Electron Configuration of Atoms in the Periodic Table

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INTRODUCTION

As a high school chemistry teacher, my biggest challenge is a pretty difficult one. It is my job to facilitate the understanding of a sometimes-confusing subject while keeping it interesting; and also not to reduce chemistry to mundane memorization. I chose this seminar, "Technology and the Discipline of Chemistry," for two reasons. First, it is my belief that teaching is a dynamic profession. With the millennium fast approaching, it is imperative that teachers take responsibility for seeking out the current and upcoming technology in their subject areas. This responsibility may fall into the professional development requirement that is built into most teacher contracts. However, I feel that additional opportunities for personal growth, like the Houston Teachers Institute, are substantive and important ways to accomplish that growth. Finally, the students I teach are in honors chemistry; therefore I am always looking for ways to enrich their understanding of the subject. The use of computers in my curriculum would only enhance my students, almost all of who have a comprehensive knowledge of and access to computers in their homes.

In education courses, future teachers are taught about visual learners, auditory learners, and tactile learners. In my short experience in education I have found that as students progress from elementary school to middle school to high school, they tend to be visual and tactile learners. Case in point – what students don't like drawing pictures, cutting and pasting, and those messy, sticky, gooey science projects done by hand? The instruction during these early years is geared toward this type of learner. However, the high school experience is a different matter. So many high school teachers, myself included, have made the lecture the focal point of our curriculum and instruction. While the truly auditory learner is in heaven, lecturing can sometimes glaze over even the eyes of the most talented students. To relieve this slightly, there are lab experiments and projects. In the day-to-day theory instruction, however, computers can be an invaluable tool to greater comprehension by offering an alternative to pure lecturing.

With that said, my next step was to choose a topic of chemistry that could best be illustrated by the use of computers. I chose electron configuration because it is a topic in chemistry that lends itself to complex visualization.

OBJECTIVES

By the end of this curriculum unit, students will be able to:

- Comprehensively **articulate** the evolution of the atomic model in order to direct their thinking to associated electron configuration;

- **Demonstrate** their knowledge of atomic orbitals;
- **Describe** the shapes of s, p, and d orbitals;
- **State** the number of orbitals and the maximum number of electrons in the s, p, d, and f sublevels; and
- **Relate** electron configuration to the periodic table and elemental properties.

All these objectives will be accomplished from computer interaction with selected websites as well as from lectures.

THEORY

In the history of chemistry, it is a surprising but true fact that many of the advances were made by physicists. The theory behind electron configuration is not an exception. It is important for chemistry students to understand and be aware of the evolution of atomic theory. From it came the research and ideas that led to electron configuration, and thus, a better understanding of the atoms that define chemistry itself. It is vital that before introducing the electron configuration portion of this unit, students must understand the following.

Atomic Model Theory in a Nutshell

This section sequentially discusses those people who were instrumental in developing the modern atomic theory from which electron configuration can be derived. There are more that were not included; however, the theories discussed represent the major contributors to our modern atomic theory.

John Dalton

He was the first to develop an atomic model based on four basic tenets:

- All matter is composed of extremely small indivisible particles called atoms.
- All atoms of a given element are alike, in mass and other properties, but atoms of one element differ from the atoms of other elements.
- Compounds are formed when atoms of different elements unite in fixed proportions, also called the *law of multiple proportions*.-
- A chemical reaction involves a rearrangement of atoms. No atoms are created, destroyed, or broken apart in a chemical reaction.

While the basic postulates Dalton advanced are true, the subsequent discovery of subatomic particles soon disproved the indivisible atom theory.

J.J. Thomson

He discovered electrons and came up with the plum pudding atomic model. Negatively charged electrons (raisins) were stuck in the middle of a lump of positively charged protons (the pudding). Neutrons were not mentioned, as they had not yet been discovered. However, Thomson's model didn't articulate the number of protons, their location in the atom, or the way atoms formed ions.

