Transcript: Chemical Reactions and Stoichiometry

Unit 5, Video 1 – Chemical Reaction Indicators

We are beginning a unit on chemical reactions so it is important for us to be able to tell when a reaction has taken place. Every time 2 substances are mixed together, that does not necessarily mean that a chemical reaction has taken place. The only true way to tell is to run some type of analytical test, which we don't always have access to.

There are some things that we can look out for that will indicate that a reaction has taken place. The most common indicators of chemical change are color change, temperature change, emitting light, the formation of a gas – and you would recognize this through bubbles – if you were to mix to things together and bubbles form that would indicate that a reaction has taken place.

And then the last common indicator is the formation of a precipitate. A precipitate is a solid that forms during a chemical reaction. Imagine you poor two liquids together and a thick cloudy substance starts to form and settle out at the bottom. That would be a precipitate.

These are the indicates of chemical change to help you recognize when a reaction has taken place.

> Suggested Question (Time 1:17)
> How can you tell when a reaction is producing a gas?

Once you have determined that a chemical reaction is taking place, the next step is to try to figure out what is being produced. Start by looking at your reactants. What substances did you put together to form the chemical reaction? Try to predict what substances you are going to produce.

Then product testing comes into play. Product testing are simple tests that you can run to see if certain chemicals are being produced by a reaction. We are going to focus on product testing for three gases. The first one is hydrogen.

If a chemical reaction is being run in a test tube and hydrogen is being produced, you can take a burning splint – I mean a piece of wood that you have lit on fire, that is a burning splint – and put it down into the test tube. If hydrogen is being produced in that test tube, you will hear a very distinct popping sound. It is like a miniature explosion taking place from the hydrogen, because hydrogen is a flammable gas. It only happens once. It is a very quick, very distinct popping sound.

So the product test for hydrogen is to put a burning splint into a test tube and listen for a popping sound. The popping sound confirms that hydrogen is being produced. The lack of a popping sound would indicate that some other gas is being produced.

Suggested Question (Time 2:44)
True or False: The popping sound that is heard when a burning splint is put into a test tube where hydrogen is being produced lasts for several seconds.

The next product test is for oxygen. If you have a test tube where you think oxygen is being produced, this product test will confirm that. Start with a glowing splint. This is a splint that you caught on fire, but then you blew out the fire and it is still really hot and it is glowing.

If you put a glowing splint into a test tube where oxygen is being produced, the flame will reignite because oxygen is one of the things that helps a flame burn.

The product test for oxygen is to put a glowing splint into a test tube. If the flame reignites that is confirmation that oxygen is being produced.

Suggested Question (Time 3:28)
True or False: The production of oxygen in a test tube will cause a glowing splint to reignite.

The final test is for carbon dioxide. The product test for carbon dioxide is to bubble it through lime water. Lime water is the common name for calcium hydroxide, and it is a clear substance.

You start with a clear liquid of lime water. If you bubble carbon dioxide through it, the solution will turn cloudy. That is confirmation that your reaction is producing carbon dioxide.

I will mention that the setup for this product test would be a little more complicated in that you would have a flask where the reaction was being run. The flask would have a tube in it and the tube would need to go down into the lime water.

The product test for carbon dioxide is to bubble it through lime water. If the solution turns cloudy, you are very likely producing carbon dioxide by your chemical reaction.

Product testing is a great way to help you figure out what is being produced by your chemical reaction. Stick around as we start to look at how you can take all of that information and put it into a chemical equation.

Suggested Question (Time 4:47)
True or False: The product test for carbon dioxide involves a glowing splint.

Unit 5, Video 2 – Chemical Reactions and Balancing

Now that we have seen how to recognize when a chemical reaction has taken place, let's take a look at how we represent them. Recognize that a chemical reaction involves the reorganization of atoms. Bonds are going to be broken and new bonds are going to be formed.

This is the reaction for when hydrogen and oxygen combine to form water. The reaction is written two different ways: you have the molecular view – trying to demonstrate what is actually happening with the molecules, and then the way we would write this reaction.

Notice in the molecular view that you start with two hydrogen molecules – and a molecule means two hydrogen atoms who are chemically bound together. Those two hydrogen molecules are combining with an oxygen molecule. These bonds are breaking and these bonds are breaking. New bonds are formed between the oxygen and the hydrogen atoms.

There are several things that you should be aware of. One is that the number of atoms remains balanced on both sides of the equation. I have 1, 2, 3, 4 hydrogen atoms on the left and 1, 2, 3, 4 hydrogens on the right. There are two oxygens on the left and two oxygens on the right. The reaction is balanced.

This arrow indicates that a reaction has taken place, and we typically read the arrow as "yields." Hydrogen plus oxygen yields water.

The things on the left hand side of the equation are what you start off with. Those are your reactants. That is what you had before the reaction took place. The right hand side of the equation represents your products. That is what you have after the reaction has taken place. Your reactants will yield your products.

Suggested Question (Time 1:53)
True or False: Bonds must be broken and new bonds formed in a chemical reaction.

It is also very common for a chemical reaction to indicate states of matter, meaning whether or not something is in the liquid, solid, or gas phase.

Here you can see that the states of matter are indicated after the substance that they are referring to. They are listed in parenthesis. Whenever you see an "s" in parenthesis, that tells me that the substance is a solid. An "l" tells me that it is a liquid, "g" tells me that it is a gas, and then "aq" indicates that whatever that substance is it is dissolved in water. It is an aqueous substance.

Suggested Question (Time 2:31)
How should water vapor appear in a chemical equation?

Chemical equations are the simplest and most common way for representing a chemical reaction. The equation tells me what atoms are present - every time you see a capital letter that is a new atom. They tell me the arrangement of the atoms – every time you see a space, then you have seen a separation between two molecules. You have methane molecules and oxygen molecules in the reactants. Carbon dioxide molecules and water molecules in the products for this reaction.

There are two different types of numbers present in a chemical reaction. The first one is a subscript. A subscript is a small number that comes after an atom and it tells you how many of that atom are present in that molecule. For example, methane has 4 hydrogen atoms bonded to 1 carbon atom. If there is not a subscript after an atom, you can assume that it is a 1.

Subscripts only apply to the element that they are directly following. Subscripts come from crossing down your charges which we saw in the last unit. Once you have crossed down those charges and determined your subscript, they are fixed. You cannot change a subscript to balance an equation.

In order to balance an equation, you need coefficients. Coefficients are larger numbers that come before molecules and they apply to the entire molecule that they are in front of. This is telling me that there are two whole water molecules.

If you are trying to balance an equation and you are trying to count up the number of atoms on each side, you would multiply subscripts by coefficients. For example, in the case of water, there are 2 times 2, which is 4 hydrogens, and then 2 times 1, which is 2 oxygens.

So your subscripts cannot be changed. They come from crossing down your charges. And then your coefficients are added in order to balance the equation – in order to get the same number of each type of atom in the products and in the reactants.

One of the most common mistakes that I see with students, is that they will try to take a subscript and bring it across the arrow. They'll say: Oh there were 4 hydrogens over here on the left, I'm going to put a 4 right there. That is incorrect. You do not bring subscripts from the products to the reactants. You use coefficients to balance the equation.

Suggested Question (Time 4:59)
In the chemical reaction just shown, how many oxygen atoms were present in the reactants ($2O_2$)?

So lets focus a little bit more on this concept of balancing equations. The reason that we have to balance equations is because the Law of Conservation of Mass says that mass can be neither created nor destroyed. A chemical reaction cannot create atoms or destroy atoms. It simply rearranges them. All of the atoms that were present at the beginning of the reaction are present at the end of the reaction, and we balance the reaction to account for that.

Recall that we cannot change subscripts in order to balance an equation. The equations that are in front of us have already had the correct subscripts put into place and we cannot change those. However we are going to be adding coefficients. Coefficients are added to equations in order to balance the number of atoms on each side of the reaction.

For this first reaction, I have underlined the places where we are allowed to add numbers. You can add coefficients directly in front of any molecule present however you cannot change any of the other subscripts that are already in place.

One of the strategies that I use when I first start balancing equations is to list all of the atoms that are present underneath the arrow. Then on the left hand side of that arrow I'm going to list how many atoms are in the reactants. There are 2 nitrogen atoms in the reactants and only 1 nitrogen atom in the products.

I'm going to do the same thing for hydrogen. There are 2 hydrogen atoms in the reactants and 3 hydrogen atoms in the products. My goal is to make these numbers match. I want to have the same number of nitrogens on the left and on the right. And the same number of hydrogens on the left and on the right.

At this point, you just pick an element and start trying to fix it. I'm going to start with the nitrogens. For the nitrogens, I have 2 on the left and only 1 on the right. I'm trying to find a number that I can put right here that is going to by multiplied by my current number of nitrogens – because remember we multiply coefficients times subscripts. What times 1 will match my 2 over here? I'm going to be putting a 2 right here for nitrogen.

That changes my number of nitrogens on the right hand side of the equation to a 2. It also changes my hydrogens. Now I have 2 times 3, which is 6 hydrogens on the right. I have fixed nitrogens but I have messed up the hydrogens. Which is fine - I just need to add another coefficient somewhere to try to fix the hydrogens.

I look on the left hand side and I see that I have 2 hydrogens already in place in that molecule. What coefficient can I put here that is going to be multiplied by that 2, to match 6? 2 times 3 is equal to 6, so a 3 right here will change this number to a 6, and my equation is balanced. I have 2 nitrogens on both sides and 6 hydrogens on both sides.

There is not a number in front of this first nitrogen, which is fine. A person will look at that and recognize that that is like a coefficient of a 1. So a blank for a coefficient means that the coefficient is a 1.

The beauty of balancing equations is that you can always check yourself. You can always take a step back, look at what you've written, and make sure that the number of each type of atom on both sides of the equation matches. I have 2 nitrogens on the left and 2 nitrogens on the right. 3 times 2 is 6 hydrogens on the left, and 2 times 3 is 6 hydrogens on the right. My reaction is balanced.

I also want to point out that it is a very common situation to have a 2 for an atom on one side and a 3 on the other side. When that happens, you're going to need to go up to a 6. Multiply each side by whatever it needs to go up to a 6.

