# Partial Pressures & Vapor Pressure: Crash Course Chemistry #15

Crash Course: Chemistry
https://youtube.com/watch?v=JbqtqCunYzA
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**=====Introduction (00:00)=====**

Imagine, if you will, a state dinner at the White House. These are not small intimate affairs. 300 or so adults mingle in close quarters yet somehow it's a quiet gathering where people behave gently, move around slowly, and almost never run into a wall. However, as a group they do take up a lot of space, they give off a lot of heat, and lots of things get moved around. In these ways the room is changed by their very presence.

Now picture a five-year old's birthday party; a very different situation. You would not call that quiet or formal. But that is not enough. Imagine the most popular five-year old in the world. His birthday party; 300 kindergartners in one room for several hours, shoving each other, running like crazy, and definitely banging into walls. Their bodies are smaller but because of their faster motion they collectively take up just as much of the room and give off just as much heat as the slow moving adults do. And things get shoved around just as much.

Now imagine we have a surprise for those 300 adults at the state dinner all very formal and tucked in. There are 300 five-year olds in the next room, bouncing off the walls and suddenly they drop the wall next to them. First of all those kids are going to run all around the room, squeezing in between the more sedate adults, that will cause the adults to move apart a little, maybe even bumping in a wall now and again. But because the adults are moving more slowly, they'll get in the way of the kids running and slow them down some, and that might even cause the kids to run into the walls a little less often.

The most important thing here is that the overall bouncing against walls, will be equal to the bouncing of the adults plus the bouncing of the kids when they were in separate rooms.

And you will not be shocked to find that we are currently in and amongst an elaborate analogy to do with gases.

*Intro Music*

**=====Theory of the Atom (01:48)=====**

It shouldn't surprise you to learn that this particular party was started by John Dalton, the English teacher, the science teacher from England. He wasn't an English teacher, he was English, who in 1803 was the first person to use real science to figure out what atoms are and how they behave. His theory included a few misconceptions, but it contained enough facts to be super useful.

Daltons's atomic theory began when he expanded French chemist's Joseph Louis Proust's Law of Definite Proportions, to develop the Law of Multiple Proportions, which says elements combine in simple, whole number ratios of their masses. But here's the cool part. That exact same research also led to an important gas law.

**=====Adding Up The Pressures (02:34)=====**

Dalton studied gases basically by mixing them together. He was measuring how much of one element would react with a given amount of another. But as he mixed those gases he noticed something else. The total internal pressure of his container was always equal to the sum of the individual pressures of the gases that he had added. Upon recognizing that this happened every single time no matter what gasses were used or what amounts were added, Dalton stated his discovery as the Law of Partial Pressures. As long as the gases don't react chemically, the total pressure exerted by a mixture of gases is equal to the sum of the pressures that the individual gases would exert if they were alone.

Here's an example; scuba tanks often contain a mixture of oxygen for breathing, obviously, and helium which helps prevent decompression sickness because it's released from the blood more readily than nitrogen in air. That allows divers to return to the surface more quickly with less risk of gas bubbles forming in their blood.

We're going to use some approximations here so don't yell at me if you know exactly how big this tank is. But we're gonna say that it's ten liters, and we'll to say that it contains 4 moles of helium and 1.1 moles of oxygen gas. The temperature in this room is about 22 degrees Celsius or 295 Kelvin. So what's the total pressure inside the tank?

In order to solve that we need to know the pressure exerted by each gas individually and we can find that with the Ideal Gas Law. Starting with the helium, we don't know the pressure so we'll solve for P. The volume is 10 liters, and we have 4 moles of helium. R is always the same, and the temperature is 295 K. The calculations should show that the helium would have a pressure of 980 kilopascals, if it were alone.

Now for the oxygen, all the numbers are the same except for the moles which is 1.1 for the oxygen. According to the calculations, the oxygen alone would have a pressure of 270 kilopascals

The Law of Partial Pressures says that the total pressure is equal to the sum of the individual pressures, so in this case 980 plus 270 equals 1250 kilopascals, or 1.25 megapascals. Eaaaaasy peasy.

This additive property of pressures is closely related to the fact that mixing gases combines their particles, thus increasing the total moles of gas present.

The ratio of moles of the individual gases in a mixture to the total number of moles is called the mole fraction. And of course, mole fraction gets to have its own little esoteric symbol to represent it, the lower case Greek letter, chi. Chi sub 1, the mole fraction of an individual gas equals n sub 1, the number of moles of that gas, divided by n sub total, the total number of moles in the mixture.

And because the total number of moles is the sum of the moles of all the gases, we can also say that chi sub 1 equals the moles of any one gas divided by the sum of the moles of all the individual gases.

You see how that last formula looks a lot like the one for partial pressures? That's because they're basically the same thing. Through the ideal gas law, the number of moles is directly related to the pressure of gas it exerts, as long as the volume and temperature remain constant.

So you may have figured this out already rather than having to calculate individual pressures first every time as we did with the scuba tank or calculate the individual moles in the opposite situation, we can often calculate what we need directly from what we already know. Let's give it a try.

