Stoichiometry: Chemistry for Massive Creatures - Crash Course Chemistry #6

Crash Course: Chemistry
https://youtube.com/watch?v=UL1jmJaUkaQ
https://nerdfighteria.info/v/UL1jmJaUkaQ

=====Intro (0:00)=====

By now, you're probably starting to see how chemistry can change your view of the world. Chemistry explains everything you can see, how it looks, the way it feels, why it behaves the way it does--by describing everything that you can't see. It helps us understand the biggest stuff in the universe by helping us understand the tiniest.

And that's why chemistry can be kind of hard to understand sometimes -- because we are, on a chemical scale, huge. Chemistry traffics in infinitesimal particles, but we are made of quadrillions of those things. They are the building blocks of mass; we are literally massive.

So mass is how we massive beings tend to understand the world; in our day-to-day dealings with substances, we need to have some sense of how much of it there is before we can use it or predict how it's going to act.

For example, chemistry will be happy to tell me that the atomic structure of the sugar in this packet is 12 carbon atoms, 22 hydrogen atoms, and 11 oxygen atoms in every molecule. But I don't have any idea how many molecules of sugar I want to put in my tea! Or how that one molecule will react with other chemicals in my body.

To understand that kind of stuff, I need to know the mass of the sugar that I'm dealing with. In other words, I need to measure it. And that, is why there's stoichiometry - the science of measuring chemicals that go into and come out of any given reaction. In Greek, it literally means measuring elements, and, in essence, it allows us to count up atoms and molecules by weighing them.

Stoichiometry, yes, contains a fair bit of math, but it's one of the most important decoders that we have as chemists. It's what we use to translate from the very small to the very big, to parley the stuff that we can't see into the stuff that we can. And because of that, chemists use it all the time. Including, yes, for sweetening your tea.

Ow... hot. It's... it's quite hot.

[intro music]

=====Atomic Mass Units (2:05)=====

Now if you've been with me for a couple of weeks, and I do hope you have, you're probably thinking to yourself, wait, wait, now, don't... don't we already have a way of measuring elements? And you're right. We do, the real coin of the realm when it comes to measuring stuff in chemistry is relative atomic mass. The average atomic mass of all of the naturally occurring isotopes of a given element.

So for example, all of the natural carbon on earth occurs as one of three and only three isotopes: C-12, C-13, and C-14. They all have six protons, but the number of neutrons varies. And these isotopes show up on our planet in totally different proportions. So the relative atomic mass of carbon is a weighted average of these three masses, which comes out to 12.01.

But 12.01 what? Well, remember when we're talking about units of measurement, we're talking about arbitrary talk. Mostly, units, except for the ones that we use to measure time, aren't based on any real, objective value. We just pick a unit, like the kilogram, and we agree for a standard on what a kilogram is, and then we run with it.

The same goes for atomic mass. We measure atomic mass in atomic mass units. We made them up, and the value for a single amu is -- bear with me now -- 1/12th of the mass of an atom of carbon-12. Why? Funny story.

Until the mid-1800s, chemists in different parts of the world used different yardsticks for measuring elements. One of the most intuitive and therefore most common was to use the smallest, simplest element, hydrogen, as a base line.

But in the 1850s, some chemists, led by German Wilhelm Ostwald, proposed using oxygen instead; they preferred oxygen mainly because it combined readily with so many other elements, so they figured it would be easier to determine the weights of lots of compounds.

So a bunch of guys stroked their beards, agonized over this for years, until in 1903, they decided that atomic weight, as it was called, should be measured in 1/16ths of an oxygen atom.

Until... in 1912, isotopes were discovered and chemists realized that you can't talk about an element like it's all the same thing! It turned out there was an oxygen-16, and an oxygen-17, and an oxygen-18! And suddenly, everyone was walking around like, "I don't know how much this such weighs anymore!"

