

Key for Periodic Trends in Properties of Elements II

The AVEEs for elements in the second period are shown in the following table. Explain the trend in these values. Compare this trend to that for the first ionization energies, IE_1 for these elements (see previous worksheet for the trend in IE_1). What is the difference in the trends? Explain why the trends are different.

element	Li	Be	B	C	N	O	F	Ne
AVEE (kJ/mol)	520	900	1170	1400	1820	1890	2310	2730

- The AVEE increases across a period for the same reason that ionization energies generally increase across a period: this is because the value for Z_{eff} increases across a row as the atomic number increases. The first ionization energies for these elements show a general increase in IE_1 , but there are a few places where the value of IE_1 decreases (from Be to B, and from N to O). The AVEE values do not show this effect because it averages the ionization energy of all of an element's valence electrons.

The AVEEs for the first four halogens are shown in the following table. Explain the trend in these values.

element	F	Cl	Br	I
AVEE (kJ/mol)	2311	1590	1530	1350

- The AVEE decreases as we go down a group because the valence electrons are further from the nucleus.

The electron affinities for elements in the second period are shown in the following table. Elements without a listed value do not have a measurable electron affinity. Explain the trend in these values. Offer an explanation for why Be, N, and Ne do not have measurable electron affinities.

element	Li	Be	B	C	N	O	F	Ne
electron affinity (kJ/mol)	60	—	27	122	—	141	328	—

- Setting aside Be, N, and Ne, the trend for the other elements in the second period is for the electron affinity to increase as we move across the row. The principal reason for this is an increase in Z_{eff} , which increases the energy of attraction between the nucleus and the electrons. For the three elements without a measurable electron affinity, Be has a filled $2s$ orbital, which means the electron must enter a less stable $2p$ orbital, N has one electron in each of its three $2p$ orbitals, which means the electron must pair up with another electron, and Ne has a filled $n = 2$ shell, which means the electron must enter a new shell.

The atomic radii for elements in the second period are shown in the following table. Explain the trend in these values.

element	Li	Be	B	C	N	O	F	Ne
atomic radius (pm)	167	112	87	67	56	48	42	38

- The atomic radii of elements decrease across a row. As we move across the row, the charge on the nucleus increases, which means that Z_{eff} increases. This increases the attraction between the nucleus and the electron, which pulls the electron in more closely.

The atomic radii for the first four halogens are shown in the following table. Explain the trend in these values.

element	F	Cl	Br	I
atomic radius (pm)	42	79	94	115

- The atomic radii decrease as we go down a group because the valence electrons are further from the nucleus.

The ionic radii for cations and anions of elements in the second period are shown in the following table. Explain the trend in these values.

ion	Li ⁺	Be ²⁺	B ³⁺	C ⁴⁺	N ³⁻	O ²⁻	F ⁻
ionic radius (pm)	76	45	27	16	146	140	133

- When comparing these ionic radii to the atomic radii it is clear that cations are smaller than their neutral atomic form and anions are larger than their neutral atomic form. Cations are smaller than their atoms because the loss of their valence electrons gives them an electron configuration of $1s^2$, which means their outermost electrons are now in a $n = 1$ shell instead of an $n = 2$ shell. In addition, having fewer electrons than protons, means the positive charge is distributed over a smaller number of negative charges, pulling them in. Anions are larger than their atoms because of the repulsion between electrons and because the positive charge on the nucleus is now spread out over more electrons.