Ernest Rutherford

He and his assistants bombarded thin metal foils with alpha (positively charged) particles. Most of the particles went through the foil with a few being deflected. Rutherford expected this because he had hypothesized that the nucleus of an atom, which is positively charged, makes up only a small part of the volume of an atom. Consequently, the particles that passed through the foils did so because the small number of electrons made for an atom with a lot of empty space. Also this model made researchers wonder, What keeps the atom from collapsing if the negatively charged electrons were attracted to the positively charged nucleus?

Niels Bohr

Bohr, who was a student of Rutherford, proposed the next widely accepted atomic model. To extend Rutherford's 'wide open space' atomic model, Bohr found that electrons did not just float around the nucleus of an atom, but had fixed paths and fixed energies. No electron was ever in danger of falling into the nucleus. His model recognized that electrons had energy levels, and regions around the nucleus where the electrons were likely to be moving. His energy levels of electrons were analogous to the rungs of a ladder. The lowest rung was like the lowest energy level. Ladders must be ascended or descended by climbing from rung to rung. Similarly, electrons can go to higher or lower energy levels. Just as one cannot stand between the rungs of a ladder, electrons cannot 'stand' between energy levels. The correct amount of energy must be gained or lost to go up or down an energy level. This amount of energy, Bohr said, is a *quantum* of energy, or the energy to move electrons to the next level. Bohr also proposed that the higher an electron was on the energy ladder, the larger the diameter of its fixed path.

Max Planck and Louis de Broglie

Planck and de Broglie are mentioned here not as atomic theorists, because neither had one. They did, however, both propose concepts that helped the modern atomic model evolve and later led to the development of the electron microscope.

During this time physicists had been experimenting with light and its relation to atomic emission spectra. Atomic emission spectra are the pattern of frequencies obtained when light emitted by atoms is passed through a prism. Atomic emission spectra are different for every element and consist of only a few select frequencies of light. Classic physicists had a difficult time reconciling this information, as they felt that any spectrum of the emitted light should be continuous. Max Planck advanced a mathematical equation that helped explain the atomic emission spectral frequencies. They were directly related to the amounts of energy involved.

De Broglie introduced the concept that all matter exhibit wavelike properties. This proposal led to a new method of describing the motion of subatomic particles, atoms, and molecules, called *quantum mechanics*. De Broglie also derived a mathematical equation that made it possible to describe the wavelengths of all matter. Of course, subatomic particles have the most measurable wavelengths, but de Broglie's equation does allow for computation of the almost undetectable wavelengths of large objects.

Erwin Schrodinger

The atomic models up until Schrodinger's research had been physical models. He was able to extend Bohr's quantum energy theory to write and solve a mathematical equation describing the location and the energy of an electron in a hydrogen atom. This equation led to the modern description of atoms used today, the quantum mechanical model. Like the Bohr model, the quantum mechanical model uses quantized energy levels for electrons. Unlike the Bohr model, it does not focus on electrons having a direct path around the nucleus of an atom. Rather, it is a mathematical determination of where there is a *high probability* of finding electrons on a particular atom.

ATOMIC ORBITALS

In the quantum mechanical model of the atom, the energy level of electrons are designated by the principal quantum numbers (n). These are assigned certain values: n = 1, 2, 3, 4, 5, 6, 7, and so on. Within each principal energy level, the mathematics of quantum mechanics gives several electron cloud shapes. This is because the quantum mechanical model of an atom divides the energy levels into sublevels. Each sublevel corresponds to a different cloud shape. Each cloud shape can be calculated from a mathematical expression called the atomic orbital. These orbitals represent the area where there is a high probability of finding electrons. Different atomic orbitals are denoted by letters. The *s* orbital defines a spherical cloud of electrons, the *p* orbitals define dumbbell shaped clouds, while the *d* orbitals have a mixture of different shapes. The principal quantum number always equals the number of sublevels within that principal energy level. The maximum number of electrons that can occupy a given energy level is given by the formula $2n^2$.

