# Periodic Trends

**Unit 7:** Periodicity

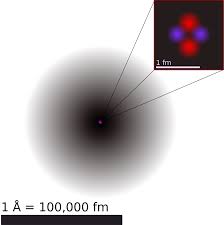
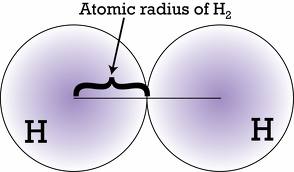
### Knowledge/Understanding Goals:

* ionization energy, electronegativity, electron affinity, atomic radius, ionic radius

### Notes:

## Atomic Radius

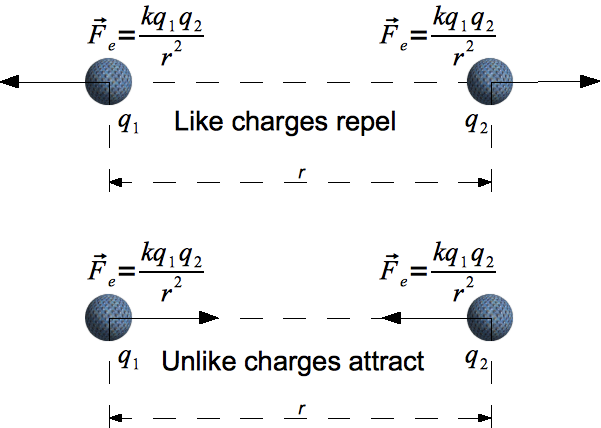
atomic radius: the average distance from the nucleus to the outermost electrons in an atom. The atomic radius is a measure of the “size” of the atom, typically measured in angstroms ([Å](http://en.wikipedia.org/wiki/%C3%85))



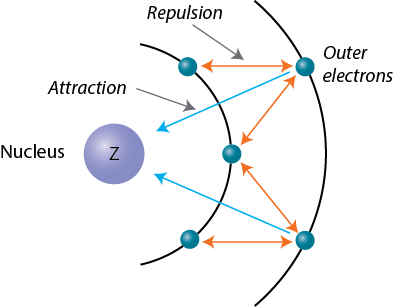
Factors Influencing Radius:

Energy levels: The more energy levels occupied by electrons, the larger the radius as you are increasing the physical distance from the nucleus to the valence level.

Nuclear Charge: The larger the nuclear charge (greater # of protons), the smaller your radius as electrons are pulled in tightly due to high electrostatic (charge) attraction.

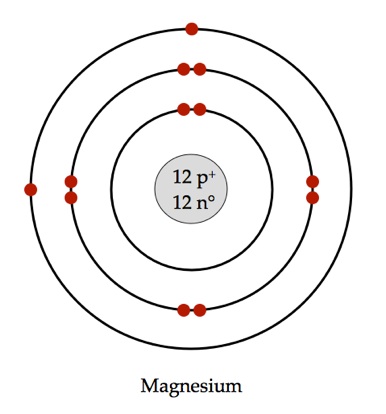


Shielding: when core electrons in the lower energy levels (closer to the nucleus) shield (block) some of the nucleus’ positive charge. This causes the outer electrons to be held less tightly by the nuclear charge, and the atom gets larger.

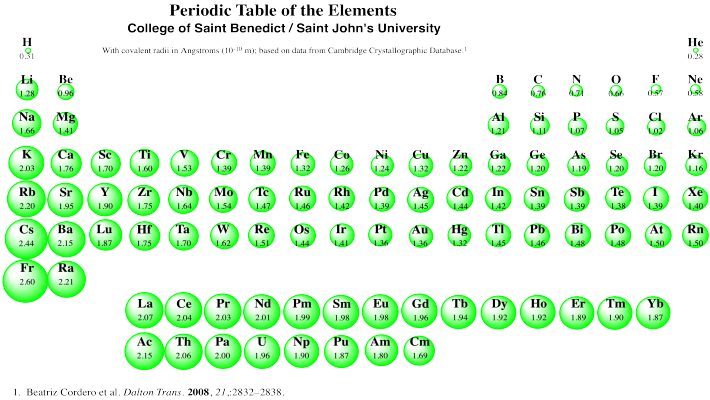


* Shielding and nuclear charge can be combined into the term ***effective nuclear charge***; which is the charge electrons actually feel from the nucleus after shielding.

Example calculations:



**Radius Trend**



* Atoms of elements get larger as you move down a column, Why?
* Atoms of elements get smaller as you move to the right within the same period, Why?



* periodic trend: any measurable property that increases or decreases according to the element’s position on the periodic table.
* Atomic radius tends to get smaller as you move up and to the right on the periodic table, and larger as you move down and to the left.