Let's take a look at our second example. Once again I have listed the atoms that are present beneath the arrow. I'm going to go through and count what is present on each side. I have 3 carbons on the left and only 1 carbon on the right. I have 8 hydrogens on the left and 2 on the right. With oxygens, there are 2 on the left. Then I have 2 here and 1 here for a total of 3 on the right.

One strategy that works well for me, is that whenever an atom is split – whenever an atom is found in two places on one side of a reaction, I'm going to approach that atom last. So I'm going to be dealing with oxygens last. This is a combustion reaction, which we will study in a later video. Typically for combustion reactions, you want to start with your carbons, then go to your hydrogens, and take care of your oxygens last.

Let's begin with our carbon atoms. I have 3 on the left and only 1 on the right. I'm going to put a 3 right here to balance my carbons. That changes this number to a 3, however it also changes our oxygens. Now I have 3 times 2, which is 6 oxygens here plus this 1 right here. So now I have a total of 7 oxygens on the right.

Now that carbon is balanced, let's focus on hydrogen. I have 8 on the left and only 2 on the right. What number can I multiply by 2 to make it go up to 8? That would be a 4. I'm going to put a 4 right here. That changes my number of hydrogens on the right hand side to an 8. However, it also changes my oxygens. My oxygens – there are 3 times 2 which is 6 here, plus 4 here. There is a total of 10 oxygens now on the right hand side.

Carbon is fixed. Hydrogen is fixed. Now let's look at the oxygen. I have 10 on the right and 2 on the left. A 5 right here will change this 2 to a 10 and complete my balanced equation.

Once again, you should always step back and check yourself. I have 3 oxygens on the left, 3 on the right. I have 8 hydrogens on the left, 4 times 2 is 8 hydrogens on the right. 10 oxygens on the left. Then I have 3 times 2 which is 6 oxygens here plus 4, 10 on the right. This is a balanced reaction.

Suggested Question (Time 11:38)
True or False: When an element shows up in multiple places on the same side of a chemical equation, you should balance that element first.

Let's look at another example. This time there is a reaction that has a polyatomic, and notice that the polyatomic stays the same on both sides of the reaction. I have nitrate on the left and nitrate on the right.

When that is the case, when you have a polyatomic that stays together on both sides of the reaction, you can simplify the way you think about the different things that are present in the reaction.

Once again, I have listed the elements that are present underneath my arrow. Notice that instead of listing nitrogen and oxygen separately, I listed the polyatomic together as one thing. That makes this reaction much simpler to balance. You can do that anytime your polyatomic stays together on the left and on the right.

Now, just like before, I'm going to go through and count the number of each type of atom present on the left and on the right.

Starting with calcium, there is 1 calcium on the left and 1 on the right. Then I'm going to nitrate. I'm looking at the whole thing. There are 2 of these nitrates – that subscript that is outside of the parenthesis tells me the number of nitrates. There's 2 nitrates on the left and 1 nitrate on the right, 1 sodium on the left, 1 sodium on the right, 1 chlorine on the left, and 2 chlorines on the right.

It really simplifies it if I keep the polyatomic together. Now I see that the first thing I need to fix is my nitrates. I'm going to do that by putting a 2 right here. That changes this number of nitrates to a 2 and it changes this number of sodiums to a 2.

Now I have fixed my nitrates but it has changed my sodiums. I can fix that by putting a 2 right here. That changes this to a 2 and it changes this to a 2.

Now my entire equation is balanced. I can take a step back and check. I have 1 calcium on the left, 1 on the right, 2 nitrates on the left, 2 on the right, 2 sodiums on both sides, and 2 chlorines on both sides. This is a balanced equation.

Suggested Question (Time 13:49)
Fill in the blank: When ________________ stay the same on both sides of a reaction, it is simpler to consider them as one particle rather than as individual elements.

I have one more example for you. I'm just trying to show you some of the unusual things that you may run into when you're balancing equations. This is another combustion reaction that contains carbon, hydrogen, and oxygen.

Once again, my oxygens are split on the right hand side. They show up in 2 places on the right, so I'm going to tackle those last.

Let's start by listing the number of atoms that are present on each side of the reaction. For carbon, I have 2 on the left and 1 on the right, hydrogen has 6 on the left and 2 on the right, and then oxygen has 2 on the left and a total of 3 on the right.

I'm going to balance carbon first by putting a 2 here. That changes the total carbons on the right to a 2 and it changes the total oxygens on the right to a 5. Then I go to balance hydrogens. I have 6 on the left and 2 on the right. Putting a 3 right here, 3 times 2 is 6, will fix the hydrogens. However, it also changes the oxygens. Now I have a total of 4 oxygens here and 3 here, giving me 7 oxygens on the right hand side.

Remember I told you that we were going to look at an unusual case. This is a case where I'm running into a problem balancing because I have an odd number of oxygens on the right and an even number on the left. In order to balance this, I would really need a 3.5 right here, which is not a good thing. You don't want to leave a decimal in your coefficients.

Technically I do have a balanced equation at this point. All of the number of atoms on each side match. However, I don't like that I have a decimal as a coefficient. In order to fix that, when you have a .5 as a decimal, you can go through and multiply all of the coefficients by 2 and it makes them all whole numbers. So this one becomes a 2, this one becomes a 7, this one becomes a 4, and this one becomes a 6. Now I have a balanced equation where everything is whole numbers.

That concludes our introduction into chemical reactions and balancing equations. Stick around as we start to look at some of the common types of chemical reactions.

Suggested Question (Time 16:20)
True or False: Coefficients should be whole numbers.

Unit 5, Video 3 – Diatomic Elements

This is going to be a short video that discusses diatomic elements. The diatomic elements are elements that will never be found by themselves in nature. The atoms that make up these elements are more stable bonded to an identical atom of itself, then they are when they are alone.

There are seven diatomic elements. They are hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine, and iodine.

For us, that means that when you see these elements in a chemical reaction, if they are not bonded to another atom, they should get a 2 for a subscript.

That means that this is not just one oxygen atom found by itself – it is a molecule made up of two atoms. Two oxygen atoms because oxygen is one of the diatomic elements. It is more stable bonded to an identical atom of itself then it is alone.

I want to show you a couple of ways that you can memorize the diatomic elements. First of all, recognize the shape that they make on the periodic table. They make an apostrophe seven, where the apostrophe is hydrogen. The seven starts with nitrogen which is element number seven, and then there are seven elements total. There are seven total diatomic elements.

Another way that you can memorize these is with a pneumonic. The one that I like the best is "I Brought Clothes From Old Navy Home." So its I brought clothes from Old Navy home to represent iodine, chlorine, fluorine, oxygen, nitrogen, and hydrogen.

Suggested Question (Time 1:41)
True or False: Carbon is more stable bonded to another atom of carbon then it is alone.

Let's look at some examples of chemical equations that contain diatomic elements. Hopefully this will reinforce some of the concepts that we saw in the last video about chemical equations.

Oxygen is one of our diatomic elements. Here is a case where oxygen was alone in the chemical reaction. It was not bound to a different element. Oxygen got a 2 for a subscript because it is a diatomic.

Over here, oxygen was not by itself, so it doesn't necessarily get a 2. In this case, when oxygen is part of a compound, it's subscripts come from crossing down charges.

Magnesium oxide does not have any subscripts because magnesium has a charge of plus 2, oxygen is minus 2. Those charges are balanced out so there were no subscripts needed.

You are only looking at this "2" concept with your diatomics when they are by themselves.

Suggested Question (Time 2:44)
True or False: Oxygen is a diatomic element. It will always have a "2" for a subscript in a chemical equation.

In this next reaction, lithium is by itself, however lithium is not one of the diatomic elements, so it is perfectly fine for lithium to be by itself.

Fluorine is by itself, but it is a diatomic, so a 2 for a subscript has been put onto fluorine. Notice that there is a 2 for a subscript for fluorine on the left side of the reaction, but not on the right. Because we never bring subscripts across the arrow. This 2 came from the fact that fluorine is a diatomic. The lack of a 2 over here comes from crossing down charges.

The fact that there are a different number of fluorine atoms on each side of the reaction was fixed by balancing the reaction using coefficients.

Your subscripts come from recognizing your diatomics and making sure that they have a 2. Your subscripts are sometimes part of a polyatomic ion and you don't want to take those away. Or your subscripts come from crossing down your charges. But once your subscripts are in place, they are fixed. You do not change subscripts to balance an equation. You add coefficients to balance the equation.

Suggested Question (Time 3:58)
True or False: Fluorine does not have a “2” for a subscript in LiF because those subscripts come from balancing out the charges.

Let’s look at one more example. This is the chlorate ion. The chlorate ion is ClO3 with a charge of minus 1. We have paired it with sodium. Sodium has a charge of plus 1, so NaClO3 does not need any additional subscripts because the charges have balanced out.

On the right hand side of the equation, I’ve got sodium with a charge of +1 with chlorine, which has a charge of -1. Cross those down and you do not need any additional subscripts.

Then I have oxygen by itself. Oxygen is a diatomic so it has a subscript of a 2. Once all of those subscripts are in place, you cannot change them. We fix any issues with the number of atoms by adding coefficients. These coefficients have been put into place to balance the reaction. But your subscripts are coming from crossing down charges, part of your polyatomic ion, or recognizing your diatomics and giving them a “2.”

Now that we know where all of the number in a chemical reactions come from, let’s take a look at some of the common types of chemical reaction.

Suggested Question (Time 5:21)
Fill in the blank: Once subscripts are in place, they should not be changed. _______________ are added to balance the equation.

Unit 5, Video 4 – Synthesis and Decomposition Reactions

In this video, we start talking about the common types of chemical reaction and we are going to focus on synthesis reactions and decomposition reactions.

The first type of chemical reaction is a synthesis reaction, which is sometimes called a combination reaction. A synthesis reaction takes place when two or more reactants come together to form just one product. A quick way to recognize a synthesis reaction is that there is no plus sign on the right hand side of the reaction, because there is only one product.

In the examples below, we are going to predict the products for some synthesis reactions. Notice that in each case, I start with multiple reactants.

In the first example, I'm starting with magnesium and oxygen. They are going to come together to form just one compound. The first thing I'm going to do is list the atoms that are present, which in this case are magnesium and oxygen.

Anytime you create a compound when you are predicting products, you should be aware that you need to cross down your charges. You want to check the periodic table. You'll find that the charge of magnesium is +2, the charge of oxygen is -2. Those do not need to be crossed down because they balance out.

