



Real Gases: Crash Course Chemistry #14

Crash Course: Chemistry

<https://youtube.com/watch?v=GIPrsWuSkQc>

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====Introduction (0:00)====

One of the things that teachers who aren't me tend to say is, you need to learn all this stuff because someday you'll be out there in the "real world" and you'll need to know how to do all the things that adults do. First of all, I don't really believe this because I think you should be learning just 'cause learning is fun. But I also remember when I was kid hearing that and thinking that this real world couldn't be that much tougher than high school, because it was pretty tough.

But looking back, I can tell you that life indeed does get a bit more complicated after 12th grade with jobs, and bills, and raising kids, and taxes and I haven't done any of the raising kids but I've seen it and it looks really hard. And then there's always knowing that there's always more work waiting to be done.

But the truth is while it's more complicated, it's not really that much harder than high school. Most adults rarely have to take pop quizzes, or final exams, or ask permission to go to the bathroom, wear orthodontics, have a curfew, live on an allowance, find a date for prom, or shower in a room full of people that they don't like, all while trying to build your first real lasting relationships.

Yes, "real life" is tough but the ideal world of teenagers: not ideal. Turns out that the biggest problem isn't so much that the pressure increases in "real life", you just have to learn a few new tricks to deal with it. And the same goes for gases.

(Crash Course Theme plays)

Last week we discussed how you can identify certain qualities of gas by using the Ideal Gas Law. $PV = nRT$. It states that the product of the pressure and the volume of a gas is equal to the number of moles of the gas which we call n times universal gas constant called R times the temperature of the gas. R is always the same and if we know three of the other four things we can calculate the fourth one. But the problem with that equation is that true to its name, it only works under hypothetical ideal circumstances.

In chemistry, as in all things, there's the best of all possible worlds, which is usually easier and more fun to think about and then there's real life.

To understand and describe how gases behave in real life, we need a few new tricks. And it was dutch scientist Johannes Diderik van der Waals who figured out those tricks for us.

Van der Waals was born in 1837 in the Netherlands. The first of 10 children. His father was a carpenter and working class kids weren't allowed to go to the fancy school that led to university. Instead, he had to settle for advanced primary education, which ended when he was 15. He didn't let that define him though, and instead of becoming a carpenter as most people would have in his position, he became a teacher's apprentice. He worked hard and in time, went from apprentice to teacher and eventually became the director of the whole school. During those years, van der Waals continued educating himself as well as his pupils, and he took advantage of every class he could possibly take. Over the course of many years he ended up taking more than enough courses in math and physics to earn a degree, but as he hadn't gone through the proper channels as a student, he wasn't allowed to have one.

So, why would he bother to study so hard with no obvious reward? Well, first of all, he just really loved physics. But, more importantly, he was trying to answer a question.

See, at that time, scientists only knew how to analyze the behavior of gases with the ideal gas law, just like you, at this point. Most of those calculations produced very little error but some scientists had

noticed that when a gas was in a situation that crowds its particles, high pressure, low temperature, just high density, it starts to deviate from what the ideal gas law predicts. They couldn't figure out why sometimes they got the wrong answer. And van der Waals was determined to figure it out.

====The ideal gas law has to be corrected for volume and pressure (3:26)====

Finally at the age of 36, after trying for roughly two decades van der Waals was awarded a doctorate in theoretical physics, and even more exciting, by that time he had figured out the solution to the puzzle of real gas behavior.

In his doctoral thesis, he introduced his new gas law, which we now call the 'van der Waals Equation,' and explained why gases don't behave quite the way everyone thought they should. And in 1910 when he was 73 years old, he was awarded a Nobel Prize for his work, and I'm glad they hurried it along because you can't get a Nobel Prize after you die.

So, what exactly is the problem with this ideal gas law? Well, if the two sides of the ideal gas law were equal all the time and for all gases, as we would expect for something called a law, then PV divided by nRT should equal 1 as any fraction in which the numerator and denominator are equal.

This relationship can be shown graphically, but when you work with real gases in a real lab it becomes apparent that the relationship isn't constant at all, much less equal to one. Instead, it changes with increasing pressure, and the amount that it changes is different for different gases.

So much for the whole constant thing. Okay, let's just start over with the ideal gas law, and we'll work from there. We create a closed system by sealing gas in a container so the amount of gas is constant. R is always constant, and for the purpose of this discussion, we'll assume that we're holding the temperature constant. So nothing here can change except for P and V . If the ideal gas law worked in all situations, any increase in pressure would cause the volume to decrease at an equal rate, thus keeping the product of the two constant. But as we saw on our graph, that's not really what happens.

So...what does happen? Well let's ask our buddy, Johannes van der Waals.

====Einstein was the bomb(5:02)====

Van der Waal's solution depends on the existence and interaction of molecules. Now, we take that for granted today, but remember the existence of even atoms was disputed until Einstein proved it in 1905.

Few scientists in 1873 were prepared to accept the theory that required atoms not only to exist but also to form molecules that interacted in complicated ways.

So the problems with the ideal gas law *do* depend on interactions between molecules, specifically, two simple facts:

First: real gases aren't some sort of massless, mystical vapors. They're made of real, physical particles that take up real, physical space. And the space that they take up deducts from the amount of space that's actually available inside the container they're in.