\*Make sure you can explain why using effective nuclear charge and occupied energy levels in your discussion.

* Nearly all of the other periodic trends can be tied back to radius, and thus effective nuclear charge and energy levels.

**Ionic Radius**

Because most of the space that an atom takes up is outside the nucleus (where the electrons are), changing the number of electrons (making an ion) changes the size of the atom.

***Cations:*** If you take away electrons, the ion gets smaller. This means ions with a positive charge are smaller than the neutral atom and also smaller than an atom of the neutral element with the same number of electrons. This is because the positive ions have more unshielded positive charge, which pulls the electrons closer.

***Anions:*** If you add electrons, the ion gets larger. This means ions with a negative charge are larger than the neutral atom and also larger than an atom of the neutral element with the same number of electrons.

|  |  |  |  |
| --- | --- | --- | --- |
| **Element** | **Radius of neutral atom (pm)** | **Charge of ion** | **Radius of ion (pm)** |
| O | 73 | −2 | 126 |
| F | 72 | −1 | 119 |
| Ne | 71 | 0 | — |
| Na | 186 | +1 | 116 |
| Mg | 160 | +2 | 86 |

## Electronegativity

electronegativity: (χ) a comparative measure of the ability of an atom of an element to attract valence electrons from a different atom.

The concept of electronegativity was first proposed by Linus Pauling. Originally, Pauling assigned the elements in period 2 the values of Li = 1.0; Be = 1.5; B = 2.0; C = 2.5; N = 3.0; O = 3.5; and F = 4.0.

Pauling later defined the electronegativity difference (Δχ) in terms of the bond dissociation energies (energy required to break a bond):



Where DAB, DAA, and DBB are the bond dissociation energies (expressed in ) of an A−B bond, an A−A bond, and a B−B bond, respectively. (The 96.485 comes from the fact that electronegativity is defined in terms of electron volts (eV). 1 eV = 96.485 .)

How can EN be related back to radius?



\*Note the electronegativity of Noble Gases = 0

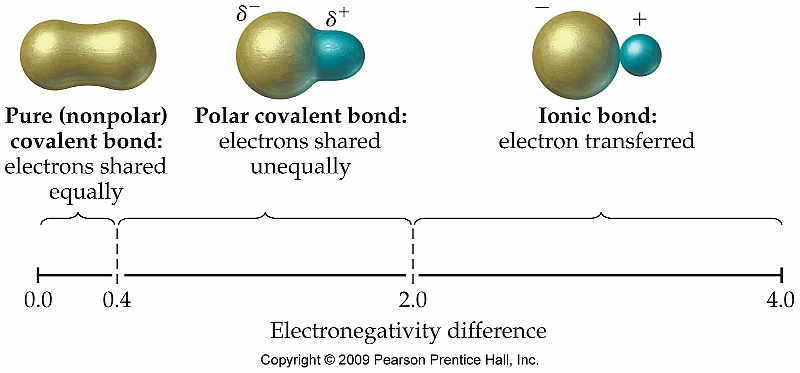
## Ionic vs. Covalent Bond Character

The main application of electronegativity values is the prediction of bond character between atoms. This will be discussed more in the polarity unit.

***Ionic Bond:*** Electrons are exchanged (given and received) due to a large difference (>1.7) in electronegativity between the bonding atoms.

***Polar Covalent Bond:*** Electrons are unevenly shared due to a slight difference (1.7-0.35) in electronegativity between the bonding atoms.

***Non-Polar Covalent:*** Electrons are evenly shared due to equal (<0.35 difference) electronegativity between the bonding atoms.

******Note that this bond characteristic scale does NOT have finite “cut-offs”, but is more of a continuum or spectrum from “more covalent” to “more ionic”.

All chemical bonds exhibit some covalent character and all chemical bonds between dissimilar atoms also exhibit some ionic character. We define a bond as being “ionic” or having “ionic character” if the bond character is at least 50% ionic.

The electronegativity difference (Δχ) can be used to predict how a chemical bond behaves. According to another of Pauling’s formulas, the percent ionic character of a chemical bond is given by the formula:



A bond with 50% ionic character corresponds with a Δχ of approximately 1.7.

## Ionization Energy

ionization energy: the amount of energy that it takes to remove an electron from an atom (making it into an ion). Ionization energy is a measure of how tightly an element holds onto its electrons.

How can IE be related back to radius?



In general, as energy sub-levels fill with electrons, they become more stable. This means the total energy of each electron in the sublevel decreases. As the electrons move to lower energy states, more energy is required to remove them, resulting in higher ionization energy values.

