

Unusual electron configurations

Transition metals: half-filled and fully-filled subshells have extra stability. d^5s^1 $d^{10}s^1$

Examples: Cr and Cu

Transition metal ions: s and d energy levels change order => all remaining s and d electrons move into the d subshell.

Example: Ti^{2+}

Periodic Trends

1. Size (radius), Figs. 7.20, 7.21 (p. 255, p. 257)

a. Depends on n and screening:

Larger n means larger orbital size.

=> expect 2nd row elements to be larger than 1st row elements, etc.

Smaller n orbitals (inner electrons) screen higher n orbitals (outer electrons) from the nucleus.

Screening: $e^- - e^-$ repulsions between the inner electrons and the outer electrons partially cancel the pull of the nucleus on the higher n electrons.

=> reduces the attraction force of the nucleus from its actual charge (Z) to an effective charge (Z_{eff})

=> allows higher n orbitals to expand away from the nucleus

=> expect 2nd row elements to be larger than 1st row elements, etc.

b. Depends on effective nuclear charge (Z_{eff}):

More protons (Z) means greater nuclear charge.

=> greater attractive force on electrons

=> electrons are pulled in closer

=> atoms get smaller

=> expect atoms to get smaller going to the right on the periodic table.

Example:

Which atom is larger, Li or K?

Which atom is larger, Na or P?

General trends:

Going down the periodic table

Going right across the periodic table

These are general trends that hold mainly for s- and p-block elements.

c. Depends on ionic charge:

Number of electrons vs. number of protons for cations (+) and anions (–).

Which is larger?

F or F[–]

Na or Na⁺

Rank in order from largest to smallest:

Na⁺

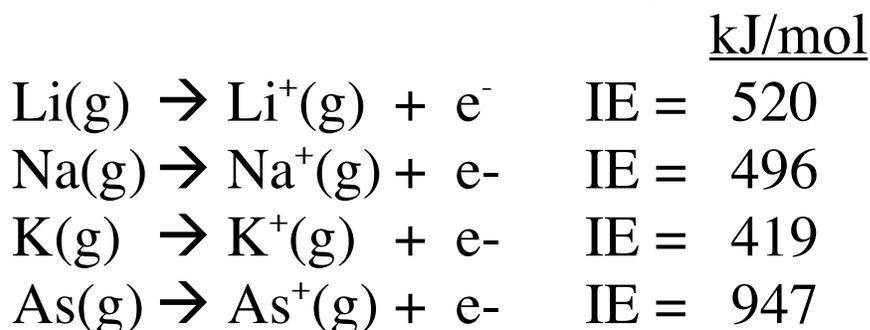
F[–]

Ne

2. Ionization Energy

Ionization energy (IE): the energy required to remove an electron from an atom or ion, $\Delta E = IE$, *generally* depends on size (Fig. 7.22, p. 259).

Measured for atoms in the gaseous state,



General trends:

Going down the periodic table

Going right across the periodic table

3. Electron Affinity

Electron affinity (EA): the energy required to add an electron to an atom or ion.



(negative energy means the process releases energy, other textbooks use a positive sign instead)

Not much of a general trend.

Bonding

Bonding orbital = atomic orbital overlap =
shared electron density = chemical bond.

The most simple bonding orbital (bond) is to
share the electrons between the nuclei.

Examples:



Only use valence atomic orbitals to form
bonding orbitals (only use valence electrons to
form bonds).

Covalent bond: a bond formed by shared
electrons, the bonding in molecular compounds.
(as opposed to ionic bonds in ionic compounds,
and metallic bonds in metals).

Electronegativity: χ (“chi”), the ability of an element to attract bonding electrons to itself, the ability of an element to pull electrons in a bond toward itself.

Unequal electronegativity means unequal sharing of the electrons.

Examples:



Figure 8.6. Values for electronegativity (p. 289)
Metals have low values. (don't want electrons)
Non-metals have high values. (want electrons)

F is the most electronegative element ($\chi = 4.0$).

Electronegativity decreases going down the periodic table (away from F) and going to the left (away from F).

Bond classification scheme

Pure covalent: no difference in χ ($\Delta\chi = 0$).

Polar covalent: some difference ($\Delta\chi$ moderate).

Ionic: large difference ($\Delta\chi$ large).

Really a continuum (Fig. 8.5, p. 288)

Practice:

Which element is more electronegative, C or O?

Which element is more electronegative, I or Cl?

Lewis structure

A picture that shows the covalent bonds for a molecule.

Does not show accurately the location of all the electrons in the molecule.

Does not show accurately the shape or bonding angles of the molecule.

Example: H₂O

Drawing and understanding Lewis structures are starting points for Chapter 9 (CHEM 102) and for organic chemistry. You need to practice drawing and understanding them until they are easy for you. (Sections 8.2-8.4, 8.7-8.8)

General rules (see handout):

1. Count the valence electrons.
2. Arrange the atoms. Usually a central atom is surrounded by ligands (outer atoms or groups). The outer groups are the more electronegative atoms; the central atom is usually the least electronegative atom because it will be forced to share electrons with the outer groups. H is never a central atom.
3. Distribute bonding pairs between atoms (for each bond, 2 electrons are shared by the bonded atoms). *Note that C prefers 4 bonds, N prefers 3 bonds, O prefers 2 bonds, and halides (F, Cl, etc.) and H prefer 1 bond.*
4. Give as many of the leftover electrons as possible to the outer atoms, but remember the Octet Rule: For 2nd row elements, only 4 pairs (8 electrons) can be touching each atom. Other elements are not required to follow the Octet Rule. *For 1st row elements, only 1 pair.* Why do we give the leftover electrons to the outer atoms before giving them to the central atom?

5. Give all remaining electrons to the central atom. Note the possible exceptions to the Octet Rule.

6. Modify the structure using the Octet Rule and the formal charges to help.
 - Use the Octet Rule to make multiple bonds by sliding electron pairs from outer atoms to central atom.
 - Reduce formal charges to decide when to make multiple bonds.
 - Note the exceptions to the Octet Rule.

Examples:

