

## Transcript: Bonding

### Unit 4, Video 1 – Ionic Bonding and Nomenclature

This is Rebecca from ChemistryIsMyJam.com. We are going to dive right into Unit 4 and talk about ionic bonding.

As the name suggests, ionic bonding has to do with ions. Let's take a look at how and why ions form.

Recall the octet rule. The octet rule tells that atoms want to have 8 electrons in their outermost energy level.

Sodium has 2 electrons in its first, 8 in its second, and 1 electron in its outermost energy level. We call that a valence electron. Chlorine has 7 valence electrons.

Sodium can give chlorine its 1 valence electron. Let's take a look at how things will appear when that takes place. That electron is now over here on chlorine, leaving an empty energy level for sodium. The outermost energy level for sodium will disappear because it is no longer necessary. That leaves the next energy level down as a full outer energy level for sodium.

You've created a situation where both atoms have satisfied the octet rule. You have also created two ions. Sodium has the same number of protons in its nucleus that it started with however it just lost an electron. It has got the same number of positive things but one less negative thing. Sodium will have a positive 1 charge. Chlorine just gained an electron so chlorine will have a negative 1 charge.

These are 2 oppositely charged particles that will be attracted to one another. That is the basis for an ionic bond. Ionic bonds are forces of attraction between oppositely charged particles.

Suggested Question (Time 1:54)

How many valence electrons will the sodium ion have?

The key to forming an ionic bond is that electrons have to be transferred. It can be one electron like the example we just saw or it can be multiple electrons being transferred.

Ionic bonds are usually made up of elements on the opposite side of the periodic table. They will contain a metal from the left side of the periodic table and a nonmetal from the right side of the periodic table.

Compounds formed from ionic bonds are usually solid crystals at room temperature. In the example you see here, there is an alternating series of positive and negative charges. By alternating between the positive and the negative, you maximize the attraction between particles that have opposite charges. You also minimize the repulsion felt between particles that have the same charge. So they line up in this alternating series that causes a solid crystal to form at room temperature.

In terms of other physical properties, ionic compounds have high melting temperatures and they can conduct electric current when they are dissolved in water.

Suggested Question (Time 3:07)

True or False: Chlorine and fluorine are both in group 17 on the periodic table. It is likely that an ionic bond would form between them.

Let's take a look at how ionic compounds are named. Throughout this unit we are going to be talking about both molecular and ionic compounds. But for now, we are focusing on ionic so we are going to stay on the top part of this flow chart.

The first question you would want to ask yourself when you see a compound that you want to name is: Is there a metal? If there is a metal, then you are likely dealing with an ionic compound and you would want to follow the set of rules that we are about to cover.

You would also want to ask yourself: Is there a transition metal, Sn, or Pb? Because these are a special case that we are going to cover in a minute. We are going to start with the simplest case where there is not a transition metal, the element Sn, or the element Pb. So we are going to go to here. When that

is the case, you list the cation, which is the positive charge, followed by the anion ending in "ide."

4:05

So here are some examples of some ionic compounds that do not contain a transition metal, Sn, or Pb. These are the simplest type of ionic compounds to name. These are binary ionic compounds because they are only made up of two elements: a metal and a nonmetal.

We are going to name them by writing first the name of the cation – the positive substance – and then the name of the anion, and we are going to make the anion end in "ide."

For example, NaCl. Na is sodium. Cl is chlorine, but we want to make chlorine end in "ide." So NaCl is sodium chloride.

$\text{Li}_2\text{O}$  is lithium oxide. We made oxygen end in "ide."

KI is potassium iodide. We have named the cation and then made the anion end in "ide."

Suggested Question (Time 5:05)

Name the following binary ionic compound:  $\text{CaF}_2$

We are going to find that charge is very important in ionic compounds. Understanding the charge of the individual ions helps you predict how many of each ion are present.

We saw in a previous video that there is a trend on the periodic table that helps you predict the charge atoms take when they become an ion. The alkali metals in group 1 took a charge of +1, these are +2, +3, + or -4, -3, -2, -1, and 0.

Then you have this section in the middle – the transition metals. The transition metals can take multiple charges. Because of that there needed to be a separate naming system for the transition metals that indicated the charge of those ions.

So when you see a transition metal in an ionic compound and you are trying to name that compound, you are going to put the charge of the cation – the charge of the positive ion in roman numerals.

6:10

Since we will be using roman numerals, here you have a quick review of what the roman numerals look like for the numbers 1 through 6.

We are going to be naming this compound that you see here,  $\text{Fe}_2\text{O}_3$ , and I want to go through those questions that we saw on the flow chart and kind of go through my thought process to name this compound.

The first question that I would ask myself: Is there a metal? Yes. Iron is a metal. It is on the left hand side of the periodic table. That means that it is an ionic compound, so then I go to the next question.

Is there a transition metal, Sn, or Pb. Yes. Iron is a transition metal, so that tells me that I need to list the cation, list the charge of the cation in roman numerals, and then list the anion ending in "ide." That is what I have done here.

However, we were not given directly the charge of the iron. We had to figure it out. It is often necessary to use the charge of the anion to figure out the charge of the transition metal.

Next, I'm going to walk you through my thought process for determining the charge of the iron. I want to show you where this 3 came from.

First of all, I want to point out that the overall compound should have a neutral charge. Otherwise, a charge would be written. So when I start figuring out the charges of the ions, they need to total up to a zero charge.

We know the charge of the oxygen. Oxygen is in group 16 on the periodic table and I expect ions to have a -2 charge if they are from that group. There are 3 of those oxygens, so I'm going to say -2 times 3.

That gives me a total negative charge of -6. We also said that the total charge of the compound should be neutral. So anything I do here needs to add up to zero.

We can then figure out what the total positive charge should be. What minus 6 equals zero? That would be +6, so I need a total positive charge of 6 to balance out the negative 6 charge. That positive 6 gets distributed over those 2 irons. So now I'm going to say 6 divided by 2. Each one of the irons should have a charge of +3. That's where the 3 came from.

We can quickly look at one more example. Consider another version of iron oxide. It is still iron "something" oxide but the charge of the iron has changed.

Oxygen continues to have a charge of minus 2. There is only 1 oxygen this time so there is a total negative charge of -2. The whole thing needs to add up to zero, meaning that there must be a total positive charge of +2. That +2 gets distributed this time between only 1 iron. So that iron in this case has a charge of +2. This would be iron(II) oxide. There would be a two for the roman numeral.

Suggested Question (Time 9:28)

Name the following binary ionic compound:  $\text{CuF}_2$

Let's shift gears and talk about how to write formulas for ionic compounds. In this case, you are given the name and you are trying to figure out the formula. I mentioned earlier that the charges are important. The charges help you figure out how many of each ion need to be present.

So you have a periodic table here with the charges written. Lithium is right here in group 1. It has a charge of +1, so I am writing that above it. Oxygen is in over here in group 16 with a charge of -2. I have written that above. If you will cross down those charges. By crossing down those charges, the 2 comes to here, the 1 comes to here.

