Chem1311Ch9Ep1 Transcript

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Hello and welcome to the first episode of Chemical bonding.

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In this episode.

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We will learn to draw Lewis dot structures.

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And we will learn to classify chemical bonds based on the electronegativity of the elements involved.

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The valence or outer shell electrons are the ones involved in bond formation and therefore are going to be of special interest to us.

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Remember that the group that a representative element belongs to is going to determine the number of valence electrons the atom will possess.

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We will consider only these valence electrons.

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When writing down Lewis dot symbols and structures.

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As you can see in this periodic table, the Lewis dot symbols for any element consists of the element symbol surrounded by a dot for each valence electron that it has.

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Ionic bonds are what we call the electrostatic force that holds together ionic compounds.

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This electrostatic force is due to the charge that is induced by an electron transfer from one element.

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To the other.

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Covalent bonds are what we call a shared pair of electrons between 2 atoms.

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This first diagram shows lithium and fluorine sharing a pair of electrons.

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So that they both achieve a noble gas configuration.

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This next diagram shows the transfer of electrons between lithium and fluorine.

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To form an ionic bond.

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The Lewis structure of ionic compounds is written in brackets and the charges are written next to the elements symbol.

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Up until now we have assumed that all metal-nonmetal combinations will form ionic bonds.

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And that all nonmetal-nonmetal combinations will form covalent bonds.

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However, this assumption is only partly true.

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In this episode we will explore an objective-ish way of determining whether we should expect a covalent or an ionic bond.

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Write the Lewis dot symbols to show the formation of.

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Aluminum oxide from its elements.

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Pause the video.

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Write down your answers.

00:04:10

And come right back.

00:04:22

Welcome back.

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In this example, we will start with the Lewis dot symbols for aluminum and oxygen by drawing 3 electrons around aluminum and six electrons around oxygen based on their chemical groups.

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The two aluminum atoms will donate a total of 6 electrons to the three atoms of oxygen.

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And that will result in Al_2O_3, aluminum oxide.

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And don't forget the brackets.

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Lattice energy is defined as the energy needed to vaporize a mole of solid ionic compound into ions.

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This energy will be proportional to the charge of the ions.

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And it will be inversely proportional to the distance between the ions.

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Those will be the two major factors.

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We can see this is true when we compare the lattice energies of two magnesium compounds, magnesium fluoride.

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And magnesium oxide.

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We can see that the lattice energy for magnesium oxide is greater.

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Because the oxygen has a larger charge than fluorine does.

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We can also see that the lattice energy for lithium fluoride is larger than that of lithium chloride, because chlorine is a much larger atom than fluorine is.

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If you must predict which compound is expected to have a larger lattice energy.

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You should base your prediction on these two properties.

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Charge and size.

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The larger the charge.

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The larger the lattice energy.

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The smaller the size, the larger the lattice energy.

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As you might expect, the melting point of an ionic compound is related to the compound's lattice energy.

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You can think of the lattice energy as.

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Energy keeping the.

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Atoms together into the crystal.

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This is of course not a perfect relationship.

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But if the compounds are related, the trends will be evident.

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For instance.

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These three graphs show the relationship between the lattice energy and melting temperature of compounds having the same cation.

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Notice that as the lattice energy goes up, so does the melting temperature.

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A covalent bond is a chemical bond in which one or more electron pairs are shared by two atoms.

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The reason why atoms share electrons is the same reason why they accept electrons from another atom.

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To fill up their outer shell.

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In this example.

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Two fluorine atoms go from having 7 electrons in their outer shell.

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To having 8 electrons.

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It's what you call.

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A win-win.

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And this is the Lewis structure.

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For a fluorine molecule.

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The covalent bond in fluorine and in any other molecule can be represented by two dots between the two atoms being bonded or by a line for clarity.

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The electrons that are not shared by the atoms are referred to as lone pairs of electrons.

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I know it's an oxymoron.

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But this doesn't seem to annoy me so much.

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A shared pair of electrons.

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Is a single covalent bond.

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The Lewis structure of water shows that oxygen shares a single covalent bond with each of the two hydrogen atoms.

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But it's also possible for atoms to share 2 pairs of electrons.

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These are called double bonds.

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And are represented by either four dots or two straight lines.

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And if two atoms share 6 electrons, then it is called a triple bond.

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4 electron pairs, however, cannot be shared between two atoms because the geometry does not work out for that.

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The length of a covalent bond is measured from nucleus to nucleus.

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Of the two atoms.

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The bond length will depend on the two atoms sharing electrons.

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The larger the atom, the longer the bond.

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But if the same 2 atoms are considered.

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Triple bonds.

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Will be shorter than double bonds and double bonds will be shorter than single bonds.

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Here we have.

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2 examples of single and double bond lengths.

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Notice that will both the carbon-oxygen and for the nitrogen-oxygen bonds.

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The double bond is shorter than the single bond.

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And here we have two examples.

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Of a comparison of all three types of bonds.

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In both instances, the triple bond is the shortest.

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Here we have a comparison between an ionic and a molecular compound which will allow us to make some generalizations.

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Notice that the melting and boiling points of ionic compounds tend to be much higher than those of molecular compounds.

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So, are the energies required for the two processes.

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And that's just a little common sense.

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Furthermore, ionic compounds tend to be more soluble in water than covalent compounds.

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A polar covalent bond is one in which the electron pair is not shared equally and ends up spending more of its time around one atom than the other.

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The bond between hydrogen and fluorine is an example of such a bond.

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The fluorine atom has a greater attraction for the shared electron pair than the hydrogen does.

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So that the electron pair will spend too much time with fluorine inducing a partial negative charge.