Electron Configuration Using the Periodic Table

As stated before, the energy levels are divided into sublevels of orbitals. The first level contains one sublevel. This is an *s* orbital and contains only two electrons. The second level has two sublevels containing an *s* orbital and three *p* orbitals, respectively. Each of the four orbitals can hold a maximum of two electrons; the second energy level may have a total of eight electrons. The third energy level has three sublevels, containing, respectively, one *s* orbital, three *p* orbitals, and five *d* orbitals. Again, each orbital holds only two electrons, so the third energy level accommodates eighteen electrons. This goes on through each energy level. Within a given energy level, the *s* orbital is the lowest energy sublevel, *p* is higher, *d* is higher, and *f* is the sublevel with the highest energy. See Figure 1.

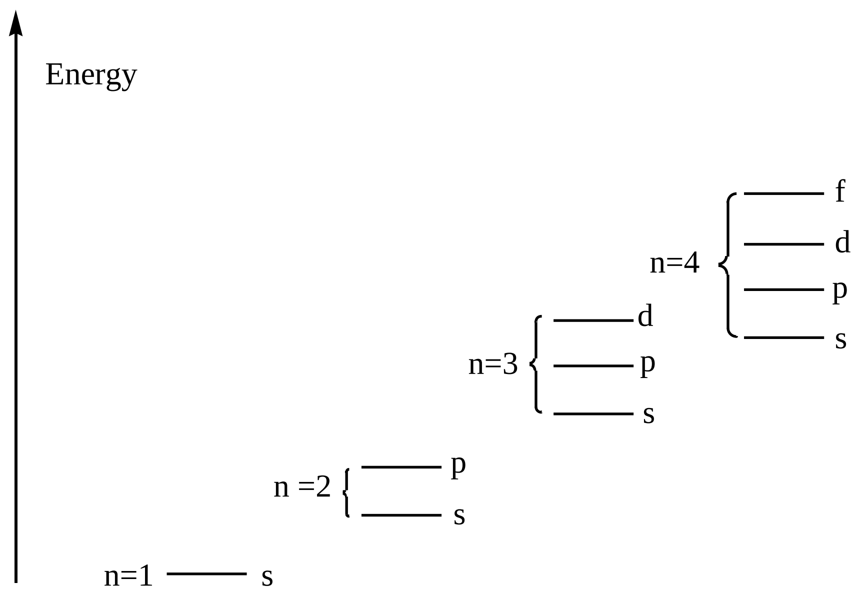

Figure 1

The relationship between the *s, p, d,* and *f* sublevels and the periodic table may be observed by noting that the table can be divided into blocks. One of the blocks is two elements wide (Groups 1A and 2A); another is six elements wide (Groups 3A through 8A); a third is ten elements wide (the Group B transition elements); and the fourth is fourteen elements wide (the inner transition elements). These correspond, respectively, to the elements whose last electrons go into the *s* orbitals, the *p* orbitals, the *d* orbitals, and the *f* orbitals. It can be seen from Figure 2 that there is relationship between the arrangement of elements on the periodic table, the order of filling the various energy levels, and the maximum number of electrons possible in the various sublevels.

Three main guidelines govern the filling atomic orbitals within the energy levels. They are the Aufbau principle, the Pauli exclusion principle, and Hund's rule.

The Aufbau Principle

Electrons enter orbitals of lowest energy first. The various orbitals within a sublevel of a principal energy level are always of equal energy. Yet the range of energy within a principal energy level can overlap the energy levels of a nearby principal energy level. As a result, the filling of orbitals does not follow a simple pattern beyond the second energy level. For example, the 4*s* orbital is lower in energy than the 3*d*.

The Pauli Exclusion Principle

An atomic orbital may only hold two electrons. To occupy the same orbital, two electrons must have opposite spin. Spin is a property of electrons that may be clockwise or counterclockwise.