Notice that I did not cross down the sign of the charge. I only crossed down the charge itself. I did not put a -2, I just put a 2. By crossing down the charges, now I have 2 of these +1 lithiums and 1 of these -2 oxygens. That makes the compound balance out to zero. By crossing down the charges you end up with a neutral compound.

Let's look at one more example. Suppose you are trying to write the formula for magnesium chloride. Start by writing the symbols for the elements. You've got magnesium and chlorine in this compound.

Then write the charges of those elements above it. Magnesium is in group 2. It has a charge of +2. Chlorine is in group 17. It has a charge of -1. Then you want to cross down those charges. The 2 comes to here. The 1 comes to here but you do not have to write the 1. The 1 is assumed so if you write  $\text{MgCl}_2$ , that is a correct answer.

11:19

The final thing that you want to consider when writing formulas for ionic compounds is that the subscripts should be reduced as much as possible. In the first example that you have here – magnesium oxide – magnesium is Mg, oxygen is O.

I'm going to write their charges above it. Magnesium is in group 2, so it has a charge of +2. Oxygen is in group 16 having a charge of -2. Those charges already add up to 0. When that is the case, you don't need to cross them down.

If I had crossed those down I would have had a 2 here and a 2 here, which is then reducible to a 1 and a 1. The correct answer is MgO. You want to reduce the subscripts as much as possible.

In the second example, copper(II) sulfide, copper is Cu which is located right here on the periodic table. That is a transition metal so I needed to give you the charge of the copper. It is right there. Copper has a charge of +2. Sulfur is in group 16. It has a charge of -2. That is a +2, -2 situation. It already adds up to zero. No subscripts are necessary for this compound. CuS is the correct answer.

That concludes our discussion of ionic bonding and nomenclature for ionic compounds. Stick around as we move on to polyatomic ions.

Suggested Question (Time 13:05)

Write the formula for iron(III) chloride.

## Unit 4, Video 2 - Polyatomic Ions

This is Rebecca from ChemistryIsMyJam.com. In this video we are going to be talking about polyatomic ions.

Poly means “many,” atomic, so “many atoms,” and then ions indicates that there is a charge. Polyatomic ions are groups of atoms that stay together and have a charge.

The example that you have in front of you is the nitrate ion. The nitrate ion always has a nitrogen with three oxygens around it, and the whole thing has a charge of minus 1. So the nitrate ion is  $\text{NO}_3$  with a charge of minus 1.

Because polyatomic ions are ions, they participate in ionic bonding. In the previous video, we looked at binary ionic compounds. Those are ionic compounds that only have 2 elements. Whenever you see an ionic compound that has more than 2 elements, a polyatomic ion is likely present.

Suggested Question (Time 1:03)

True or False: Polyatomic ions often participate in ionic bonding.

The polyatomic ions are one of those things that never really go away in chemistry. It is helpful if you take the time to get familiar with them and maybe even memorize them.

There are a lot of different lists available of common polyatomic ions and I'll put a link in the description to one list.

One strategy for memorizing the polyatomic ions is to memorize the “ate” version of the ions. That would be the ions that end in a-t-e. All of the ones that you have on the screen in front of you are “ate” polyatomic ions.

I have them listed in order of charge. Notice that the first three have a charge of minus 1, these two have a charge of minus 2, and the last one has a charge of minus 3. That is helpful if you will sort them like that while you are memorizing them. Because, you need to know the atoms that are present as well as the charge. All of that is important when you are memorizing polyatomic ions.



Phosphate will always be 1 phosphorus, 4 oxygens, the whole thing has a charge of minus 3. All of that information is important.

Suggested Question (Time 2:16)

What is the charge of the nitrate ion?

Once you have memorized the "ate" version of the ion, it is very easy to switch to the other versions of the polyatomic ion.

All of these are derivatives of the chlorate ion. Chlorate is  $\text{ClO}_3$ , so it is 1 chlorine, 3 oxygens, and it has a charge of minus 1. You would need to memorize the "ate" version of the ion.

Once you know the "ate" version, small changes in the name change the number of oxygens present. For example, to go from chlorate to chlorite, the number of oxygens went from 3 to 2. So the number of oxygens drops by 1.

I do want to point out that an "ate" ion will not always have 3 oxygens. That is why you have to memorize this "ate" version, and then to go from the "ate" to the "ite," you just drop the number of oxygens by 1.

To go from chlorite to hypochlorite, another oxygen goes away.

Going back to the chlorate version, to go from chlorate to perchlorate, an oxygen gets added.

So if you know the "ate" version, it is easy to change the number of oxygens to get to the other polyatomic ion.

Another example. If you know that sulfate is  $\text{SO}_4$  with a charge of -2, then you can easily figure out that sulfite is  $\text{SO}_3$  with a charge of -2. The number of oxygens dropped by 1. The charge didn't change. The sulfur didn't change. The only thing that changed is the number of oxygens.

Suggested Question (Time 4:04)

Nitrate is  $\text{NO}_3^{-1}$ . What is the formula for the nitrite ion?

In this unit, we are learning about both ionic and covalent bonds. Polyatomic ions participate in ionic bonds, because they themselves have a charge. They

are attracted to other charged particles. Ionic compounds are attractions between charged particles.

I do want to make a distinction thought. The bonds within the polyatomic ion itself, are covalent bonds. These are bonds between nonmetal atoms where electrons are being shared instead of transferred. So you have covalent bonds between the atoms of the polyatomic ion, but then the whole thing will ionically bond to something else to form an ionic bond.

That wraps up our discussion of polyatomic ions. Stick around as we begin looking at covalent bonding.

Suggested Question (Time 5:12)

True or False: The bond between nitrogen and oxygen in the nitrate ion is an ionic bond.

## Unit 4, Video 3 – Covalent Bonding

This is Rebecca from ChemistryIsMyJam.com. In this video, we are going to look at covalent bonding.

Recall that atoms bond in order to gain full outer energy levels. The octet rule tells us that most atoms are satisfied with 8 electrons in their outermost energy level.

Atoms can either transfer or share electrons in order to achieve full outer energy levels. In the case of covalent bonding, atoms will share electrons.

Take fluorine for example. Fluorine is a diatomic element, meaning that it is often found bonded with itself. Let's take a look at the type of bond that forms between two fluorine atoms.

Each individual fluorine atom, before it bonds, has 7 valence electrons. Each one wants to have 8. If fluorine will come together with another fluorine atom and each one will share an electron with the other, they will each have 8 valence electrons – satisfying the octet rule and forming a covalent bond.

In previous videos, we have looked at ionic bonding where an electron has been transferred. That works well when you have a metal and a nonmetal, however it does not work well in a case like fluorine where you have two nonmetal atoms coming together.

In the case where you have two nonmetal atoms coming together, a covalent bond will be formed where electrons are shared.

Suggested Question (Time 1:37)

True or False: Electrons are shared when covalent bonds form.

The key to covalent bonding is that electrons are shared. Covalent bonds are formed between nonmetal atoms which are mostly found on the right hand side of the periodic table.

