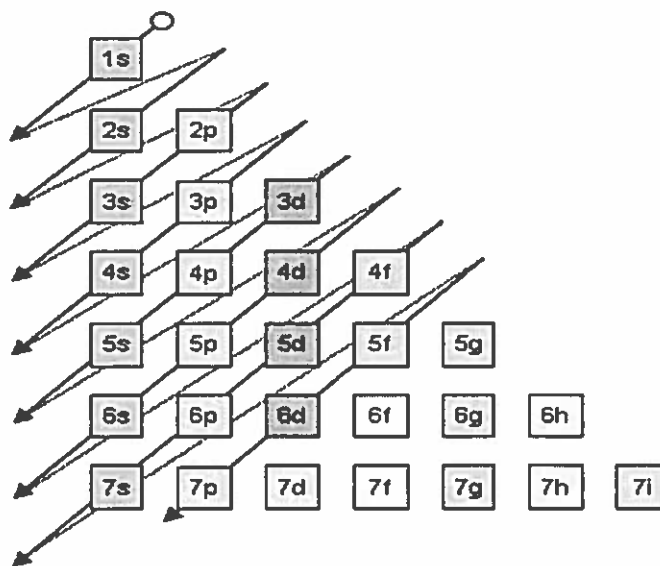


energy: the "Aufbau principle". The periodic table follows the Aufbau principle very closely - if you know where the s, p, d, and f blocks are, and know that the Aufbau principle makes you read the periodic table from left to right, then writing electron configurations is easier than you think.

Periodic Table of the Elements
Orbital Shell Blocks

1A	Orbital Shell Blocks																8A			
1s	2s		s p d f										3s 3p			4s				
← 2s →	← 2p →												← 2p →			← 2s →				
← 3s →	← 3p →												← 3p →			← 3s →				
← 4s →	← 4p →		← 3d →										← 4p →			← 4s →				
← 5s →	← 5p →		← 4d →										← 5p →			← 5s →				
← 6s →	← 6p →		← 5d →										← 6p →			← 6s →				
← 7s →	← 7p →		← 6d →										← 7p →			← 7s →				
																	← 4f →			
																	← 5f →			

Alternatively, many students like to use the following "road map"⁸ to write configurations using the Aufbau principle. Read the map starting at the circle, and follow the arrow.



Remember that an atomic orbital can only hold two electrons, and you're all set. You can also use an electron configuration to identify an atom, or identify the valence energy level and

⁷ Helmenstine, Anne Marie. "What Are Element Blocks?" *About.com Education*. N.p., 29 Nov. 2014. Web. 18 Dec. 2016.

⁸ Bigler, Jeff. "Electron Configurations." *Electron Configurations*. N.p., n.d. Web. 18 Dec. 2016.

number of valence electrons in that atom. Also, we operate under the assumption that all the lowest possible energy levels are filled up to the proper number of electrons, hence the common term "ground state electron configuration". An electron configuration where an electron or two seem to lie in higher than normal orbitals is an "excited state electron configuration". Ions will have electrons added to or removed from the highest orbital.

Examples:

1. Write the ground state electron configuration for an atom of each of the following elements:

- 1 e⁻ a) Hydrogen $1s^1$
 7 e⁻ b) Nitrogen $1s^2 2s^2 2p^3$
 22 e⁻ c) Titanium $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$
 78 e⁻ d) Platinum $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^8$
 107 e⁻ e) Bohrium $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2 5f^{14} 6d^5$

2. Which element is denoted by each ground state electron configuration? What is its valence energy level and how many valence electrons are there?

	element	valence energy level	# valence e ⁻
a) $1s^2 2s^2 2p^6 3s^2$	Mg	3	2
b) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$ $2+4=6$	Se	4	6
c) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^1$	Cs	6	1

3. Isoelectronic species have similar electron configurations. Which of these are isoelectronic?

- a) Li^+ , H^- , He $Li^+ = 1s^2$ $H^- = 1s^2$ $He = 1s^2$
 so Li^+ , H^- , He are all isoelectronic

- b) Ca^{2+} , Ne, S^{2-} $Ca^{2+} = 1s^2 2s^2 2p^6 3s^2 3p^6$ Ca^{2+} and S^{2-} are isoelectronic
 $Ne = 1s^2 2s^2 2p^6$
 $S^{2-} = 1s^2 2s^2 2p^6 3s^2 3p^6$

