substances can be phase changed into gases, but for many this happens only at very high temperatures passed their boiling point. H_2O boils into a gas at 373 K at standard pressure. Iron gas forms at 3023 K, and that is remarkably hot! For an example that doesn't, think about wood, or certain plastics. They might catch fire, or decompose before they can change into the gas phase.

There's also the idea of "ideal gases" which are the theoretical perfection of gases, but they are not real. They are like super heroes. We know that super heroes are fake - but we accept them because it makes for a good movie, or in this case, makes it easier to understand gases in general.

Ideal gases: are perfect, make believe, super hero gases. They are the idea of gases, but there are no examples of ideal gases, none actually exist. Real gases are real, they exist on the periodic table and in the air. Ideal gases are used to help explain what gases are. They are the gases of the KMT (below). They are a "model" of gases, not actual gases.

Ideal gases are models, or conceptualizations of gases. Real gases can turn into liquids, because they are real. Ideal gases are perfect gases and can't ever become liquid or solid.

A real gas acts most ideally when it is at high temperature, and it is at low pressure. This is because at high temperatures, any particle collisions are strong enough that the gas particles bounce off of each other rather than stick to form a liquid. At low pressure they collide less frequently, and have a much less chance of forming into a liquid. The biggest flaw a gas can have is to become a liquid! That means a real gas most closely follows the Kinetic Molecular Theory when it's at high temp & low pressure.

When comparing two real gases at the same conditions, the one with the smaller particles is more ideal.

Helium is the most ideal of the real gases, because they are the smallest particles

Carbon dioxide is "more ideal" than octane, when both are at the same temperature and pressure, because the CO_2 particles are smaller than the C_8H_{18} particles.

There are some gases that are atomic— they exist as single atoms, which are the noble gases.
Other gases are diatomic, or paired, like the HONClBrIF twins: H_2, O_2, N_2, Cl_2, Br_2, I_2, and F_2 . Bromine is a liquid at room temperature, and iodine is a solid, but both become gases are relatively low temperatures.

Some gases are compounds, like CO_2, SO_3, methane CH_4, etc.

And there are a few odd gases that we'll learn about like: ozone O_3.

What's the formula for laughing gas? (He, He, He!)

The 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 (KMT) of gases states that gases...

1. Gases are made up of small particles such as atoms or molecules.
   The volume of gas particles is considered to be negligible.
2. Gas 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. Gas particles have no attraction for or any repulsion for any other gas particles.
   This is not quite true either, but the attraction and repulsion they have for one another is small, and unless the temperature is crazy cold, this attraction has almost no real effect on gases.
4. Gas particles move fast, and only in straight lines. They are in random, constant, straight line motion
   The particles of a gas cannot spiral, make loops, or hover in place either.
5. All particle collisions are elastic; when the gas particles hit each other all of their energy is transferred, none is lost. In theory, when the particles have collisions that will transfer energy between particles, and the total energy of the gas system will remain constant.
   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 between particles result in pressures being exerted.
   The more collisions the higher the pressure. The stronger the collisions, the higher the gas pressure too.
7. Gas particles are separated by vast distances from each other relative to the size of the gas particles. In theory, they can be compressed indefinitely, the gas will remain a gas.
   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, like points in mathematics, but that's silly.

Measuring Gases

We will measure gases 4 ways in our class, the gas pressure, the gas volume, the gas temperature, and the number of moles of gas.

1. Volume: Volume is measured in liters of space, or milliliters of space (mL) Converting from these units means knowing that 1 Liter = 1000 milliliters
2. Temperature: 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.
   Zero Kelvin means absolute zero.
3. Number of Moles: Just what is says, how many moles of gas are present.
4. Pressure: Pressure is measured in 4 different units that you know. All of these have unlimited SF.
   101.3 kilo-Pascals (kPa) has been decided to be standard pressure in metric units.
   This is equal to pressure in other units, such as:
   1 atmosphere (atm)
   760 millimeters of mercury (mm Hg) measured in an old style barometer, or
   14.7 pounds per square inch (psi).

The relationship between Pressure, Volume and Temperature of gases.

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

Gases three main measures, pressure and volume and temperature are all in different relationships with each other at the same time. We will explore the three of these now.

The combined gas law, at right, connects all three of these variables. The starting conditions of pressure multiplied by the volume, which is then divided by the Kelvin Temperature is equal to the new conditions of the gas, pressure X volume, then divided by the Kelvin Temperature.

We always use Kelvin because having a negative number, or a zero as a denominator will wreck our math. Units for pressure or volume can be which ever you choose, as long as they are the same on both sides of the equal sign.

Gas math problems have a lot of words, there are six variables and you need to be told five of them. Sometimes the letters STP are used for two at once, standard temperature and standard pressure. They are in your reference tables, any units will work, but read the problem, if you start or end at STP, but there are psi or have mm Hg on the other side, use the same units.

