Passing Gases: Effusion, Diffusion and the Velocity of a Gas - Crash Course Chemistry #16

Crash Course: Chemistry
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=====Introduction (00:00)=====

Welcome to another edition of CrashCourse! Today--what is that smell? Oh, thanks for the rotten eggs. Who is this--who--why are we doing this? I mean, are going to be using this to illustrate some kind of scientific principle? Because we probably could have, like, used some, like, some lilacs or something that wasn't egg.

So, like, we're saying, "I can smell it when I'm this close, but if I were farther away, I couldn't smell it." Now, eggs produce hydrogen sulfide when they start to go bad, which smells like sulfur--that's what we think of when we think of the "sulfur smell." It takes time for the molecules of hydrogen sulfide to make their way to my nose 'cause obviously gases don't travel from one place to another instantly.

It take some time since, as we've learned over the past couple of weeks, gases have real-life constraints on how they move here in the non-ideal world. These are the rules that determine, say, how fast a leaking tire goes flat; or why helium balloons don't last forever; and, of course, how, where, and when we smell stuff.

One key to understanding the behavior of a gas is its velocity--that is, the speed and direction that it moves. And THAT depends on a few important physical factors. You ready to learn more? Let's do it.

That smell IS coming from the eggs, isn't it?

[intro music]

=====Net Velocity vs. Average Velocity (01:17)=====

So. "Velocity of a gas": what does that even mean? I mean, it's easy to understand the velocity of a car. You can find the velocity by dividing the distance that the thing traveled by the time it took to do it, and then tell what direction it's going, and then you're done.

But the velocity of gases is harder to describe because they move in more than one direction at a time and the overall shape and volume of a gas cloud can change constantly. To understand gas velocity, we have to know what factors affect it, and how. As you might have guessed, there's math associated with this.

There are two ways of studying and talking about gases. First, we can consider them as a collection of atoms or molecules that act together as a system. When we do that, we often talk about the net velocity, or how fast a sample of gas moves from one place to another, like the rotten egg smell wafting its way to my nose. At other times, we focus on the individual atoms and molecules that make up the gas. In that case, we use the average velocity of each of the particles. That's the statistical mean of the speeds of all the individual atoms or molecules in the system. There are always faster ones and slower ones, but it all evens out in the mean.

The net velocity of a gas in any one direction will always be lower than the average velocity of its molecules because the overall motion of the gas is hindered by collisions among the individual particles. As with most things in chemistry, and also with life, how a gas moves is more complex than it at first appears.

The individual particles in a gas never move at exactly the same speed or direction for very long, instead, they bounce around like crazy- bumping into the walls of the container and each other over and over.

It's sort of like a group of kids in a hallway, they may be moving fast individually, but they bounce around a lot and get sidetracked. They always make it to the lunchroom eventually but the group's overall speed is lower than the speed of any individual kid.

So. What makes a gas move faster or slower? Well we already know that gas atoms and molecules move faster when the temperature increases. We can understand why that happens by remembering what temperature actually is: Temperature is a property of matter that is proportional to the average or mean kinetic energy of all the atoms or molecules in the system. We sometimes think of temperature merely in terms of measuring hotness - not that kind of hotness! The hotness or coldness of a material, and it is related to that, but when you get right down to it temperature is really just a way of expressing average kinetic energy.

The reason a stove burns you, is that the fast moving particles of the burner make the particles in your hand move so fast that they tear apart your cells and tissues. It's just a transfer of kinetic energy. And what does kinetic energy have to do with velocity? Everything. And, it also describes the relationship between a particle's velocity and its mass. So, a good understanding of kinetic energy will help a lot as we study the motion of gases.

So first, the formula for kinetic energy is one half m v squared. And as you can see, any changes in the kinetic energy are directly linked to changes in velocity. If we rearrange the formula, we find that the velocity of a body equals the square root of two times the kinetic energy, divided by its mass. This little math exercise highlights two important points.

One - because the mass is in the denominator, we can tell that it is inversely proportional to the velocity - meaning that bigger masses move more slowly than smaller masses that have the same kinetic energy. And two - the velocity is proportional to the square root of the mass - meaning that a fairly large change in mass is required to make a significant change in the velocity.

So, now we can show how the velocity of a gas relates to both its temperature, and therefore its average kinetic energy; and also to the mass of its particles. But, how do we put that all together? This is where the Scottish chemist Thomas Graham comes in.

=====Effusion (04:41)=====

In 1846, Graham published his research on the motion of gas particles. Which he did by passing gases, if you will, through a porous barrier. The process by which gases travel through an orifice or opening is known as effusion. Graham was interested in how fast the gases pass through the barrier. But this is called the rate of effusion, not the velocity. That's because here we're measuring the amount of gas that passes through at any given period of time, not the distance that it moves in that time.

We might measure amount in terms of moles or in terms of volume, but the rate of effusion never uses distance. One more thing before we move on, don't let the symbols confuse you. Notice that lower case v stand for velocity and capital V stands for volume - the perils of symbols.

=====Graham's Law (05:24)=====

Thomas Graham measured the rates of effusion for various gases and his results fit perfectly with what we already know - the more massive a gas is, the more slowly it moves. From his observations, Graham developed a formula - now known as Graham's Law of Effusion - for comparing the rates of effusion of different gases.

It states: Under identical conditions, the ratio of the rate of effusion of gas a to the rate of effusion of gas b is equal to the ratio of the square root of the molar mass of gas b to the square root of the molar mass of gas a.