4. Which of these are excited state electron configurations? In all cases, identify the atom.

- | | atom |
|----------------------------------------------------------------------|------|
| a) $1s^2 3p^1$ * excited state should be: $1s^2 2s^1$ | Li |
| b) $1s^2 2s^2 2p^3 3s^1$ * excited state should be: $1s^2 2s^2 2p^4$ | O |
| c) $1s^2 2s^2 2p^6 3s^2 3p^1$ ground state | Al |

For elements nearer the bottom of the periodic table, it is very tedious to write out the full, unabbreviated electron configuration. There is a shortcut: find the symbol of the noble gas coming immediately before your element, and write it in square parentheses, i.e. [Ne] for Al and [Kr] for Mo. Start at that noble gas and follow the Aufbau principle as normal. This type of configuration is called a noble gas core electron configuration, or sometimes just called the "shortcut" or "abbreviated" configuration."

Examples:

1. Write the noble gas core ground state electron configuration for:

- a) Manganese $[\text{Ar}] 4s^2 3d^5$
 b) Palladium $[\text{Kr}] 5s^2 4d^8$
 c) Tantalum $[\text{Xe}] 6s^2 4f^{14} 5d^3$
 d) Meitnerium $[\text{Rn}] 7s^2 5f^{14} 6d^7$
 e) Radon (no, it isn't [Rn]) $[\text{Xe}] 6s^2 4f^{14} 5d^{10} 6p^6$

2. Identify the element.

- a) $[\text{Kr}] 5s^2 4d^{10} 5p^6$ Xe
 b) $[\text{Ne}] 3s^2 3p^1$ Al
 c) $[\text{Ar}] 4s^1$ ~~Ca~~ K
 d) $[\text{Xe}] 6s^2 4f^{14} 5d^7$ Ir

Electron Configuration: Exceptions to Aufbau's Principle

In several cases, the correct electron configuration looks like it represents an excited state configuration when it actually doesn't. The noble gas core electron configuration for lanthanum is $[\text{Xe}] 6s^2 5d^1$. For any element in the "lanthanide series" beginning with Ce (so, elements 58 - 71), electrons fill the 4f block in order after that *and include that 5d¹*. For example, the noble gas core electron configuration for cerium is $[\text{Xe}] 6s^2 4f^2 5d^1$, and for dysprosium, $[\text{Xe}] 6s^2 4f^9 5d^1$.

The "actinide" series (elements 90-103) works in exactly the same way, except that they include 6d¹: actinium is $[\text{Rn}] 7s^2 6d^1$, curium is $[\text{Rn}] 7s^2 5f^7 6d^1$, and fermium is $[\text{Rn}] 7s^2 5f^{10} 6d^1$.

The electron configurations of chromium and copper do not exactly follow the Aufbau principle either, largely because there is no other way to try to explain their observed charges. Chromium has the electron configuration $[\text{Ar}] 3d^5 4s^1$, and copper is $[\text{Ar}] 3d^{10} 4s^1$. This is because a half filled or a completely filled inner energy level is more stable; also, chromium and copper do have observable charges of +1, which is understandable if there is a 4s electron that is removable. This exception also applies to any elements in the chromium or copper group.

Examples:

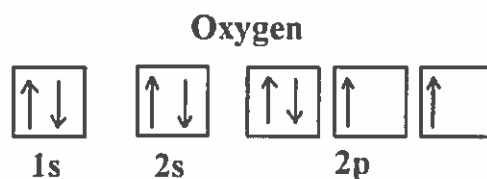
1. Write the noble gas core electron configuration for each of the following elements:

- a) Molybdenum $[\text{Kr}] 5s^1 4d^5$
- b) Gadolinium $[\text{Xe}] 6s^2 4f^7 5d^1$
- c) Silver $[\text{Kr}] 5s^1 4d^{10}$
- d) Americium $[\text{Rn}] 7s^2 5f^6 6d^1$
- e) Gold $[\text{Xe}] 6s^1 4f^{14} 5d^{10}$
- f) Seaborgium $[\text{Rn}] 7s^2 4f^{14} 5d^4$

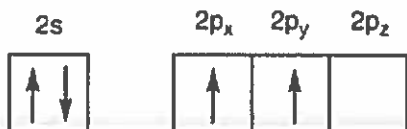
Orbital Notation

Another way of representing an electron configuration is by drawing the orbital notation. The orbital notation uses arrows pointing up or down to stand for electrons; the up or down directions denote spin ($+\frac{1}{2}$ for up, $-\frac{1}{2}$ for down), and either a short horizontal line, a circle, or a box to represent an atomic orbital. Only two arrows, or electrons, can be in a given orbital and must have opposite spin (remember, Pauli exclusion principle). Moreover, if electrons are entering, one at a time, into orbitals of equivalent energy (i.e. into the three orbitals of the 2p sublevel), then they enter individually into each of these orbitals until they are forced to pair up. This is called "Hund's rule." Orbitals of the same energy level and sublevel are called "degenerate orbitals."

