15.1: Interacting Electrons, Energy Levels, and Filled Shells

In fact, electrons do interact with each other. In the previous chapter, we made arguments that these interactions should be smaller than the interaction with the nucleus. Because electron probability clouds are spread out, and outer shell clouds only have relatively small overlap with inner shell clouds, often, especially when viewing inner shells, you can approximate them as just lowering the net effective charge of the proton. That is, if you look at a Sodium atom, it has 11 electrons. The first 10 electrons will fill up the $1s$, $2s$, and $2p$ states. That leaves the outermost electron in the $3s$ state. Because there isn't a whole lot of probability for that $3s$ electron to be found where the inner electrons are usually found, you could approximate the situation for that outer electron that it's orbiting a ball of charge with a net charge of +1 (in atomic units), neglecting the fact that that charge is made up if +11 in the tiny nucleus and −10 in the outer electron cloud. However, even though interactions between electrons are secondary to the interaction between each electron and the nucleus, they are there, and they do ultimately have a lot of influence as to how elements at different places on the periodic table behave.

One of the primary effects of electron interactions is that the s, p, and d orbitals for a given value of n are not at exactly the same energy. In a Hydrogen atom—or any ion that only has one electron— they are, to a fairly good approximation. If there is more than one electron, however, the electron-electron interactions modify the energies of these states. In general, levels with higher l will be higher energy states than levels with lower l but the same n. In the absence of something external (such as a magnetic field), levels of different m but at the same n and l will still have approximately the same energy. Sometimes, you will find levels with a higher n but a lower l to be at a lower energy level than levels with a lower n and higher l. For instance, the $4p$ states tend to be filled before the $3d$ states. This isn't always a hard and fast rule; sometimes you will see the states filled out of the "standard" order. The interactions between electrons make the entire system a many-body system, and many-body systems are often notoriously difficult to solve in Physics.

For the most part, atoms are "happiest" (if you will allow for some anthropomorphization for purposes of discussion) if the number of electrons equals the number of protons. If there is one too many electrons, the ion will generally be happy to give away one of its negative electrons to the first positive charge that goes along. Likewise, if there is one too few electrons, the ion has an extra positive charge, and will tend to snap up any spare electrons in its vicinity.

However, this is not the only consideration for atom happiness. Atoms also like to have a filled shell. That is, Helium is more chemically stable than Hydrogen, because whereas Hydrogen only has one of two possible electrons in the $1s$ state, Helium has entirely filled the $n=1$ shell by placing two electrons in the 1s state. Likewise, Neon, with 10 electrons, has filled up both $1s$ states, both $2s$ states, and all six $2p$ states, making it a very chemically stable element. The elements down the right column of the Periodic Table are called "noble gasses". They are so called because they are chemically stable, and don't tend to interact with other atoms or form molecules. (They're noble, and thus above it all, or some such. Doubtless sociologists of science love to tear apart this nomenclature to display cultural bias in scientists.) The reason they are so stable is that each one of these noble gasses is an element that has just completely filled a set of p orbitals. (The one exception is Helium. It has completely filled the $n=1$ shell, where there are no p orbitals.) Ne has completely filled its set of $2p$ orbitals. Ar has completely filled its set of $3p$ orbitals. Kr has completely filled its set of $4p$ orbitals. And so forth.

You can get a first guess at the chemical properties of an element by comparing how close it is to a noble gas. If an element has just one or two electrons more than a noble gas, the easiest way for it to be more like a noble gas would be for it to lose an extra electron. Elements like these are more apt to form positive ions than negative ions. An example is Sodium. Sodium has atomic number 11. The first 10 electrons fill up the $1s$, $2s$, and $3p$ orbitals; that is, they're like a Neon inner core. Then, just outside that, is a single $3s$ electron. If Sodium loses that electron, then it is electrically positive, but now it has a happy noble-gas-like electron configuration. In contrast, Chlorine has 2 electrons in the $3s$ shell and 5 electrons in the $3p$ shell. All it needs is one more electron to have a full $3p$ shell, giving it the electronic configuration of Krypton. If you put these two elements together, each Cl atom will tend to take away an electron from each Na atom, leaving the Cl a negative ion and the Na a positive ion. Those two ions then will have an electrostatic attraction towards each other as a result of their opposite charges. The result is a crystal, Sodium Chloride, more commonly known as salt. In this case, the bonds holding the crystal together are "ionic bonds". In most molecular bonds, an electron is shared between elements. In this case, however, the Sodium is so eager to get rid of an electron and the Chlorine is so greedy for another one that effectively the electron transfers all the way across from the Na to the Cl.