

***Final Exam is scheduled for:**

Wednesday, May 21

8:00-11:00am

Snowden Auditorium (W-1-88)

-Let Dr. Sevian know by April 25 if you have a conflict with one of your other final exams.

Comparing Elements

In the same period:

- use the Z_{eff} argument because Z_{eff} increases as you move across (left to right) the periodic table.
- for example, the Z_{eff} for **Li** = $3-2= +1$ and the Z_{eff} for **Be** (the next element to the right of Li) = $4-2= +2$.
- this is the case because the charge in the nucleus increases as you move further right on the periodic table but the core electrons that shield the nucleus' charge remain the same.
- the amount of valence electrons increases when you move left to right on the periodic table

In the same group:

- use the # of shells argument because the # of shells increases as you move down a group, meaning that the radii become larger.
- the Z_{eff} remains the same

Electron Affinity

- an element's **electron affinity** is the energy associated with adding an electron to that element.
- this is measured in kJ/mol (See Lecture 21 Slides, bottom of page 2 for general trends in electron affinity in periodic table)
- the more negative the electron affinity, the more exothermic. An element with a really negative electron affinity means that it has a strong attraction to electrons. Examples of elements that are very negative electron affinity include **F**, **O**, and **Cl**.

CHAPTER 8

Ionic vs. Molecular Bonds (What we already knew)

Ionic Bonds: -occurs between ions

- ions are positively or negatively charged because of losing or gaining electrons respectively
- in a solid state, ions are arranged into regulating, repeating, alternating cations (+) and anions (-) to make a lattice structure (See Fig. 2.23b in book)
- materials formed by ionic bonds are solid at room temperature and are strong. Therefore, they have high melting points.
- remember that Coulomb's Law predicts that force of attraction between two oppositely charged ions depends on **BOTH** magnitudes of charges (direct relation) **AND** distance separating them (inverse relation).

Covalent Bonds (Review)

- Co** = sharing, **valent** = valence electrons
- occurs between neutral atoms within the same molecule.
- force of attraction that holds atoms together comes from each atom's nucleus attraction towards neighboring atom's electrons as well as its own.
- the covalent bonds that hold together atoms inside molecules are weaker than ionic bonds because in ionic bonding the + and - ions have strong attractions for each other, while in covalent bonding the atoms are neutral so what holds them together is the fact that the electrons that are shared between two atoms are attracted to both atoms

Gray Area

- there is a gray area between covalent and ionic bonding because covalent bonds can take on ionic character when electrons aren't shared equally, and then you have a partially + charged atom on one side of the bond that is held by an additional attraction (beyond covalent bonding) to a partially - charged atom on the other side of the bond
- before we get to the gray area, let's clearly define the two extremes: ionic and covalent

Lattice Energy In Ionic Compounds

- lattice energy** is referred to the amount of energy in kJ/mol it takes to break apart the bond.
- the **lattice energy** is equal and opposite in numerical value to the exothermic energy produced when the same mol of ions form compounds
- See Lecture 21 slides, bottom of page 9 for period and group trends in lattice energy
- *Lattice energy was stressed in lecture

Lewis Dot Structure Model for Explaining Covalent Bonding in Molecules

- use for the first three periods of the periodic table
- predicts molecular shapes
- theory is that valence electrons are distributed as either pairs of electrons that are shared by two atoms (shared) or pairs of electrons that belong to a single atom (unshared)
- Lecture 21 Slides, pages 8 and 9 are good visual examples of how to make Lewis structures.

Group Problem 7

- this was given to us at the end of lecture as homework to bring to class and turn in our group's answer on Thursday

Draw these Lewis structures:

- NO_2^+
- CN^-
- SCN^-
- O_3

- 1) Which two of the above structures are isoelectronic?
- 2) What do you notice that is special about isoelectronic species?

*isoelectronic = same # of valence electrons and same # of atoms to distribute them around.