For example, the orbital notation for oxygen ($1s^2 2s^2 2p^4$) is⁹:



In oxygen, there are three boxes for the 2p sublevel, since each box stands for each allowable value of m_l (-1, 0, or 1), whereas there is only one box for each of 1s and 2s because you can have only one orbital in each of those sublevels. You can draw an orbital notation using the noble gas core as well.



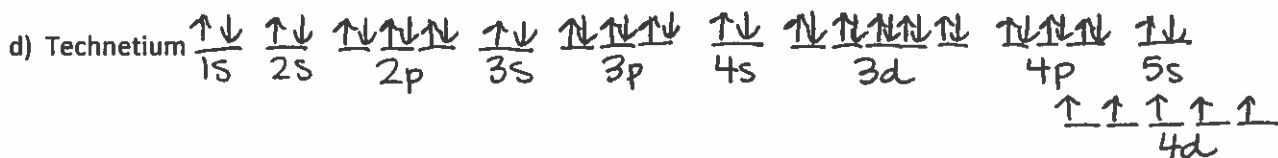
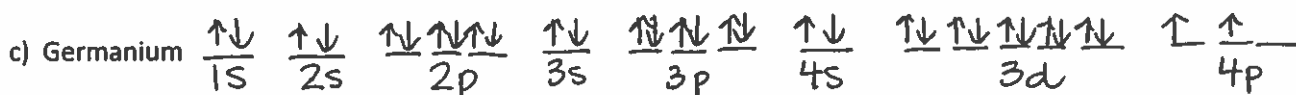
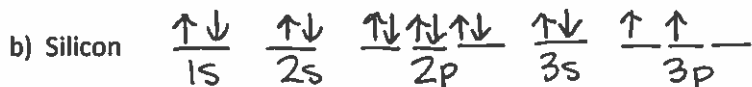
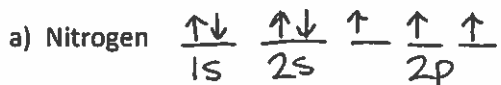
As with electron configurations, it's possible to write orbital notation showing only the valence orbitals and electrons. The image on the left is the valence orbital notation for carbon.¹⁰

⁹ "Hund's Rules." *Chemistry LibreTexts*. N.p., 21 July 2016. Web. 18 Dec. 2016.

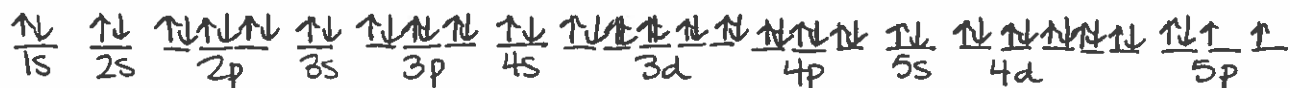
¹⁰ "Hybridization." *OChemPal*. N.p., n.d. Web. 19 Dec. 2016.

Examples:

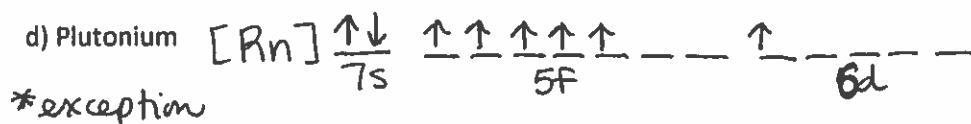
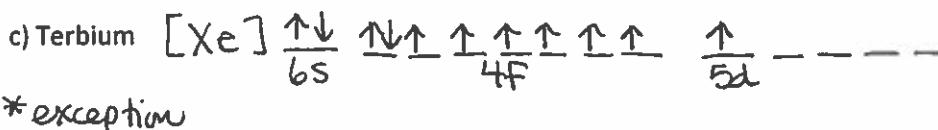
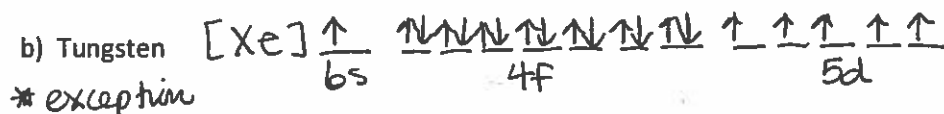
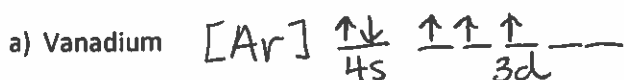
1. Draw the unabbreviated orbital notation for each of the following atoms.



e) Tellurium



2. Draw the noble gas core orbital notation for each of the following atoms.



4.3 Practice and Homework

1. Are each of the following sets of quantum numbers allowable or not allowable? For the ones that are not allowable, state the mistake(s).

- a) $n=4, l=2, m_l=0, m_s=-\frac{1}{2}$ allowed
 b) $n=2, l=-1, m_l=1, m_s=+\frac{1}{2}$ not allowed; l cannot be negative so m_l is incorrect too
 c) $n=3, l=0, m_l=-2, m_s=-\frac{1}{2}$ not allowed; $m_l=0$ if $l=0$
 d) $n=5, l=-3, m_l=4, m_s=+\frac{1}{2}$ not allowed; l cannot be negative so m_l is incorrect too
 e) $n=4, l=2, m_l=-1, m_s=-1$ not allowed; m_s can only be $+\frac{1}{2}$ or $-\frac{1}{2}$

2. Write the unabbreviated electron configurations for the following elements:

- a) Sodium $1s^2 2s^2 2p^6 3s^1$
 b) Iron $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$
 c) Bromine $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$

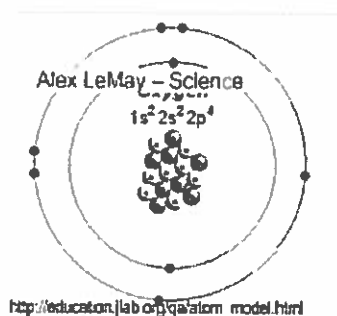
3. Write the noble gas core electron configurations for the following elements:

- a) Cobalt $[Ar] 4s^2 3d^7$
 b) Silver $[Kr] 5s^1 4d^{10}$ (see section on exceptions)
 c) Europium $[Xe] 6s^2 4f^6 5d^1$ (see section on exceptions)

4. Determine what elements are denoted by the following electron configurations, and indicate the valence energy level and number of valence electrons.

- a) $1s^2 2s^2 2p^6 3s^2 3p^4$ Sulfur; $n=3$ is valence energy level; $2+4=6$ valence e^-
 b) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1$ Rubidium; $n=5$ is valence energy level; 1 valence e^-

This means that the net positive charge from the nucleus to the valence electrons is 5.



Now comparing this to oxygen, oxygen has one extra electron and one extra proton, but no new core electrons. The addition of that extra proton in the nucleus is going to increase the net positive charge and thus increase the attraction of the electrons to the nucleus, as can be seen by a Z_{eff} of 6: $Z_{\text{eff}} = 8 - 2 = 6$.

The greater Z_{eff} is the more attraction there is between the electrons and the nucleus implying there is a stronger pull on the electrons from the nucleus.

Xiue (Michelle) Xie, UC Davis

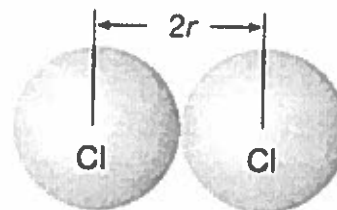
Two elements within the same group are going to have the same Z_{eff} , however the attraction from the nucleus is not the same due to **shielding**. One way to imagine shielding is to think of being at a concert or theater performance. If you are in the front row, you have full access of sight and sound to the artists, however as the number of rows and spectators, energy levels, between you and the stage increases your view and quality of sound decreases. This is shielding. As the number of core electrons increase the net positive charge felt on the valence electrons from the nucleus is going to decrease. Shielding decreases moving down a group (increase in core electrons), but remains constant within a period (core electrons remain the same).



in short, Z_{eff} is used to explain a trend seen within a period and shielding is used to explain a trend within a group.

Atomic Radius

Atomic radius is one-half the distance between the two nuclei in a molecule consisting of two identical atoms. Notice the trend is for atomic size to increase as you go down a column, and for atomic size to decrease as you move across a row.



As one moves across the 2nd period of the table, energy of the valence electrons remains approximately constant. The effective nuclear charge, however, increases. The electrons are being pulled toward the nucleus more strongly on the right side of the