This was so crazily disruptive that it took another 50 years of strokey-beard meetings for everyone to decide to use another standard -- carbon-12. Like oxygen, carbon is common, and kind of promiscuous, when it comes to what it bonds with, and since it has twelve protons and neutrons, the mass of other, similar elements would be expressed as some fraction of it.

So, since 1961, science has pegged one amu as 1/12th of an atom of carbon-12. Which means that carbon has a relative atomic mass of 12.01 amu. Oxygen, 16 amu, and hydrogen, 1.008 amu. So that's how we way atoms.

=====Moles (4:43)=====

But, none of this solves my tea sweetening problem. Like, I don't know how many amus of these molecules together are going to make this taste good to me, or how many other molecules of sugar I can consume while maintaining my slim yet robust physique. This doesn't happen by itself, you know.

And in order to make these calculations and predict reactions, I first need to be able to convert the atomic mass of this sugar, into a standard amount of substance. Not weight, not volume, just purely, objective amount of stuff. You heard me, stuff. That, my friends, is what moles are for. Not those moles, though those are nice-looking moles.

A mole is arguably the most important unit in all of chemistry, because it allows us to express a chemical's atomic mass in terms of grams. And to define what a mole is, no matter what it's a mole of, we use our old standby, carbon-12. There are 6.022×10^{23} atoms in 12 grams of carbon-12 and by definition, that number of anything is a mole of that thing. That's a lot, and it is known as Avogadro's number, one of the most important constants in chemistry, and although Avogadro isn't the one that arrived at this number, it's named in his honor because he used this basic principle of comparing amounts of substances to first weigh atoms and molecules.

=====Molar Mass (5:59)=====

So there are this many carbon atoms in a mole of carbon-12 and there are the same number of anything in a mole of anything else. Like a dozen roses is twelve roses, but a mole of roses is 6.022×10^{23} roses, which would be enough roses to cover the surface of the earth quite deep. A mole of sand would be 6.022×10^{23} grains of sand and if they were each one millimeter long, a mole of them would stretch one hundred quadrillion kilometers. So you get the picture, it's a big number, but in chemistry the thing to remember is this: a mole of any element contains 6.022×10^{23} atoms of that element no matter what. This is what lets us translate number of atoms into grams. It lets us weigh elements.

Alright, follow me here. One mole of carbon-12 contains 6.022×10^{23} atoms and weighs 12 grams, right? So one mole of oxygen also contains 6.022×10^{23} atoms but because oxygen atoms are more massive it weighs 16 grams and you'll recall that oxygen's relative atomic mass is 16 amus. The number of atoms per mole remains the same, but the mass of a mole depends on the average mass of the element. This simply means that one mole of any element equals its relative atomic mass in grams.

So now you've got it, 1 mole of hydrogen weighs 1.008 g, a mole of iron is 55.85 g, and a mole of natural carbon is 12.01 grams. This is known as an element's molar mass. And now that we know the molar mass of elements we can calculate the molar mass of any compound. All we have to do is add up the molar masses of its component elements.

So for instance, the formula for this sugar or sucrose is $C_{12}H_{22}O_{11}$. One mole of sucrose, by definition contains 6.022×10^{23} molecules, and since each molecule contains 12 carbon atoms and 22 hydrogen atoms and 11 oxygen atoms, then one mole of sucrose contains 12 moles of carbon, 22 moles of hydrogen, and 11 moles of oxygen. Multiply the number of moles of each element by its molar mass and add them all up, that's the molar mass of the whole compound.

See, the mole is like our chemical Rosetta Stone, with it, we can translate anything from the level of atoms and molecules to the level of grams and kilograms. And we can use it to describe not only elements and compounds, but reactions, and you don't need a lab full of samples to do it, just a pencil and a calculator.

To get back to my tea problem, let's say, y'know, hypothetically, that I'm watching my weight, so I want to know what it'll take for me to burn a certain amount of sugar that I consume. That's a reaction!--and it's a pretty simple one, my body uses sucrose by combining it with oxygen to create energy plus CO_2 and $H_2$0 as waste.