We have correctly predicted that the product is magnesium oxide. The last thing that you should do is balance the equation. I have 2 oxygens on the left and only 1 on the right, so I'm going to put a 2 here. Then I need a 2 here to balance out the magnesiums. That is a correct chemical equation. It is a synthesis reaction because there is only one product, and we have correctly predicted that product.

Suggested Question (Time 1:50)
Fill in the blank: Synthesis reactions form when two or more reactants combine to form ______ product.

Let's look at the next example. In the next example, we have lithium and fluorine. They're going to come together to form lithium fluoride. Anytime we make a compound, we want to check the charges. Lithium has a charge of +1.

Fluoride has a charge of -1. When you cross those down, you don't need any additional subscripts.

However, we do need to balance. I'm going to put a 2 here to balance out the fluorines, and then I need a 2 here to balance out the lithiums.

Suggested Question (Time 2:19)
True or False: Charges should be considered when determining the subscripts for each compound.

My last example is a synthesis reaction involving aluminum and sulfur. I'm going to go ahead and create a compound that has those two elements. I do need to check the charges for that compound.

Aluminum has a charge of +3. Sulfur has a charge of -2. They don't match so I need to cross down the charges, to put a 3 here for the sulfur and a 2 for the aluminum. So Al_2S_3 gives me aluminum sulfide.

Then I need to balance the reaction by putting a 2 here and a 3 here.

These are all synthesis reactions. They all have 2 or more reactants combining to form 1 product. When you predict what that 1 product is, you need to make sure you cross down your charges.

Suggested Question (Time 3:10)
True or False: The subscripts for each element must match for the reactants and the products.

Decomposition reactions are like the opposite of synthesis reactions. You start with 1 reactant and it breaks apart into 2 or more products. These are a little bit more difficult to predict because you can have multiple products.

3:30

There are some trends that help us to predict the products that will be formed in a decomposition reaction. These are some guidelines. In these guidelines, anytime you see a capital M, that stands for some metal, and anytime you see a capital NM, that stands for a nonmetal.

The first thing that we can predict is the metal carbonates. If you have a metal carbonate, it is going to break apart into a metal oxide plus carbon dioxide.

You still need to check your charges whenever you create these new structures.

The metal oxide, you would want to find the charges on the periodic table and then cross them down. Carbon dioxide is not going to change. It's charges are already correct.

4:19

Metal hydrogen carbonates, or bicarbonates, are going to break apart in a predictable fashion to form a metal oxide with water and carbon dioxide. This metal is not going to change. You just need to check the charges when you form the metal oxide.

4:39

Metal hydroxides are going to break apart to form a metal oxide plus water.

4:45

Metal chlorates are going to break apart to form a metal chloride plus oxygen.

4:52

And then anytime you have an acid, an acid typically starts with an H, that acid is going to decompose to form a nonmetal oxide plus water.

Let's look at some examples.

Suggested Question (Time 5:07)
What type of reactant will decompose to form a metal oxide plus water?

Here are some examples of decomposition reactions where we are going to predicting our products, and we are going to be using these rules to help us predict our products. Notice – I can recognize that these are all decomposition reactions because there is only 1 reactant. These things are going to break down into 2 or more products.

5:30

The first example is potassium carbonate, so I'm going to go to this rule about the carbonates. Potassium is a metal, so I'm looking at a situation where a metal carbonate is going to break down into a metal oxide plus carbon dioxide.

In this case, my metal is potassium and the reaction is going to form potassium oxide plus carbon dioxide. I can predict that it is going to be KO plus CO_2.

However, I need to check the charges on that potassium oxide that I just created. Potassium has a charge of +1. Oxygen has a charge of -2. Those charges are going to need to get crossed down so that I get K_2O.

The last thing that you would want to do is make sure that the reaction is balanced. One nice thing about the metal carbonates – the fashion in which they break down, your reaction should always be balanced. But it is a good idea to check.

Suggested Question (Time 6:34)
In the decomposition of lithium carbonate, the products would be lithium oxide plus __________.

Our second example is potassium hydroxide, so I'm going to go to this rule about the hydroxides. A metal hydroxide is going to break apart into a metal oxide plus water. This time my metal is potassium, so it is going to form potassium oxide plus water.

Keep in mind, every time you form a compound, you want to check your charges. Potassium is +1, oxygen is -2, you want to cross those down to get K_2O.

Suggested Question (Time 7:08)
In the decomposition of sodium hydroxide, the products would be sodium oxide plus __________.

This last example is for the decomposition of sodium chlorate. I'm going to go to this rule about the chlorates. In this case sodium is my metal. A metal chlorate is going to decompose to form a metal chloride plus oxygen.

When I go to predict my products, sodium is my metal, is going to be with chloride this time. I need to check my charges on that. Sodium has a charge of

+1, chloride has a charge of -1 so there would be 1's for subscripts. Plus oxygen. Oxygen is a diatomic so it gets a 2.

The last thing you would want to do is balance that equation. This equation needs to have a 3 here and a 2 here to balance the oxygens. Then another 2 here to balance the rest of the equation.

These rules that we have really help us to predict the products of decomposition reactions.

Now that we have seen how synthesis and decomposition reactions work, stick around for the next video where we look at single replacement reactions and the activity series.

Suggested Question (Time 8:31)
True or False: It is necessary to fill in correct subscripts and coefficients for decomposition reactions.

Unit 5, Video 5 – Single Replacement Reactions

Let's continue our discussion of chemical reactions by looking at single replacement reactions.

A single replacement reaction takes place when a single element in a compound gets replaced. An element by itself reacts with a compound and that element is going to replace something in the compound.

The element that was by itself in the reactants is now part of the compound in the products.

The element that starts off by itself in the reactants can either be a metal or a nonmetal. It can either take a positive or a negative charge when it goes into the compound. Because of this, you end up with 2 general formulas for single replacement reactions.

In the top case, A is a metal. A would be something on the left hand side of the periodic table that normally takes a positive charge in an ionic compound. So, when it reacts with this compound, it is going to replace the positive element. Remember that we always write positive elements first. A is trying to replace B. The positive metal is trying to replace the positive portion of the compound. In this case, you can see that A has replaced B. Now A is part of a compound with C and B is leftover by itself.

In the bottom case, D is the element that is going to be doing the replacing. In this case, D is a nonmetal. When it goes into a compound, it is going to take a negative charge. In this case, D is going to try to replace the negative portion of the compound, which would be the thing that is written second. In this case, D is trying to replace C. On the products side you can see where B and D are forming a compound, and C has been left by itself.

Suggested Question (Time 2:04)
True or False: A metal by itself will attempt to replace the metal in the compound in a single replacement reaction.

Let's look at example of a single replacement reaction and predict the products. Starting with the top one, I can see where lithium is reacting with sodium chloride.

I recognize this as a single replacement reaction because I have an element by itself reacting with a compound. This element is trying to replace a single element within this compound.

In this case lithium is a metal, so lithium is going to try to replace the metal in the compound. Lithium is trying to replace sodium. When lithium replaces sodium, lithium is now paired with chlorine.

Anytime you write a formula for a compound, you should check your charges. Lithium has a charge of +1, fluorine has a charge of -1. There are no additional subscripts needed for this compound.

What happened to the sodium within the compound? That's the remaining portion of the reaction. Sodium is left over. Sodium is now by itself. It was bumped out of the compound by the lithium.

You should ask yourself anytime you write an element by itself, is that element a diatomic? In this case, no, sodium is not a diatomic, so it is fine without any additional subscripts.

The final thing you want to look at is to make sure that the reaction is balanced. In this case, it is.

This is a single replacement reaction where a single element within a compound was replaced.

> Suggested Question (Time 3:41)
> True or False: It is important to use ion charges and the list of diatomic elements to confirm subscripts when predicting products in single replacement reactions.

Before we continue with the examples, recognize that every time an element and a compound come in contact with each other, that does not mean that a reaction is going to take place. In the case of a single replacement reaction, we have a tool that helps us to determine when the reaction takes place. That tool is called the activity series.

There is an activity series for metals, when you have a metal trying to replace another metal. There is an activity series for nonmetals, specifically for the halogens, when a halogen is trying to replace another halogen.

The way that the activity series works is that an element higher in the series can replace an element lower in the series.

Let's look at some examples of how you can use this to determine whether or not a reaction will take place.

Suggested Question (Time 4:35)
True or False: According to the activity series for metals, magnesium can replace barium in a single replacement reaction.

In the example we looked at earlier, lithium was trying to replace sodium. Lithium is here in the activity series while sodium is here. Lithium is higher in the chart than sodium. This reaction will take place the way it's written. Lithium has the ability to replace sodium.

4:58

In the second example, the element by itself is fluorine. Fluorine is a nonmetal, so fluorine is trying to replace the nonmetal in this compound. Fluorine is trying to replace iodine.

We want to look at the other version of the activity series – the one that is for nonmetals. You want to find fluorine and find iodine. As long as the replacing element is higher in the series the reaction will take place. In this case, fluorine is higher than iodine so this reaction will take place.

Now lets predict the products. Fluorine comes in and replaces the iodine. Now you have magnesium paired with fluorine instead of with iodine. That is magnesium fluoride.

I do want to point out that this subscript did not come across the reaction. This subscript came from crossing down your charges. Magnesium has a charge of +2, fluorine has a charge of -1. You would cross those charges down to get MgF_2.

The iodine was leftover so the iodine is alone in the products side now. The iodine has a 2 because iodine is a diatomic.

When you write your compounds, you want to cross down your charges to get the subscripts. And then when you write an element by itself, you do want to check to see if it is a diatomic. The last thing you would want to do is confirm that the reaction is balanced – and it is.

6:35

In the next example, the element by itself is lead. Lead is a metal. Lead is trying to replace calcium. The other thing that you could do is make sure that you are finding things within the same series. Lead and calcium are both in the activity series for the metals. Lead is right here and calcium is right here.

Lead does not have the ability to replace calcium. Lead is lower in the series. It will not replace calcium.

In this example, it is very appropriate to simply write "No Reaction." This reaction would not take place. You do not need to predict the products.

In the next reaction, you have iodine as the element by itself. Iodine is a nonmetal so it is trying to replace the nonmetal in the compound. Iodine is here in the series. Chlorine is here. Iodine does not have the ability to replace chlorine because it is lower in the series. This is another example where you would write "No Reaction."