In terms of physical properties, covalent compounds have low melting points and low boiling points. They tend to have poor electrical conductivity.

Covalently bonded atoms form molecules, or molecular compounds. These are individual, discrete units.

What do I mean by them being individual, discrete units? Let's compare formulas of covalent versus ionically bonded compounds to see the difference.

Here, on the right, you have a water molecule, which is made up of covalent bonds. Water will always be 1 oxygen with 2 hydrogens. The chemical formula gives me the actual number of atoms present.

In the case of an ionic compound, like sodium chloride, the Na and the Cl gives me the ratio of the atoms present. In an ionic compound, you have a series of positive and negative ions that are lining up to try to maximize the attractions between charged particles.

So in the case of a molecule, the chemical formula is giving you the actual atoms present. In the case of an ionic compound, the chemical formula is giving you the ratio of the ions present.

So these are a discrete, individual unit, whereas these form a crystal structure.

Suggested Question (Time 3:14)

True or False: The formula for an ionic compound indicates the actual number of atoms present in the compound.

This is the flow chart that I introduced earlier in this series for how to name compounds. The first thing you want to ask yourself is: Is there a metal present?

For the case of covalently bonded compounds, the answer here is usually no. Typically, there is not a metal present for molecular compounds. They are made up of 2 nonmetals. When you go to name those, you want to use prefixes in order to name your molecular compounds.

3:45

Here you have the prefixes that are necessary in order to name molecular compounds. The prefixes for the numbers 1 through 10 are mono, di, tri, tetra, penta, hexa, hepta, octa, nona, and deca. It would be a good idea to memorize those prefixes.

Prefixes are added at the beginning of each element in the compound and they tell you the number of each of those atoms present in the molecule.

The molecule will need to end in "ide," just like in ionic compounds.

The one rule to keep in mind is that "mono" can be omitted from the first element in the compound, but cannot be omitted from the second element in the compound.

You will find that writing formulas for molecular compounds is simpler than ionic compounds because the charges do not matter. You simply go by the prefixes to determine the number of each element present.

Suggested Question (Time 4:49)

True or False: The prefix indicates the charge of the element.

Let's look at some examples of how molecular compounds are named. We are going to be using the prefixes that we were just introduced to, and the prefixes are related to the subscripts. The prefixes are going to indicate the number of each type of element present.

In the case of  $N_2O_5$ , there are 2 nitrogens and 5 oxygens. The prefix for 2 is "di," so we have dinitrogen. The prefix for 5 is "penta" so it is pentoxide. The compound needs to end in "ide" and the prefixes tell me the number of each element present.

In the case of carbon dioxide, there is 1 carbon and 2 oxygens. Remember the rule that "mono" can be omitted from the first element but not from the second element. I didn't need to write monocarbon here. It is just carbon. The first element does not have "mono." There is 1 carbon and then 2 oxygens - carbon dioxide.

In this example, there is 1 of each - 1 carbon and 1 oxygen. I am allowed to omit "mono" from the first element but not from the second element. I don't need "mono" here but I do need it here - carbon monoxide.

Notice that they all end in "ide" and the prefixes tell us the number of each element present.

Now that we have had a video introducing both ionic and molecular compounds – stay tuned for the next video where we will compare the 2 and use that flow chart to name a series of mixed compounds.

Suggested Question (Time 6:44)

Name the following compound:  $\text{SF}_2$

## Unit 4, Video 4 – Ionic Versus Covalent Bonding

Previously, in this unit, we have had separate videos covering ionic and molecular compounds. Now, I would like to take the two and put them side by side so that you can really see the differences between them.

One of the first major differences comes in their structure. This is a typical ionic compound. This would be an image of something like sodium chloride, where you have positively charged sodium ions and negatively charged chloride ions. They are going to line up side by side in an alternating pattern that maximizes the attraction between the positive and the negative particles.

For an ionic compound, the formula NaCl is giving you a ratio of the number of atoms present. It is telling you that there is one sodium ion for every one chloride ion. The formula is a ratio rather than the actual number of atoms present.

Contrast that with a molecular compound. Here you have water, which is a typical molecular compound. Water is made up of molecules and each molecule contains exactly one oxygen and two hydrogen atoms. So for a molecular compound, the chemical formula is giving you the actual number of atoms present. The molecule is an individual discreet unit.

Whereas with an ionic compound, it is not a molecule, it is a series of positively and negatively charged ions. And the formula is giving you the ratio of the charges present.

Suggested Question (Time 1:43)

True or False:  $\text{CaCl}_2$  is an ionic compound composed of exactly one calcium and two chloride atoms.

Let's take a look at the properties of ionic and molecular compounds.

Ionic compounds are made up of a metal and a nonmetal, whereas molecular compounds are typically made up of nonmetals. This is crucial information that helps us identify the type of compound that you're looking at. If you see a compound that contains a metal, from the left hand side of the periodic table, and a nonmetal from the right, it is ionic. However, if it only contains nonmetals, it would be molecular. Recall that your nonmetals are mostly found

on the right side of the periodic table, with the exception of hydrogen which is on the left.

In terms of physical properties, ionic compounds are solid crystals at room temperature. They have high melting and boiling points, and the subscripts in the compound are giving you the ratio of the ions present.

Molecular compounds typically have low melting and boiling points. They are made up of molecules, which are individual, discreet units. And the subscripts give you the actual number of atoms present.

Suggested Question (Time 2:59)

True or False: Based on the properties of ionic and molecular compounds, sodium chloride should have a high melting point.

When you have a compound that you need to name, there are a series of questions that you need to ask yourself. There is a printable version of this flow chart in the description.

The first question you should ask: Is there a metal present. If there is a metal then it is an ionic compound. You're going to end up listing the cation followed by the anion ending in "ide."

If there is not a metal present, then it is a molecular compound and you should use prefixes to indicate the number of each atom present in the compound.

3:35

Let's look at a mixed list of ionic and molecular compounds. I'm going to use the questions that we just saw to identify the type of compound, and then to name the compound.

In the first example, I see  $\text{CaCl}_2$ . The first question you should ask yourself: Is there a metal present? Yes. Calcium is a metal so this is an ionic compound.

The next question I would ask myself: Is there a transition metal, Sn, or Pb? There is not. So I'm going to simply name the cation, calcium, and then make the anion end in "ide." So calcium chloride.

In the second example, I ask myself: Is there a metal present? In this case, both sulfur and oxygen are found on the right hand side of the periodic table.

There is not a metal present. This is a molecular compound. In the case of a molecular compound, you should use prefixes to name the compound.

Remember that there is a rule that tells that you can eliminate "mono" from the first element but not from the second element. The prefix for 1 is "mono" but I did not need to put monosulfur because sulfur is the first element. The prefix for 2 is di, so this becomes sulfur dioxide. It still needs to end in "ide."

The third example, is there a metal present? Yes. Barium is a metal. This is an ionic compound. Then you would ask yourself: Is there a transition metal, the element Sn, or the element Pb? In this case there is not, so I simply list the cation, which is barium, and then in this case, the anion is a polyatomic ion, so I'm not going to have to change the name this time. This is barium sulfate. If I made it barium sulfide, that would be barium with the element sulfur.

In the fourth case, you've got Pb, which is lead, and chlorine. Pb is a metal so this is an ionic compound. The next question I would ask myself: Is there a transition metal, sn, or Pb. There is. This is one of those special cases where you need to list the charge of the cation in roman numerals. Typically you would look at the charge of the anion to figure out the charge of the cation. Chlorine is in group 17 on the periodic table, meaning that it has a charge of negative 1, and there are 2 of those chlorines. So the lead must have a charge of +2. So I've used the roman numeral II to let me know that this is lead(II) and then the chlorine needs to end in "ide." Lead(II) chloride.

In the last example, I ask myself, is there a metal. No there is not. Carbon and chlorine are both nonmetals so it is a molecular compound. I need to use prefixes to name it. There is only one carbon. The prefix for 1 is "mono" but I do not have to put "mono" for the first element. The prefix for four is "tetra" so this is carbon tetrachloride.

If you would like to see more details about naming ionic and molecular compounds, check out the playlist listed in the description. You can go back to the video that is specifically on ionic or molecular compounds.

That concludes our comparison of ionic and molecular compounds. Stick around as we learn to draw Lewis Structures.

Suggested Question (Time 7:33)

Name the following compound:  $\text{Al}_2\text{O}_3$

## Unit 4, Video 5 – Lewis Structures

In this video we are going to focus on Lewis Structures. You have an example of the Lewis Structure for carbon dioxide here. You can see that the Lewis Structure is a drawing that shows me where the atoms are located, what other atoms they are attached to, the type of bond between the atoms, and then where any valence electrons are located on that molecule. There is a great deal of information that we can get from a Lewis Structure.

This is a series of rules that we are going to go over that helps you to draw Lewis Structures. There is a printable version of this in the description.

I'm going to start by quickly talking through the rules, but then the rules will make more sense when we apply them to an example in just a moment.

The first thing that you do is total the valence electrons. You can determine valence electrons from the location of an element on the periodic table, and you would want to add those up for all of the elements present in the molecule.

Once you have done that, if there is a charge, a positive charge gets subtracted from your total, because a positive charge means that you have lost electrons. A negative charge gets added to your total.

Once you have totaled your valence electrons, you are going to arrange the atoms. Typically, carbon goes in the middle, or whichever atom you have the least of goes in the middle.

At that point, you complete octets by making sure every atom has eight electrons around it. The exceptions are that hydrogen can only have two atoms around it, and boron is often happy with just six atoms around it.

Once you have completed the octets, you compare what you have drawn to the number of valence electrons that you calculated at the very beginning. If you have too many electrons drawn, you make double or triple bonds. If you have too few electrons drawn, any extra electrons can go on the central atom.



Suggested Question (Time 2:03)

For a water molecule ( $\text{H}_2\text{O}$ ), which element should go in the middle of the Lewis Structure? The hydrogen or the oxygen?

Here is our first example where we will be drawing the Lewis Structure for water, which has 2 hydrogens and 1 oxygen. The first thing that we are going to do is total the valence electrons.

Hydrogen – there are 2 of them in the molecule, so that is where that 2 came from. Hydrogen is group 1 and has 1 valence electron. Everything in group 1 has 1 valence electron. 2 times 1. There are 2 valence electrons due to the hydrogens in water.

For oxygen, there is 1 oxygen within the molecule. To get the valence electrons, count over – skipping the transition metals. So 1, 2, 3, 4, 5, 6. Oxygen is in the sixth group over and has 6 valence electrons. 1 times 6 gives me 6. My total valence electrons is 8.

There is not a charge for this molecule. That is it for my total valence electrons. You do not need to add or subtract anything from that total.

The next step is to arrange the atoms. If you will remember from the rules, you want to put the least prevalent atom in the middle. You have 2 hydrogens and 1 oxygen. Since there is less oxygen, it should go in the middle.

So I have an oxygen in the middle with a hydrogen on each side and I have drawn single bonds in between them.

In this case, we only have 2 atoms attached to the central atom, so everything ended up in a straight line. However, if there were more atoms, you could have put them on the top or on the bottom.

The next step is to complete the octets for the elements that are present. What I mean by that is making sure all of the elements have a total of 8 electrons around them. You can see that I have added 2 sets of electrons to oxygen. That is because each one of these bonds counts as 2 electrons. There are 2 electrons there, 2 electrons there. That got me up to 4. Oxygen needs 8 electrons so I added 2 pairs of electrons around it.

The electrons are going to come in pairs like this. You can see how they will be placed together in a pair.

Hydrogen was an exception to the octet rule. Hydrogen is a very small element. It only has 1 energy level. Hydrogen can only have 2 electrons bonded to it. This bond fills up this hydrogen. It does not need any extra electrons added to it. I have completed all the octets for the elements within that molecule.

Finally, you want to compare what you have drawn to the number of valence electrons that you started with. We started with 8 valence electrons. What we have drawn is 2, 4, 6, 8 valence electrons. Keep in mind that bonds count as 2.

Our valence electrons match the electrons we have drawn. This is an accurate Lewis Structure for water.

It does not show me bond angles. It does not show me how the structure might be bent. But it does give me the arrangement of the atoms and where any extra electrons are located.

Suggested Question (Time 5:28)

How many valence electrons are in a molecule of  $\text{CCl}_4$ ?

Let's look at another example that uses some different aspects to the rules. We are going to start by totaling the valence electrons for carbon dioxide, which has 1 carbon and 2 oxygens.

I have 1 carbon. Carbon has 1, 2, 3, 4 valence electrons. 1 times 4 is 4. There are 2 oxygens. Oxygen has 1, 2, 3, 4, 5, 6 valence electrons. 2 times 6 is 12. Add them together. We should have a total of 16 valence electrons.

The next step is to arrange the atoms. Carbon goes in the middle. There are less carbons than there are oxygens. So it makes sense to put it in the middle, with an oxygen on each side, and a single bond between it and each of the atoms.

Then you want to complete the octets. Each of these atoms is comfortable with 8 electrons around it. For the oxygens, they started with a bond – that is

2 electrons – so they need 6 more electrons around it. Those electrons should be put into pairs. Each of the oxygens got 3 pairs of electrons around them.

The carbon had 2 bonds attached. So that is 2, 4 electrons. It needs 4 more to get up to 8. A pair here and a pair here gets carbon up to 8 electrons. We have completed the octets for each of the atoms.

Next you want to compare what you have drawn to your original number of valence electrons. Keep in mind that bonds count as 2 electrons. These are covalent bonds. A covalent bond is a pair of shared electrons. What I have drawn is 2, 4, 6, 8, 10, 12, 14, 16, 18, 20. I have drawn 20. I should only have 16.

I have too many electrons in my Lewis Structure. To fix that I'm going to need to make double or triple bonds. To make a double bond, if I'm going to put a double bond right here, I do that by removing a pair of electrons from each side of it. So to put one double bond here, I'm going to remove a pair of electrons from here and a pair of electrons from here.

