Gas BASICS

Gases are the most fun phases 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 3023K, and that is remarkably hot!

We also will introduce 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 they 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. Here goes.

Ideal gases: are perfect, make believe, super hero gases. They are the idea of gases, but they aren't 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 the gases of the KMT (below). They are a "model" of gases, not actual gases.

Ideal gases are models, or conceptualizations of gases, they are not real. They are always gases, no matter what conditions you put them under. Real gases can turn into liquids, because they are real. Ideal gases are perfect (unreal!).

A real gas acts like an ideal gas when it is at high temperature, and it is at low pressure. This is because at high temperatures the particles, when they do collide, are colliding so strongly that they bounce off of each other rather than stick (form liquid). At low pressure they collide less frequently. The biggest flaw a gas can have is to allow itself to become liquid! That means a real gas most closely follows the KMT 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 many gases that are atomic (exist as single atoms such as the noble gases.)

Many others are diatomic (paired, like some of the HONClBrIF twins: H_2, O_2, N_2, Cl_2, and F_2)

and some gases are compounds (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 of gases states that gases...

1. Are made up of small particles such as atoms or molecules. The volume of gas particles is considered to be negligible.

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 fast, and only in straight lines. They are in random, constant, straight line motion
   There is 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. In theory, when the particles have collisions that will transfer energy between particles, but 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 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. 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, 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. <u>Volume</u>: Volume is measured in liters of space, or milliliters of space (mL) Converting from these units means knowing that 1 Liter = 1000 milliliters

2. <u>Temperature</u>: 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. <u>Number of Moles</u>: Just what is says, how many moles of gas are present.

4. <u>Pressure</u>: Pressure is measured in 4 different units that you know. All of these have unlimited SF.

   kPa or kilo-Pascals is the primary metric unit. It's standard is 101.3 kPa

   atm or atmospheres is next. Normally, a "whole atmosphere" is pressing on you.

   mm Hg or millimeters of mercury. The original barometric measure. 760 mm of Hg is normal

   psi or pounds per square inch. The American system still uses this. Normal is 14.7 psi

The relationship between Pressure, Volume and Temperature of gases.

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.

Pressure and Volume of gases. Imagine a balloon that you blow up, it is of certain firmness (pressure) and size (volume). You put an amount of gas into this balloon that is measurable and constant. Your air is in the balloon, and it's made of matter, which cannot be created or destroyed.
If you sit on this balloon gently, the balloon gets smaller, and the pressure goes up. (you knew that) Since it is still the exact same gas, as the volume shrinks, the pressure must increase. If you could glue your hands onto the balloon, and pull it to a bigger size, you can understand as the balloon volume increases, the pressure would decrease (more room = less collisions = lower pressure).

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

Since the balloon holds one sample of gas, your sample, when you multiply the pressure X the volume, it equals the gas sample's GAS CONSTANT. Every sample of gas has a different gas constant. If your balloon is 2.50 liters in size, and has pressure of 1.50 atm, this balloon has a gas constant of…

$$1.50 \text{ atm} \times 2.50 \text{ L} = 3.75 \text{ L}\cdot\text{atm}$$ which is a sort of weird unit (we love those!).

If we were to change this to algebra, every gas follows the equation of: P X V = Constant

If this balloon were to be sat on, and the volume was squished to 1.25 Liters, what is the new pressure of this gas sample? This gas always has the same constant, 3.75 L·atm, so, to do this, we set up the math this way:

$$P \times 1.25 \text{ L} = 3.75 \text{ L}\cdot\text{atm} \qquad P = 3.75 \text{ L}\cdot \text{ atm / by } 1.25 \text{ L} = 3.00 \text{ atm}$$

As the volume decreased, the pressure increased, because P and V are INVERSELY PROPORTIONAL.

Since this will always be true, we can change the simple P X V = a constant to…

Boyle's Law (a name that is important, but you don't have to remember) $P_1V_1 = P_2V_2$

Which shows that the...

original pressure X original volume of any gas will be equal to a new pressure X it's new volume
or, PRESSURE AND VOLUME ARE INVERSELY PROPORTIONAL.