That concludes our discussion of single replacement reactions. We've seen how you can recognize them, how you can predict the products, and how you can use the activity series to determine whether or not the reaction will take place. Stick around as we move on to double replacement reactions.

Suggested Question (Time 8:16)
True or False: According to the activity series for nonmetals, iodine can replace chlorine in a single replacement reaction.

Unit 5, Video 6 – Double Replacement Reactions

As we continue looking at types of chemical reactions, let's focus on double replacement reactions, which are often called metathesis reactions.

Double replacement reactions take place when two elements within a compound are replaced. The ions within the compound get swapped.

You have the general formula here. If you have a compound that is made up of AB and a compound that is made up of CD. Remember that the positive ion is always written first, followed by the negative ion.

In the products, A and D end up together. The positive ion from here ends up with the negative ion from here. In the other compound is the positive ion from here and the negative ion from here.

Both elements within the compound get replaced and the ions are swapped.

> Suggested Question (Time 1:00)
> True or False: The compounds in double replacement reactions always have the positive ion written first, followed by the negative ion.

We're going to use this general formula to predict the products for some reactions. Keep in mind any time you form new products, you should always put the cation first followed by the anion, and you should cross down your charges to make sure that the subscripts are correct. Finally, you should always make sure that the reaction is balanced using coefficients.

> Suggested Question (Time 1:22)
> True or False: It is not necessary to balance double replacement reactions.

In my first example, magnesium oxide is my AB and potassium chloride is my CD. It's going to produce A and D, the magnesium and the chloride are going to come together to form a compound. The formula for magnesium chloride is $MgCl_2$ and that comes from crossing down the charges. Magnesium has a charge of +2, chloride has a charge of -1, you would cross those down to get $MgCl_2$. And you should do that anytime you form a new compound to make sure that the subscripts are correct.

The next portion is C and B coming together. The potassium and the oxygen are going to come together to form a compound. That's going to give us potassium oxide, K_2O. The 2 once again came from crossing down charges. Potassium has a charge of +1, oxygen has a charge of -2, cross those down to get potassium oxide.

The final thing that you should do is make sure that your reaction is balanced by adding coefficients. In this case, a coefficient of a 2 is needed here to balance out the number of potassiums and chlorides.

Suggested Question (Time 2:38)
True or False: The subscripts on the products side of a double replacement reaction come from copying the subscript on the reactants side of the reaction.

In my second example, AB is sodium bromide and CD is calcium sulfide. A and D are going to come together, so sodium and sulfide are going to come together to form a compound. You would once again cross down the charges. Sodium has a charge of +1, sulfur has a charge of -2, cross those down to get Na_2S.

Finally, the C and B are going to come together. Calcium and bromine are going to come together to form calcium bromide. Calcium has a charge of +1, bromine has a charge of -1, cross those down to get $CaBr_2$.

The last thing that you should do is to confirm that the reaction is balanced. In this case, a coefficient of a 2 is needed here in order to balance the sodiums and the bromides.

Suggested Question (Time 3:35)
True or False: The subscripts in a double replacement reaction come from crossing down the charges of the individual ions in the compound.

My final example uses one of our polyatomic ions. OH is the polyatomic ion hydroxide. That is taking the place of the B. The potassium is A, a polyatomic is B. That is going to stay together in the reaction. I have potassium hydroxide as AB, and sodium chloride as CD. A and D are going to come together. Potassium and chloride are going to form the compound potassium chloride. My charges, potassium is +1, chlorine is -1, no additional subscripts are needed for that compound.

Finally, the C and the B. Sodium and hydroxide are going to come together to form a compound. The sodium has a charge of +1, the hydroxide has a charge of -1. No additional subscripts are needed. This reaction is balanced the way it is written.

That wraps up a quick introduction to double replacement reactions. However, stick around. We are going to look at specific types of double replacement reactions called precipitation reactions. We are going to look at solubility rules and net ionic equations in the next video.

Suggested Question (Time 5:06)
True or False: Polyatomic ions generally break apart into their individual atoms in a double replacement reaction.

Unit 5, Video 7 – Precipitation Reactions and Net Ionic Equations

The last video focused on double replacement reactions. These were reactions were compounds swapped ions. The general formula was AB plus CD yields AD plus CB. If you would like to see more of that, there is a link to the playlist in the description.

This video is going to focus on a specific type of double replacement reaction. We are going to look at precipitation reactions.

0:31

A precipitation reaction is one that forms a precipitate, and a precipitate is a solid that forms during a chemical reaction. Very often in double replacement reactions one of the products is something that is not soluble in water, so it forms a solid. It comes out of the solution and forms a solid.

Suggested Question (Time 0:53)
A _______________ is a solid that forms during a chemical reaction.

We have a tool that helps us to determine when a precipitate will form. That tool is the solubility rules.

Here is one version of the solubility rules. This is a list of ions that tend to be soluble.

For example, chlorides, bromides, and iodides tend to be soluble unless they are paired with silver, lead or mercury.

This is a list of ions that tend to be insoluble, meaning they would form a solid.

These ions tend to not form a solid. They're soluble in water. They dissolve in water.

These ions tend to form a solid. A precipitate would form.

Suggested Question (Time 1:39)
True or False: When a substance is soluble in water, a precipitate forms.

Let's look at an example. This is sodium fluoride.

Notice all of the rules deal with anions. These are negatively charged ions that are found at the end of the compound.

In my example, sodium fluoride, I'm trying to find a rule about the fluorides. That rule is here. It tells me that all of the fluorides are soluble, except group 2, lead II, and iron III.

Because sodium is not one of the exceptions, then sodium fluoride is soluble. We would put (aq) beside sodium fluoride to let the reader know that this was an aqueous compound. A precipitate did not form. It was soluble. It did dissolve in water.

Suggested Question (Time 2:29)
True or False: The solubility rules focus on the negatively charged ion in a compound.

Consider another example – silver chloride. Once again, focus on the second ion in the compound. The chlorides. Here is the rule about the chlorides. It says the chlorides are soluble except silver. Silver chloride is one of the exceptions so that kicks it over to the other side. Silver chloride is insoluble. Something that is insoluble does not dissolve in water. A solid would form. You would put the state of matter to be (s) for solid.

Suggested Question (Time 3:04)
When a precipitate forms, ____________ should be listed as the state of matter.

Another example would be lithium hydroxide. Once again, focus on the anion. We are looking for a rule about the hydroxides. That rule is here. It says your hydroxides are insoluble except group I, strontium, barium, and ammonium. So you should ask yourself: Is lithium in group I, or is it strontium, barium, or ammonium. Lithium is in group I on the periodic table. Lithium hydroxide is one of the exceptions. Even though you find the hydroxides here, the fact that it is one of the exceptions throws it back over to here. Lithium hydroxide is soluble. You would put (aq) to indicate that it was aqueous. It does dissolve in water. A precipitate does not form.

Suggested Question (Time 3:56)
True or False: When a precipitate does not form, ___________ should be listed as the state of matter.

Now that we have seen how the solubility rule work, let's take a look at some precipitation reactions. These look like double replacement reactions. The first thing that you should do is to predict the products of the reaction, following this general pattern.

In this case, I've got lead nitrate. Lead is my A, nitrate is my B. With potassium iodide. That's my C and D. I can see here that the lead, which is my A, is going to end up with iodide. On the right hand side of the reaction, I should have lead iodide plus what's left over is potassium nitrate.

Now that you have the products predicted, it is time to go to the solubility rules and determine if either one of these compounds would form a precipitate.

Focus on the second ion in the compound. Find a rule about the iodides. It tells me that all chlorides, bromides, and iodides are soluble except silver, lead, and mercury. Lead is one of the exceptions. This compound is insoluble. It would form a precipitate and I have indicated that by listing an (s) for the state of matter.

Now we focus on the other compound. Find a rule about nitrate. It says that all nitrates, acetates, and ammoniums, and all group IA salts are soluble. And there are no exceptions. All nitrates are soluble. This compound does dissolve in water and you would like (aq) as the state of matter.

At this point, my precipitation reaction is complete. You can see that an aqueous solution plus another aqueous solution formed a solid solution plus another aqueous solution.

In a lab environment, these aqueous solutions would be liquid solutions that you would pour together, and you would actually see a cloudy substance start to form and fall out of the solution. That cloudy substance would be your precipitate. That is the solid that is forming during the chemical reaction.

Suggested Question (Time 6:14)
True or False: A chemical reaction must begin with a solid as a reactant in order to produce a solid product.

Let's do another example. This time barium hydroxide is combined with potassium sulfate. This is a double replacement reaction and we are going to predict the products.

Barium is going to end up with sulfate – remember that A ended up with D. We have barium sulfate. The other product is going to be potassium with hydroxide.

At this point, you are ready to go to your solubility rules and determine which one of these might be a precipitate. Remember to focus on the second ion in the compound.

We want to find a rule about the sulfates. Here it is. It says all sulfates are soluble except for calcium, strontium, barium, mercury, lead, and silver. So this is one of the exceptions, so it is insoluble. We need to list solid as the state of matter.

The other compound, potassium hydroxide. We should find a rule about the hydroxides. It says all hydroxides are insoluble, except for group I. Potassium is in group I, so this is one of the exceptions. Potassium hydroxide is soluble, and we need to put (aq) to indicate that it has dissolved in water.

Suggested Question (Time 7:38)
True or False: Potassium hydroxide is soluble.

The precipitation reactions that we just wrote are considered molecular equations. A molecular equation shows you the compounds that are present and the state of matter.

There is a lot of information given to you there. The fact that barium hydroxide is aqueous, means that those ions – the ions that make up barium hydroxide, have broken apart and dissolved into the water.

A complete ionic equation shows that process. You can see that barium hydroxide, because it is aqueous gets broken apart into barium and hydroxide ions.

When you write a complete ionic equation, you should include the charges for each of the ions present. The number of ions present is also reflected. There were 2 hydroxide ions for every 1 barium ion. Here there are 2 hydroxide ions. That subscript has become a coefficient.

Potassium sulfate was also aqueous, so in a complete ionic equation that shows up as 2 potassium ions, and 1 sulfate ion.

Barium sulfate is our solid. In a complete ionic equation, barium sulfate – your solids – stay together.

Potassium hydroxide, once again is aqueous so we break it apart into its ions. Notice that the reaction is balanced, so there needed to be 2 of each of those to balance the reaction.

