Prediction of Element Properties Based on Periodic Trends

If you've ever taken a chemistry course in person, you've undoubtedly seen a large periodic table poster plastered somewhere in the room. You've probably also learned the about qualities and characteristics of different elements.

But the beauty of the periodic table is that it's organized in a specific fashion for a reason. The periodic table is arranged to make certain characteristics easy to deduce for every element, as long as a few core trends are understood.

Overview

- In this lesson, you'll be able to look at the periodic table and deduce the relative
- Atomic size and Ionic Size
- Ionization Energy and Electron Affinity
- Electronegativity
- Metallic Character
- and the Melting Point for each element.

Atomic Size and Ionic Size

If I were to tell you to describe the physical properties of atoms and molecules, where do you think you'd start? Perhaps you'd tell me that they make up everything, they can be in different states or something of that sort. However, I'd reckon that you'd include in there somewhere that they're small- really, really small. While this is certainly true, not all atoms are created equal. The periodic table can actually be read in terms of atomic radius.

If we want to look for the radius of an atom, we need its size. Intuitively, we can think of this radius as reaching from the center of its nucleus to the edge of the furthest orbital in its electron cloud. However, because atoms are so hard to observe, we can’t measure them directly in this fashion. There is a clever workaround to this, though. If we were to place two atoms directly next to each other, it's more feasible to measure the distance between the two nuclei, and then half it. This will give us a rough idea of what the radius is for an atom, and therefore, its size.

Definition

Atomic radius can be found by placing two identical atoms next to each other and halving the distance.

If we did this for every element on the periodic table, we’d find that **the radius of atoms increases as we move from the right to the left across each period, and will also increase as we move down each group**. We know that this is true because as we go down each group, we're adding additional orbitals to the electron cloud, and therefore expanding the atomic size.

So we've established a rule: *generally, more electrons means a larger radius*.

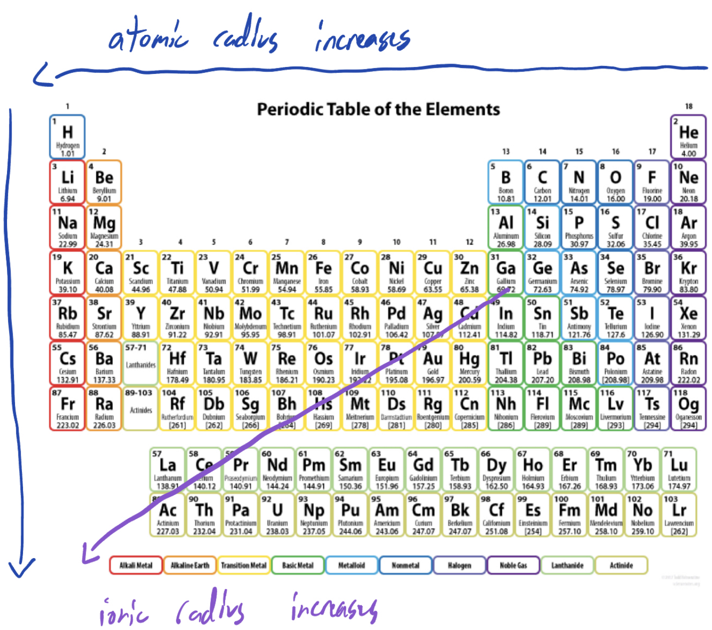
Let's try to apply this rule to ions. If you recall, we get an ion when we take an atom and either remove or add electrons. For example, we know that table salt, NaCl, is just two ions in a complex: Na^+ and Cl^-. (Remember that a negative charge implies an additional electron, and a positive charge implies a missing electron.) So what happens when we look at these ions under our new rule? As it turns out, it holds true. The atomic radius shrinks as Na is converted into Na^+, and Cl's radius increases as it is converted from Cl to Cl^-.

From this, we can derive our second rule: *cations decrease the atomic radius, and anions increase the atomic radius*.

If you recall, each element has an ion that it gets converted to most often. Therefore, if we apply our two rules together, we find another trend, **the ionic radius is proportional to the number of electrons**.

> Definition
>
> **Ionic radius** is proportional in size to the number of electrons within the ion. More electrons mean a larger radius because of additional orbitals, therefore, cations are smaller than the atomic radius and anions are larger.

To visualize this trend, here it is applied to the periodic table.



Drawn periodic trend for atomic and ionic radius.

Ionization Energy and Electron Affinity

For the next set of periodic trends, recall that atoms are significantly more stable when their valence electron shells are full. This means that when atoms are extremely close to having a full outer shell, they will more readily give up or accept electrons to get there. Let's define our Ionization Energy and Electron Affinity before we go further.

Definition

Ionization Energy is the amount of energy required to remove one electron from an atom. A low value means it is incredibly easy to remove an electron. A high value means it takes a large amount of energy to do so, making it difficult.

Definition

Electron Affinity is the energy released when one electron is added to an atom. A high electron affinity means that an atom more readily accepts an electron, while a low electron affinity means it is less so.