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And not enough time with hydrogen, resulting in a partially positive charge.

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The atom’s ability to attract the electron pair toward itself is called electronegativity.

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It follows the same periodic trend as electron affinity, but is not the same thing.

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Electron affinity can be measured.

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Whereas electronegativity is assigned a value in comparison to fluorine which has the highest assigned value of four.

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This periodic table.

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Shows the electronegativities of common elements.

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There are no units for electronegativity.

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Also, this table is part of your reference materials, so you don't need to remember this trend.

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You will have it available to you.

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As you can see in this graph, the trend is very clear for representative elements, but it gets a little murky with the transition metals.

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I know what an unexpected turn of events, right?

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There is an important role that electronegativity plays in bond formation.

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Namely, the difference in the electronegativities of the two atoms involved determines the type of bond they will have.

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If the electronegativities are equal.

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The bond will be covalent.

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If they differ by two or more.

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The bond will be ionic.

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And if they differ by less than two.

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It will be polar covalent.

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You can think of the types of bonds.

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As being part of a continuum.

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Going from covalent.

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All the way.

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To ionic and not simply on and off.

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Here are two examples of compounds with bonds.

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Yearning to be classified.

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Pause the video.

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Write down your answer.

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And then come right back.

00:19:06

Welcome back.

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Your answers should be based on the difference in the electronegativities.

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The first compound, hydrogen chloride, has a polar covalent bond.

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Because the difference in the electronegativities of hydrogen and chlorine is 0.9.

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The second compound, potassium fluoride, is ionic.

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Because the difference in electronegativity of potassium and fluorine is 3.2.

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And the third example.

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Is covalent.

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Because both carbon atoms have the same electronegativity.

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Writing Lewis structures will require for you to follow a few basic guidelines.

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First you will draw the structure connecting the atoms with single bonds and placing the least electronegative atom in the center.

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We will refer to this atom as the central atom.

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Then count the valence electrons of all the atoms in the structure.

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Add an electron for each negative charge and subtract an electron for each positive charge.

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You will then count the number of bonds you have drawn and subtract 2 electrons from the total for each single bond.

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Next you will add enough lone pairs to each atom except hydrogen to complete their octets.

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Be careful not to draw more electrons than your total number of valence electrons.

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If you are short of valence electrons, then you will draw an extra bond for each electron pair that you are short.

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Of course, life is better with examples.

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So let's draw the Lewis structure of nitrogen trifluoride.

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The first step is to draw the skeletal structure with single bonds.

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Next, we need to Add all the valence electrons.

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There are five for nitrogen and seven for each fluorine, bringing our total to 26 electrons.

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Since there is no charge.

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That is, the number of electrons we have to begin with.

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The structure we drew has three bonds.

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So we are going to subtract 6 from the initial 26 and now our total has dropped down to 20 electrons.

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Each fluorine will need 6 electrons to complete its octet.

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That is, 18 of the 20 electrons we had leaving just two.

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The nitrogen requires 2 electrons to complete its octet, and that happens to be exactly what we have left, so we're done.

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In this Case, no additional bonds were needed because we had sufficient electrons to complete all the octets.

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Now we are told to try nitric acid.

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The skeletal structure is drawn with four bonds.

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But I have an issue with it.

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You see, nitric acid is a strong acid.

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And therefore, will ionize completely.

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We could draw the Lewis structure, but this molecule doesn't really exist, so I refuse.

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Instead, we will draw the structure of nitrate.

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On principle.

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This is way more interesting anyway.

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The nitrogen has five electrons, and each oxygen has six electrons.

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Valence electron that is.

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Plus one more electron due to the negative charge for a grand total of 24 valence electrons.

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The structure we drew has three bonds, so we must subtract 6 electrons.

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From the original 24.

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And so we are left with 18 electrons.

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Each of the three oxygens needs 6 electrons to complete their octet, so that's all 18.

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Nitrogen needs 2 electrons.

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But we don't have them.

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Because we are short one electron pair, we must draw one more bond.

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We made one of the single bonds into a double bond and now we start again.

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This time we have drawn 4 bonds, so we subtract 8 electrons from the original 24, leaving us with only 16 electrons.

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The nitrogen already has a full octet.

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The single bonded oxygen need 6 electrons each, so that's twelve of the 16 electrons we had, leaving just 4 electrons.

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And the double bonded oxygen needs only four.

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Which we do have.

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So this is the Lewis structure for nitrate.

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Next. Next in the list is carbonate.

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We will draw the skeletal structure, which requires 3 bonds.

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Each oxygen has 6 valence electrons.

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And the carbon has four.

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That is 22 electrons.

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Plus two more electrons for the negative 2 charge.

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Giving us a total of 24 electrons.

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We already drew three bonds, so we'll have to subtract 6 electrons from the 24, leaving us 18.

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Each oxygen is going to need 6 electrons.

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And the carbon needs 2 for a total of.

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20 electrons that we need, but we only have 18.

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That means that we are short 2 electrons.

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Because we are short 2 electrons.

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We will need to draw an additional bond.

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And then start the counting again.

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We have 6 electrons from each oxygen and full form carbon. Same as before, plus two more from the negative 2 charge for a total of 24 electrons.

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This time we drew four bonds, so we have.

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used 8 of the 24 electrons, leaving us with 16.

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The carbon has an octet already.

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So the double bond that oxygen needs 4 electrons.

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And each of the single bonded oxygens need 6 electrons.

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For a total of 16 electrons.

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That's all we need, great.

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Whoa, what is this?

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Why are some electrons larger than others?

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Oh, apparently we went over a budget and.

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Had to buy used dots.

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Well, alright then.

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And that's all there is.

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There isn't anymore.