Suggested Question (Time 9:20)
True or False: Complete ionic equations break solid substances apart into their ions.

In a complete ionic equation there are often ions that show up on both sides of the reaction exactly the same. These are ions that did not participate in the chemistry. We call them spectator ions. They were simply watching the chemical reaction take place.

9:41

There is another type of equation called a net ionic equation. A net ionic equation leaves out those spectator ions. Once you've written the complete ionic equation, you start crossing out any ions that stay the same on both sides.

For example, hydroxide shows up here and here. There are 2 of them, 2 hydroxide ions on each side of the equation – so we can cross those out. There are also 2 potassium ions on each side of the equation. We can cross those out.

Notice that the sulfate ion is present on the left, but on the right sulfate is part of barium sulfate – it is part of a compound. That is different. I can't cross out sulfate and I can't cross out barium.

What remains is called the net ionic equation. The net ionic equation shows you the chemistry that actually took place. The other ions that are not present in a net ionic equation are spectator ions. They simply watched the chemistry take place. They did not participate.

Suggested Question (Time 10:56)
Ions that are not included in a net ionic equation are called ____________________. They did not participate in the chemical reaction.

So just to recap: The equation that you write when you are doing your double replacement reaction and you identify your precipitate. That is your molecular equation.

From your molecular equation, you can write a complete ionic equation. To do that, any substance that is aqueous gets broken apart into it's ions. Any substance that is solid or liquid stays together.

Then you start looking for things that can be cancelled out. Cancel out anything that is exactly the same on both sides of the reaction.

What remains is your net ionic equation.

That completes double replacement reactions and precipitation reactions. Stick around as we look at our last type of reaction, which are combustion reactions.

Unit 5, Video 8 – Combustion Reactions

The final type of reaction for us to cover in this series is a combustion reaction. In this video we are going to look at what a combustion reaction is, what the general formula looks like, what is a hydrocarbon, and some tips for balancing combustion reactions.

0:18

To begin with, a combustion reaction takes place when a hydrocarbon burns in the presence of oxygen and it produces water plus carbon dioxide.

So these are called combustion reactions because you actually have burning taking place.

The beauty of a combustion reaction is that the products are always the same. You will always produce water plus carbon dioxide. The only thing that changes about this reaction is your hydrocarbon and how it is balanced.

Suggested Question (Time 0:51)
True or False: Combustion reactions always produce carbon monoxide plus water.

So what is a hydrocarbon? A hydrocarbon is a compound made up of carbons and hydrogens covalently bonded together. Hydrocarbons follow this general formula.

The formula is C (n), H (2n+2). This means that for a normal hydrocarbon, if you know the number of carbons, you can predict the number of hydrogens.

This is methane. Methane has 1 carbon. In the formula you would plug a 1 in for n. To figure out the number of hydrogens, 2 times 1 is 2, plus 2 is 4. Methane is expected to have 4 hydrogens.

This is hexane. Hexane has 6 carbons. We would be plugging a 6 in for n. 2 times 6 is 12, plus 2 is 14, meaning we expect hexane to have 14 hydrogens.

Suggested Question (Time 1:49)
How many hydrogens will a hydrocarbon have if it has 8 carbons?

Predicting the products of a combustion reaction is incredibly easy. You always have your hydrocarbon plus oxygen, and every single time you're going to product water plus carbon dioxide.

2:04

Lastly, we need to discuss the balancing of combustion reactions. I'm going to show you 2 examples. One is going to work fairly simply and the second is going to require a little tweaking.

In the case of combustion reactions, notice that on the right hand side oxygens are split, and anytime an element is split on one side, I always recommend balancing that element last.

So for a combustion reaction, I recommend balancing the carbons first, then the hydrogens, and the oxygens last.

Suggested Question (Time 2:39)
True or False: The oxygen in a combustion reaction should be balanced last.

In this first example, there is 1 carbon on the left and 1 on the right. That's balanced. There are 4 hydrogens on the left and only 2 on the right, so we'll need to put a 2 right here. And then in terms of the oxygens, there are 2 on the left. On the right hand side there are 2 here plus 2 here for a total of 4, so I'm going to need to put a 2 right here to complete the balancing of the reaction. That one was fairly straightforward.

3:09

For our next reaction, I'm going to recommend that you write in pencil because we are going to be doing some tweaking along the way.

To begin with, there are 6 carbons on the left, so I'm going to need to put a 6 here to balance those. Then there are 14 hydrogens on the left so there should be a 7 here to balance the hydrogens.

Once I balance the carbons and the hydrogens, I run into a problem with the oxygens. On the left hand side I have 2 oxygens. That's an even number. On the right hand side I have 7 here plus 12 here for a total of 19, which is an odd number. So you run into a problem balancing the number of oxygens. It makes you need to put a 9.5 here to make it balanced.

And hopefully that makes you cringe a little bit. You should not have a half of an atom or a half of a molecule in your balanced equation. You should have whole numbers.

At this point, we are balanced, but we have this .5 that we need to fix. In order to fix that, you can go in and multiply all of your coefficients by 2.

I'm going to take this 1 up to a 2. The 9.5 is going up to a 19. The 7 is going to a 14 and the 6 is going to a 12. At that point all of my numbers are whole numbers and my reaction is balanced.

This is a pretty common problem for combustion reactions. Some combustion reactions are very simple to balance. Some you run into this problem with the oxygens where you have to go in and multiply all of the coefficients by 2. But the beauty of your combustion reactions – your products are always water and carbon dioxide.

That wraps up our series on chemical reactions. Stick around as we start to talk about the mole.

Suggested Question (Time 5:22)
True or False: The coefficients in balanced chemical equations should be whole numbers.

Unit 5, Video 9 – The Mole and Using Avogadro's Number

This video is going to introduce a concept called the mole, which is a number that is going to help us bridge the gap between the microscopic world of atoms, and the real world where we can actually measure things on a scale.

Let's begin with an example that shows why we need this concept of the mole. If you consider an aluminum can. An aluminum can has a relatively small mass. An empty can would be about 14 grams.

14 grams of aluminum is roughly this many atoms of aluminum. It is a huge number of atoms because atoms are so, so tiny.

A chemist is not going to report something in terms of numbers of atoms because it is too easy to make a mistake when you have to deal with that many zeros.

Instead, a chemist would divide this number by another giant number and report "moles of atoms" instead of atoms.

> Suggested Question (Time 1:02)
> True or False: It is common to report scientific data in units of atoms because the numbers are rather small.

So what is the mole? The mole is just a number. When I say the word dozen, you think 12. When I say the word mole, you should think 6.022 x10^23. It is called Avogadro's number and it is used by chemists as a bridge between the microscopic world and the laboratory world.

In the example I just gave, we could take this many atoms of aluminum, and divide it by the number for the mole, and then we would say instead of this many atoms, we have 0.52 moles of aluminum. It is a much friendlier number to deal with.

By counting in units of moles, it allows us to deal with much more reasonable numbers.

> Suggested Question (Time 1:49)
> What is the value for Avogadro's number?

Avogadro's number is a conversion factor that can help you convert between something that is very very small, like an atom, to a number that is more reasonable to deal with.

You're not limited to using it for atoms. You could use Avogadro's number to count anything that was really tiny. You could count atoms, you can count particles, you can count molecules. You are not limited to just using Avogadro's number for atoms. Just like you are not limited dozen to count eggs. You can count 4 dozen people and it still means 4 times 12 number of people.

> Suggested Question (Time 2:33)
> True or False: Avogadro's number can be used for conversions involving atoms, particles, and molecules.

Here is an example where we are going to use Avogadro's number as a conversion factor. It says, how many moles are in 1.2x10^21 atoms? That is a huge number of atoms.

When you set a problem like this up, you should always start with the number and the unit that you are given in the question: 1.2x10^21 atoms.

Then set up a fraction. So times some fraction.

Next focus on the units. You want the unit that you started with to go on the bottom of your fraction. That way it will cancel out. So on the bottom of my fraction, I have atoms. I'm going from atoms to moles. There are 6.022x10^23 atoms per 1 mole.

This is a conversion factor that could be flipped if I need it to be, but because I'm starting with atoms, I need atoms to be on the bottom.

At this point, you would multiply by anything that is on the top and divide by anything that is on the bottom to get your final answer. 0.0020 moles.

> Suggested Question (Time 3:48)
> True or False: When making conversions, the unit you start with should go on the top of the next fraction so that it will cancel out.

If you're trying this example in your calculator and you got a different number than me, I want to take a minute and look at how you should put this into your calculator and show you some common mistakes that students make.

It is very easy when you are dividing by an exponent to end up putting it into your calculator wrong and you'll end up with the right numbers in your answer but the decimal is in the way wrong position. I want to show you some ways that you can avoid that.

One possible way to do that is to make sure that you put Avogadro's number in parenthesis. You can put 1.2x10^21 divided by – open your parenthesis – 6.02x10^23 – and then close your parenthesis and hit enter.

If you don't do that, your calculator is going to divide by this part of the number and then multiple by this part of the number, and your answer is going to be off by a factor of 23. Your decimal is going to move 23 places in the wrong direction.

Another way that you can do that – which is my personal favorite – is to use this EE function that is on your calculator. Many calculators have it underneath the comma.

The EE function does the x10^ portion for you. So to put this into my calculator using the EE function, I would press 1.2 and then I would hit second EE 21. That has entered this number for me. Then I'm going to divide by 6.022 – I'm going to use the EE – second EE 23, and hit enter. If you put it in that way, your calculator will not make the error of dividing by this and then multiplying by that. Your calculator will do the math correctly if you use that EE function.

Just to recap, you want to start with the number given in the problem. Whatever number you have here needs to go here. Avogadro's number is a conversion factor that can go between something tiny and a mole.

Suggested Question (Time 6:07)
When dividing numbers that have exponents, it is important to use parenthesis, or use the _____ function on the calculator.

Let's look at another example. This time we're going to start with moles and convert to atoms. Notice that the pattern that we are going to follow to answer the question is the same. Always start with the number and the unit that you are given in the question.

This says how many atoms are in 3.9 moles. I'm going to start with 3.9 moles. I'm going to set up a fraction and then I'm going to think about my units. The unit that you start with goes on the bottom of the next fraction.

If you're comparing this to the example you just did notice that this fraction is flipped now. And the reason for that is that I need moles to cancel out.