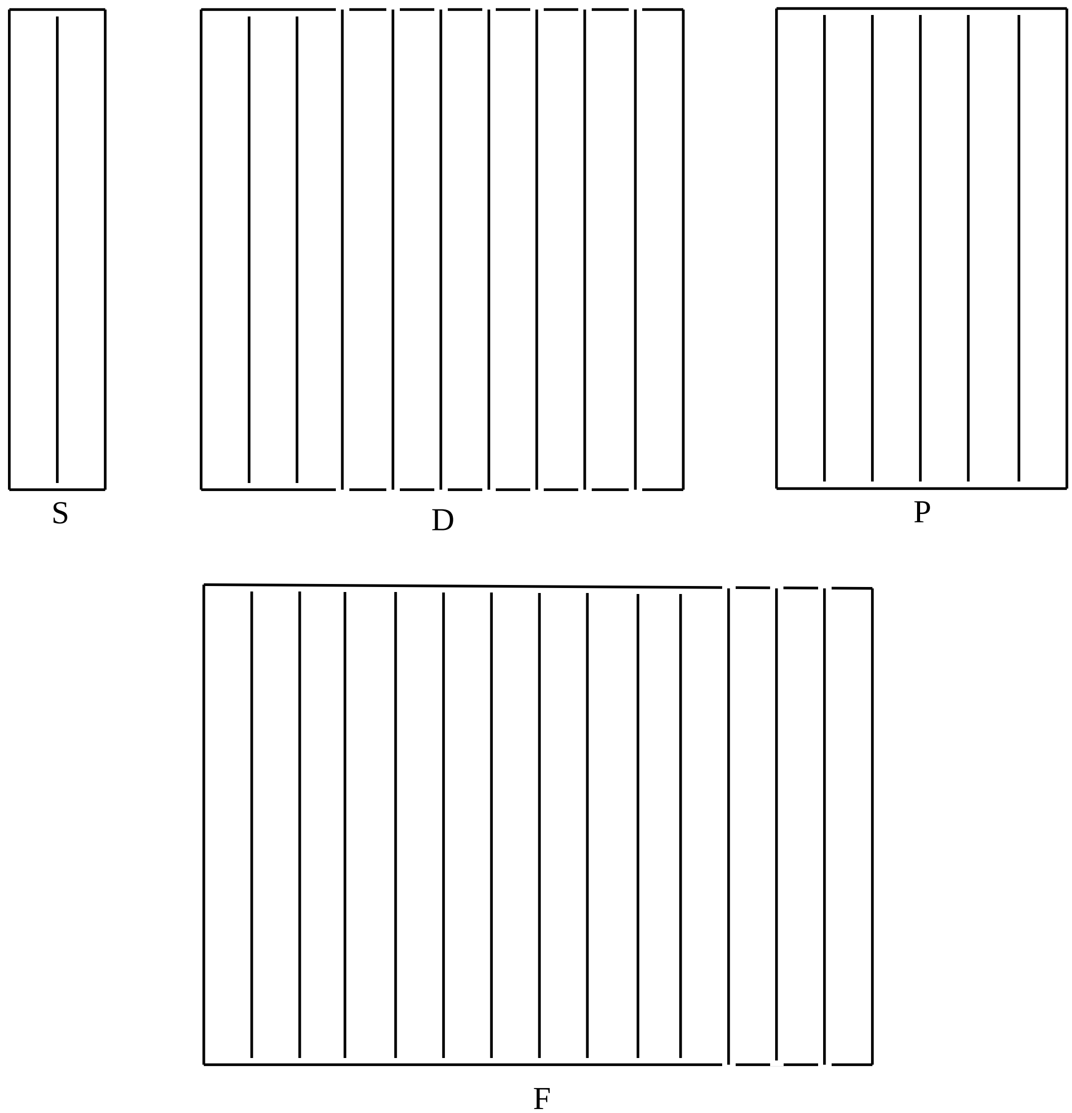

Figure 2

Hund's rule

When electrons occupy orbitals of equal energy, orbitals must be singly occupied with electrons having parallel spins. Second electrons are then added to each orbital so that the two electrons in each orbital have opposite spins. The arrows denote the direction of the electron spin. This is shown in Figure 3.

