

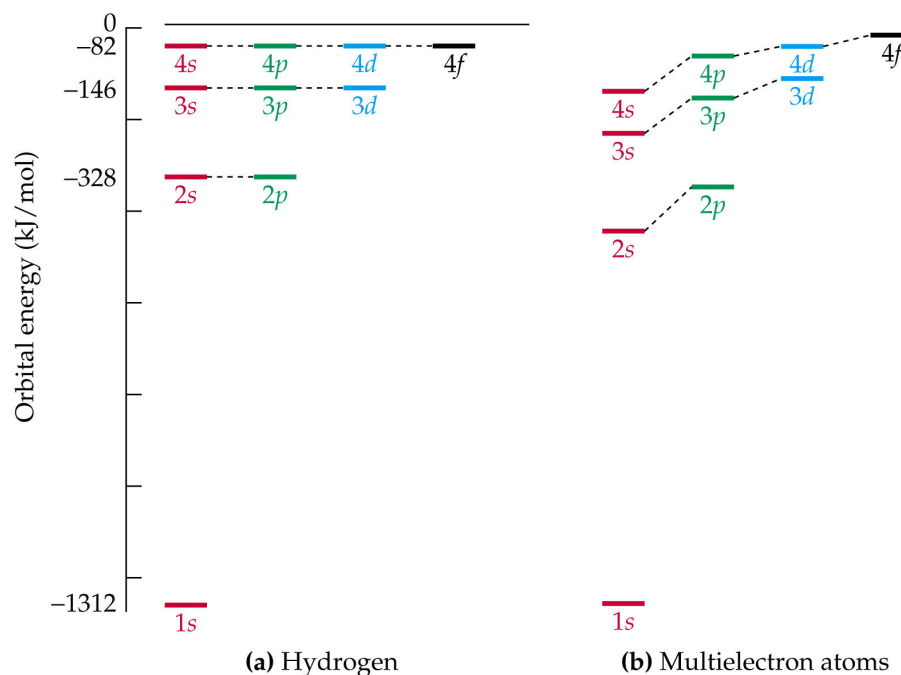
AP Chemistry

Chapter 8 Outline – Electron Configurations, Atomic Properties, and the Periodic Table

Diamagnetism and Paramagnetism

Elements that contain unpaired electrons exhibit a property known as **paramagnetism**. Paramagnetic materials are attracted to external magnets because the magnetic field produced by the "spinning" unpaired electron is not cancelled by the antiparallel opposite spin of another electron. Elements in which all the electrons are paired are known as **diamagnetic** and are not attracted to magnets (most are actually slightly repulsed by magnets). A third type of electron spin orientation is known as **ferromagnetic**. These are substances (including iron, nickel, cobalt, neodymium and certain alloys) that have all their spins aligned in small clusters of atoms within the substance that are known as domains. These are typically the types of materials that are known as permanent magnets.

The energy levels of the hydrogen atom vs a many-electron atom (see diagram below).



Representing the Electrons in an Atom

Spectroscopic Notation:

A short-hand notation has been developed to represent the electrons in a given atom. An example might be $1s^2$, $3p^4$, or $4d^3$, and so on. The co-efficient number refers to the principal quantum number n , the letter corresponds to the azimuthal quantum number ℓ , and the superscript designates how many electrons exist in that particular orbital.

For instance, the electrons of the lithium atom (as represented in **spectroscopic notation**) are:



Lithium has 2 electrons in the s-orbital of the 1st energy level and 1 electron in the s-orbital of its second energy level.

This brings up a point about which orbitals are allowed in a particular energy level. Each energy level can have the following orbital values (which you should be able to see why from the quantum numbers listed earlier)

$n = 1$ s

$n = 2$ s, p

$n = 3$ s, p, d

$n = 4$ s, p, d, f

$n = 5$ s, p, d, f, g

Each new energy level builds one more new orbital from the previous one.