Avogadro's number tells me that there are $6.022x10^{23}$ atoms per mole. Then I'm going to multiply by that number this time and I will get a final answer of $8.6x10^{21}$ atoms.

Notice that my unit for my answer is in atoms because moles has cancelled out.

You can use this concept of the mole to convert from anything that is tiny: molecules, formula units, anything that is tiny to 1 mole.

Stick around as we look at another conversion factor that uses the mole and helps us to convert between moles and grams.

Suggested Question (Time 7:44)
When converting from moles to atoms, what unit should be on the bottom of the fraction?

Unit 5, Video 10 – Mole Conversions

Let's continue our discussion of moles by looking at another conversion factor where moles appear. Remember that moles help us to go from really really tiny levels – like atoms and molecules – up to something that we can measure on a scale. In this video, we're going too see how we can use the concept of the mole to convert from atoms to grams – something that can be measured in the lab.

The new conversion factor that I want to introduce comes from the periodic table. We need to rethink the way we look at this average atomic mass. That is now going to be thought of as a molar mass. That is the amount of mass in one mole of a substance. So the average atomic mass from the periodic table is going to be a conversion factor to help us convert between grams and moles.

In the example of nitrogen, nitrogen has a mass of 14.01 grams per 1 mole. That is a conversion factor that we can flip however we need it. So you can write it this way, where grams is on top and moles is on the bottom. Or you can write it this way, where moles is on top and grams is on the bottom.

Suggested Question (Time 1:17)
What are the units for molar mass?

Let's look at an example where we are going to convert between moles and grams using the molar mass off of the periodic table.

The question is how many grams of lead are in 1.42 moles of lead?

You always want to start with the number and the unit that you are given in the question. In this case, I have started with 1.42 moles of lead.

And then you would set up a conversion factor where you are going to have the units you start with cancel out, and the units you are trying to go to on the top. What I mean by that – we started with moles of lead. In my conversion factor, I need moles of lead to be on the bottom. That way moles of lead is going to cancel and I am converting to grams of lead.

This number came off of the periodic table. That is the average atomic mass for lead.

At this point I can see that moles is going to be cancelling out and my answer is going to be in grams of lead, which is what I was hoping for.

I am ready to put things into my calculator. When you do this you always want to multiply by what is on the top and divide by what is on the bottom.

So in my calculator, I'm going to say 1.42 times 207.2 divided by 1, which will give me 294 grams of lead. Moles has cancelled leaving my answer in grams.

Suggested Question (Time 2:43)
Where can the conversion factor between moles and grams be found?

Let's look at another example, but this time we are going to use the conversion factor the other way. This time we are going to convert from grams to moles.

The question is how many moles of zinc are in 5.78 grams of zinc?

Once again you want to start with the number and the unit that you are given in the question. So start with 5.78 grams of zinc.

Then you want to set up a conversion factor and in your conversion factor, whatever unit you started with needs to go on the bottom. So this time I have put the molar mass on the bottom, per 1 mole on the top. This time when I multiply across and divide down, I would put in my calculator 5.78 times 1 divided by 65.38 giving me .884 moles of zinc as my final answer.

And notice that my answer units match the answer on top that didn't cancel out. These units have cancelled out.

Suggested Question (Time 3:47)
When converting from grams to moles, what unit should go on the bottom of your conversion factor?

Here is another example, but this time we are going to tack on what we saw in our last video. In our last video we learned how to convert from moles to atoms or from atoms to moles. Now we are going to go all the way from grams to atoms.

Every calculation that we do in the concept of stoichiometry has to go through moles. We are going to go from grams to moles, and then from moles to atoms. It is going to be a 2 step process.

So here is the question. How many atoms are in an aluminum can with a mass of 14.1 grams? So if you take an average aluminum can, it has a mass of 14.1 grams.

Follow your process. Always start with the number and the unit that you are given in the question. I'm starting with 14.1 grams of aluminum.

Next, whatever unit I have here needs to go on the bottom of this fraction. So I'm going to use the mass from the periodic table to convert from grams of aluminum to moles of aluminum. Grams needed to go on the bottom so that grams would cancel out.

Now that I'm in moles, I'm going to use Avogadro's number – 6.022x10^23 – to go from moles to atoms. Notice that since moles was on the top here, it needed to go on the bottom here so that it would cancel out.

At this point everything in pink is going to cancel. Grams has cancelled. Moles has cancelled. Leaving me in units of atoms.

I'm ready to put this into my calculator. I'm going to multiply by everything on the top and divide by everything on the bottom. I'm going to say 14.1 times 1 times 6.022x10^23 divided by 26.98 and divided by 1. That gives me 3.15x10^23 atoms of aluminum.

I know that is a big number, but you are counting atoms. They are really really small so there is a whole bunch of them.

Just a few things to reiterate about the process. You always start with the number and the unit from your question and then whatever unit you start with has to go on the bottom of the next fraction. Whatever unit is on the top of one fraction goes on the bottom of the next fraction.

And every step is going to involve moles somehow. We are always either going to be going to moles or going from moles.

Suggested Question (Time 6:29)
Every step (or fraction) in the calculation involves the unit of ___________ somehow.

Here is a similar example but this time we are going the opposite direction. This time the question is how many grams are in a sample of copper containing 1.25x10^24 atoms?

This time we are starting with atoms and going to grams. It looks similar except for my fractions are going to be flipped. You always want to start with the number and the unit that you are given in the question. 1.25x10^24 atoms.

Atoms is what I started with so I need atoms of copper to be on the bottom of the next fraction. The only place you can go from atoms is to moles, using Avogadro's number as a conversion factor.

At that point, atoms has cancelled out and I am in units of moles. So now I'm going to have moles on the bottom of my next fraction and I'm going to go from moles to grams. When I go from moles to grams, I use the average atomic mass from the periodic table. So 63.55 grams per mole is now the molar mass of copper.

Take a step back and make sure that your units make sense. Atoms has cancelled out. Moles has cancelled out. I've gotten to my desired unit of grams.

I'm ready to put this in my calculator by multiplying by everything on top and dividing by everything on the bottom. But you should be cautious because you have an exponent on the bottom. You want to be very careful how you put that in your calculator.

I like to use the EE function. Here is my whole process for putting this in the calculator. I'm going to say 1.25 EE – so second EE – 24. That's 1.25x10^24 times 63.55 and then I'm going to divide by 6.022 EE – so second EE – 23. Hit equals and you should get 132 grams of copper.

So as you are going through this process, you want to make sure each fraction goes through moles. The tools that you have are Avogadro's number, which can go from moles to atoms, and you have the molar mass off of the periodic table, which can go from grams to moles.

And then if you will follow my process for whatever unit you have here goes on the bottom here, and then whatever unit you have here goes on the bottom here. That will help you figure out which direction the conversion factors need to be, because remember they are flappable.

You always want to take a step back, make sure that your units cancel properly to give you your desired unit, and then always be careful when you have exponents in the denominator. You want to make sure that you either use parenthesis in your calculator or use this EE function that I showed you.

That finishes up for us how we convert between atoms to moles and moles to grams. In the next video we are going to take a deeper look at this concept of molar mass, and what to do if you're calculating the molar mass of a compound rather than just an element.

Suggested Question (Time 10:00)
True or False: Converting from atoms to grams only requires one conversion factor.

Unit 5, Video 11 – Molar Mass

Molar mass is a number that we can get from the periodic table and we can use that number to convert back and forth between moles and grams, or grams and moles. And remember that a mole is something that allows us to convert between the microscopic world of atoms up to the laboratory world where we can measure things on a scale.

I introduced the concept of molar mass in my previous video, but I want to spend a little more time on it now. Because in the previous video, we were looking at the molar mass of single elements. Now I want to look at the molar mass of whole compounds.

For example, if you were asked to find the molar mass of nitrogen dioxide, NO_2, you would want to find nitrogen and oxygen on the periodic table and add their masses up.

There was 1 nitrogen in NO_2. It has a mass of 14.01. 1 times 14.01 is 14.01.

There are 2 oxygens in NO_2. So this 2 right here is going to get multiplied by the molar mass for oxygen, or the average atomic mass for oxygen. 2 times 16 gives me a 32.

When I add those 2 together, I have a total molar mass for this compound of 46.01 grams per mole.

So when you're dealing with molar mass, the units are grams per mole.

Suggested Question (Time 1:33)
Molar mass is a value for each element that can be found on the ______________________.

This is a conversion factor. It is something we can use to convert back and forth between grams and moles.

Because it is a conversion factor, you can write it as 46.01 grams per 1 mole. Or you can flip it where the 46.01 is on the bottom and think of it as 1 mole per 46.01 grams.

Just to recap so far, we're adding up the numbers – the average atomic masses- off of the periodic table. These are going to have a unit of grams per mole, and it is a conversion factor that is flappable.

We are going to spend a few minutes practicing calculating molar mass and then we will take a look at how you can use it as a conversion factor.

Suggested Question (Time 2:20)
True or False: The units of molar mass are grams per mole.

Here is some molar mass practice.

To begin with, find the molar mass of chlorine. Remember that chlorine is a diatomic, meaning that it has a 2. So the subscript is a 2 – I need to multiply 2 by the number for chlorine off of the periodic table. So I said 2 – that's the subscript – times 35.45, which is the molar mass off of the periodic table.

That gives me 70.90. That is in units of grams per mole.

2.54

The next one. Find the molar mass of calcium nitride. Calcium, there are 3 of them – so 3 times the number for calcium off of the periodic table, which is 40.08 – is 120.2. That's for calcium.

And then for nitrogen, there are 2 nitrogens times 14.01 off of the periodic table, gives me 28.02.

Add these 2 together. Calcium nitride has a total molar mass of 122.2 grams per mole.

3.29

Finally we have calcium phosphate. Calcium phosphate has 3 calciums and 2 phosphates. This 2 gets distributed inside the parenthesis. You're going to see that there is going to be 3 calciums, 2 phosphorous's, and then 2 times 4 which is 8 oxygens.

So when you add that up and use the numbers off of the periodic table: 4 times 40.08 gives me 120.2. 2 times 30.9 gives me 61.94, and then remember that these numbers got multiplied. So there were 8 oxygens times the 16 off the periodic table gives me 128. Add these 3 together. Calcium phosphate has a total molar mass of 310.1 grams per mole.