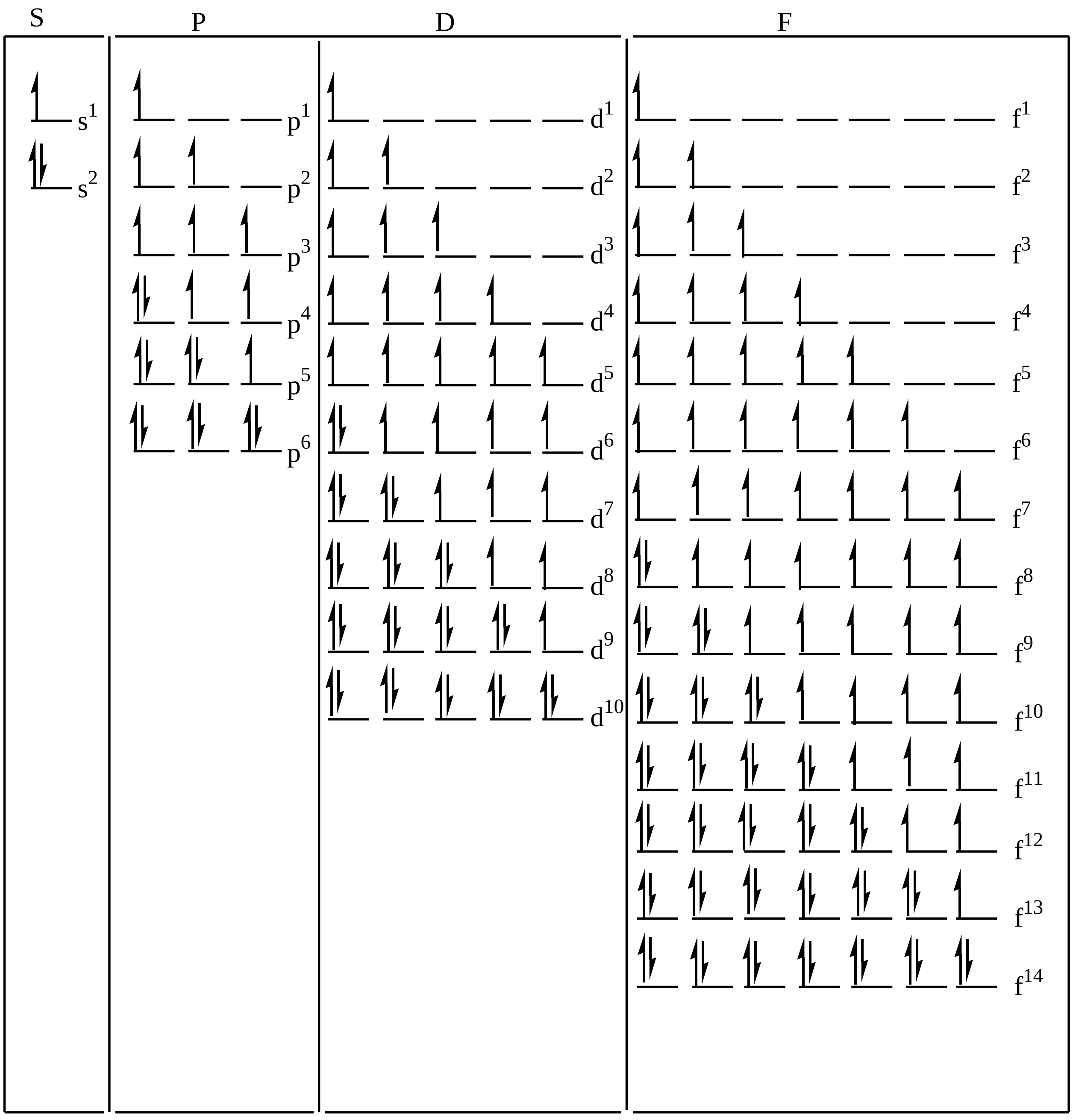

Figure 3

The electron configuration of any atom may be written by reading the periodic table from left to right in order of increasing atomic number and by correlating the areas in the chart with the various sublevels. This relationship is illustrated by the chart below.

