Chem1311Ch7Ep4 Transcript

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Hello and welcome to the 4th episode of Theory and electronic structure of atoms.

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Previously, in theory, and electronic structure of atoms.

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We define Pauli exclusion principle.

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We learned the relationship between quantum numbers and the resulting orbitals.

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And we applied the terms shell, energy level, subshell, and orbital accurately.

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In today's episode.

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We will apply the Aufbau principle and Hund's rule to writing electron configurations.

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And we will differentiate between electron configurations and orbital diagrams.

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The Aufbau principle simply states that electrons fill the lowest energy orbitals available.

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That means that it is easy to predict which orbital will be occupied if you know the number of electrons.

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This diagram.

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Shows the relative energy of some of the orbitals.

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Notice that being in the lower shell does not necessarily mean a lower energy level.

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Notice, for example, that 3d orbitals have a higher energy level than the 4s orbitals.

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That the 4d.

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Orbitals have a higher energy level than the 5s orbital.

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When an atom has a partially field of PDF or G subshell hunds rule.

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should be kept in mind.

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It means that when filling up a subshell, you place electrons in every orbital spinning in the same direction before you pair them up.

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The order in which the electrons in an atom fell in the orbitals does not need to be memorized.

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It can be figured out from a quick to draw diagram.

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There are various versions of this diagram and this is my favorite because it's the easiest to draw.

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Start by writing the s subshells in a column from 1s all the way down to 8s.

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Next, Skip 2 lines and to the left of 3s, write the p orbitals starting with 2p all the way down to 7p.

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Next, skip two more lines.

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And left of the 4p write the d subshells from 3d all the way down to 6d.

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Notice that every time we skip 2 lines.

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And we start with a 1s, 2p, 3d. Next will be the 4f (spoiler.)

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And finally skip two more lines.

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And left the 5d write down f subshells from 4f to 5f.

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To read the order of the orbitals.

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Start at the top.

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And read from left to right just like you would read a newspaper.

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So you have 1s.

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Followed by 2s.

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Followed by 2p and 3s.

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Followed by 3p, then 4s.

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Followed by 3d, 4p, and 5s.

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And so on.

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A reminder, due to the number of orbitals in each type of subshell.

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The maximum number of electrons for an s subshell is 2.

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For a p subshell 6.

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For a d subshell, 10 electrons.

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And for an f subshell, 14 electrons.

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Also, this may seem random right now, but it will be extremely useful later all noble gases except for helium and with a full p subshell.

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Hopefully we have not forgotten how to do this.

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Pause the video.

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Write down your answer and then come right back.

00:05:51

Welcome back.

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There is no need for bad language.

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I will show you where these numbers came from.

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If you can imagine the 3p subshell.

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That is room 3P. It has three orbitals or beds.

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The orbitals will have m_l values from -1.

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All the way to 1.

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We will have six sets of quantum numbers.

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Because there are six electrons.

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Here's the first electron.

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n and l were given to us when 3p was specified.

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And m_l we got from the little diagram I just drew.

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The electron is spinning clockwise, so m_s +½.

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Hund’s rule states that the next electron must go to an empty orbital, so we placed it on.

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m_l is equal to 0 and he has to spin clockwise, so another +½.

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The next electron goes to the last empty orbital and it is now time to begin pairing up electrons.

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The 4th electron goes back to the orbital with m_l value of -1, but the spinning must be counterclockwise.

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Then the 5th electron.

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And finally the 6th electron.

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So as you can see, each one of those numbers, each one of those quantum numbers, describes the electron that we just drew.

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An electron configuration shows how the electrons in an atom are distributed in the different orbitals.

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It uses a shorthand notation.

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It shows the value of n as a number.

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And the value of l as a letter.

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The number of electrons in that subshell is written as a superscript.

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The Aufbau principle must be adhered to when writing electron configuration.

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An orbital diagram uses a slightly different format to provide this information.

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It shows each orbital as a straight line.

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And each electron as an arrow.

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Pointing up or down depending on the direction of spin.

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Orbital diagrams are especially useful when writing down the values of the quantum numbers for a particular electron.

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Most atoms will have both paired and unpaired electrons.