Next we will see how we can use this molar mass as a conversion factor to go back and forth between grams and moles.

Suggested Question (Time 4:31)
How many hydrogen atoms are in $(NH_4)_2S$?

Here is an example question. How many grams of calcium chloride are in 0.325 moles of calcium chloride. You want to start always with the number and the unit that you are given in the question – so 0.325 moles of calcium chloride is the first thing that I'm going to write.

At this point, I want to switch from moles to grams. I can do that using the molar mass off of the periodic table. That is a conversion factor that allows me to make that switch.

I'm going to calculate the molar mass. I find calcium off of the periodic table. I have 1 calcium in my formula, so 1 times 40.08 is 40.08. I have 2 chlorines in my formula, so 2 times the number for chlorine off of the periodic table – 35.45 – gives me 70.90. Calcium chloride has a molar mass of 110.98 grams per mole.

That is a conversion factor that can be switched. I can write it as 110.98 grams per 1 mole on the bottom, or I can have 1 mole on the top divided by 110.98 grams.

The process of: whatever unit I start with here goes on the bottom here – is what is going to dictate how I write that conversion factor. Since I started with moles, I want moles to cancel out. I'm going to write this conversion factor so that moles is on the bottom and my 110.98 grams is on the top.

At that point everything is set up so that my units are going to cancel – moles cancels and I end up with grams left over. My final answer is going to be in grams of calcium chloride.

I'm going to take my calculator and multiply by everything on the top, divide by everything on the bottom. 0.325 times 110.98 divided by 1 gives me 36.1 grams of calcium chloride.

If this had been a question where we had started with grams and were calculating to moles, I would've ended up with this conversion factor being flipped. This is just an example of how you can use molar mass to convert from moles to grams or from grams to moles.

Suggested Question (Time 7:11)
True or False: When converting from grams to moles, the unit of moles should be on the bottom of the conversion factor.

Unit 5, Video 12 – Empirical and Molecular Formulas

This video is going to focus on empirical versus molecular formulas. These are two types of formulas that can be used to represent a compound.

The difference between the two is that an empirical formula is always going to be reduced as much as possible. The numbers are in their lowest possible form.

Whereas a molecular formula is not necessarily in the lowest possible form. A molecular formula is going to count the absolute number of atoms present.

0:32

Let's look at some examples. This is a molecular formula – N_2O_4. Notice that 2 and 4 are both divisible by 2. You could reduce those numbers to get to a lower form. So the empirical formula for this molecular formula is NO_2, where both of the subscripts have been reduced. They have both been divided by 2.

Here is another example. C_6H_{14} is a molecular formula. It is not in the lowest possible form. Both of those numbers are divisible by 2. So the empirical formula for this molecular formula is C_3H_7.

A final example is $C_6H_{12}O_6$ where all of these numbers are divisible by 6 – which would give an empirical formula of CH_2O.

> Suggested Question (Time 1:26)
> What is the empirical formula for the following molecular formula? C_6H_6

Here is an example where we have been asked to calculate an empirical formula. The question is: A compound contains 38.67% carbon, 16.22% hydrogen, and 45.11% nitrogen. Find the empirical formula.

The first thing – step 1 – is to write those percentages in units of grams. What you are doing here is assuming that you have a 100 gram sample. So 38.67% carbon becomes 38.67 grams of carbon, 16.22% hydrogen becomes 16.22 grams of hydrogen, Notice that I am not moving any decimals to go back and forth between percentages. I'm simply taking away the percent sign and writing grams instead. That is step 1 for finding an empirical formula.

Step 2 is to convert from grams to moles. I'm going to take this number of grams, divide it by the molar mass off of the periodic table to find the number of moles for each of these atoms. So I've converted to moles for carbon, hydrogen, and nitrogen – in all cases by dividing by the molar mass off of the periodic table.

Step 3 is to divide all of the mole values by the smallest mole value. Here are my mole values: 3.22, 16.09, and 3.22. Which number is the smallest? In this case 3.22 is the smallest. So I'm going to divide all of those numbers by 3.22.

I did it for carbon and got a 1. For hydrogen, 16.09 divided by 3.22 gives me 5. And for nitrogen, 3.22 divided by 3.22 gives me a 1.

These numbers that I have found become subscripts in your equation. There is going to be 1 carbon, 5 hydrogens, and 1 nitrogen in your empirical formula.

In this case, all of these numbers were whole numbers, so they were perfectly fine to go into the empirical formula just as they were. However, if one of them had not been a whole number, you could multiply all of them by something to get them up to a whole number.

Suggested Question (Time 3:59)
True or False: While calculating empirical formulas, it is necessary to move the decimal when writing percentages as grams.

Now that we have seen how to calculate an empirical formula, let's look at how we can take an empirical formula and turn it into a molecular formula.

This is the process we are going to be following. You take the molar mass of the molecular formula – and this is usually given to you in the question – and you divide it by the molar mass of the empirical formula, which you very often have to calculate.

This is going to give you some whole number and that is the whole number that you're going to multiply your empirical formula by.

Let's look at an example. A compound has a molar mass of 220.1 grams per mole and an empirical formula of CO_2. What is the molecular formula?

Let's take a peak at the information we are given. This compound – the molecule – has a molar mass of 220.1 grams per mole. That is the molar mass of the molecular formula. It is not super obvious the way its written, but that is the molar mass of the molecular formula. That is going to be my top number in this equation.

Then they gave me the empirical formula. I can use this information to calculate the molar mass of the empirical formula.

At this point, I'm going to go to the periodic table and calculate the molar mass of the empirical formula. Carbon dioxide has 1 carbon and 2 oxygens, giving it a total molar mass of 44.01. That is going to be my bottom number in this equation.

My top number was given to me in the problem. My bottom number I have calculated. I've plugged those in and divide them.

This should give you a whole number. My whole number – the number that it gave me – was a 5. This is the number I'm going to multiply my empirical formula by.

My empirical formula was CO_2. There was 1 carbon times 5. In my molecular formula, there are 5 carbons. My empirical had 2 oxygens, times 5. In my molecular formula, there are 10 oxygens.

I have an empirical formula that I have multiplied by some whole number to get my molecular formula. This is the equation that you want to use to figure out what to multiply your empirical formula by to get your molecular formula.

This, in my opinion, is one of the kinder examples because they have given you the empirical formula. Next what we are going to look at is an example where we have to calculate the empirical formula and then calculate the molecular formula.

Suggested Question (Time 6:42)
True or False: The molar mass given in the question is usually the molar mass for the molecular formula.

Here is an example where we are not given an empirical formula. In some situations you are required to determine the empirical formula before you can calculate the molecular formula.

Here is the question. Nicotine contains 74.0% carbon, 8.70% hydrogen, and 17.3% nitrogen. It has a molar mass – nicotine, the molecule – it has a molar mass of 162.2 grams per mole. What is the molecular formula for nicotine?

So, we are going to follow the steps for writing an empirical formula. Step 1 is to write your percents as grams. Each one of these percentages, I'm taking the % sign off and writing them as grams. I now have 74 grams of carbon, 8.7 grams of hydrogen, and 17.3 grams of nitrogen.

Step 2 for writing an empirical formula is to convert your gram values to moles. For carbon, hydrogen, and nitrogen I have divided by the molar mass off of the periodic table to get these mole values. 6.16 molse of carbon, 8.63 moles of hydrogen, and 1.23 moles of nitrogen.

Step 3 for an empirical formula is to take whichever mole value is the smallest – in this case nitrogen – and divide all of the mole values by that value. Since nitrogen was the smallest, I have taken each of the mole values and divided them by 1.23. 6.16 divided by 1.23 gives me a 5. 8.63 divided by 1.23 gives me a 7, and then 1.23 divided by itself gives me a 1.

These are the values that become the subscripts in my empirical formula. My empirical formula for nicotine is C_5H_7N.

It goes on to tell me that nicotine has a molar mass of 162.2 grams per mole. I am going to be turning this empirical formula into a molecular formula that must have this molar mass.

Here is the equation I want to use. I want to take the molar mass of the molecular formula and divide it by the molar mass of the empirical formula.

The molar mas of the molecular formula was given to me in the question. The molar mass of the empirical formula I can calculate because here is my empirical formula.

I can go to the periodic table, add up the masses of 5 carbons, 7 hydrogens, and 1 nitrogen. That has a molar mass of 181.12 grams per mole. That is my bottom number.

Then I take the molecular formulas mass, divide it by the empirical formulas mass, and at this point I should get a whole number. In this case my whole number is a 2.

That is the number that I'm going to multiply my empirical formula by. Here was my empirical formula. It had 5 carbons, 5 times 2 is 10. My molecular formula has 10 carbons.

The empirical formula had 7 hydrogens, times 2, my molecular formula has 14.

Nitrogen just had 1 in the empirical formula, times 2. The molecular formula has 2 nitrogens.

The molecular formula for nicotine – the actual molecule – has 10 carbons, 14 hydrogens, and 2 nitrogens. The empirical formula for nicotine if you were to ratio those down to the smallest value would be C_5H_7N.

Suggested Question (Time 10:45)
Consider a molecule that has an empirical formula of CH_4 and a molecular formula of 32. What is the molecular formula for the molecule?

Unit 5, Video 13 – Mole Ratios and Stoichiometry

In previous videos, we have seen how we can use the concept of the mole to convert back and forth between grams and moles or moles and grams. Now we want to build on that – we are really getting to the heart of stoichiometry where we are going to be using mole ratios to switch substances within a chemical reaction.

What do I mean by that? Suppose you are given information about oxygen, but you are trying to calculate information about water. That is switching substances. In order to do that, you have to have a balanced chemical reaction.

This is a combustion reaction. It has already been balanced for you.

The mole ratios, these coefficients, become conversion factors that allow us to switch between substances. Mole ratios can be used to help convert between substances in a chemical reaction.

Some examples of mole ratios. There are 5 moles of oxygen – that comes from the coefficient – also from the coefficient.

That is a conversion factor that can be flipped. I can write that with oxygen on top, or I can write that with water on top. Just depending on how I need my units to cancel. Let's look at some examples.

Suggested Question (Time 1:23)
The ____________________ in a balanced chemical equation provide the numbers for a mole ratio.

Here is my first stoichiometry example where I am converting from moles of one thing to moles of something else within a chemical reaction.

