Atoms and electrons: Electronic Structure: Student Review Notes

Electronic Configuration

Aufbau principle: This comes from the German word "aufbauen" which means "to build up." So, the aufbau principle is quite literally, the building up principle and it states that as electrons are added around a nucleus, they occupy the lowest energy state available.

Notation:

Energy levels, n, go from 1 to 7 (like the periods of the periodic table),

Sublevels have the names s, p, d, and f

Orbitals can each hold 2 electrons.
- s-sublevel has 1 orbital (2 electrons)
- p-sublevel has 3 orbitals (6 electrons)
- d-sublevel has 5 orbitals (10 electrons)
- f-sublevel has 7 orbitals (14 electrons)

You need to know how to **apply the aufbau principle to generate the electronic configuration of an atom or ion**. The electronic configuration shows the total number of electrons, the electrons on each energy level and within each sublevel. Writing these out is very easy if you understand that all you need to do is walk across the periodic table and add electrons.

Take a look, electrons are added to s and p orbitals when you go over the first two and last six groups. Here the energy level, n, goes as the period. Electrons go into d-orbitals when you are over the transition metals and the energy level goes as n-1. Finally, you put electrons in f-orbitals for the lanthanides and actinides and here the energy level here goes as n-2. So, for example, **walk across the periodic** table and generate the electronic configuration for bromine, element 35.

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$
(35 electrons)

s (n) — d (n-1) — p (n)

n																		
1	$1s^1$																	$1s^2$
2	$2s^1$	$2s^2$											$2p^1$	$2p^2$	$2p^3$	$2p^4$	$2p^5$	$2p^6$
3	$3s^1$	$3s^2$											$3p^1$	$3p^2$	$3p^3$	$3p^4$	$3p^5$	$3p^6$
4	$4s^1$	$4s^2$	$3d^1$	$3d^2$	$3d^3$	$4s^1$ $3d^5$	$3d^5$	$3d^6$	$3d^7$	$3d^8$	$4s^1$ $3d^{10}$	$3d^{10}$	$4p^1$	$4p^2$	$4p^3$	$4p^4$	$4p^5$	$4p^6$
5	$5s^1$	$5s^2$	$4d^1$	$4d^2$	$4d^3$	$5s^1$ $4d^5$	$4d^5$	$4d^6$	$4d^7$	$4d^8$	$5s^1$ $4d^{10}$	$4d^{10}$	$5p^1$	$5p^2$	$5p^3$	$5p^4$	$5p^5$	$5p^6$
6	$6s^1$	$6s^2$	57 $5d^1$	$5d^2$	$5d^3$	$6s^1$ $5d^5$	$5d^5$	$5d^6$	$5d^7$	$5d^8$	$6s^1$ $5d^{10}$	$5d^{10}$	$6p^1$	$6p^2$	$6p^3$	$6p^4$	$6p^5$	$6p^6$
7	$7s^1$	$7s^2$	89 $6d^1$	$6d^2$	$6d^3$	$7s^1$ $6d^5$	$6d^5$		$6d^7$									

f (n-2)

lanthanides	58 $5d^1$ $4f^1$	$5d^0$ $4f^3$	$5d^0$ $4f^4$	$5d^0$ $4f^5$	$5d^0$ $4f^6$	$5d^0$ $4f^7$	$5d^1$ $4f^7$	$5d^0$ $4f^9$	$5d^0$ $4f^{10}$	$5d^0$ $4f^{11}$	$5d^0$ $4f^{12}$	$5d^0$ $4f^{13}$	$5d^0$ $4f^{14}$	71 $5d^1$ $4f^{14}$
actinides	90 $6d^2$ $5f^0$	$6d^1$ $5f^2$	$6d^1$ $5f^3$	$6d^1$ $5f^4$	$6d^0$ $5f^6$	$6d^0$ $5f^7$	$6d^1$ $5f^7$	$6d^0$ $5f^9$	$6d^0$ $5f^{10}$	$6d^0$ $5f^{11}$	$6d^0$ $5f^{12}$	$6d^0$ $5f^{13}$	$6d^0$ $5f^{14}$	$6d^1$ $5f^{14}$

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Long Notation: Long notation shows all the electrons. For example, germanium has 32 electrons and the electronic structure would be written as: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^2$

Short-hand Notation: Short-hand notation for the electronic structure of an atom or ion uses the previous noble gas to denote all electrons up to that point. So, if we stick with germanium as the example, the short-hand notation would be:$[Ar]4s^2 3d^{10} 4p^2$

Ions: Atoms ionize by gaining or losing electrons. Remember that you can only change the number electrons to affect the charge of an atom. If you change the number of protons, you haven’t formed an ion, you’ve changed the element. Cations are positively charged--they’ve lost electrons. Anions are negatively charged--they’ve gained electrons. Simply walk across the periodic table until you get to the total number of electrons for the ion. For example, Ca^{2+}, a cation, has lost 2 electrons. The electronic configuration of Ca^{2+} is: $1s^2 2s^2 2p^6 3s^2 3p^6$. Notice that Ca^{2+}has the same electronic configuration as argon. In such a case, the two species are said to be **isoelectronic**. **Very important**: **Transition metals ionize by losing their s-electrons first**. This is something of a contradiction since the d-electrons go on after the s-electrons but remember that the s-electrons come off first. For example, the electronic configuration of $Fe^{2+}$is:$[Ar]4s^0 3d^6$

Exceptions to the aufbau principle: These occur because of the stability of a half-filled sublevel and the stability of a completely filled sublevel. Electrons will fill out of the expected order so that either of these configurations may be achieved. This is a pretty good rule of thumb for the transition metals. It holds in a general way for the lanthanides and actinides but not completely.

