Electron Configuration Worksheet

For the Teacher

http://www.kb-creative.net/

Purpose and Philosophy

The electron configuration worksheet is a general-purpose worksheet for helping high school students understand electronic structure in atoms. With this background they can better understand the structure of the periodic table and the nature of chemical bonds.

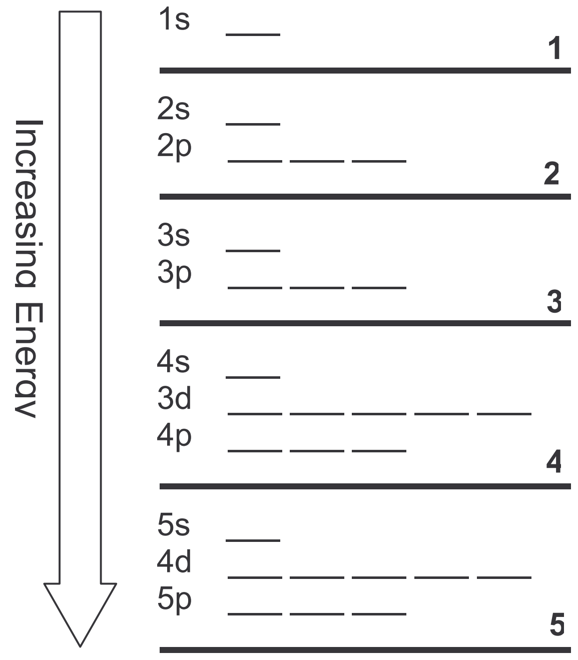
The design of this worksheet tries to meet these principles:

- High school students must understand (and not just memorize) the structure of the periodic table and the nature of chemical bonds.
- High school students should become familiar with modern ideas about the electronic structure in atoms.
- Even though the rules for electronic structure (Hund's rule, the *Aufbau* principle) can seem complicated and can only be understood with advanced theories (quantum mechanics), the basic rules are simple, and can be easily learned.

Basic Usage

The worksheet has space for students to fill in the electron configuration for six atoms or ions. Each space looks like this:

Atom or Ion: ___________________________
Protons: _________ # Neutrons: _________



There is a set of rules for filling in the electron configuration. Here are the instructions, copied from the worksheet:

How to use this worksheet

See the example for Oxygen at the left. This sheet works up to Xenon (atomic number 54).

1. Fill in the name of the atom or ion
2. Fill in the number of protons
3. **If** you are working with an isotope (e.g., Carbon-14), then put the number of neutrons
4. Fill in the electrons:
   a. Indicate electrons with an up arrow (↑) or down arrow (↓) for "spin up" and "spin down" electrons
   b. Add two electrons *maximum* for each underline, one spin up and one spin down (e.g., ↑ or ↑↓ are OK)
   c. Add electrons from lowest energy to highest energy
   d. For underlines that are at the same energy (e.g., 2p), add the "up" electrons first, then "down" electrons

Pauli Exclusion Principle

Aufbau Principle

Hund's Rule

Instructions numbers (1) and (2) are straightforward. For number (3), the number of neutrons is only important if you are focusing on a particular isotope (e.g., Carbon-12 vs. Carbon-14). Otherwise, the number of electrons is determined entirely by the number of protons, since charges must balance in a neutral atom.

The tricky rules are in item number (4). These rules can be understood only with quantum mechanics. Rules (b), (c) and (d) correspond to famous principles, as shown above. In the box below are some suggestions for explaining these rules.

Suggestions for explaining the rules

- For "spin up" and "spin down" (rule **a**) explain that electrons are like tiny magnets, and the magnet can point up or down.
- For rule **b** (two electrons for each underline), explain that no two electrons can be in the same state. So, for example, if you already have a "spin up" electron in the 1s level, then to add another electron it has to be different in some way. The only way it can be different is to be a "spin down" electron. (There are three "flavors" of *p* levels, called p_x, p_y and p_z: the three underlines are different states. There are five flavors of *d* levels.)
- For rule **c** (lowest to highest energy), explain that if an electron is in a high-energy level it will lose energy eventually and drop to the lowest possible energy state. It does this (usually) by emitting light. The lowest-energy state is called the "ground state."
- For rule **d**, explain that because the electrons repel each other, and because of a quantum mechanical effect involving spins, it is easier to fill in all the "up" slots first on the same energy level, and then the other slots. (*Note:* Choosing the up arrows first is arbitrary. I could have said to fill in the down arrows first, but *one* of them has to come first.)

