Ideal Gas Problems: Crash Course Chemistry #13

Crash Course: Chemistry
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What does your perfect world look like? An end to genocide, corn dogs for dinner everyday, David Tennant at my house for board games, and Carl Sagan would be there too, he'd be immortal... and not dead. We could cure cancer with puppies and vomit wouldn't exist. Unfortunately, sometimes we need to vomit to survive, David Tennant doesn't know I exist, and while scientists are experimenting with dogs that can diagnose cancer through scent, puppies are not helping to cure the disease.

Gases, likewise, don't live in a perfect world either. It'd be great if their particles always fulfilled the assumptions of the ideal gas law, and we could use PV=nRT to get the right answer every time. But the ideal gas law, much like our culture, has really unrealistic expectations when it comes to size and attraction. Namely, it assumes that particles do not have size at all, and that particles never attract each other. Well I don't know about you, but it's hard for me to take up no space no matter how much I want to disappear and attraction, so far, that's never gone away. So the ideal gas "law" often becomes little more than the ideal gas estimate when it comes to what gases do naturally.

(Intro)

=====Earlier Gas Laws =====

(01:14)
You remember the ideal gas law, it's a combination of three related laws that were discovered by a variety of scientists. Let's review real quick. Boyle's law was first published by Robert Boyle in 1660, but it was actually discovered by two of his contemporaries, Richard Towneley and Henry Power. It says that the product of the pressure and the volume of a gas is always constant as long as the temperature remains the same. Boyle's law requires a closed system where the amount of gas is constant (like in my balloon).

Next is Charles' law, which was discovered by Jacques Charles and actually named for him, though it was first published by Joseph Louis Gay-Lussac, who wasn't a scoundrel (Yes, I'm looking at you Robert Boyle). Charles' law, which also requires a closed system, states that the volume of a gas divided by its temperature gives a constant as long as the pressure is held steady. And then there's Avogadro law, discovered by resident house elf, Amedeo Avogadro, and this one does not require a closed system. In fact, it's all about changing the amount of gas (like - like I just did there). It says that the volume if the gas divided by the amount, results in a constant as long as the temperature and pressure are held steady. The more gas you have at a given temperature and pressure, the more space it takes up, and vice versa. That's pretty intuitive, right?

=====Mendeleev to the Rescue ===== (02:29)

The trouble is that it's kind of inconvenient to use three different laws to figure out what one sample of gas is doing. So eventually the three laws combined into one by, surprise, our old friend Dmitri Mendeleev. He pulled all the elements together into the periodic table and it turns out he pulled all the gas laws together too. Smart guy! The main thing Mendeleev did was to calculate just one constant, R, also called the universal gas constant, which incorporates the various constants from all the gas laws. With a little rearranging, we now have the ideal gas law that we know and love.

R equals 8.3145 liter kilo-Pascals per Kelvin mole. But don't let the crazy units scare you. It looks complicated but it's really just a measurement label like meters or Newtons and it helps us out a lot in our calculations, which is mostly what we'll be doing today.

This is all pretty cool, but remember that it's called the *ideal* gas law for a reason: it only works perfectly when gases behave ideally, with the particles not taking up too much space and not being attracted to each other. It's a bit of an impressive regime, but under normal conditions most gases come close enough to ideal behavior that we get very little error.

=====Large Size + Attraction to Others===== (03:36)

It only breaks down badly in conditions like high pressure or low temperature or high density that shove the particles so close together that they take up a large proportion of the available space. This also causes any attraction between them to be magnified, just like how I didn't think that Alison Cayne was that attractive until she became my lab partner, and then she was right there next to me all the time, and I couldn't handle it.

Johannes van der Waals came up with a way to correct the equation for real gas behavior, and we're going to explore that soon, but it's important that we know first how to use the base equation. And besides, it's close enough most of the time without the correction factor, so why make it harder than it needs to be.

===== STP and the Ideal Gas Law=====

(04:12)

One last thing; remember that STP (or Standard Temperature and Pressure) is zero degrees Celsius and one hundred kilo-Pascals, but gas law equations are done in Kelvins, not degrees Celsius so we'll call it 273.15 Kelvin and 100 kilo-Pascals. STP is sort of a baseline level that we use so scientists everywhere can compare gas behaviors under the same conditions, and as we'll see, it's useful in calculations too.

So, we've got this fancy-shmancy new law just lying around doing nothing. Let's figure out some stuff! I wonder how much space 1.00 mole of an ideal gas takes up at STP. That would be its volume, and we can calculate it like this: We know that PV=nRT. We know that we want 1.00 mole. We know what R is, since it's constant, and since we're at STP, we know the temperature and pressure too. Plug in all the numbers, leaving the volume as a variable (since that's what we're trying to find out). By doing the math (I'm doing it in my imaginary calculator here) we find that the volume equals 22.71105675. And because of the big, scary unit on R, the kilo-Pascals, moles, and Kelvins all cancel out, leaving us with liters, which is exactly what you'd expect for a volume unit. This confirms that the problem was set up correctly, so that's nice. Rounding our answer to the correct number of significant digits and adding the unit gives a final answer of 22.7 liters per 1.00 mole of gas.