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So when we talk about the volume in the ideal gas law, we need V not to be the volume of the container, but the volume of the container *minus* the space that the particles themselves are taking up.

So we have to create a correction for volume, and that correction is going to include n , because it's proportional to the number of molecules in the container, but also b , an experimentally determined constant that's different for every type of gas.

Obviously gases with large atoms or molecules would need more correction than those with small particles, and with that V is reduced to a more accurate value.

Second: even though some people didn't want to accept it, particles in real gases are attracted to each other. Even more so when they're close together. When they're crowded, they tend to stick closer together and therefore hit the sides of the container less often, decreasing the internal pressure.

So, when we're figuring out the number for P , we have to account for the decrease in pressure that's caused by those intermolecular attractions.

To correct the pressure, we start out with n and V because the amount of attraction depends on both of those factors. In addition, they're both squared, for complicated reasons, having to do with how particles interact. Just trust me -- and van der Waals. You should trust him because he's much more trustworthy than I am.

The we include another experimentally determined constant that is again different for every gas. Some types of atoms and molecules are more attracted to each other than other kinds are, and the degree of attraction determines the value of a .

So, we add all of that to P and the formula ends up looking like this. That's the corrected pressure *times* the corrected volume *equals* nRT . And yeah, that is a long way to go to get to a corrected equation! I don't know about you, but I'm not gonna do all that work and then not use it, so let's try it out.

Suppose we had a balloon with a volume of ten liters. And, I know that that sounds like a lot, but it's only about this much - big balloon.

How much pressure do you think the gas is exerting on the insides of the wall of the balloon? Well, let's try calculating it both ways. We will try it first with the ideal gas law.

Our volume is 10.0 Liters, let's say I've put exactly 0.500 mole of nitrogen gas in this, R is always the same and the temperature here is, let's just say, 21.50, which converts to 294.65 Kelvin. If we do the calculation, round appropriately, we find that the answer is 122 kilopascals, okay?

But what does the van der Waals Equation tell us? We have to plug in all the same numbers, plus the two correction factors. The pressure correction factor a for nitrogen gas is 139 kilopascals per liter square, mole square, and the volume correction factor b is 0.039 liters per mole.

Do the math, round correctly, and you have 122 kilopascals. Well, that is disappointing. No difference at all. Stupid van der Waals!

No, wait a minute! We talked about this. The Van der Waals equation is most important and most relevant at low temperatures and high pressures. So, how can we increase the pressure inside this balloon? How can we get the gas particles closer together?

We know that pressure is caused by the gas molecules bumping

against the inside of the container. We can make them bump around more often by increasing the temperature, or by crowding the same molecules into a smaller container.

It's hard to keep track of the temperature inside this balloon, so let's try increasing the pressure by using a smaller balloon. Or, into a ping-pong ball!

Now, this contains about 1.00 Liters of gas, and obviously I probably can't squeeze this much gas into a container this size cause the ping-pong ball would explode... Probably. I'm not trying.

We could certainly calculate how much pressure there would be if we did squeeze all of this gas into this container.

Starting with the ideal gas law, plug in all the same numbers as before, except change the volume to 0.100 liters. Do the calculation and you'll find that it gives us 12,200 kilopascals of pressure.

Okay, that's not too surprising. We have the same amount of gas in a container that's one hundredth the size of the first one. So it makes sense that it would exert 100 times as much pressure.

=====van der Waals Equation (9:38)=====

Now we've got to try it out with the van der Waals equation.

If we plug all the appropriate values and correction factors into the van der Waals equation and do the calculation, we find that the gas in the small container actually exerts a pressure of 11,800 kilopascals.

So it really does matter! That's a difference of 3,000 kilopascals from what we would expect using the ideal gas law alone! And that kind of mistake isn't not a huge problem when you're just blowing up balloons, but it definitely would be a problem if you were taking measurements in a lab or planning a mission to the bottom of the Marianas Trench.

=====Never Give Up! (10:08)=====

Pretty impressive stuff for a guy who wasn't allowed a normal education because society was like, "Oh, your dad's a carpenter..."

I'm totally impressed by all that van der Waals has taught us about real gases, but I think it's also important to remember the lesson we can learn from his life. Don't give up on your dreams. Seriously. You never know where they're going to take you.

Thank you for watching this episode of Crash Course Chemistry. If you were paying attention, you learned that the constants in the gas law aren't really all that constant when stuff gets real, that the ideal gas law has to be corrected for volume because atoms and molecules take up space and for pressure because they're attracted to each other. And that Einstein was behind a lot more of what we know today than most people realize. You also learned that the correction factors of the van der Waals equation really do matter but only under specific conditions and that you should never give up just because things are hard or people think your ideas are crazy. Because, in the end, they may just turn out to be right.

This episode of Crash Course Chemistry was written by Edi Gonzalez. The script was edited by Blake de Pastino and myself. And our chemistry consultant was Dr. Heiko Langner. It was filmed, edited, and directed by Nicholas Jenkins. Our script supervisor and sound designer is Michael Aranda. And our graphics team is Thought Cafe.

That didn't go as far as I thought it was going to.