Here's an example for filling in electrons for Oxygen, the example shown on the worksheet. Oxygen has eight electrons (to balance the charge on its eight protons). Here is how to add the electrons, one by one:

1) Add an up electron (rule **d**) to the lowest-energy level (rule **c**). That's the *1s* level:

1s ↑__

2s ___
2p ___ ___ ___

2) Add a down electron (rule **b**) to the *1s* level:

1s ↑↓

2s ___
2p ___ ___ ___

3) Add an up electron (rule **d**) to the lowest-energy level (rule **c**). The *1s* level is full (rule **b**), so add it to the *2s* level:

1s ↑↓

2s ↑__
2p ___ ___ ___

4) Add a down electron (rule **b**) to the *2s* level:

1s ↑↓

2s ↑↓
2p ___ ___ ___

5) Add an up electron (rule **d**) to the lowest-energy level (rule **c**). The *1s* and *2s* levels are full (rule **b**), so add it to the *2p* level:

1s ↑↓

2s ↑↓
2p ↑_ ___ ___

---

6) There are more underlines available at the *2p* level, so add another up electron (rule **d**) to the *2p* level:

1s ↑↓

2s ↑↓
2p ↑_ ↑_ ___

---

7) There is one more underlines available at the *2p* level, so add another up electron (rule **d**) to the *2p* level:

1s ↑↓

2s ↑↓
2p ↑_ ↑_ ↑_

---

8) No more up arrows can be added to the *2p* level (rule **b**), so add a down arrow:

1s ↑↓

2s ↑↓
2p ↑↓ ↑_ ↑_

Finished! There are eight electrons, so that's the electron configuration for Oxygen.

Beyond Basic Usage

The worksheet can be used to understand

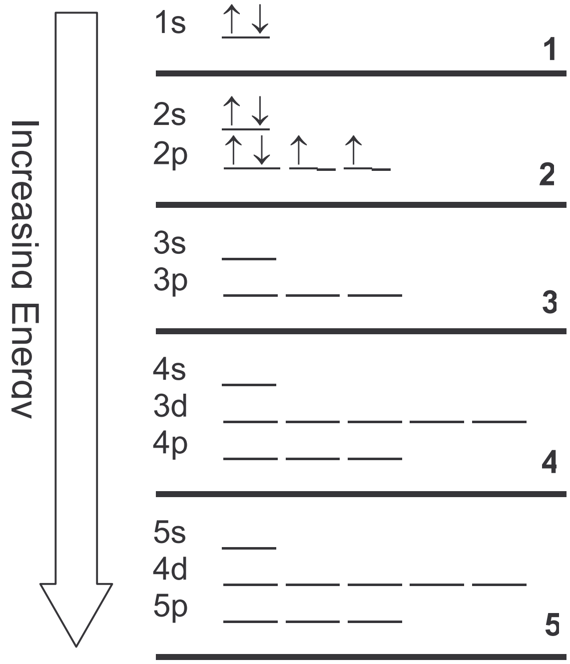
- The periodic table
- The octet "rule" and chemical bonding
- Atomic spectra
- Magnetism

Each of these topics is explained below.

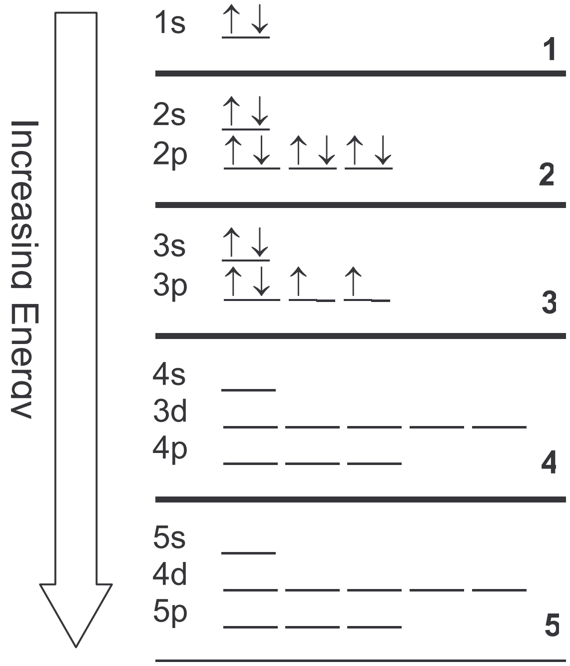
The Periodic Table

Compare the electron configurations for Oxygen and Sulfur:

Atom or Ion: Oxygen
Protons: 8 # Neutrons: ________



Atom or Ion: Sulfur
Protons: 16 # Neutrons: ________



Notice that the *2p* level in Oxygen looks exactly like the *3p* level in Sulfur. More importantly, the combination *2s* + *2p* in Oxygen looks exactly like the *3s* + *3p* level in Sulfur. These are the **valence electrons** in these atoms.