Now, let's try to think about this practically for a moment. Our atoms don't particularly care how many shells they have. Instead, they care much more about whether or not the valence shell is full. For example, let's consider Oxygen.

Example

Oxygen has a noble gas electron configuration of [He] $2s^2$ $2p^4$. This means that it's only two electrons away from a full valence shell... two electrons away! (This is even more apparent if you're looking at a periodic table.) This means that if we add two electrons, we would fulfill our 'full shell' requirement.

Now, what if we were to take away electrons instead? We would have to get rid of four electrons to make 2s2 our full valence shell. This isn't the optimal way to get there, as we want to get to our full valence while doing the least amount of work.

Through this, we can deduce that Oxygen would much rather add electrons than take them away. By our definitions, we can now assume that it has a high electron affinity and high ionization energy, meaning it is easy to add electrons and hard to take them away. Again, this is so that we can fill our valence shell with maximum efficiency.

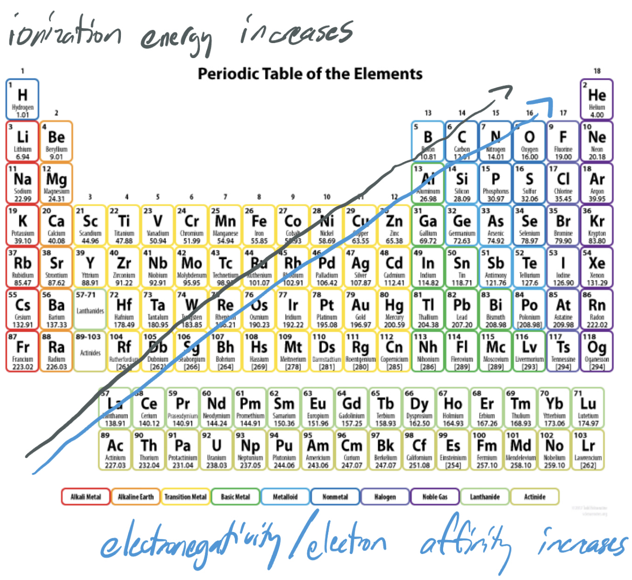
We could easily show you a trend for ionization energy and electron affinity. However, this much more intuitive approach will help you during the AP exam to visualize what is going on with each element. We can do one more example with sodium.

Example

Sodium has a noble gas electron configuration of [Ne] $3s^1$. This means that if we wanted to get a full valence electron shell, we would just have to take away one electron. This means that sodium has a low first ionization energy- as it takes only a small amount of energy to remove one electron from sodium. This is also why we see Na readily forming into the ion Na+.

Let's take this a step further. Taking away one electron from sodium gives it a full valence shell, meaning its ionization energy is low. However, what if we wanted to take away another electron? This would break the full valence shell rule, making the atom much more unstable. Obviously, sodium doesn't want that. Therefore, the second ionization energy (removing a second electron) is much higher.

Hopefully, you're starting to get a grasp of how this intuitive method makes sense. If we were to do this with every element, we'd see something like this.



Drawn periodic trend for ionization energy, electronegativity and electron affinity.

Electronegativity

Electronegativity and electron affinity are very closely related. While electron affinity is the actual energy released when we add an electron, electronegativity is the actual characteristic of an atom wanting to have an electron added to it.

Definition

Electronegativity is the tendency to attract electrons towards an atom. In other words, the higher an element's electronegativity, the more attractive it is to electrons.

Therefore, we can intuitively say that elements that have a high electron affinity will have a high electronegativity. This logically makes sense: if an element wants to have electrons added to it, it's highly electronegative. So when this electron is added, a high level of energy will be released, indicative of a high electron affinity.

The trend for electronegativity will be the same as the trend for electron affinity, as pictured in the graph in the previous section.

Metallic Character

You've probably already covered that there are metals and nonmetals on the periodic table. As you might've gathered by the previous covered periodic trends, metals typically lose an electron when in a reaction. (Recall the sodium example we discussed earlier!) Just as electronegativity describes how readily an element accepts electrons, metallic character describes how readily an element gives up an element.

> Definition
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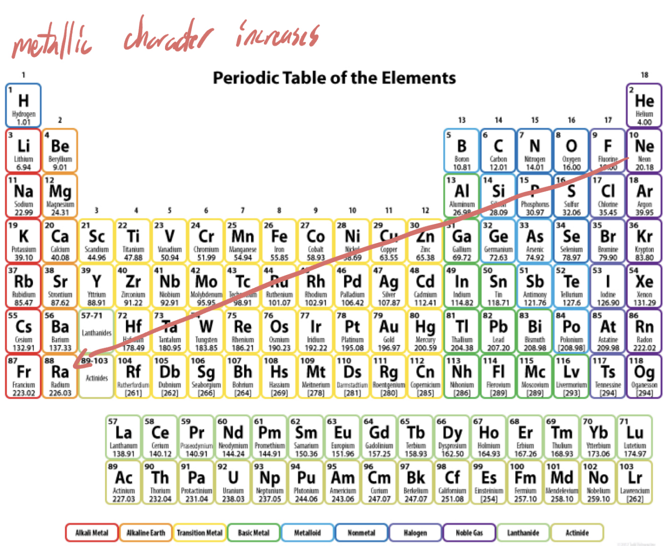
> **Metallic character** is the opposite of electronegativity. It describes the readiness of an atom to give up an electron.

