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Chromium electron configuration chemguide

THE 3D AND 4S ORBITAL FILLING ORDER This page analyzes some of the problems with the usual way of explaining the electronic structures of the d-block elements based on the order of filling of the orbitals d and s. I'm grateful to Dr. Eric Scerri of UCLA, who noted these problems to me and provided me with some useful academic papers that I otherwise wouldn't have been able to accomplish. The way orbital filling order is normally taught gives you an easy way to work on the electronic structures of the elements. However, problems are thrown when you come to explain various properties of transition elements. This page takes a closer look at this, and offers a more accurate explanation that avoids problems. The commonly taught version This section is just a summary of the way it is currently taught. It takes away from things you've probably already read elsewhere in Chemguide. You shouldn't find anything unknown. The order to fill orbitals Electrons fills low-energy orbitals (closer to the core) before filling them with higher energy. When there is a choice between orbitals of equal energy, they fill the orbitals in a singular way as far as possible. The diagram (not scaled) summarizes the energies of orbitals up to level 4p. Rarity is the position of 3D orbitals. They are displayed at a slightly higher level than the 4 - and so it is orbital 4s that will be filled first, followed by all 3D orbitals and then 4p orbitals. Similar confusion occurs at higher levels, with so much overlap between energy levels that 4f orbitals are not filled until after 6, for example. I just want to focus on the fourth period. The beginning of the fourth period Everything is simple to this point, but the 3-level orbitals are not all full: 3D levels have not yet been used. But if you refer back to the orbital energies, you will see that the next lowest energy orbital is the 4s – so it fills up first. d-block elements d-block elements are considered as elements in which the last electron to be added to the atom is in an orbital d. (Actually, that turns out not to be true! We'll go back to that in detail later.) The electronic structures of the d-block elements are shown in the following table. Each additional electron you add usually enters a 3D orbital. Most chemistry books and chemistry teachers try to explain the pauses in pattern in chromium and copper - but not very convincingly. I'll go back to that later too. To make the table seem less complicated, I'm using [Ar] to represent 1s22s22p63s23p6. Making positive ions from block ions this is probably the most unsatisfactory thing about this approach to the electronic structures of blocking elements throughout the chemistry of transition elements, orbital 4s behaves like orbital outer and higher. The inverted order of 3d and 4s orbits only seems to apply to the construction of the atom in First. In all other aspects, 4s electrons are always electrons to think first. 4s electrons are first lost followed by one of the 3D electrons. What about this version? Evidence of ion formation This last bit about the formation of ions is clearly unsatisfactory. We say orbitals 4 have lower energy than 3D, so orbitals 4 are filled first. We know that 4 electrons were first lost during ionization. The lost electrons first come from the highest energy level, farther away from the influence of the nucleus. So orbital 4s should have a higher energy than 3D orbits. These statements are directly opposed to each other. They both can't be right. When talking about ionization energies for these elements, it is spoken in terms of 4s electrons as the outer electrons that are protected from the core by inner 3D levels. We say that the first ionization energies do not change much through the transition series, since each additional 3D electron more or less screens the 4s electrons of the extra proton in the nucleus. Explanations around ionization energies are based on 4 electrons that have more energy, and are therefore removed first. Where is the error that is usually taught? The usual way to teach this is an easy way to figure out what the electronic structure of any atom is - with some strange cases to learn like chromium or copper. Problems arise when you try to take it too literally. It's the way structures work – no more than that. The error is in the diagram we started with: Let's draw this diagram, and then assume it works for all atoms. In other words, suppose that the energies of the different levels will always be the ones we draw in this diagram. If you stop and think about it, that must be wrong. As you move from element through the periodic table, additional protons are being added to the core, and additional electrons around the core. The various attractions and repulsions in atoms are bound to change as you do - and it's those attractions and repulsions that govern the energies of different orbitals. This means that we need to rethink this on the basis that what we have drawn above is not likely to seem the same for all elements. The elements to the argon There is no problem with these. The overall pattern we have drawn in the above chart works well. Potassium and calcium The pattern is still working here. Orbit 4s has lower energy than 3D, and so it fills up below. This fits fully with potassium and calcium chemistry. The elements of block D And now everything is wrong! For reasons that are too complicated to enter this level, once the scandium is reached, the energy of 3D orbitals becomes slightly lower than that of the 4, and this remains in the rest of the transition series. So instead of working the electronic structure of scandium by imagining that you just launched the electron in a calcium atom, with the electron entering a 3D orbital because the 4 is already full, you really need to look at it more carefully. Remember that, actually, for Sc via Zn 3D orbitals have the least energy – not the 4. So why isn't scandium [Ar] 3d3 instead of [Ar] 3d14s2? Doing Sc3+ Imagine you are building a scandium atom from proton boxes, neutrons and electrons. You have built the core from 21 protons and 24 neutrons, and now you are adding electrons around the outside. So far 18 electrons have been added to fill all levels up to 3p. Essentially you have made the Sc3+ ion. Doing Sc2+ Now you will add the following electron to do Sc2+. Where will the electron go? The 3D orbitals in the scandium have lower energy than the 4, so the next electron will move to a 3D orbital. The structure is [Ar] 3d1. Doing Sc + You would expect the next electron to enter a lower 3d orbital energy too, to give [Ar] 3d2. But it doesn't. You have something else to think about here too. If you add another electron to any atom, you are required to increase the amount of repulsion. Repulsion increases the energy of the system, making it less energetically stable. Obviously it helps if this effect can be kept to a minimum. 3D orbitals are quite compactly arranged around the core. The introduction of a second electron into a 3D orbital produces more repulsion than if the next electron entered orbital 4s. There is no very large gap between the energies of 3d and 4s orbits. Reducing repulsion more than compensates for the energy needed to do so. The most energy-stable structure for SC+ is therefore [Ar] 3d14s1. Doing sc put the final electron, to make an atom of neutral scandium, needs the same kind of discussion. In this case, the lowest energy solution is the one that the last electron also enters level 4s, to give the family structure [Ar] 3d14s2. Summary In each of these cases we have looked at, 3D orbitals have the lowest energy, but as we add electrons, repulsion can push some of them towards the highest level of 4s energy. If you accumulate the scandium atom from scratch, the last electrons to enter are the two 4s electrons. These are electrons at the highest level of energy, so it stands to reason that they are first removed when the scandium forms ions. And that's what happens. 4s electrons are also clearly the outer most outer electrons, thus defining the radius of the atom. The lowest energy 3D orbitals are within them, and will contribute to screening. There is no longer any conflict between these properties and the orbital filling order. The electronic structures of two more elements of block D The difficulty with this approach is that it cannot be used to predict the structures of the rest of the elements of the transition series. In fact, what you have to do is look at the actual electronic structure of a specific element and its and then after about what should be happening in terms of the energy gap between 3d and 4s orbitals and repulsions between electrons. Electrons.