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If the atom has any unpaired electrons, thanks to hund’s rule, most of them do.

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It will be paramagnetic and it will be attracted to a magnetic field.

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If the atom has only paired electrons.

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It will be diamagnetic and will be repelled by magnetic fields.

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What is the maximum number of electrons that can exist in n is equal to 3?

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That's the third shell.

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Pause the video.

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Write down your answer and then come right back.

00:10:56

Welcome back.

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Remember that the third shell has three subshells or rooms.

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S, p, and d subshells.

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Hold 1, 3 and five orbitals or beds respectively.

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Because each of these nine orbitals holds 2 electrons.

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The answer is 18 electrons total.

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Write the four quantum numbers of each of the 8 electrons of an oxygen atom in the ground state.

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Pause the video.

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Write down your answers.

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And then come back.

00:12:01

Welcome back.

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This question does not require you to write down the electron configuration.

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Or the orbital diagram, but they are easy to do and can be very helpful.

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So, following the Aufbau principle.

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Oxygen 8 electrons are placed two into 1S, 2 in the 2s and four in the 2p subshells.

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Here's an orbital diagram.

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In the making.

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The first electron goes into the 1s subshell.

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So n is equal to 1.

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l is equal to 0.

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And m_l has to be 0.

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And it has a.

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Clockwise spin so +½.

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So those the second one. But notice that m_s is determined by the direction of the arrow, so this is -½.

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The 3rd and 4th.

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Electrons will go in the 2s subshell.

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For the 5th electron.

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We will need the 2p subshell.

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For the 6th electron, we need to remember Hund's rule and place it on the next empty orbital.

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The same applies to the 7th electron.

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And the 8th electron pairs up.

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In the 1st 2p orbital.

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The one with the value of m_l of -1.

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One thing that surprises many students is that the shape.

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Of the periodic table.

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Is determined by the quantum model of the atom.

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The four sections of the periodic table are refered to as the s block.

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The p block.

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The d block.

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And the f block.

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If you are thinking why is helium in the P block when this does not even have a p subshell?

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My advice is: look the other way and don't bring that up.

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Ever again.

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Helium is the noble gases’, little sister and her brothers are very protective. JK.

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The elements in the s block all have electron configurations that end with an s subshell.

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The electrons in the p block all have electron configurations that end with a p subshell.

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Except for you know who.

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The elements in the d block all have electron configurations that end with a d subshell.

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And the elements in the f block all have electron configurations that end with an f subshell.

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There is a time saving convention that can be used to write the electron configurations of elements.

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That is most practical for elements that have moderate to large number of electrons, and it relies on the noble gases and the fact that they all have configurations that end with a full p subshell.

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Sulfur has 16 electrons.

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We will look for a noble gas that has close to 16 electrons but does not go over 16.

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Neon has 10 electrons and so this one will work.

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Because Neon is the 2nd noble gas, its electron configuration will end with 2p6.

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So if you write down neon in brackets.

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It's less if you have written down.

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$1s^2 2s^2$ and $2p^6$ already.

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When we consider the order in which orbitals are filled, we don't need to start from the 1s. We can start from the place where the chosen noble gas left off.

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Neon has 10 electrons.

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So when we write down neon in brackets, we have already accounted for all the orbitals that are found in neon.

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After the neon, we have the 3s subshell, which fits 2 electrons and that is 12 so far.

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Then follows

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The 3p subshell which fits up to six.

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So we place the last four electrons there.

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And done.

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Now Palladium, which has 46 electrons.

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The noble gas closest to 46, which does not go over is Krypton.

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The 4th Noble gas.

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Which means its electron configuration ends with a full 4p subshell.

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Since it's the 4th Noble gas 4p.

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That means we will skip all the subshells up to the four P subshell.

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And instead we will write down kripton in brackets, Kr.

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The next subshell is 5s.

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And it brings us up to 38 electrons.

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Next is the 4d where we can place the remaining 8 electrons.

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And done.

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The electron configurations of the elements in the periodic table will give us an insight into their chemical properties.

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We are going to be looking first at the elements in the first period.

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A hydrogen with its single electron.