To complicate matters even further, in a many electron atom, because of the interaction between the electrons (the **shielding effect**) and the nucleus (**effective nuclear charge Z^***), the energy levels overlap one another in such a way that the increase in energy for those atoms are:

1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p

Note that even though the 3d sublevel is higher in energy than the 4s, when writing spectroscopic notation, energy levels are grouped together by n number, not by the increase in energy. As an example of this Arsenic (As) would be written as $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^3$ instead of $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^3$

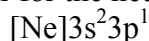
Condensed Spectroscopic Notation

Another way to represent the electron configuration of an atom is called **condensed spectroscopic notation**. Condensed spectroscopic notation is a shorthand method for writing the spectroscopic notation (i.e. $1s^2 2s^2 \dots$) for the electrons in an atom. The rules are:

- Identify the noble gas that occurs prior to the element you are writing the notation for.
- Place that noble gas in brackets (e.g. [Ar])
- Proceed from that point on using normal methods for writing spectroscopic notation.

Ex. Write the condensed spectroscopic notation for Aluminum.

Solution: The last occurring noble gas before aluminum is neon, so Ne is placed in brackets, then you continue on up to aluminum. Remember that aluminum has a total of 13 electrons (equal to its atomic number for the neutral atom).



On Your Own:

Try writing the condensed spectroscopic notations for oxygen and bromine.

Orbital Box Diagram

Yet another way of representing electron configuration is known as the **orbital box diagram**. The advantage of this method is that one can see the actual physical pairing of the electrons. In the orbital box diagram, a box is drawn for each energy level and orbital orientation, then electrons are placed in the box using an up arrow to represent one spin orientation and a down arrow to represent the other spin orientation. By **Hund's Rule**, you must ensure that electrons in an orbital have each of their orientations filled by electrons individually before they are paired (this reduces electron repulsion and is energetically favored). Also, by the **Aufbau Principle** electrons are always added to the lowest energy level first (like jelly beans in a jar would be filled from the bottom up).

To write the orbital box diagram for an element use the following rules:

- a. One box is drawn for each orbital orientation (m_l) in a sublevel

s = 1 box

p = 3 boxes

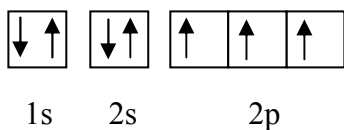
d = 5 boxes

f = 7 boxes

b. Each box can contain a maximum of two electrons. One with an up arrow spin and one with a down arrow spin.

c. Fill sublevel boxes with single electrons before doubling them up and keep the spin orientation the same in each singly filled box.

As an example the element nitrogen would be represented as:

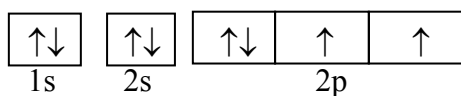


Note that the $2p_x$, $2p_y$ and $2p_z$ sub-orbitals are filled singly before pairing as per Hund's Rule.

Filled energy levels are very stable and are indicative of noble gases (note the filled s and p orbitals).

Half-filled orbitals are also fairly stable and sometimes an atom will migrate an electron to a higher energy level to achieve a half or fully-filled orbital. For instance chromium migrates a 4s electron up to the 3d to half-fill its 3d level and copper migrates a 4s electron up to the 3d to completely fill its 3d orbital. This is the exception and not the rule though.

Ex. Write the orbital box diagram for oxygen.



Note that instead of 2 pairs of 2 electrons in the p boxes there are 2, 1, 1, in accordance with rule c.

On Your Own:

Try writing electron box diagrams for phosphorus (P) and vanadium (V)

Ionization Energy:

By definition, ionization energy is the energy required to remove an electron from an isolated atom. Recall from earlier chapters that metals like to lose electrons to become like the noble gases that came before them (cations). Nonmetals like to gain electrons to become like the noble gases that occur after them (anions). Since metals like to lose electrons anyway, it doesn't take much energy input to remove an electron. In addition, elements with valence electrons (s and p electrons in the highest energy level that are responsible for most bonding) that are far away from the nucleus are not held as strongly by the protons as are those with valence electrons that are close in. Nonmetals like to hold onto their electrons so their ionization energy is higher than that of metals. Also, those with valence electrons close to the nucleus are held particularly strongly, and have higher ionization energy than those that don't. Therefore the trend on the periodic table for ionization energy is that ***ionization energy tends to increase as you go right and up on the periodic table.***

Ionization Energy - Energy required to remove an outer electron from an atom in its gaseous state.