The valence electrons are the electrons in the **valence shell**: the highest-energy occupied shell. The **shells** are indicated by the bold lines. So in Oxygen, the valence shell is all the electrons between the bold lines labeled by a bold **2**.

The valence electrons are the electrons that are involved in chemical bonds, so by comparing the configurations for Oxygen and Sulfur, you can deduce that Oxygen and

Sulfur should have a lot of similarities in their chemical properties. Sure enough, that is true, and it is why Oxygen and Sulfur are in the same column in the periodic table.

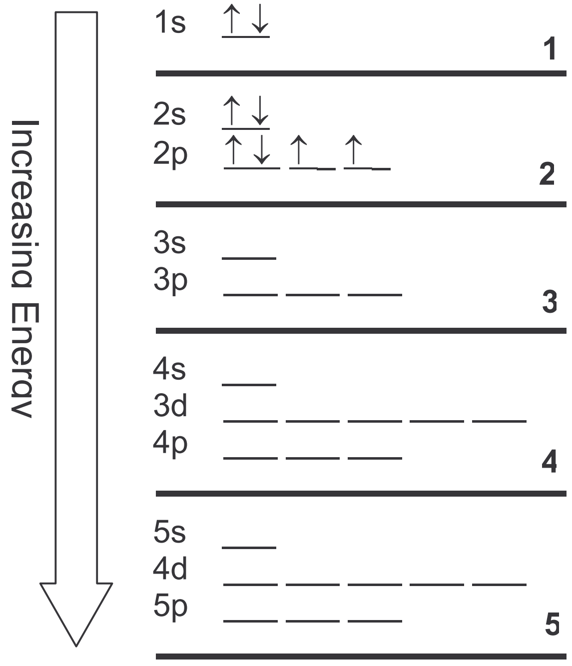
Using this approach, students can figure out where different atoms ought to be in the periodic table. The only tricky thing is the *3d* and *4d* levels. The transition metals have electrons in these levels, and it is what gives the metals their unique properties. Once the *d* level is filled, such as in Se and Br, the atoms act a lot like the atoms above and below them in the periodic table. So, even though Se and Br have electrons in the *3d* level, they act a lot like S and Cl.

The Octet "Rule" and Chemical Bonding

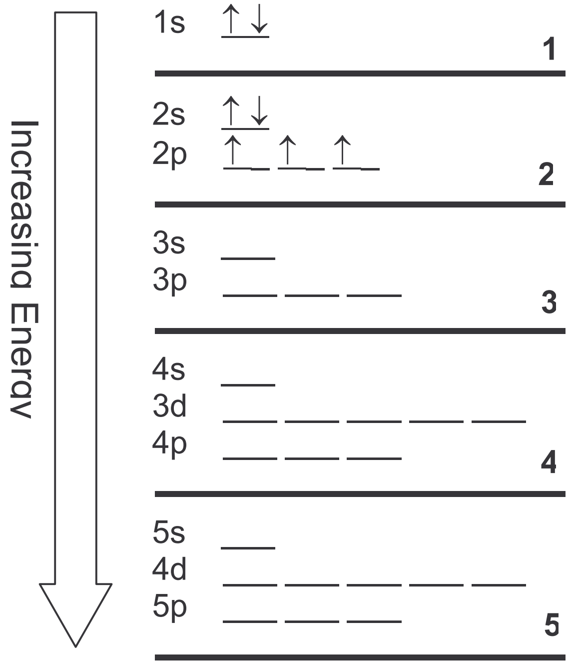
The chemistry of the atoms on the left and right sides of the periodic table are often explained in terms of "octets" of electrons. Atoms try to fill an "octet" of valence electrons, and this explains chemical bonding. This is the octet "rule." I put the word "rule" in quotation marks because it doesn't apply in all cases. It breaks down for metals and for Hydrogen. Still, it's a *very* useful rule, even if it's just a "rule."