The air that we breathe is about 21% oxygen or 21 parts oxygen, and 100 parts air. What is the partial pressure of $O_2$ in air at a total atmospheric pressure of 97.8 kPa? We can substitute moles for parts and use the mole fraction formula to solve this problem. Plug in 21 moles of oxygen for the individual gas and 100 moles for air for the total amount, and then put in the total atmospheric pressure, 97.8 kPa and do the equation. With correct rounding, that will give you a partial pressure of 21 kilopascals for oxygen.

Sometimes however, gases mix together in ways that aren't so predictable. For instance, one way to collect a gas is by bubbling it through a column of water to trap the gas at the end of the column. This is called collecting a gas over water.

The only problem is liquid water constantly is giving off a small amount of water vapor. That's why if you leave a glass on a desk for long enough, there will eventually be no water in it. There's always going to be a few molecules with enough kinetic energy to escape the liquid. The amount of vapor that's released depends on the temperature of the water. The more heat energy it has, the more vapor we get. Like all gases, the water molecules move around a lot, sometimes bump into the sides of the container, thus creating pressure. This is called the water's vapor pressure. The water molecules mix with the gas that is being collected and when that happens, as good ol' John Dalton taught us, total pressure in the column equals the pressure of the collected gas plus the vapor pressure of the water.

So to know how much gas we really collected, we have to subtract the vapor pressure of the water from the total pressure. This gives us the pressure exerted from the collected gas and from that, we can calculate the moles of gas present.

=====Mixing Vinegar & Baking Soda (07:15)=====

Here's how it's done. Here I have a tub of water and a graduated cylinder that I have put into the water and then filled with water, so now you can see, it's got no gas in it, just water. And we are going to capture gas in here and see how much gas we can capture.

Here I have a bottle and it's sealed with this little tube. This is the only way for gas to get out of here and in here right now I have vinegar, an aqueous solution of acetic acid. This is baking soda; more properly sodium hydrogen carbonate, or sodium bicarbonate. I think we all know what will happen when this comes in contact with vinegar. Now to prevent that from happening before I want it to happen, I made it a little boat for it. And the boat is theoretically, I hope, going to sit on top of the level of the vinegar, and float there without spilling. Yeah, yeah, yeah. I did it.

Before the big action scene, let's talk about what's actually happening chemically speaking because as you know, I love to speak chemically.

The acetic acid and sodium bicarbonate combine to form sodium acetate, carbon dioxide and water. The sodium acetate dissolves readily in water and will stay that way. The carbon dioxide, on the other hand, is the source of all our fun. It's a gas and it forms so many little bubbles in the surrounding liquid that it's almost a foam expanding quickly and dramatically.

As the carbon dioxide escapes through the foam, it goes through the tubing and bubbles through the column of water so we get double the bubbles; double the fun. In the end we will be able to calculate how many moles of $CO_2$ are produced by this reaction. Let's get to it.

**=====Collecting Gas Over Water (08:54)=====**

It going to take more than two hands to do this so I have a lab assistant with me today. This is Michael Aranda. Alright Michael put the tube into the cylinder there, and I get to do the fun part, which is to shake this and make bubbles. Bubbles! Little more, little more.

So we have collected exactly, about 9, almost exactly 90 milliliters, like 90.5 milliliters of carbon dioxide. You can go away now.

We also know because I know that the atmospheric pressure is 101.9 kilopascals. Remember we also need to know what the vapor pressure of the water was that we have to subtract. The amount of vapor pressure that the water gives off depends on its temperature, and there's a table we can look at. The water is at 19 degrees Celsius. According to the table, the vapor pressure is 2.2 kilopascals.

So now finding the pressure of the carbon dioxide alone is easy. We simply subtract 2.2 kilopascals from 101.8 to give us a pressure of 99.6 kilopascals for the carbon dioxide. Finally we have to make sure the pressure inside the graduated cylinder is the same as the atmospheric pressure and to do that, we just have to make sure that the levels of both of the liquids are the same. In doing that I can see that it is almost exactly 90 milliliters of carbon dioxide in the cylinder.

Now we plug everything into the Ideal Gas Law. P is 99.6 kilopascals, V is 0.090 liters, n is what we are trying to find, R is always the same, and the temperature is 292 Kelvin. One quick calculation and we find that we have collected 0.0037 moles of $CO_2$. At a molar mass of 44 grams per mole that's .16 grams of carbon dioxide.

I am excited by our success here! That might not sound like a lot but collecting a fairly substantial amount of gases and being able to make accurate measurements of them with a simple apparatus like that is pretty impressive. For me, it sounds like time for party. *Blows party blowers*

**=====Summary (11:00)=====**

Thank you for watching this episode of CrashCourse chemistry. If you weren't late to the party, you learned that John Dalton built both his theory of the atom and his Law of Partial Pressures on the foundation laid by Joseph Louis Proust and that you can add up the pressures of mixed gases just as you can do with their amounts. You also learned that the chemical reaction that occurs with vinegar and baking soda, how to collect a gas over water, and how to use that technique to figure out exactly how much of the gas you have.

*credits and music*