We will see soon that this law, Boyle's Law, is PART of the COMBINED GAS LAW. The Combined Gas Law will solve ALL of our gas problems in high school chemistry. There are two more laws to review, then we can COMBINE them into one law (which is on the reference table S.)

The Gay-Lussac Law of Gases (another name you don't need to remember) states 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.

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

Imagine your balloon again, the one you blew up on the last page. It's 2.50 Liters and firm (1.25 atm).
If you were to heat it up, now the same amount of gas particles move faster, and so they collide more, and they release more heat (the balloon temperature goes up). If the balloon is put into the fridge, and gets colder, the particles are moving slower and collide with less energy, so the pressure decreases.

If pressure increases, so does the temperature. If pressure decreases so does the temperature. ALSO…
If temperature increases, so does the pressure. If temperature decreases, so does the pressure.

Note this now:
Pressure & Temperature are DIRECTLY PROPORTIONAL so they are in a division relationship.

Pressure & Volume are INVERSELY PROPORTIONAL, so they are in a multiplication relationship.
($P_1V_1 = P_2V_2$)

The final "law" is Charles' Law of Gases (another name you don't have to remember!) states that…

The volume and the temperature of a gas are DIRECTLY PROPORTIONAL.
As one variable increases, so does the other. Or, as one variable decreases, so does the other.

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

Imagine your balloon again, the one you blew up twice already! It's 2.50 Liters and firm (1.25 atm).
If you were to squish it by sitting on it, now the same amount of gas particles have less room, and so the balloon would get smaller and there would be more collisions, increasing the temperature. And if you glued your hand to the balloon, and stretched it out, the volume would increase, and there would be less collisions, making the temperature decrease.

If temperature increases, the volume increases too. If temperature decreases, so does the volume. ALSO…
If volume increased, it was due to higher temperature. If the volume decreased, that would be due to lower temperature of the gas.

The Combined Gas Law Combines these 3 Laws into ONE.

It 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.
Mathematically, parts of this equation are inversely proportional (P and V) and parts are directly proportional (P and T, and V and T) all at the same time.

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

P_1 is original pressure, P_2 is new pressure. V_1 is original volume, V_2 is new volume. Any units are fine for these two variables. T_1 and T_2 are the original and new temperatures, but these MUST be in KELVIN.

This formula can handle ANY gas problems in our class. Any one, and all of them.

The reason for the KELVIN only rule is this: gases exist at temperatures called zero centigrade or even at negative centigrade temperatures. When a denominator in a formula is a zero, or negative number, the math falls apart. Kelvin has only the one "true" zero. Kelvin has NO negative numbers. Kelvin is a metric unit. (centigrade is too, but the zero and negative business is too messy). Fahrenheit is not metric and just never good in our class.

All gas math problems work in this formula. Certain things come up though. For instance, if the problem has constant temperature, you can pick ANY temperature (as long at it is Kelvin and not zero) for both T_1 and T_2. Or you can cancel temperature out.

Constant pressure gives us just Charles' Law. Constant volume gives us just the Gay-Lussac Law.
Cancel them out, or you can make up any pressure or volume, and keep them the same on both sides.

THINK… A real gas problem might read like this… A weather balloon is filled to 39.0 liters at 2.75 atm and is at 285 Kelvin. When it rises up into the air to1000 feet the temperature drops to just 265 Kelvin and volume decreased to 34.5 liters, what is the new pressure inside the balloon?