Example problem one…

Your balloon holds 15.6 liters of helium gas at STP, when it rises into the night air, the temperature drops down to just 255 Kelvin, and the volume shrinks to 12. 3 Liters. What is the new pressure? (note, you are given five of six variables, you solve for the last one, make sure you put the numbers and units in the right place, and then it's just a simple cross multiply and divide to solve for the missing variable.

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

$$\frac{(1\text{ atm})(15.6\text{ L})}{273\text{ Kelvin}} = \frac{(P_2)(12.3\text{ L})}{255\text{ Kelvin}}$$

$$(1\text{ atm})(15.6\text{ L})(255\text{ K}) = (P_2)(12.3\text{ L})(273\text{ K})$$

$$\frac{(1\text{ atm})(15.6\text{ L})(255\text{ K})}{(12.3\text{ L})(273\text{ K})} = P_2$$

$$1.18\text{ atm} = P_2$$

That is as hard as these problems might get, which is not bad. What is important to see is what might happen if the problem says, for example, pressure temperature remains constant. Or one of the other variables remains constant.

You can always choose the same number with unit to use on both sides of the = sign, these will cancel each other out, or you can cancel part of the combined gas law formula out, which is what you will see on the next page. Either way, the same answers come out.

Always use the same units on both sides of the = sign, and always use Kelvin, never centigrade, or worse, the "F" word (Fahrenheit!).

Mathematically we can look at this formula in parts as well. Algebra allows us to have, for example, constant temperature. That changes the math to just $P_1V_1 = P_2V_2$ (the temperature cancels out). .

The pressure and the volume of a gas are in what is called an INVERSELY PROPORTIONAL relationship. As one variable increases, the other decreases.

A problem might read like this… A gas held at constant temperature is 35.5 Liters and at 1.45 atm. If the pressure is increased to 2.78 atm what is the new volume?

Cancelling out the temperature (it's constant), we can use this formula: $P_1V_1 = P_2V_2$

$P_1V_1 = P_2V_2$ becomes $(1.45 \text{ atm})(33.5 \text{ L}) = (2.78 \text{ atm})(V_2)$ solve for V_2. → → 17.5 Liters = V_2.

The math would give us the exact same answer if we used the "whole" combined gas law and inserted (say) 305 Kelvin for T_1 and also for T_2, since temperature is constant. Either way, the answer is 17.5 Liters.

If Pressure were constant, we could omit P_1 and P_2 from the formula in the same way.

If Volume were constant, we could omit the V_1 and the V_2 from the formula as well.

Or, you could use (say) 1 atm as P_1 and P_2, or 22.4 Liters for V_1 and V_2, the math the same works either way.

Separating the parts of the combined gas law into it's "component parts", there are three gas laws inside of it. You do not have to know them by name, but they are real gas laws that have been mathematically combined into just one big equation.

$$\boxed{P_1V_1 = P_2V_2}$$

is called Boyle's Law. (that's a name you do not have to remember)

This law shows that Pressure and volume are inversely proportional. That means that as one variable increases, the other must decrease proportionally to keep the equality.

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

This is the Gay-Lussac Law of Gases (another name you don't need to remember)

This law established that the pressure and temperature of a sample of gas are DIRECTLY PROPORTIONAL. As one variable increases, so does the other. Or, as one variable decreases, so does the other.

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

This is known as Charles' Law (the third name that is not necessary to memorize)

This law established that the volume and temperature of a sample of gas are DIRECTLY PROPORTIONAL. As one variable increases, so does the other. Or, as one variable decreases, so does the other.

All gas problems can be solved with the combined gas law formula. But, if one variable is a constant, you can omit it prior to the math, or insert the same value on both sides of the equal sign. Remember, you must use Kelvin temperature only, that prevents you from getting a negative number or a zero as a denominator.

Example problems….

1. At constant pressure, a sample of gas of 22.4 Liters and standard temperature is warmed up to 365 Kelvin. What is the new volume of this gas?

You could decide to use 1 atm pressure on both sides of the equal sign, or use the smaller formula that omits the pressure before the math.

$$\boxed{\frac{V_1}{T_1} = \frac{V_2}{T_2}} \rightarrow \frac{22.4\ \text{L}}{273\ \text{K}} = \frac{V_2}{365\ \text{K}} \rightarrow (V_2)(273\ \text{K}) = (22.4\ \text{L})(365\ \text{K})$$

$$V_2 = 29.9\ \text{Liters}$$

2. At constant volume, a gas sample at standard pressure and 303 Kelvin is cooled to just 245 Kelvin. What is the new pressure on this gas?

$$\boxed{\frac{P_1}{T_1} = \frac{P_2}{T_2}} \rightarrow \frac{101.3\ \text{kPa}}{303\ \text{K}} = \frac{P_2}{245\ \text{K}} \rightarrow (P_2)(303\ \text{K}) = (101.3\ \text{kPa L})(245\ \text{K})$$

$$P_2 = 81.9\ \text{kPa}$$

All combined gas law problems can be solved with the combined gas law, or one of the smaller laws that combine together to form it. The math works out exactly the same either way.

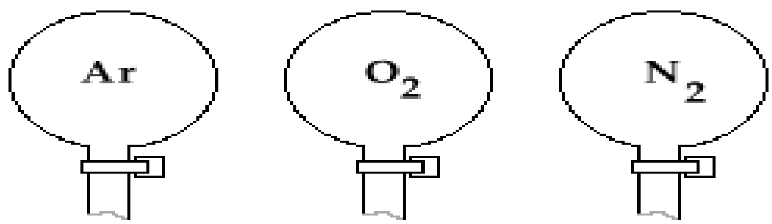
Avogadro's Hypothesis

Amedo Avogadro has that famous number named after him (6.02 x 10^{23} particles per mole). He studied gases and came up with one of the best "one liners" in chemistry history, called Avogadro's Hypothesis. It would be a law but no one can count to his number. Here's what it is, and here's what to memorize;

> **"Equal volumes of different gases at the same temperature & pressure have the same number of particles, and the same number of moles."**

Go slowly through the diagram and figure this out.

In each of these 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 6.02 x 10^{23} 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 x 10^{23}	6.02 x 10^{23}	6.02 x 10^{23}

Memorizing this will save you from doing tedious math problems. You can memorize it, or grunt out all of the math to figure out questions like this

Which sample has the same number of molecules as 3.69 Liters of carbon dioxide at 125 kPa and 305 Kelvin?

A. 1.23 Liters of N_2 at 125 kPa and 305 Kelvin
B. 3.69 Liters of CH_4 at 125 kPa and 305 Kelvin
C. 7. 38 Liters of CO at 125 kPa and 305 Kelvin
D. 1.00 liters of C_8H_{18} at 125 kPa and 305 Kelvin

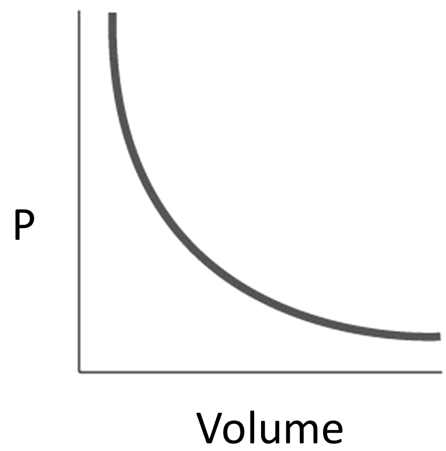
The answer is B, equal volumes of DIFFERENT GASES at the same temp and pressure have the same number of particles (or moles).
The type of gas does not matter, Avogadro does.

The Graphs showing the relationships of Pressure—Volume—Temperature of Gases.

You will have to recognize and draw graphs showing the inversely proportional relationship of P and V.

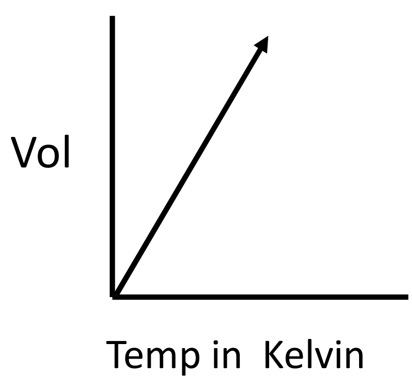
Also you will have to recognize and draw graphs showing the directly proportional relationships between P and T, and between V and T.

There are only 2 different shaped graphs to show these three relationships, that is because one shows inversely proportional, and the other directly proportional. Watch axis labels and you're good.



For Pressure + Volume, as one variable increases, the other decreases.

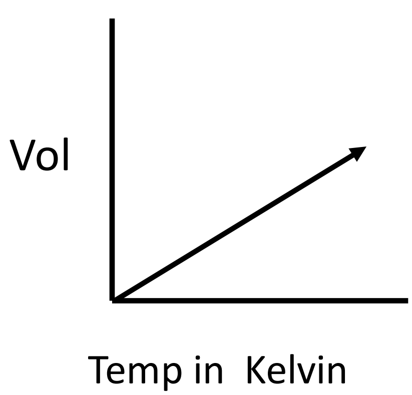
Or, as one variable decreases, the other increases.



For Pressure + Temperature, as one variable increases, the other increases as well.

Or, as one variable decreases, the other does the same.

Temperature MUST be in KELVIN.



For Volume + Temperature, as one variable increases, the other increases as well.

Or, as one variable decreases, the other does the same.
Temperature MUST be in KELVIN.