Let's take a look at how that turns out. In order to make a double bond right here, I'm going to need to remove a pair of electrons from both sides. So I'm going to be removing a pair of electrons from this oxygen and a pair of electrons from this carbon.

By making the double bond, I have reduced the total number of electrons that I have drawn by 2. I now have 2, 4, 6, 8, 10, 12, 14, 16, 18 electrons drawn, which is closer to my 16. I need to draw one more double bond in order to get me all the way down to the 16 valence electrons that I need.

This time I will put the double bond on the other side. In order to draw a double bond there, I'm going to once again remove 2 electrons from both sides of the double bond. Another pair of electrons has to leave this carbon and a pair of electrons has to leave that oxygen.

Every time you create a double bond, it is removing a pair of electrons from each side. A double bond is representing 2 shared pairs of electrons.

Now my structure has 2, 4, 6, 8, 10, 12, 14, 16 valence electrons. The correct Lewis Structure for carbon dioxide should have 2 double bonds.

Suggested Question (Time 9:19)

True or False: Adding a double bond reduces the number of electrons drawn by 2.

We are going to look at one more example. This time it will be a molecule that has a charge.

This is the polyatomic ion carbonate. Carbonate has 1 carbon, 3 oxygens, and a total charge of -2. So I added up the total valence electrons.

There was 1 carbon. Each carbon has 4 valence electrons. 3 oxygens, each with 6 valence electrons. That gave me a total of 22 valence electrons.

Then the carbonate ion has a negative 2 charge so it has gained 2 electrons. That means 2 electrons needed to be added to my total. 22 came from adding up carbon and oxygen, plus 2 because of the negative 2 charge – giving me a total of 24 valence electrons.

Next you want to arrange the atoms with carbon in the middle, three oxygens around it, and a single bond between each one.

It is also common to put ions in brackets and put the charge on the outside.

Then it is time to complete the octets. Each of these atoms should have 8 electrons around it. The carbon initially had 2, 4, 6 from the bonds, so it needs one more pair of electrons to get it up to 8.

Each of the oxygens had 1 bond, giving it 2 electrons. It needs 2, 4, 6 – three more pairs of electrons around it.

The next step is to compare what you have drawn to your original number of valence electrons. Bonds count as 2 so this structure has 2, 4, 6, 8, 10, 12, 14, 16, 18, 20, 22, 24, 26 and there should only be 24. We need to make a double bond.

The double bond could go in any of the three locations. I'm going to put mine over on the lefthand side. In order to make my double bond, I need to remove 2 electrons from each side of it. So I've removed a pair of electrons from the carbon and a pair from the oxygen. My structure now has 2, 4, 6, 8, 10, 12, 14, 16, 18, 20, 22, 24 electrons, which matches the 24 valence electrons that we started with.

The correct structure for the carbonate ion will have a double bond.

Suggested Question (Time 12:01)

How many valence electrons are present for the nitrate ion ( $\text{NO}_3^{-1}$ )?

In this video, we have seen single bonds and double bonds. There are also triple bonds. A single bond is a pair of shared electrons. A single bond is made up of 2 shared electrons. Double bonds are 4 shared electrons and triple bonds are 6 shared electrons.

Single bonds are the longest type of bond whereas triple bonds are the shortest. Each time you add a pair of shared electrons, the bond gets a little bit shorter.

Single bonds are considered the weakest pair of electrons and bond strength is a measure of the energy required to break a bond. It is easiest to break a single bond. It is hardest to break a triple bond because they are the strongest type of bond.

There really is a great deal of information that can be determined from a Lewis Structure. By understanding the type of bond present, it gives you an understanding of the length of the bond and the strength of the bond. The Lewis Structure also lets me know where there are extra pairs of electrons on the molecules.

That concludes our discussion of how to draw Lewis Structures. In the next video, we are going to look at how you can use a Lewis Structure and the VSEPR theory to determine the shape of a molecule.

Suggested Question (Time 13:38)

True or False: Single bonds are longer and stronger than triple bonds.

## Unit 4, Video 6 – Octet Exceptions

In this video, we are going to be focusing on octet rule exceptions. So we are going to be looking at Lewis Structures where the central atom has either less than or more than 8 electrons. Because following the octet rule means having 8 electrons. We need to look at some atoms that do not necessarily follow that rule.

Common examples of atoms that have less than 8 electrons are hydrogen and boron. Hydrogen is an extremely small atom. It only has 1 energy level. It can only handle having 2 electrons. You will find hydrogen with 1 single bond, and no more. That is an exception to the octet rule.

Another common example is boron. Boron is very often stable with only 6 electrons. It is very common to find boron with just 3 single bonds.

That would be boron with only 6 electrons or hydrogen with only 2. Those are exceptions to the octet rule.

Let's look at some examples and then we will come back and look at exceptions where they have more than 8 electrons.

Suggested Question (Time 1:16)

Hydrogen is an exception to the octet rule. How many valence electrons are often found on hydrogen?

We will begin by drawing the Lewis Structure for boron trifluoride,  $\text{BF}_3$ . The first thing that you should always do is total your valence electrons. As a reminder, there is a trend on the periodic table that helps you figure out how many valence electrons each atom has.

The atoms in group 1 have 1 valence electron, these have 2, skip the transition metals, 3, 4, 5, 6, 7, 8 valence electrons. In this structure, we have 1 boron. Boron is here with 3 valence electrons. 1 times 3 is 3. This structure has 3 fluorines. Fluorine is here with 7 valence electrons. 3 times 7 is 21.

Add those together. When I'm done with this structure I should have 24 valence electrons drawn.

The next step is to arrange the atoms with the least prevalent atom in the middle. There is 1 boron and 3 fluorines. Boron should go in the middle with a single bond to each of the 3 fluorines.

The last step is to compare what I have drawn to the number of valence electrons that should be present. I have drawn 2, 4, 6, 8, 10, 12, 14, 16, 18, 20, 22, 24 valence electrons. Remember that a bond counts as 2 electrons. I should have 24 valence electrons.

What I have drawn matches what I should have. This is a good Lewis structure for boron trifluoride.

Remember that boron is most of the time stable with 6 electrons. There are cases where boron has 8. If I had gotten to this last step, and I needed 2 more electrons on this structure, it would have been perfectly fine to add them right there to the boron and complete its octet.

Suggested Question (Time 3:15)

Boron is an exception to the octet rule. How many valence electrons are often found on boron?

The other type of octet rule exception would be having more than 8 electrons. This is going to apply to elements that are in period 3 on the periodic table or below. The first 2 periods of elements cannot do this, but if you're in period 3 or below, then these elements can handle having more than 8 electrons.

Remember when we did electron configuration. We called this row the 3s, this was the 3p, here was the 4s, and then this was the 3d. That energy level dropped by 1. So the elements in period 3 or below have an empty D subshell that they can use. We call this an expanded octet. It allows those elements to handle having more than 8 electrons.