Put each variable in the right place, and solve for P_2. Go slowly, cancel out the units, watch SF.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \longrightarrow \frac{(2.75\text{ atm})(39.0\text{ L})}{285\text{ K}} = \frac{(P_2)(34.5\text{ L})}{265\text{ K}}$$

If your problem reads like this:
At constant temperature, the pressure of 55.5 Liters of neon is 1.37 atm. If the pressure increases to 1.88 atm, what is the new volume? (here the temperature can be either constant, or cancelled out before you work)

Either of these set ups leads to the correct answer…

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \longrightarrow \frac{(1.37\text{ atm})(55.5\text{ L})}{273\text{ K}} = \frac{(1.88\text{ atm})(V_2)}{273\text{ K}}$$

Both of these equations lead to the same correct answer,
40.4 L

$$\frac{P_1V_1}{\cancel{T_1}} = \frac{P_2V_2}{\cancel{T_2}} \longrightarrow (1.37\text{ atm})(55.5\text{ L}) = (1.88\text{ atm})(V_2)$$

You can ALWAYS use the combined gas law, or cancel part out before the math. Either way it's okay!

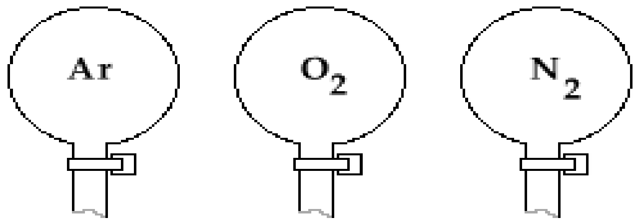
Avogadro's Hypothesis

Amedo Avogadro has that famous number named after him (6.02×10^{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 6.02×10^{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	6.02×10^{23}	6.02×10^{23}	6.02×10^{23}

Look, think, repeat after me:

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

Memorizing this will save you from doing tedious math problems. You can memorize it, or grunt out 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?

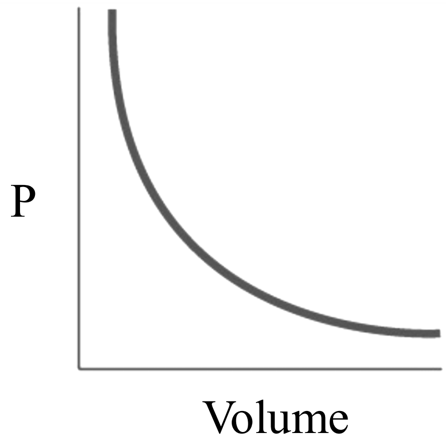
A. 1.23 Liters of N_2 at 125 kPa and 305 Kelvin
B. 3.69 Liters of CH_4 at 125 kPa and 305 Kelvin
C. 7. 38 Liters of CO at 125 kPa and 305 Kelvin
D. 1.00 liters of C_8H_{18} at 125 kPa and 305 Kelvin

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

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

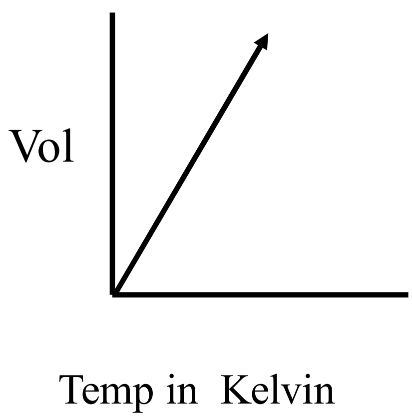
You will have to recognize and draw graphs showing the inverserly proportional relationship of P and V. Also you will have to recognize and draw graphs showing the directly proportional relationships between P and T, and between V and T.

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.

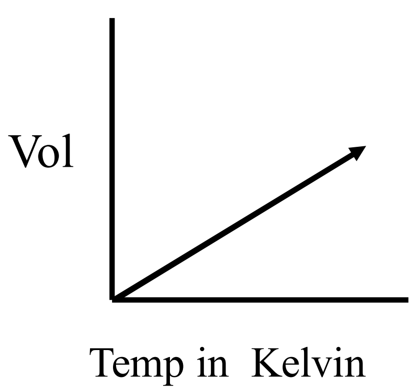
Or, as one decreases, the other increases.



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

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

Temperature MUST be in KELVIN.