Increases bottom to top: Due to outermost (valence) electrons being closer to the nucleus.

Increases left to right: Due to the increase in nuclear attraction as protons are added.

Ex. Rank the following elements by increasing ionization energy: Al, Ca, F.

Solution: Using the rule that ionization energy increase to the right and up, the answer would be

Ca<Al<F

On Your Own: Rank the following elements by increasing ionization energy:

a. Sn, Sr, S

b. Cl, Br, I

Atomic Radii:

The trend for the sizes of the atoms on the periodic table is influenced by two factors:

a. As you work your way across a row (period) on the periodic table you are filling up electrons within the same principal energy level, which means that the electrons are all being added to approximately the same distance away from the nucleus. As this is occurring, each successive element is also adding more positively charged protons to the nucleus, thus increasing the force on each electron. The combined effect causes the electrons to be pulled closer to the nucleus as you move across the row and the radius gets smaller.

b. As you work your way down a column (group or family), the principal energy level is increasing for each successive element. This means that the average distance from the nucleus is also increasing. This is similar to the different layers of an onion. Also, there are more core electrons (electrons between the valence electrons and the nucleus) causing the repulsion on the outer electrons to be greater and the pull from the nucleus to be weaker. These combined effects cause the radius to increase as you move down the group.

In summary, although there are some exceptions, the general rule for atomic radii is that:

The atomic radii of the elements decrease as you go right and up on the periodic table.

Generally speaking, a jump up or down on the periodic table has a larger effect on the atomic radii than a jump to the left or right.

Ex. Rank the following elements by increasing atomic radii; Rb, N, Ca

Solution: Based on the trends listed above the answer would be N<Ca<Rb

On Your Own: Rank the following elements by increasing atomic radii:

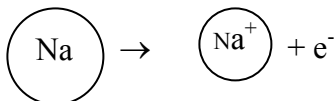
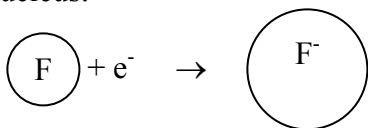
a. Al, S, K

b. Ba, Ca, Be

Ionic Radii - the radius of an ion compared to its neutral counterpart.

Anions are always larger: Due to the added repulsion among electrons

Cations are always smaller: Due to fewer electrons attracted by the same number of protons in the nucleus.



Electronegativity - The relative tendency for an atom in a bond to pull electrons towards it.

Electronegativity is a dimensionless number with the highest two being fluorine at 4.0 and oxygen at 3.5.

Differences in electronegativity ΔEN greater than about 1.7 are considered to be ionic in their bond character. A difference of zero is nonpolar covalent, and between these two numbers, the bond is polar covalent.

Increases bottom to top: Due to bonding (valence) electrons being closer to the nucleus.

Increases left to right: Due to the increase in nuclear attraction as protons are added.

(This is related to nonmetals wanting to gain electrons and metals wanting to lose electrons).

Questions:

1. Write the spectroscopic, condensed spectroscopic and orbital box representations for the elements sulfur, bromine and neon.

2. What is wrong (if anything) with the following sets of quantum numbers?

$$1, 1, 0, +\frac{1}{2}$$

$$3, 2, 1, -\frac{1}{2}$$

$$1, 0, 1, +\frac{1}{2}$$

3. Determine whether aluminum, neon and magnesium are predicted to be paramagnetic or diamagnetic.

4. Draw the shape of a 1s orbital and a $2p_x$ sub-orbital

Electron Affinity:

By definition, electron affinity (EA) is the energy involved in the process of an atom (in its gas phase) gaining an electron. Values range from 0 (for atoms that do not form stable anions in their gas phase) to large negative numbers (indicating a high “affinity” or attraction for electrons). The larger the negative number, the more exothermic the process is.

We would expect (as is the case), that elements that readily form anions (such as the halogens) will have large, negative electron affinities.

The general form of the process is:



Whereas electron affinity may be expressed in joules per atom, the more common form is in units of kJ/mol.

Based on the definitions of EA and IE, elements with high ionization energies generally have high electron affinities.

The general trend is that *electron affinity increases to the right and up on the periodic table* (i.e. the values become more negative.), however it should be noted that there are *many* exceptions to this trend.