The octet rule can be understood by looking at the electron configurations of atoms. Look at Oxygen and Nitrogen:

Atom or Ion: Oxygen
Protons: 8 # Neutrons: ________



Atom or Ion: Nitrogen
Protons: 7 # Neutrons: ________



How many valence electrons do these atoms have? Counting up the electrons between the bold lines labeled **2**, there are six for Oxygen and five for Nitrogen. How many do they need to have full valence shells? In each case, they need eight – they need an **octet** of electrons to have full valence shells.

The valence electrons can be indicated using **Lewis Dot Diagram**. In a Lewis dot diagram, valence electrons are indicated by a dot:

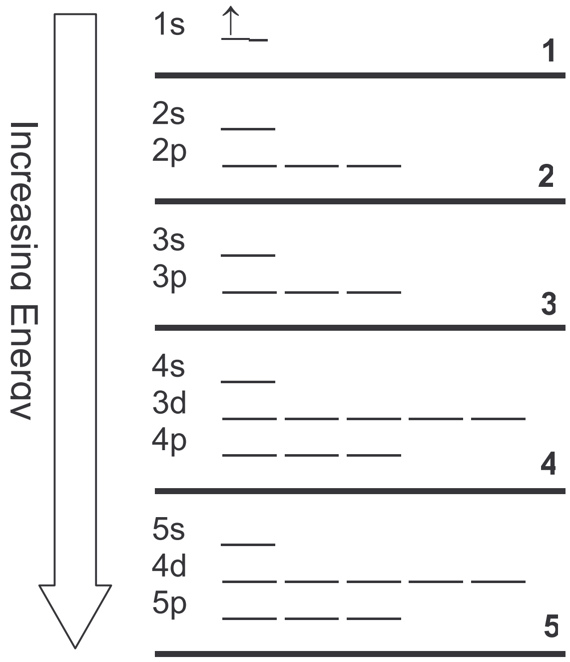
.　　　.
: O :　　: N ·
.　　　.

An atom with a full valence shell has all its dots filled in. This is the case for the **Noble Gases**, which have full valence shells and therefore do not react chemically:

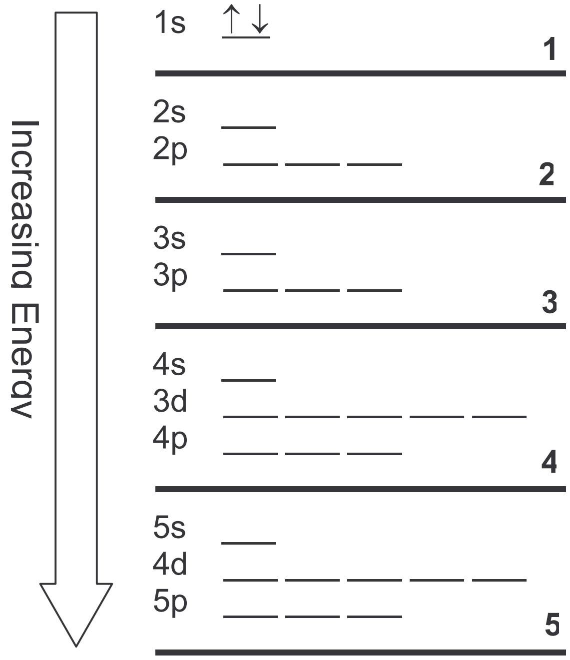
. .　　　. .
: Ne :　　: Ar :
. .　　　. .

An important exception to this rule is the Hydrogen atom. It has one valence electron, but it only needs one more electron to give it a full valence shell. This can be seen clearly from the electron configuration worksheets:

Atom or Ion: Hydrogen
\# Protons: 1 # Neutrons: ________



Atom or Ion: Helium
\# Protons: 2 # Neutrons: ________



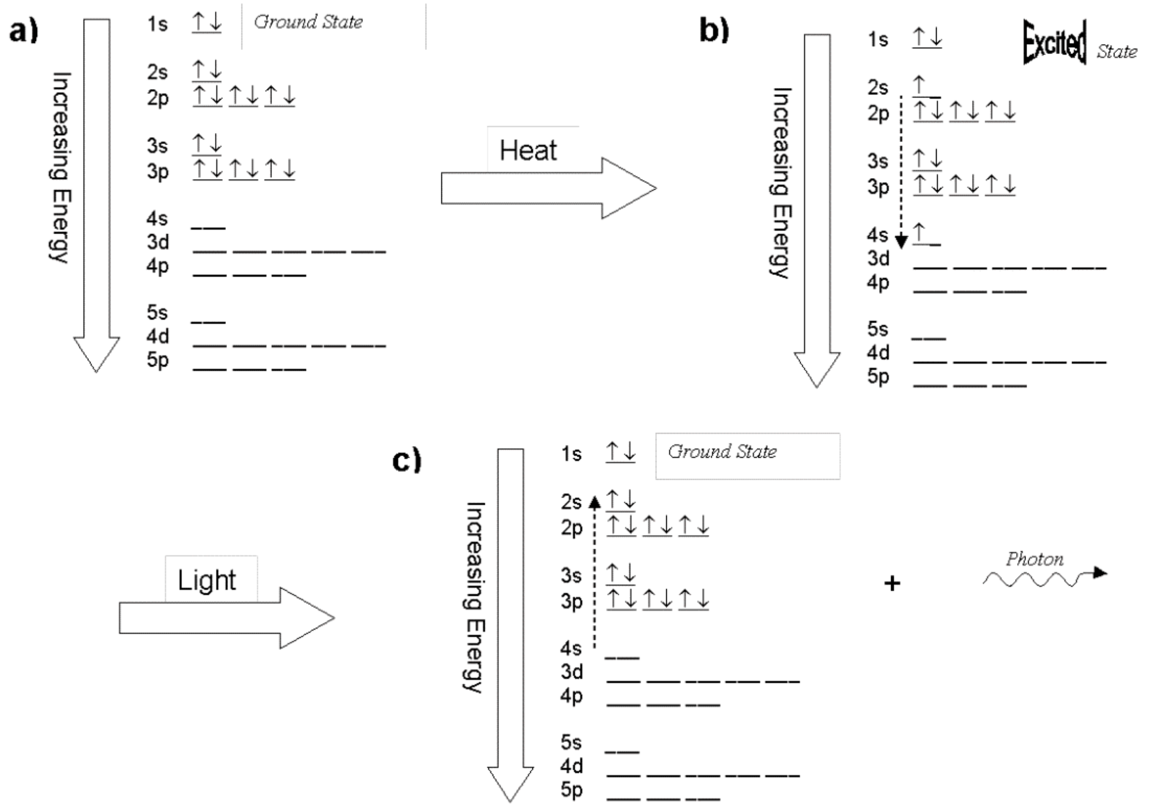
Sure enough, Helium has a full shell, and should be placed with the Noble Gases. Hydrogen has only one electron, but also needs only one more electron to have a full shell.

When atoms bond together to make molecules, they either *exchange* electrons (ionic bonds) or *share* electrons (covalent bonds) to get full valence shells. By combining the electron configuration worksheets with conventional devices such as Lewis Dot Diagrams, students can gain a better understanding of how many electrons atoms have in their valence shells and the kinds of bonds they can make.

Atomic Spectra

When atoms are excited, their electrons jump from low energy levels to higher energy levels. Then the electrons drop back down to lower energy levels. Because energy is conserved, the excess energy has to go somewhere. It is usually released as a *photon* – that is, as light.

The diagram below, which uses the electron configuration worksheet, can be used to explain the process. It can be used to explain the flame test, fluorescent lights and other phenomena and devices.



The element shown in the diagram is Argon. In the diagram, the outside energy is labeled “Heat,” but it could also be an electric discharge (as in a fluorescent light), light or some other source of energy. Here are the steps:

a) The atom is in its ground state. The electrons are in the lowest possible energy configuration. No energy can be lost, so no light can be emitted.

b) With heat (or other energy) applied, the electrons in some atoms jump up to higher-energy unoccupied levels. The atom is now in an excited state.

c) After a short time, the electron jumps back down again to a low-energy level. Because energy is conserved, the extra energy must go somewhere. It is emitted as light.

The light that is emitted from excited atoms comes only at very specific levels, because the electrons can only jump between the distinct energy levels available in the atom. For this reason, when looking through a spectrometer at excited atoms, the light shows up in bright lines or bands, rather than being smeared out, like it is in a rainbow in the sky.

If you have access to spectrometers (and there are some very cheap and effective ones available), then you can have students look at fluorescent lights, incandescent lights, and sunlight. You may also be able to have them look at ions in a flame (the flame test). They should see bright lines in the fluorescent light and the flame test, but smeared out light in the incandescent light and sunlight.