1. **Locate Groups 1A and 2A.** These two columns represent elements whose atoms have their valence electrons (outer electrons) in the *s* sublevel. (Remember there is a maximum of two electrons in any *s* sublevel.)

2. **Locate Groups 3A through 8A.** These six columns contain the elements whose valence electrons are in the *p* sublevel. (A maximum of six electrons in any *p* sublevel.)

3. **Locate the Group B transition elements.** These ten columns represent elements that have electrons in the *d* sublevel. (Once again, a maximum of ten electrons can be placed in *d* sublevels)

4. **Locate the two rows of inner transition elements at the bottom of the table.** These are elements whose electrons enter the *f* sublevel. There are fourteen elements in each row, the same number as the maximum number of electrons possible in an *f* sublevel.

Periodic Table of the Elements

1A	**2A**	**3B**	**4B**	**5B**	**6B**	**7B**	**8B**	**8B**	**8B**	**1B**	**2B**	**3A**	**4A**	**5A**	**6A**	**7A**	**8A**
H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	**La**	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	**Ac**															
			Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu	
			Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr	

STRATEGIES AND PLANS

The objective of this unit is to use technology to supplement the instruction of electron configuration. The most important tool that will be used is the computer. A classroom equipped with an overhead computer, which would allow students to see the teacher’s web interaction projected onto an overhead screen is ideal for this particular type of instruction. However, in even the most technologically advanced high school chemistry classes, this piece of equipment is not yet the standard. The next best choices are computer labs for individual student work or multiple classroom computers to enhance cooperative learning. The following represents plans for the curriculum, with instruction occurring over three class periods.

Lesson 1

Activities

This lesson will include a short introductory lecture on atomic theory evolution. A demonstration of the atomic model evolution on the computer would be appropriate here. Student participation would be an informal question and answer period following the lecture. It would also be a good idea to poll the students' prior knowledge of the Internet with respect to surfing various websites. Students could then do a guided surfing session of some of the chemistry websites that relate to this topic.

Suggested Materials

Overhead projector with computer
Individual student computers
Websites: www.chemtutor.com/struct.htm
http://ww.chem.ualberta.ca/~plambeck/che/p101/p01224.htm
http://antoine.frostburg.edu/chem/senese/101
http://www.sfu.ca/chemcai/QUANTUM/Quantum_Primer.htm

Lesson 2

Activities

Demonstration of electron configuration and orbital shapes using a computer generated representation of a periodic table. Students will be evaluated at the end of the lecture by their ability to demonstrate element electron configuration and different orbital shapes. In addition, students would have been able to visit several websites to practice these skills. This exercise does not necessarily require the use of a computer, as students could perform their evaluation orally or as a written exercise.

Suggested Materials

Individual student computers
Overhead computer with projector
Websites: www.chemtutor.com/struct.htm
http://library.advanced.org/10429/text/eleconfig/electron.htm
http://dummyweb.albany.net/~cprimus/orb/index.html
http://antoine.frostburg.edu/chem/senese/101

Lesson 3

Activities

Using information from Lessons 1 and 2, students will perform guided practices based on orbital shapes and electron configuration. Students will be able to draw and recognize these shapes as well as atomic electron configurations. During this activity, students would be encouraged to visit web sites related to the topic to further supplement their instruction.

Suggested Materials

Individual student computers
Websites: http://dummyweb.albany.net/`cprimus/orb/index.html
http://www.chem.ualberta.ca/~plambeck/che/p101/p01224.htm
http://www.intschool-leipzig.com/bailey/tutorial/index.htm
http://www.scifun.chem.wisc.edu/scifun.html

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http://www.chem.ualberta.ca/~plambeck/che/p101/p01224.htm
http://antoine.frostburg.edu/chem/senese/101
http://zopyros.ccqc.uga.edu/lec_top/quantrev/quantrev.html
http://despina.advanced.org
http://www.sfu.ca/chemcai/QUANTUM/Quantum_Primer.html
http://www.intschool-leipzig.com/bailey/tutorial/index.htm
http://scifun.chem.wisc.edu/scifun.html