an electron jumps out of the s-sublevel and into the d-sublevel to get a half filled sublevel (Cr, Mo, W)

an electron jumps out of the s-sublevel and into the d-sublevel to get a completely filled sublevel (Cu, Ag, Au)

n	n		n-1										n					
1	$1s^1$																	$1s^2$
2	$2s^1$	$2s^2$											$2p^1$	$2p^2$	$2p^3$	$2p^4$	$2p^5$	$2p^6$
3	$3s^1$	$3s^2$											$3p^1$	$3p^2$	$3p^3$	$3p^4$	$3p^5$	$3p^6$
4	$4s^1$	$4s^2$	$3d^1$	$3d^2$	$3d^3$	$4s^1$ $3d^5$	$3d^5$	$3d^6$	$3d^7$	$3d^8$	$4s^1$ $3d^{10}$	$3d^{10}$	$4p^1$	$4p^2$	$4p^3$	$4p^4$	$4p^5$	$4p^6$
5	$5s^1$	$5s^2$	$4d^1$	$4d^2$	$4d^3$	$5s^1$ $4d^5$	$4d^5$	$4d^6$	$4d^7$	$4d^8$	$5s^1$ $4d^{10}$	$4d^{10}$	$5p^1$	$5p^2$	$5p^3$	$5p^4$	$5p^5$	$5p^6$
6	$6s^1$	$6s^2$	57 $5d^1$	$5d^2$	$5d^3$	$6s^1$ $5d^5$	$5d^5$	$5d^6$	$5d^7$	$5d^8$	$6s^1$ $5d^{10}$	$5d^{10}$	$6p^1$	$6p^2$	$6p^3$	$6p^4$	$6p^5$	$6p^6$
7	$7s^1$	$7s^2$	89 $6d^1$	$6d^2$	$6d^3$	$7s^1$ $6d^5$	$6d^5$		$6d^7$									

n-2

lanthanides	58 $5d^1$ $4f^1$	$5d^0$ $4f^3$	$5d^0$ $4f^4$	$5d^0$ $4f^5$	$5d^0$ $4f^6$	$5d^0$ $4f^7$	$5d^1$ $4f^7$	$5d^0$ $4f^9$	$5d^0$ $4f^{10}$	$5d^0$ $4f^{11}$	$5d^0$ $4f^{12}$	$5d^0$ $4f^{13}$	$5d^0$ $4f^{14}$	71 $5d^1$ $4f^{14}$
actinides	90 $6d^2$ $5f^0$	$6d^1$ $5f^2$	$6d^1$ $5f^3$	$6d^1$ $5f^4$	$6d^0$ $5f^6$	$6d^0$ $5f^7$	$6d^1$ $5f^7$	$6d^0$ $5f^9$	$6d^0$ $5f^{10}$	$6d^0$ $5f^{11}$	$6d^0$ $5f^{12}$	$6d^0$ $5f^{13}$	$6d^0$ $5f^{14}$	$6d^1$ $5f^{14}$

Orbital Diagrams

Orbital diagrams show the **placement of electrons in individual orbitals including the spin of the individual electrons.** Electrons still fill up according to the aufbau principle—that tells you the progression of energy levels and sublevels. An **orbital is denoted by a box or a line** and electrons are up (spin +1/2) or down (**spin -1/2**) arrows.

Hund's Rule: The rule of maximum spin multiplicity. Hund's rule basically says that electrons don't like each other (like charges repel) and you use it when you fill electrons in an orbital diagram. Within a given sublevel, the order of filling of the orbitals is such that that there is a maximum number of half filled orbitals.

Element	Electronic Configuration	Orbital Diagram 1s	2s	2p		
Li	$1s^22s^1$	↑↓	↑			
Be	$1s^22s^2$	↑↓	↑↓			
B	$1s^22s^22p^1$	↑↓	↑↓	↑		
C	$1s^22s^22p^2$	↑↓	↑↓	↑	↑	
N	$1s^22s^22p^3$	↑↓	↑↓	↑	↑	↑
O	$1s^22s^22p^4$	↑↓	↑↓	↑↓	↑	↑
F	$1s^22s^22p^5$	↑↓	↑↓	↑↓	↑↓	↑
Ne	$1s^22s^22p^6$	↑↓	↑↓	↑↓	↑↓	↑↓

An atom with **unpaired electrons is called paramagnetic and is attracted in a magnetic field**.
From the above table, Li, B, C, N, O, and F are paramagnetic

An atom which has **all its electrons paired is called diamagnetic and it is not attracted in a magnetic field**.
From the above table, Be and Ne are diamagnetic

What can an electronic configuration tell us about the chemistry of elements?

Elements in the same group will have similar electronic configurations and orbital diagrams for their valence electrons (valence electrons are the electrons on the highest energy level). **Elements with the same number of valence electrons have similar chemical properties**.

You can also predict some chemistry: How would you expect As and Na to bond?
A rule of thumb in chemistry is that **atoms want to obtain a full valence shell (octet) when they bond**.

The electrons on Arsenic: $[Ar]4s^24p^3$ valence shell orbital diagram: 4s [↑↓] 4p [↑][↑][↑]

The electrons on Sodium: $[Ne]3s^1$ valence shell orbital diagram: 3s [↑]

From this you see that arsenic wants to gain three electrons to complete its outer octet. Sodium on the other hand, wants to lose 1 electron to achieve a complete outer octet. It's therefore reasonable to assume, based on the electronic configuration of each element, that the compound $AsNa_3$ would be a stable molecule.

Because the valence shell is similar for all the elements in a group, you would therefore also expect the Group IA elements to which sodium belongs to form similar compounds with the Group 5A elements to which arsenic belongs.