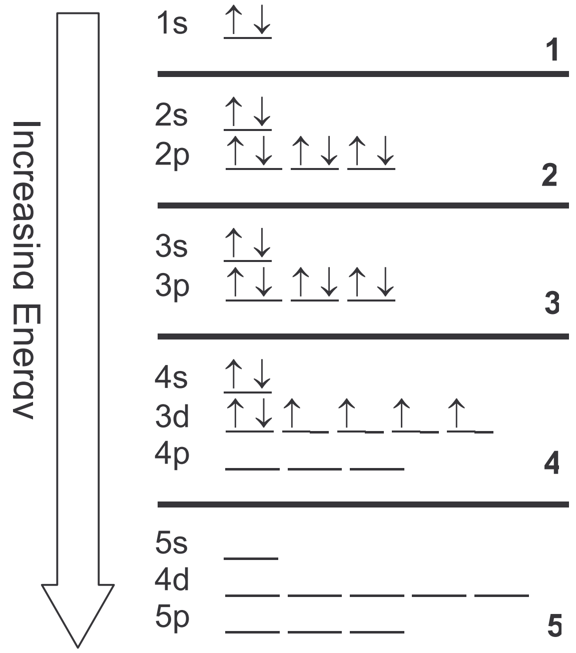
Note that our eyes reverse this process. Molecules in our eyes *absorb* light, which puts the molecules in a high-energy configuration. Then they drop to a ground-state configuration. Once again, the excess energy has to go somewhere. It goes into *chemical* energy that is used to stimulate nerves.

Magnetism

Look at the electron configuration for Iron:

Atom or Ion: Iron

Protons: 26 # Neutrons: ________



There are spin-up electrons in the *3d* level. Because the electrons act like little magnets, having four parallel spins leads to a pretty strong magnetic field for each atom (as atomic magnetic fields go). With enough iron atoms, the magnetic field can get quite strong. There are many complications, but this is the basic reason why iron is a magnetic material.

Using the Worksheet in Class

There are several ways I use the worksheet with my class:

- As a stand-alone worksheet, with a set of ions and atoms to diagram listed on a separate sheet or on the board.
- As an aid during a lecture. I lecture about the periodic table, then ask students to fill in the worksheet during the lecture so they can understand the lecture better.
- As part of a larger worksheet. (To do this, you might want to insert the *ElectronConfigEntry.gif* image file into your worksheet.)

Also, when I lecture on atomic spectra, I hand out spectroscopes as well as the electron configuration sheets. Then I present the diagram above showing the process of exciting atoms and having them emit light.

Other Questions and Caveats

I have tried to make the worksheet follow some simple rules. This means that I have omitted some complexities: atoms aren't *that* simple. However, I think the things I've omitted are really of secondary importance. I list some of them here as *Other Questions* – that is, things that students might ask about – and *Caveats* – things that students probably won't ask about, but that the teacher should know about.

Other Questions

- What do *s*, *p*, *d* (and *f*) stand for?
  - The levels are labeled by *s*, *p* and *d*. There is also a higher level that holds 14 electrons called the *f* level. These stand for "symmetric," "principle," "diffuse" and "fundamental." These terms come from long-ago efforts to make sense of atomic spectra. As far as I know, there is no reason for students today to know about this, unless they are interested in the history of science.
- Do all light sources look like bright bands when using a spectrometer?
  - No. All hot objects glow because of very basic laws of how energy is distributed in hot matter, and when you look through a spectrometer it looks smeared out, not like bright bands. The sun basically acts like a hot object, and so do incandescent lights and the heating elements in toasters and electric stoves. The red or yellow glow that wood fires give off is because of very small clumps of glowing ash, so they act like hot objects.

Caveats

- The energy levels are approximate. For example, as the *3d* level fills, electrons jump back and forth between the *4s* and *3d* levels. However, the worksheet does a pretty good "first-order" job.
- The more electrons an atom has, the harder it is to explain its properties in terms of a basic electron configuration. In my opinion, there is nothing wrong with saying

you'll focus mainly on the lighter elements in an introductory high school chemistry class. For students who want to become chemists, they should know that there is a lot of variety and surprise in the wilds of the periodic table.

- When atoms bond together, the locations of the electrons are described better using what are called "hybrid" orbitals that are combinations of *s*, *p*, *d* and *f*. If you want or need to talk about the shapes of molecules, you might need to talk about sp^3 orbitals, etc.