When you move across a period from the right side to the left side, metallic character increases because of a diminished interaction between the nucleus and the valence shell. It also increases as you move down a group. This is because additional shells are added, further distancing the valence electrons from the nucleus. This makes the nucleus' pull on the valence electron weaker, resulting in a valence electron that is more easily given up.

If we were to visualize this trend, it would look something like this.

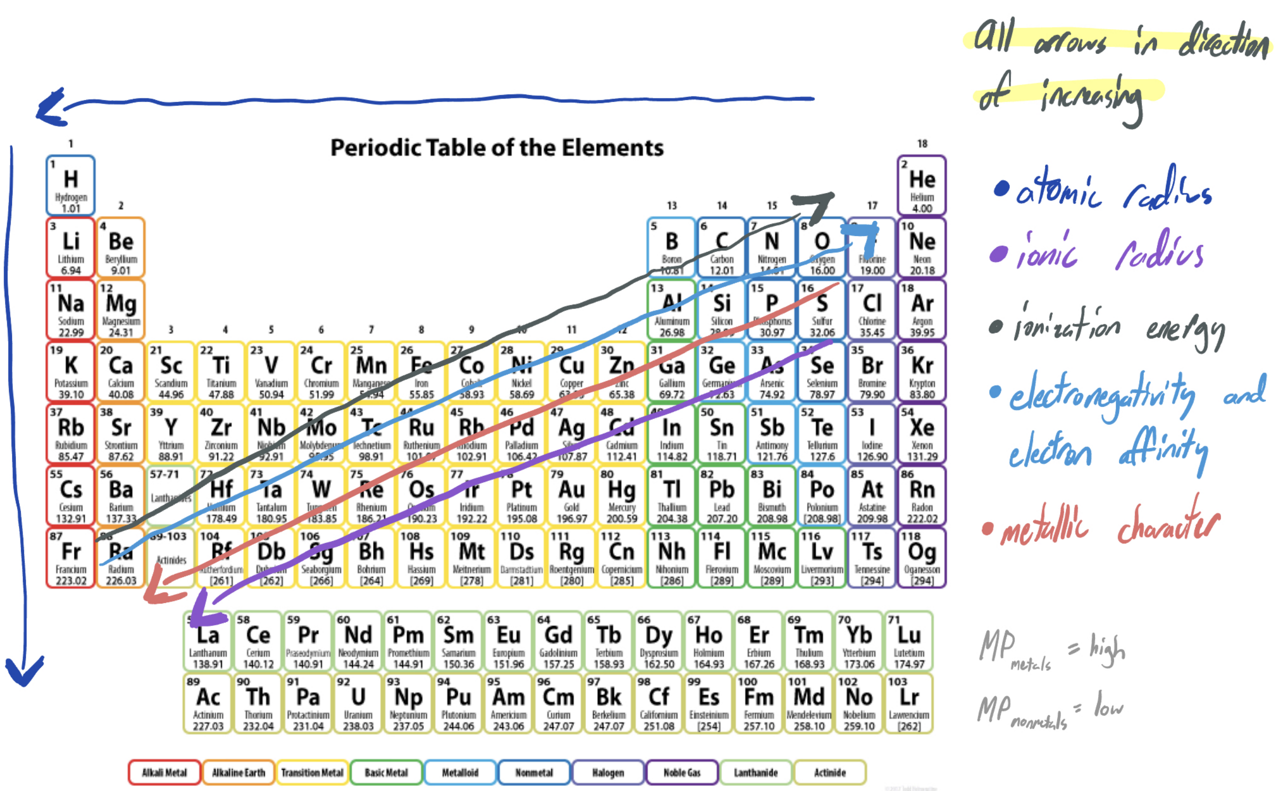


Drawn periodic trend for metallic character.

Melting Point

Lastly, we have the characteristic of melting point. While there isn't a grand trend on the periodic table like the other trends that can be easily observed, there are a few key rules that can be observed on the periodic table. Metals usually have high melting points, while non-metals usually have low melting points. These have been experimentally determined, and while these rules are somewhat intuitive, you most likely won't be tested on a melting point periodic trend on the AP exam.

Hopefully, the methods that we've described help to give you a stronger grasp on periodic trends. Being able to predict the qualities of different elements by remembering a few key trends will eventually help you to predict reaction conditions, and will prove to be useful as you advance into higher-level chemistry courses. To put everything together here are all of the periodic trends put together.



Drawn graph of the overall periodic trends.

Prediction of Element Properties Based on Periodic Trends - Key takeaways

- In this lesson, you learned how to discern an element's relative

- Atomic Size and Ionic Size
- Ionization Energy and Electron Affinity
- Electronegativity
- Metallic Character
- and the Melting Point just by looking at the periodic table.