Ionization Energy

Skills to Develop

- Describe the significance and periodic trend of ionization energy

Ionization energy is the energy needed to remove an electron from an atom or ion. Unlike atomic radii, we can and do measure ionization energies in the gas phase, when the atom or ion is not interacting with anything else. The first ionization energy, IE_1, is the energy of this reaction

$$A(g) \rightarrow A^{+}(g) + e^{-}(g) \tag{1}$$

The **second ionization energy**, IE_2, is the energy of

$$A^{+}(g) \rightarrow A^{2+}(g) + e^{-}(g) \tag{2}$$

You can also measure third, fourth, etc. All of these assume that the highest energy electron is knocked off. If you are removing core electrons, you would write them like this: IE_{1s}, to show what orbital the electron comes from.

How do you Measure Ionization Energy?

Generally you do this by shining high energy photons on the material you want to study. UV will ionize valence electrons and X-ray can ionize core electrons. You can vary the wavelength of the photons, and also measure the kinetic energy of the electrons that come off, to see what the binding energy (orbital energy) for each electron is. You can also measure the ionization energy of molecules, which is very interesting because it can tell you how the orbital energies change because of bonding.

Why does it matter?

Ionization energy tells us how likely an atom is to form a cation, and if so, what charge. In general, it tells us how tightly the electron is bound, how stable it is. It can tell us the energies of real orbitals, the effects electrons have on each other, and help us predict reactivity and properties of molecules. (We'll talk more about this when we get to bonding!)

Predicting Relative Ionization Energies

Ionization energy depends on orbital energy, which depends on the type of orbital and the effective nuclear charge. Thus, it follows predictable patterns in the periodic table. As you go down, n increases, and the energy of the orbital increases. That means the orbital is less stable, so it's easier to pull off the electron, so ionization energy decreases. (Stable, bound electrons have negative energies; electrons that aren't in an atom have 0 energy.) As you go across the periodic table, usually the type of orbital is the same, and the effective nuclear charge increases, making the orbital more stable, so ionization energy increases. But when you change subshells, the ionization energy might increase less, because the new subshell is less stable. Also, remember Hund's rule: electrons are more stable when they don't share the same orbital. So if you have to put a new electron into an occupied orbital, that also makes the ionization energy increase less.

You can predict any relative IE just by thinking about how big the Coulomb forces are. Bigger n, means bigger distance, weaker force. Removing an electron from a neutral atom is easier than removing an electron from a cation, because of the charge. Also consider effective nuclear charge and electron-electron repulsions (especially in the same orbital). In summary, mostly IE increases up and to the right, because of low shells and high effective nuclear charge.

Look at IE for yourself!

Go to Ptable's ionization energy page. You can look at first, second, third, etc. See the general trend (bigger up and right), and also notice some exceptions as you change blocks (s-block or d-block to p-block).

Outside Links

- Khan Academy: Periodic Table Trends - Ionization Energy (12 min)
- CrashCourse Chemistry: The Electron (13 min)

Contributors and Attributions

- Emily V Eames (City College of San Francisco)