Here is the question: Using the chemical equation shown, how many moles of water can be produced when 2.05 moles of oxygen are consumed?

We are given 2.05 moles of oxygen and we are trying to figure out how many moles of water will be produced.

You always want to start with the number and the unit that you're given in the question, so 2.05 moles of oxygen is the first thing that I'm going to write.

I'm in units of moles, which is great because when you are in units of moles you can switch substances within a chemical reaction.

Whatever unit I start with needs to go on the bottom of the next fraction. Moles of oxygen is what I started with, I'm putting moles of oxygen on the bottom of the next fraction.

I'm going to use a mole ratio. I'm going to go from moles of oxygen to moles of water using the coefficients as my conversion factor. This 5 became the mole value for oxygen. This 4 – the coefficient for water – became the mole value for water.

At this point, I'm ready to put things in my calculator: 2.05 times 4 divided by 5 gives me 1.64 moles of water. Oxygen has cancelled out leaving me with water as my answer.

I've used a mole fraction to switch from one thing to something else within a chemical reaction.

Suggested Question (Time 3:09)
True or false: When using a mole ratio to convert from oxygen to water, the coefficient for oxygen should go on the bottom of the mole to mole conversion factor.

Let's build on that a little bit. Suppose you are given grams of one thing and you are supposed to convert to grams of something else. This is more of a real life example, because we have scales that can measure in grams.

Here is the question: Using the chemical equation shown, how many grams of carbon dioxide – carbon dioxide is over here – can be produced when 8.45 grams of propane, C_3H_8, are consumed?

We are starting with information about the C_3H_8 and we are converting to CO_2.

You always want to start with the number and the unit that you are given in the question, so 8.45 grams of C_3H_8. At this point, whatever unit you start with goes on the bottom of the next fraction. Grams of C_3H_8.

The only place you can go from grams is to moles of the same substance. Moles of C_3H_8 goes on the top. Whenever you convert between grams and moles, your conversion factor is from the periodic table. 44.09 is the molar mass off the periodic table. That is adding up that masses of 3 carbons and 8 hydrogens.

At this point, you're in units of moles of C_3H_8. We have cancelled out the grams.

Once you get to moles you are ready to switch substances. Whatever unit I have here goes on the bottom of this fraction. Moles of C_3H_8 goes on the bottom. Moles of what I'm trying to convert to - the question wanted carbon dioxide – that goes on top.

These numbers, this is a mole ratio. In a mole ratio, your numbers come from the coefficients in a balanced reaction. C_3H_8 didn't have anything in front of it. That means it is a 1. CO_2 has a coefficient of a 3. Those numbers become the mole ratio that allows to switch from C_3H_8 to CO_2.

At this point, I'm in moles of carbon dioxide. That's great. We are in the right substance now but the question wants grams of carbon dioxide. We need one more step where whatever unit I have here goes on the bottom here.

I have moles of Carbon dioxide here. I'm going from moles to grams. Whenever you do that, your conversion factor is the molar mass off of the periodic table. I'm going to add up 1 carbon and 2 oxygens off of the periodic table.

Notice that the molar mass was on the bottom over here and on the top over here. How did we know to do that?

Follow your units. Whatever unit you have here must go on the bottom here. Whatever unit is here should be on the bottom here.

You should also be writing in your substances. C_3H_8 is going to cancel out grams of C_3H_8. Moles of C_3H_8 is going to cancel out moles of C_3H_8.

The only time you can switch substances is in a mole to mole ratio. For ever other fraction, your substances should match.

So this is a full stoichiometry question where you are switching from grams of one thing to grams of something else.

Suggested Question (Time 6:48)
True or False: It is acceptable to switch substances in the grams to mole step of a stoichiometry problem.

Here is that same example – that same set up that we just saw, but I have color coded it for you. So you can see that the only fraction containing two different substances is the mole to mole fraction.

I've got C_3H_8 in pink each time and CO_2 in blue each time. In this mole to mole fraction, that is where I switched from C_3H_8 to CO_2. Before that everything was about C_3H_8. After that everything was about CO_2.

That mole ratio is crucial for switching from one substance to another substance within a chemical reaction.

Suggested Question (Time 7:28)
True or False: The only time you can switch substances within a balanced chemical equation is when using a mole to mole conversion factor (or mole ratio).

And the beauty of this is now labs become quantitative. At this point, you can do a lab where you put two things together and you can predict the amount of something that you are going to produce. Because we now know how to switch substances.

Unit 5, Video 14 – Limiting Reactants

Let's look at the concept of limiting reactants. This is the idea that when you have a chemical reaction going on and you put two things together, one of them is going to run out first. One of them is going to limit how much of your products can be produced.

The thing that runs out first is your limiting reactant.

Here is an example. In this reaction hydrogen plus oxygen is combining to make water. When you balance that, there are 2 hydrogen molecules plus 1 oxygen molecule, combining to produce 2 of these water molecules.

If you begin this reaction with 4 hydrogen molecules and 4 oxygen molecules – hydrogen runs out first. Notice that we need 2 hydrogen molecules for every 1 oxygen molecules. Hydrogen runs out first. That limits the amount of water that can be produced.

The other thing that I notice is that there are 2 unreacted oxygen molecules after the reaction takes place. 2 of these guys didn't get used.

Hydrogen was the limiting reactant. Oxygen is the excess reactant.

What I want to show you in this video is how you can recognize a limiting reactant problem, a few steps to help you identify your limiting reactant, and then how to solve the question.

Suggested Question (Time 1:32)
The reactant that runs out first in a chemical reaction is the ________________ reactant.

Here is an example. How many grams of ammonia can be produced when 10 grams of nitrogen and 10 grams of hydrogen react? Here is my chemical reaction. Nitrogen plus hydrogen are reacting to produce ammonia. The reaction has been balanced.

Here is how I know this is a limiting reactant problem. I was given information about 2 of my reactants. I was given information about nitrogen and information about hydrogen. That is my clue that this is a limiting reactant problem. Because how do I know which one of those pieces of information to start with?

Suggested Question (Time 2:13)
True or False: Limiting reactant problems can be recognized by the fact that information is given about more than one reactant.

There are some steps that you can use to help you identify your limiting reactant.

The first thing you want to do is convert to moles. I have taken both of the pieces of information that I was given and converted them to moles. 10 grams of nitrogen divided by the molar mass off of the periodic table gives me 0.357 moles of nitrogen. 10 grams of hydrogen divided by the molar mass off of the periodic table for two hydrogen atoms gives me 4.96 moles of hydrogen. I have converted to moles.

The second thing that I want to do is to divide by the coefficients. For nitrogen, I took that number of moles and divided by the coefficient for nitrogen. I'm going to do the same thing for hydrogen. Hydrogen, I had 4.96 moles, divided by the coefficient for hydrogen, gives me a number of 1.65.

These numbers are pretty much meaningless except they can help us to identify the limiting reactant. Whichever one is the lowest is your limiting reactant.

In this case, I now know that nitrogen is limiting and hydrogen is the excess reactant. I'm not going to use these numbers for anything else, other than identifying the limiting reactant. I have not answered this question yet. All I've done is said which one is limiting.

Now, step 3. I use the limiting reactant to answer the original question. My original question was how many grams of ammonia can be produced? I'm going to start with my information about nitrogen because nitrogen was limiting.

At this point, it looks like a normal stoichiometry problem, where I'm starting with the limiting reactant. 10 grams of nitrogen, using the molar mass of nitrogen off of the periodic table to convert to moles, using a mole ratio between nitrogen and NH_3 to convert to NH_3, and then using a molar mass for NH_3 to get back to grams.

My final answer is 12.2 grams of NH_3. So step 3 looks just like a normal stoichiometry problem, but I needed to go through these other steps to figure out what to start with.

Suggested Question (Time 4:55)
True or False: After identifying which reactant is limiting, the excess reactant is used to start the stoichiometry calculations.

Unit 5, Video 15 – Percent Yield

This video is going to cover the concept of percent yield, and percent yield is really a lab concept. You are in a lab. You have predicted how much of something you are going to produce, and then you are going to measure how much you actually produced.

So you have what you predicted that you would product, and then what you actually produced. We are going to use those two things to calculate our percent yield.

This is the formula for percent yield. It is actual yield – that's what you did in the lab – divided by the theoretical yield, times 100 to make it a percentage.

In this example, stoichiometry predicted that a reaction would produce 2.74 grams. This was before you did the lab, you ran the math and you predicted how much of something you were going to produce. That is the theoretical yield. The theoretical yield comes from the calculations.

Then, you actually ran the experiment, and you produced 2.57 grams. That is the actual yield. This would come from a lab experiment, or that number would need to be given to you.

We have our theoretical yield, which we have calculated and our actual yield which we measured on a scale during the lab.

We can plug those in to the percent yield formula. My actual yield was 2.57 grams, divided by my theoretical yield of 2.74 grams, times 100. The percent yield of my experiment was 93.8%.

Let's look another example.

Suggested Question (Time 1:46)
____________________ yield is the expected yield based off of stoichiometric calculations.

Here is the question. 4.39 grams of methane were burned in the presence of oxygen, producing 9.02 grams of water. What is the percent yield of the reaction?

We are trying to find a percent yield. Go ahead and write out the formula for percent yield. That is actual yield divided by theoretical yield, times 100. Then take a look at what you are given.

We started with 4.39 grams of methane, of CH4, and we were also given how much was actually produced by the reaction – the actual yield was 9.02 grams of water.

The theoretical yield is something we need to calculate. We need to take this grams of methane that we were given, and convert it to grams of water.

I want you to notice that your actual yield and your theoretical yield should have the same unit. So 9.02 grams of water. My theoretical yield should have units of grams of water.

That kind of helps me figure out where I'm trying to go in this question.

I want to start with what I was given. I'm going to convert from grams of CH_4 to moles of CH4 using the molar mass. I'm going to use a molar ratio to convert from CH4 to H2O, and them I'm going to use the molar mass of H2O to convert to grams of H2O. At that point, I have 9.86 grams of water. I have used stoichiometry to find my theoretical yield. My theoretical yield and my actual yield have the same unit.

At that point, I can plug everything in. 9.02 divided by 9.86 times 100 gives me a percent yield of 91.4%.

Percent yield is comparing what you expected to produce according to the stoichiometry, to what you actually produced.

Suggested Question (Time 4:10)
True or False: Actual yield and theoretical yield will have the same units.