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You may remember that hydrogen sometimes donates 1 electron when forming compounds.

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Not surprising because it has one electron to donate, but it sometimes accepts 1 electron.

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Because it has room for one more in its partially filled subshell.

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Helium as a noble gas does not accept electrons.

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Because it's already full.

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There's no place for him to put any more electrons.

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Now looking at the second period.

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Lithium has a single electron in its outer shell.

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And it is unsurprising that it donates 1 electron.

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Beryllium has two electrons in its outer shell.

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And it's in the habit of donating 2 electrons.

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Boron, a metalloid sometimes donates its three outer shell electrons and sometimes accepts 5 because it has five available places for accepting electrons.

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Quick reminder to follow Hund’s rule. Follow Hund’s rule.

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Carbon, a nonmetal with room for four electrons, accepts 4 electrons.

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Nitrogen, a nonmetal accepts 3 electrons because it has room for three electrons.

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You may start to see a pattern.

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Oxygen has room for two electrons, and that's what it will accept.

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Fluorine has room for one electron, and it will gladly accept one.

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And Neon has no vacancy and therefore accepts no electrons.

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As you can see, the charges that the elements have when they form ions are explained by their electron configurations.

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Who knew?

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We will now see that the same trend also applies to the third period.

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Sodium a Group One metal has one electron in its outer shell, and it tends to donate it.

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Magnesium in Group 2 donates 2 electrons because it has two electrons in its outer shell.

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Aluminum donates 3 electrons because three electrons occupy its outer shell.

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Silicon, a metalloid accepts 4 electrons because it has room for them in its 3p subshell.

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Phosphorus, like nitrogen, has room for three electrons in its outer shell.

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So that's what he accepts.

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Sulfur, another nonmetal will accept 2 electrons to fill up its outer shell.

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Chlorine will accept one electron because that is all it needs to have a full outer shell.

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And of course, argon with no room to spare, will receive no electrons from anyone.

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Since this outer shell is completely full.

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The trend continues for every period.

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But there are six exceptions to the Aufbau principle that I need to show you and we need to see why transition metals tend to be polyvalent.

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So one last period, let's look at the 4th.

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Potassium has one electron in its outer shell, so as expected it donates 1 electron.

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Calcium has two electrons in its outer shell, and it donates them.

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Scandium is the first of the transition metals.

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It has two electrons in its outer shell.

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You see, the 3d is not part of the outer shell, only the 4th shell is the outer shell.

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But it has the option of donating an electron from the 3d subshell because it is close in energy.

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To the outer shell electrons.

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This is why transition metals tend to be polyvalent,

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They have options.

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Titanium has two electrons in the 3d subshell.

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Vanadium, as expected, has three electrons in the 3d subshell.

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And next we have one of the exceptions to the Aufbau principle.

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We expect that chromium would have 4 electrons in the 3d subshell.

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But it does not.

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Instead, it has five.

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Notice that one of the chromium’s outer shell electrons has been moved up to the 3d subshell.

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To get a half-filled subshell.

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This will also be true of the other metals in Group six of the periodic table.

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Manganese is next and we're back to normal.

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Iron is next with six electrons in the 3d subshell.

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Then cobalt.

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And nickel.

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With seven and eight electrons in the 3d subshell, respectively.

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Copper is another exception to the Aufbau principle.

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Getting a full 3d subshell by promoting one of its outer shell electrons.

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This applies to other metals in Group 11.

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And with zinc we are back to normal.

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Remember that zinc and cadmium always donate 2 electrons, to donate, any more than that, would lose the full 3d subshell.

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Copper donates either one or two electrons, but one electron is more common.

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Gallium has three electrons in its outer shell and it will donate 3 or 5 when forming ions.

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Germanium follows with four outer shell electrodes and a willingness to donate either two or four.

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Even though it is not a transition metal, neither is gallium. They do have a 3d subshell from which they can draw. Extra electrons are to be polyvalent.

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Arsenic has a tendency to accept 3 electrons to get a full outer shell.

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Selenium, as expected, will accept 2 electrons.

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And bromine will accept one electron.

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And of course, Krypton will have none of that.

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Because that's all there is.

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There isn't any more.