It is a mouthful, but it really just confirms what we deduced earlier - the rate of motion of a gas is inversely proportional to the square root of its mass. For example, if it takes 4.5 minutes for 1.0 liter of helium to effuse through a porous barrier, how long will it take for 1.0 liter of chlorine to effuse under identical conditions?

All we need to figure it out, is Graham's Law. Let's make the helium gas a, and chlorine will be our gas b. First, we have to find helium's rate of effusion - in this case, we're using Volume with a capital V. So for helium, 1.0 liter in 4.5 minutes gives the rate of 0.22 liters per minute.

So, we plug that into the main formula - chlorine gas has a molar mass of two times that of atomic chlorine - 35.5 times two equals 70.9 - and helium's molar mass is 4.00. Put those numbers in too. Careful calculation shows that Cl_2's rate of effusion under these conditions would be 0.052 liters per minute - significantly slower than helium. This, of course, is reasonable, because Cl_2 is much more massive than helium.

=====Concentration Gradient & Diffusion (06:51)=====

Graham mainly studied gases passing through orifices in a barrier, but gases aren't usually trapped with an orifice in a barrier. They are usually able to just move freely - so how do we study their motion under those conditions? When gases are allowed to move freely, they tend to move from regions of high concentration to regions of low concentration. In a sense, they move away from places where they're crowded and toward places where they have a little bit more elbow room.

The difference in concentration between two points is called a concentration gradient -- and it's kind of like a hill that matter rolls down, always moving from high to low. Gases spread out like that until they've dispersed evenly throughout the available space. This process is called diffusion.

Now, it's important to keep in mind that the particles in a gas don't work together somehow to move in a specific direction. Remember those kids making their way down the hall? They all know where the lunchroom is, and they're hungry, so even though there are a lot of distractions, they move purposely toward their destination. Gases, on the other hand, only appear to move in a specific direction. It's really just random collisions pushing the molecules apart, thus making them spread out in every direction.

In other words, the gross eggy gas that I smelled earlier didn't travel just toward my nose; its particles spread out in every direction from the eggs, and my nose simply noticed the ones that happened to have spread in THAT direction.

Because diffusion is completely free movement, calculating the motion of individual particles becomes even more complicated. Now, we're not gonna get into the super complex mathematics of all these collisions, but it is possible to make decent estimates regarding the net velocity of gases simply by disregarding the collisions and applying Graham's Law of Effusion to diffusion as well.

=====Precipitation Reactions With Gasses (08:21)=====

Here's how it works. This is a simple acrylic tube -- nothing fancy. On one side, we have a cotton ball that's soaked with concentrated ammonia, and on the other side, a cotton ball soaked with concentrated hydrochloric acid -- which is why we're doing this in a lab, and not at my desk.

Both of these substances are very smelly, because they're giving off lots of fumes -- or in other words, a lot of the molecules in the liquid are being released in gas form.

When we put the cotton balls in the glass tube and close off the ends, the gases continue to spread, but they have no where to go except farther into the tube.

Fun fact: ammonia and hydrochloric acid react together readily to form a solid--ammonium chloride. That's a precipitation reaction, but this time it's happening in a mixture of gases, not liquids.

Because the ammonium chloride is a solid at room temperature, it forms a superfine white powder where the two gases meet in the tube, so we'll be able to tell exactly how far each gas traveled before they met.

Let's try to figure out where that's going to be. First of all, remember that Graham's Law is only an estimate when applied to diffusion, but it's a good enough estimate to work for our purposes, here.

The molar mass of ammonium, which we call "gas a," is 17 and the molar mass of hydrochloric acid, "gas b," is 36. If we plug those into Graham's Law, and set hydrochloric acid's rate of diffusion at 1.0, since we are only looking for a ratio, here, we'll find that ammonia's rate of diffusion will be about 1.5 times as fast as that of hydrochloric acid. That means that the ammonium chloride should form about three fifths of the way down the tube from the NH_3 and about two fifths of the way from the HCl.

And, there it is. See that white cloud? That's the ammonium chloride powder forming. It'll eventually settle in the glass, but it's so fine that even the tiny currents that occur as the two gases mix are enough to toss it around a little.

The ammonia has indeed moved through about three fifths of the tube. Meanwhile the hydrochloric acid has only traveled through about two fifths of the tube.

So, we've proven now that Graham's Law, as an estimate at least, works. The distances that the two gases traveled were indeed proportional to their molar masses. Makes sense. That's one of the cool things about science: it always ends up making sense, once you know what you're looking for.

=====Summary (10:20)=====

And that's it for this episode of Crash Course Chemistry. Thank you for watching. If you listened carefully, you learned the difference between the net velocity of a gas and the average velocity of its particles, and that both things have a lot to do with the mass and kinetic energy of the particles. You learned about effusion and what Thomas Graham's Law of Effusion tells us about it, and how the law of effusion can also be applied to diffusion, but not as reliably. You also learned that the collisions that occur among the particles of a gas have a huge effect on both its motion and our attempts to calculate it. You learned what a concentration gradient is, and finally, you learned how to do a precipitation reaction with gases.

This episode of Crash Course Chemistry was written by Edi González. Who I have to commend for only making one fart joke in the whole script! 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 was Caitlin Hofmeister. Our sound designer is Michael Aranda. And our graphics team is Thought Café.