Suggested Question (Time 4:14)

True or False: Nitrogen is an element that can exceed the octet rule by having more than 8 valence electrons.

An example would be sulfur hexafluoride,  $\text{SF}_6$ . Let's draw the Lewis Structure.

This structure has 1 sulfur. Sulfur has 6 valence electrons. 1 times 6 is 6.

There are 6 fluorines. Fluorine has 7 valence electrons. 6 times 7 is 42 for a total of 48 valence electrons.

The next thing that you want to do is arrange the atoms with the single most atom in the middle. So sulfur goes in the middle and then there are 6 fluorine atoms around it.

The next step is to complete the octets. Each of these fluorines started with 2 electrons. 6 more were added to each fluorine. Sulfur started with 2, 4, 6, 8, 10, 12 electrons, but that is OK. Check the periodic table. Find sulfur. Sulfur is in period 3 or below. It is OK for it to have an expanded octet. It is OK for sulfur to have more than 8 electrons.

Once you have completed those octets, you want to count all of the electrons that you have drawn. Remember that bonds count as 2. I have drawn 48 electrons. I should have had 48. This is a great Lewis Structure for sulfur hexafluoride.

Suggested Question (Time 5:44)

In the Lewis Structure for sulfur hexafluoride, how many electrons does sulfur have?

This last example is going to be sulfur tetrafluoride, 1 sulfur and 4 fluorines. Each sulfur has 6 valence electrons. Each of the fluorines has 7 valence electrons. We should have a total of 34 valence electrons.

Next you want to arrange the atoms with the least prevalent atom in the middle. I have put sulfur in the middle with 4 fluorines around it.

Then it is time to complete the octets. The sulfur already had a completed octet. It had 4 bonds around it. Each bond counts as 2. That is a total of 8 electrons.

Each of these fluorines had 1 bond. That counts as 2 electrons so I have added 6 more electrons to each of the fluorines. I have completed the octets for all of the elements in the structure.

Then it is time to compare what I have drawn to the number of valence electrons. I have drawn 32 total electrons and I should have 34. There is a problem there. This structure is missing 2 electrons.

Our rules tell us that those electrons should go on the central atom, and those rules are in the description if you need them. We are going to add the extra electrons to the central atom. It looks a little strange but this is what the Lewis Structure for this structure should look like.

The last thing you want to do is ask yourself, is it OK for sulfur to have more than 8 electrons. It is because sulfur is in period 3 or below so it is perfectly fine for sulfur to have an expanded octet.

This is an appropriate Lewis Structure for sulfur tetrafluoride.

That completes our discussion of Lewis Structures. Stick around as we look at how we can use those Lewis Structures to predict the shape of a molecule. In the next video, we are going to be talking about the VSEPR theory.

Suggested Question (Time 7:56)

If extra electrons are needed on a Lewis Structure, where should they be added?

## Unit 4, Video 7 – The VSEPR Theory

Once you have learned to draw Lewis Structures, the next step in understanding a molecule is to understand the shape of the molecule.

A Lewis Structure is going to give you the arrangement of the atoms, the type of bond that you are dealing with – whether or not it is a single, double, or triple covalent bond, and it is going to give you the location of any unpaired electrons. A Lewis Structure by itself is not going to give you the shape of the molecule. That is where the VSEPR Theory comes in.

VSEPR stands for Valence Shell Electron Pair Repulsion Theory. Let's break that down. Valence shell means that you are dealing with electrons in the outer most energy level, electron pair – so those electrons are found in pairs, and then repulsions – those electrons are trying to get away from one another. Whether or not you are dealing with bonded electrons or nonbonded electron pairs, either way they are pushing away from each other. That is going to dictate the shape of the molecule.

Here is how we are going to use the VSEPR theory. In the description, you can find some printable versions of VSEPR charts. We are going to be using those charts to analyze a Lewis Structure.

When you are looking at the Lewis Structure, you want to focus on the central atom in the Lewis Structure and you are going to count domains. Domains are regions of space where electrons are located. A bond is counted as one bonding domain, even if it is a double or a triple bond. A triple bond counts as one bonding domain. A pair of electrons is considered a nonbonding domain.

Once you have figured out that information, you go to your VSEPR charts and the VSEPR charts will help you predict the shape and the bond angle.

I do want to point out that the bond angle in a VSEPR chart is an approximation. There are some things that affect that. For example, lone pairs tend to change bond angles slightly. So if you have a pair of electrons in the place of a bond, they take up a little more space than bonded electrons do.

Now that we have kind of been over the basics, let's dive into a VSEPR chart and look at some examples.



Suggested Question (Time 2:26)

True or False: Lone pairs on the central atom of a molecule do not affect the shape of the molecule.

Here you have our first look at a VSEPR chart. Most of your VSEPR charts are going to be arranged in order of total domains. The simplest case would be 2 domains and then it would increase from there.

You want to be using your VSEPR chart to analyze a Lewis Structure. We will start by looking at carbon dioxide.

Focus on the central atom. Carbon has 2 places where electrons are located on it. It has this double bond right here – that counts as 1 bonding domain. This double bond also counts as a bonding domain. There are 2 bonds total.

There are no pairs of electrons on the carbon. They would look like this. If you had a pair of electrons on the carbon, they would look like this. But you do not have that in place for the carbon.

There are 2 domains total, no pairs of electrons. Anytime you have that situation, the shape of the molecule is linear. What has happened is the electrons are spreading out as much as possible. They are pushing away from each other. They are about 180 degrees away from each other to give you that linear shape.

Suggested Question (Time 3:43)

What shape will result from a molecule where the central atom has two domains?

The next example is  $\text{BH}_3$ . Focus on the central atom. Look at the boron. There are 3 places where electrons are located. There are 3 total domains in  $\text{BH}_3$ . The boron has no pairs of electrons on it. Even if there are electrons on the outer atoms, we do not count those. We only focus on the central atom.

Any time you have 3 domains, 0 lone pairs, the shape is trigonal planar. All of these atoms are in the same plane. It is a very flat structure. Trigonal planar is the name of the shape. By having it in this trigonal planar shape, the bonds are spreading out as much as possible.



Suggested Question (Time 4:32)

True or False: A flat drawing of a trigonal planar structure is accurate because all of the bonds are in the same plane.

The next example is the first one that we have seen that has a pair of electrons on the central atom. There are 3 total domains. Let's look at those 3 domains.

This single bond is 1 domain. This double bond is 1 domain. This pair of electrons is 1 domain. There are 3 total domains, 1 pair of electrons. Whenever you have that case, the pair of electrons does not really show up in the shape, however it does contribute to the shape. That pair of electrons is taking up space. These bonded electrons are pushing away from these electrons.

That causes the shape of the molecule to be bent. Even though you don't directly see the electrons in this shape, they very much contribute to the shape. When you draw your Lewis Structures you want to make sure that you have put all of your lone pairs in the right place. Because they are going to contribute to your bond angles and to your shape.