1st ionization energy: the amount of energy it takes to remove the first electron from an atom.

2nd ionization energy: the amount of energy it takes to remove a second electron from a +1 ion. The 2nd ionization energy is always higher than the first, because energy must be added to overcome the attraction due to the positive charge as well as the energy of the electrons in the sub-level. (Radius is *always* smaller after removing first e-)

3rd, 4th ionization energy, *etc.*: the amount of energy it takes to remove a third electron from a +2 ion, a fourth electron from a +3 ion, *etc.* As the positive charge of the ion increases, the ionization energy also increases. (Radius gets smaller and smaller)

There is always a particularly large jump in the ionization energy once the atom or ion has the same electron configuration as a noble gas (a “noble gas core”). This is because removing an electron from a “noble gas core” requires removing an electron from a lower energy sub-level that has additional stability (even lower energy) because all of its orbitals are filled.

Consider the following table of ionization energies:

|  |  |  |  |
| --- | --- | --- | --- |
| **Element** | **1st Ionization Energy (kJ/mol)**  (neutral atom) | **2nd Ionization Energy (kJ/mol)**  (+1 ion) | **3rd Ionization Energy (kJ/mol)**  (+2 ion) |
| Ne | 2081 | 3952 | 6122 |
| Na | 496 | 4562 | 6912 |
| Mg | 738 | 1451 | 7733 |

In the above table, neon (Ne) has the largest 1st ionization energy, because filled s and p sublevels are more stable, meaning it takes more energy to remove electrons from them.

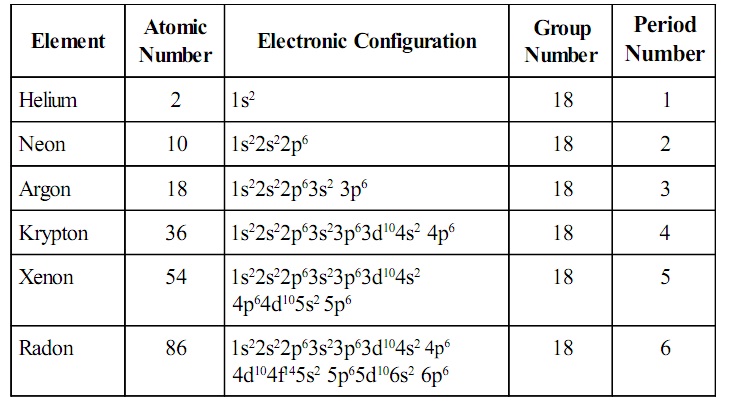
Sodium (Na) has the smallest 1st ionization energy. This is because removing one electron will remove the 3s electron entirely, leaving the remaining electrons are all in lower (lower energy and therefore more stable) sublevels. Also, those sublevels are filled—the +1 ion has a “noble gas core.” The filled sublevels are more stable (lower energy), and electrons in them require even more energy to overcome this added stability.

Magnesium (Mg) has the lowest 2nd ionization energy, because removing the second electron gives it a noble gas core. However, because the +2 ion has a noble gas core, the third electron is much more difficult to remove, giving Mg a particularly large 3rd ionization energy.

## Tying Stability to EN and IE

Hopefully at this time, you have noticed the correlation between the “stability” of the noble gas electron configuration and their inherent electronegativity and ionization energies.

Noble gas configurations are energetically favorable due to the s and p valence sublevels being completely full.



This stability is reflected in the EN and IE values for noble gases:

***Electronegativity = 0***

* Atom is no longer able to attract (add) more electrons to its valence shell, which would take it out of the stable noble gas configuration.

***Ionization Energy = Extremely High***

* Very difficult to remove (lose) electrons from its valence shell, which would take it out of the stable noble gas configuration.

## Electron Affinity

electron affinity: the energy required to remove an electron from a −1 ion. (It is sometimes also expressed as the energy released by the atom when forming the −1 ion. However, because the energy change for the atom is negative, this definition can create confusion over the sign of Eea.)

Higher electron affinities indicate higher stability of the negative ion, and therefore a stronger tendency for the atom to attract electrons.

Because the positive nucleus can attract and hold electrons, most atoms can form a stable −1 ion, even if the positive ions are more stable. However, because half-filled and completely-filled energy sublevels are more stable (lower energy), electrons in the middle and end of the p, d, and f blocks of the periodic table tend to be unable (or much less able) to form a stable −1 ion. These atoms have either very low or undefined electron affinities (“< 0” means that the −1 ion is unstable and no value can be measured).

There is a general trend toward higher electron affinities as you move to the right, toward the middle or end of each sublevel within a section of the periodic table.

## 