You can write this out as an equation, in which the reactants combine on the left to yield the products on the right. But there's a problem here: this equation doesn't reflect chemical reality.

=====Equation Balancing (8:45)=====

During a reaction, bonds are broken and new ones are formed but the number of atoms of each element remains the same. The sugar and oxygen molecules may be busted apart and mixed up but the number of each kind of atom that you start with ends up being exactly the same after the reaction. Conservation of mass, yo.

So when writing a reaction out as an equation the number of atoms of each element has to be exactly the same on both sides. Reconciling the reactants with the products is called equation balancing, and it's a good bit of what stoichiometry is all about, because from a chemical perspective an unbalanced equation is pretty useless. It doesn't tell you how much is going in and how much is coming out. Without balancing the equation it's like saying, "When a mommy and a daddy love each other very much, a baby appears and that's all you need to know," but that's not all you need to know!

So how do you do it? Not make a baby, balance an equation. I did biology last year. Well the best way is to start with the most complicated molecule, which in this case is, of course, the sucrose. For every molecule of sucrose that goes into the reaction, you know that you're gonna have twelve carbon atoms so right off the bat you know that you're gonna have to end up with at least 12 molecules of CO_2 as a product, because that's the only molecule where those carbon atoms end up.

Now let's deal with the hydrogen, because that also shows up in only one molecule on both sides of the equation so that's easier. You know that at least 22 atoms of hydrogen go into the reaction and the product contains some multiple of 2 hydrogen atoms (that's the H_2 in the water molecule), so if there were eleven water molecules produced that would balance the hydrogen with 22 hydrogen atoms on each side.

Finally, the oxygen. Since we know we have 12 CO_2 molecules and 11 water molecules as products so far, we also know that we're gonna end up with thirty-five oxygen atoms. If you look at your reactants, on the left, you see that you have 11 oxygen atoms in the sucrose molecule and 2 in the molecular oxygen, O_2. The carbon and hydrogen are balancing nicely with only one molecule of sucrose so let's leave that alone but there could be any number of paired oxygen atoms involved. Since you need 35 and you know you have 11 to start with in the sucrose you just need 24 more, which would equal 12 molecules of O_2.

And now, the equation is balanced! You know exactly what my body is producing. For every molecule of sucrose I'm metabolizing I have to inhale 12 molecules of oxygen and in return, in addition to a little sugar buzz, I'll produce 12 molecules of carbon dioxide and 11 molecules of water. This is incredibly useful in helping us to understand the proportions of chemicals as they react at the molecular level.

=====Molar Ratios (11:11)=====

But in a lab, or in life, you have to work with measurable amounts of stuff, so the last stoichiometric trick you need up your sleeve is to calculate specific masses of the reactants and products.

So for instance, how much oxygen will I need to inhale in order to burn 5 grams of sugar? To figure that out, we just need to focus on the left part of the equation, because we only need to quantify the reactants. First, convert your balanced equation into molar masses; in order to get from molecules to grams, you need to go through moles first. (NOTE: In the video, it says 348 g of O_2, it should be 384 g. This is just a typo, all calculations from there on are correct.) When you figure out the molar masses, you see that the ratio of sucrose to oxygen is actually pretty close: 384 grams of oxygen for every 342.3 grams of sucrose. Then you simply compare this ratio to the masses of reactants in your experiment, 5 grams of sugar to X grams of oxygen, and hopefully you know how to solve for X. For every 5 grams of sugar I ingest I'll need to inhale 5.6 grams of oxygen, which I happen to know is about 35 breaths' worth.

So as long as I manage to stay alive for the next minute and a half or so, I'll manage to burn off this five grams of sugar. Down the hatch!

Today, we learned about two of the most important units of measure in chemistry, atomic mass units and moles. We also learned how to calculate molar mass and how to balance a chemical equation and finally, we talked about how to use molar ratios to calculate the amount of stuff that goes in and out of a reaction.

(credits and endscreen)