Suggested Question (Time 5:42)

True or False: The bent shape of silicon dioxide results from the lone pairs of electrons on the sulfur repelling the bonded electrons.

Here we are continuing our VSEPR chart but moving on to 4 domains. The first example we are going to look at is methane which is  $\text{CH}_4$ . You have a carbon in the center with 4 hydrogens around it. Focus on the central atom, so you are looking at that carbon. That carbon has 4 bonds around it. That would be 4 total domains. The carbon also has 0 pairs of electrons on it. There are no nonbonding electrons on the carbon.

Anytime you have 4 domains and no lone pairs, the shape is tetrahedral, and a tetrahedral looks like this. Recognize that the tetrahedral shape is very 3 dimensional. These 2 bonds are in the plane of the paper. This one is going back and this one is coming out towards you. The tetrahedral shape is very 3 dimensional allowing the electrons to spread out as much as possible.

6:47

The next example is ammonia which is  $\text{NH}_3$ . You have a nitrogen in the center with 3 hydrogens bonded to it and a pair of electrons. Let's count the domains on that nitrogen. Each of the bonds counts as 1 domain, so there are 1, 2, 3 bonding domains and the pair of electrons also counts as a domain. There is a total of 4 domains. One of them is a lone pair. The shape is trigonal pyramidal.

This trigonal pyramid originated from the tetrahedral. Whenever you've got a total of 4 domains, the parent shape is a tetrahedral, but then a portion of it is missing. This pair of electrons right here is like taking away that bond to give you the trigonal pyramid that you see here.

I point that out because I want to make sure that you recognize that this is also a very 3 dimensional shape. If you were to make a model of a trigonal pyramid and set it down on a table, it is going to stand up on the table. It is not going to sit flat the way our planar structure did. It is going to stand up on the table and be very 3 dimensional. By being very 3 dimensional you are spreading the electrons out as much as possible, which is really the heart of the VSEPR theory.

8:18

Our last example is water. Water has an oxygen in the center, two hydrogens bonded to it, and 2 pairs of electrons on the oxygen. There are a total of 4 domains, 2 lone pairs on the oxygen. Go back to your parent tetrahedral, it is like you took away 2 places. This one is gone and this one is gone, leaving you with a bent shape. Water is a bent structure.

Suggested Question (Time 8:51)

The Lewis Structure for  $\text{PH}_3$  has 3 single bonds and a pair of electrons. What shape do you expect this molecule to have?

I want to spend just a moment introducing a concept called hybridization because it is related to the VSEPR theory. The orbitals within an atom in their normal positions would not be able to get into the bond angles that the VSEPR theory predicts. Scientists believe that these orbitals come together and they hybridize to form new hybrid orbitals that allow for the bond angles we see within molecules.

For example, here is an s orbital and 3 p orbitals. This s orbital and one of those p orbitals come together and hybridize to form 2 sp hybrid orbitals. These sp hybrid orbitals are different shapes from their original orbitals, and they are at different energies. The energy of this new orbital is somewhere in between the energy of the original orbitals. Which is good, because originally this orbital had a higher energy. So by coming together forming these hybrid orbitals, that new hybridized orbital is lower in energy and has a shape that allows for the bond angle that is required for the VSEPR theory.

What I have shown you is just one example of a hybridized orbital. There are many others out there but I wanted to introduce the idea so that we could look at how hybridization relates to bond angles.

Suggested Question (Time 10:25)

True or False: The shape of the original orbitals within an atom allow bonded electrons to spread out as much as possible.

The type of hybridization present depends on the number of domains. Once again we are focusing on that central atom and we are counting the domains – the regions of space around the central atom where electrons are located.

For carbon dioxide, there are 2 domains. Anytime there are 2 domains, you are looking at an sp hybridized situation. sp hybridized will give you a bond angle of 180 degrees.

Anytime there are 3 domains, like in the example of  $\text{BH}_3$ . That I sp<sup>2</sup> hybridized and the bond angle should be roughly 120 degrees.

When there are 4 domains, it is sp<sup>3</sup> hybridized and the bond angles should be around 109.5 degrees.

I do want to point out that these bond angles are approximations. They are most accurate when all of the domains around the central atom are bonding domains.

For example, this is methane. Methane has a carbon with 4 bonding domains around it. Those angles are pretty close to 109.5 degrees.

Contrast that with ammonia. Ammonia is  $\text{NH}_3$ . The nitrogen has 3 bonded hydrogens and 1 nonbonded pair of electrons. Ammonia still has 4 total

domains. It is still  $sp^3$  hybridized even though one of those domains is a pair of electrons.

However, that pair of electrons takes up more space than those bonds do. The bond angle is actually a little bit lower than this expected 109.5. The bond angles in ammonia are more like 107.5.

Recognize that these bond angles are approximations and that lone pairs will lower the bond angles.

That concludes our introduction to the VSEPR theory. Up next we are going to look at polar covalent bonds.

Suggested Question (Time 12:45)

A molecule has a central atom with two single bonds and one double bond. What is the expected bond angle for this molecule?

## Unit 4, Video 8 – Polar Covalent bonds

In this video we are going to be looking at the concept of polarity, and polarity is very much dependent on the concept of electronegativity. So let's begin with a brief reminder of what electronegativity is.

Electronegativity is the idea that if 2 atoms are bonded together, one of them has the ability to hog the electrons. An atom with a higher electronegativity can pull electrons closer to itself.

On the periodic table, fluorine has the highest electronegativity. Atoms that are close to fluorine have higher electronegativity values. Atoms that are far away from fluorine tend to have lower electronegativity values.

Suggested Question (Time 0:48)

Which of the following atoms has the ability to pull electrons closer to itself: carbon or nitrogen?

With that in mind, let's look at this concept of polarity. A polar bond is a bond that has a positive end and a negative end. For example, hydrogen and chlorine. They are pretty far apart on the periodic table. Chlorine is very close to fluorine. Chlorine has a high electronegativity. Hydrogen has a low electronegativity.

Chlorine is pulling the electrons to itself. That gives chlorine a partially negative charge and hydrogen a partially positive charge. That makes this polar bond. There is a large difference in electronegativity making it a polar bond.

Think about the word polar. Polar implies that there are two poles, a positive end and a negative end.

There are a couple of different ways that we indicate polar bonds. One way is through these little delta symbols. This delta negative means that there is a partially negative charge on this end of the molecule. Delta positive means that there is a partially positive charge on the other end of the molecule.

It is not a fully positive charge or a fully negative charge, because the electron hasn't totally been transferred. However, the electron is shifted towards the chlorine giving a partially negative charge and a partially positive charge.

The other way that we can indicate a polar bond is with an arrow. The arrow is always going to be pointing at the negative side. The arrow will point at the more electronegative element. The arrow also will have a little line right here that makes it look like a plus sign on the other side. Those are the 2 ways that we indicate polar bonds.

The HCl bond is very polar because there is a difference in the electronegativity values of the 2 atoms present.

Contrast that with the chlorine molecule where 2 chlorine atoms are bonded together. Chlorine and chlorine are the same atom. They have the exact same electronegativity value. There is no difference between the way these 2 share their electrons. This is a nonpolar bond.