'But wait!' you cry. 'Didn't we learn last time that 1.00 mole of any gas is a volume of 22.4 liters?' Yes! I am astounded that you remembered that so well. In fact, 1.00 mole of a gas will take up slightly different amounts of space depending on what the pressure is. I said, just now, that it was at STP, which is a pressure of 100 kilo-Pascals. The 22.4 liter number is for a gas at 1 atmosphere. While one hundred kilo-Pascals and one atmosphere are similar, they are not the same. In fact, 1 atmosphere is 101.325 kilo-Pascals, and I assure you that if you did the calculation using 101.324 kilo-Pascals of pressure, you would get 22.4 liters. But you don't have to trust me, you can do the calculation yourself.

===== The Hindenburg Disaster =====

(06:17)

So instead of doing that, which would be boring, let's make things a little more interesting. Like bigger! You know, like super big, like way... bigger. Let's make it Hindenburg big. The Hindenburg was a zeppelin, a flying machine that's held aloft by a lighter-than-air gas, but, unlike a hot air balloon or a blimp, has a rigid metal frame to support its shape. It was nearly as big as a cruise ship and many times larger than the Airbus A380, today's largest flying passenger vehicle. It was built in Germany in 1936 and made ten trips to the US and seven to Brazil in its first year of service.

The Hindenburg annoyed Hitler and the Nazis. For one thing, it was named for Paul von Hindenburg, the president of Germany before Hitler took over. Propaganda Minister Joseph Goebbels tried to make the Germans name it "the Hitler", but they refused, thankfully. And second, the Hindenburg was fancy new technology that the Nazis wanted to use both militarily and as propaganda. The makers did not want that to happen. They did allow it to be used for one huge propaganda mission, called the Die Deutschlandfahrt, or Tour of Germany (no giggling!), but after that, it was used only for passenger and freight service. But none of that is the reason that we still talk about the Hindenburg today. Sadly, it's most famous for catching on fire and crashing while it was trying to land in Lakehurst, New Jersey on May 6th, 1937.

=====Helium vs. Hydrogen ===== (07:32)

That's bad enough under any circumstances, but it also happened to be filled with hydrogen gas, H_2, which is highly flammable. The original plan was to fill it with non-flammable helium, but helium was extremely expensive at that time, and because hydrogen gas has a molar mass roughly half that of helium, it provides more lift too, so the zeppelin company felt that using hydrogen was an acceptable risk, but the result was the loss of 36 lives. No one knows how the fire started, but it's clear that it ended in tragedy.

So let's do some calculations to build our understanding of what the Hindenburg was all about. First, let's figure out how many moles of H_2 were aboard the ship when it left Germany. We know that the hydrogen bags had a total volume of 211,890,000 L, so we can start there. We don't know what the atmospheric pressure was that day, but we can assume that it was close to average atmospheric pressure, about 100 kPa, and the internal pressure would be about the same. So the volume is $2.1189 \text{ X } 10^8$ L, n is what we are trying to find out, R is always the same (8.3145 liter kilo-Pascals per Kelvin mole) and average early May temperatures rage around 10° Celsius which is 283.15 Kelvin. With this information we can now solve for the amount of hydrogen gas. Doing the math and rounding properly we find that n equals $9.00 \text{ x } 10^6$. All the units cancel except for moles which we add to the end. So the Hindenburg started out with $9.00 \text{ x } 10^6$, that's 9 million moles of hydrogen gas. That is a lot.

So how much extra could the Hindenburg carry because it was using hydrogen instead of helium? 9 million moles of H_2 times its molar mass, 2.016 grams per mole calculates to a little over 18 000 kilograms, or 18 metric tons of gas. On the other hand helium's molar mass it 4.003 grams per mole. So if they'd used 9 million moles of that, the gas would have had a total mass of over 36 000 kilograms. So with hydrogen instead of helium the Hindenburg could carry 18 metric tons more.

Here's another interesting question: It's warmer in New Jersey than in Germany. On an average day in May the outdoor temperature there would be around 18° Celsius or 291.15 Kelvin. If we assume that the volume and amount of gas in the Hindenburg were constant, how much did the internal pressure increase? We're looking for P, so we leave that as variable. The amount and the volume have to stay the same as before, R is always the same and our new temperature is 291.15. Using those numbers we find that the new pressure is 102.822. Given the right number of significant digits and the right unit, the final answer becomes 103 kPa.

=====Making Fire with Cotton and Your Fist=====

(10:12)

So looks like we've gotten pretty good at these calculations, time for some fun! This is a fire piston. Like everything around us it contains air and it has a plunger that allows you to compress that air very quickly. The sudden but extreme increase in pressure will cause the temperature inside to rise equally suddenly and extremely. Of course hot air looks just like cool air, so that's kind of boring on its own. But if I put a small piece of cotton inside to provide fuel, watch closely. And... boom! Did you see that? The inside of the piston actually got so hot that it ignited the cotton. Solid prove that increased pressure does indeed lead to increased temperature. Pretty cool!

=====Summary=====

(10:51)

Thank you for watching CrashCourse Chemistry! If you were paying attention, you learned that taking up space and attraction to others cause just as much trouble for gases as it does for us. That Mendeleev came to the rescue once again by combining three separate gas laws into one simple equation and how to use that equation. You also learned a lot about the amazing Hindenburg and its tragic end, why the stuff that makes party balloons cool might have made the Hindenburg cool too, and how to make fire with cotton and your fist.

=====Credits=====

(11:17)
This episode of CrashCourse Chemistry was written by Edi González, the script was edited by myself and Blake de Pastino. Our chemistry consultant was Dr. Heiko Langner, it was filmed, edited and directed by Nicholas Jenkins. Our script supervisor is Caitlin Hofmeister. Our sound designer is Michael Aranda and our graphics team is Thought Café.