A polar bond has a difference in electronegativity. A nonpolar bond does not.

Suggested Question (Time 3:15)

Consider a bond between carbon and fluorine. On which atom would it be appropriate to include a delta positive symbol?

The examples we just looked at were all polar bonds. Two atoms with one bond in between them. Now I want to scale that up a little bit and look at polar molecules. I want to make sure that we are distinguishing between the two. Because you do have cases where a molecule has polar bonds within it, but the molecule itself is nonpolar. Now we are looking at the entire molecule and trying to determine if there is a positive end and a negative end.

Your first example is water. Water has an oxygen with 2 hydrogens attached to it. It is a bent structure because of some pairs of electrons that are located up here on the oxygen.

Oxygen is a very electronegative atom. Hydrogen is not a very electronegative atom. The oxygen has the ability to pull the electrons closer to itself.

Oxygen is going to have a partially negative charge, leaving these hydrogens with a partially positive charge. Water is a polar molecule.

The same situation is happening for ammonia. Ammonia has this nitrogen, which is fairly electronegative, and these hydrogens which are not. The nitrogen is pulling the electrons to itself. The hydrogens are partially positive and the nitrogen is partially negative. That has been indicated with these arrows. These are polar molecules.

All of the examples at the bottom are nonpolar molecules, even though they have polar bonds. Consider the carbon-oxygen bond. Carbon has a lower electronegativity than oxygen is. This oxygen is pulling the electrons closer to itself. However, this oxygen is doing the same thing. You end up with 2 regions of negative and a region or partially positive in the center. That symmetry within the atom cancels out the polarity of the atom.

Symmetry can cancel out polarity. The same thing happens in boron trifluoride. There is a polar bond, in fact there are 3 polar bonds. The fluorine is much more electronegative than the boron. The fluorine is pulling electrons to itself. However, a planar structure has a lot of symmetry in it. That negativity gets spread evenly throughout the molecule and the polarity gets cancelled out.

The same thing happens in a tetrahedral structure. Even though there are polar bonds within the structure, the polarity gets cancelled out by the symmetry.

When you look at a structure, keep an eye out for 3 shapes: linear, trigonal plan, and tetrahedral are 3 shapes that will very often cause polarity to be cancelled out. That is only going to happen when the outside atoms are identical.

When you look at a molecule and you are trying to determine if the molecule is polar, first look for polar bonds. If there are no polar bonds present, the molecule is not polar. However, if there are polar bonds, you may have a polar molecule.

The second thing that you want to look for is symmetry. Because if you have symmetry – if you have a linear, a trigonal planar, or a tetrahedral shape – that symmetry is going to cancel out the polarity of the bonds.

The last thing I want to mention is why we care about polarity. Polarity affects so many different things about a molecule and how it interacts with other

molecules. It affects things like your solubility – whether or not your molecule will dissolve in certain solvents depends on polarity. It affects boiling point, melting point, viscosity, and quite a few other physical properties of the molecule.

Polarity is a big deal. It doesn't go away in chemistry. Understanding whether or not a molecule is polar or nonpolar is crucial to understanding how molecules interact with one another.

That concludes our discussion of polar covalent bonds. Stick around as we look at metallic compounds and wrap up our unit on bonding.

Suggested Question (Time 7:58)

True or False: Symmetric molecules are often nonpolar even though there are polar bonds within the molecule.

## Unit 4, Video 9 – Metallic Bonding

Let's finish up our discussion on bonding by talking about metallic compounds. The metallic bond is very different from the ionic and the covalent bond in that the electrons are delocalized. Specifically the valence electrons have the ability to move freely between all of the atoms present.

In this illustration, the positive signs are representing atoms that are tightly packed together so you can see the positive nuclei are tightly packed together and then the electrons are moving freely throughout the structure.

This is often called a "sea of electrons" and it explains many of the properties that we see with a metal. Metals are malleable, which means that you can bend them. They are ductile, which means that you can draw them into wires. If there were strong covalent or ionic bonds between these atoms, you would not be able to bend them or draw them into wires. This sea of electrons allows motion between the atoms.

Metals are also great conductors of electricity, because electricity results from moving electrons. The movement of these electrons allows for electricity to transfer.

Metals also have luster, which means that they are shiny. The electrons on the outer surface of the metal serve as a barrier that prevents light from entering in as well as it normally would. That causes the light to bounce off and gives the metal the shiny appearance that we are used to.

So many of the properties that we are familiar with with metals are explained by this sea of electrons – the idea that the valence electrons for metallic atoms are delocalized.

Suggested Question (Time 1:57)

What property of metal allows them to be drawn into wires?

Whenever other substances are added to a pure metal, an alloy is formed. Here you have a pure metal where nothing has been added. You should notice that there is an arrangement to the atoms. There is a pattern to the way that they are arranged in the structure.

The first type of alloy that I would like to show you is a substitutional alloy. In a substitutional alloy, atoms from the pattern have been removed and a new atom has been put into their place. In a substitutional alloy, atoms have been substituted into the original pattern or the pure metal.

Bronze is an example of a substitutional alloy. Bronze is made up of mostly copper atoms with tin atoms substituting some of those coppers. This will change many of the physical properties of the substance. Copper and bronze will have different colors, different hardnesses, they will be different in the way that they bend. They just function differently.

An alloy has different properties than the pure metal that it originated from.

The other type of alloy is an interstitial alloy. In an interstitial alloy, the original pattern has not been disrupted. However, other atoms have been inserted into the gaps for the pattern. Steel is an example of an interstitial alloy where the majority of the atoms are made up of iron and there are carbon atoms inserted into the gaps between the pattern. The properties of steel are very different than the properties of iron. And there are many different types of steel out there that depend on the amount of carbon that has been added or if any other atoms have been added.

You can also have alloys that are a combination of substitutional and interstitial alloys. In this example, the blue atoms are substitutional – they have replaced one of the original atoms. The green atoms are interstitial – they have been inserted into the gaps.

One thing that I would like to point out is that the size of the atom helps determine what type of alloy will form. In the case of a substitutional alloy, the original atom and the substituted atom tend to be fairly close in size. However, in the case of an interstitial alloy, the original atom tends to be much larger than the atom that has been inserted into the gaps.

Suggested Question (Time 4:43)

Consider an alloy that is made up gold and platinum atoms. Is this alloy likely to be substitutional or interstitial?

This has been a very brief introduction to alloys. I do want to point out that there are entire courses and curriculums available for understanding metals and understanding alloys.

It is crucial that an engineer who is designing a highrise, choose the appropriate type of steel to make sure that the building can sway but not collapse. Or someone designing a bridge has to choose the right metals to make sure that it can carry a certain amount of weight.

Depending on what you want to major in, an in depth understanding of alloys may be in your future.

That concludes our unit on bonding. Stick around as we start to learn about stoichiometry.

Suggested Question (Time 5:39)

True or False: The properties of alloys are usually very different from the properties of pure metals.