GY 111 Lecture Note Series
Elemental Chemistry

Lecture Goals:
A) Basic Atomic structure (Chemistry 101)
B) The Periodic Table

Reference: Press et al. (2004), Chapter 3; Grotzinger et al (2007), Chapter 3

A) Basic atomic structure (with apologies to my Chemistry colleagues)

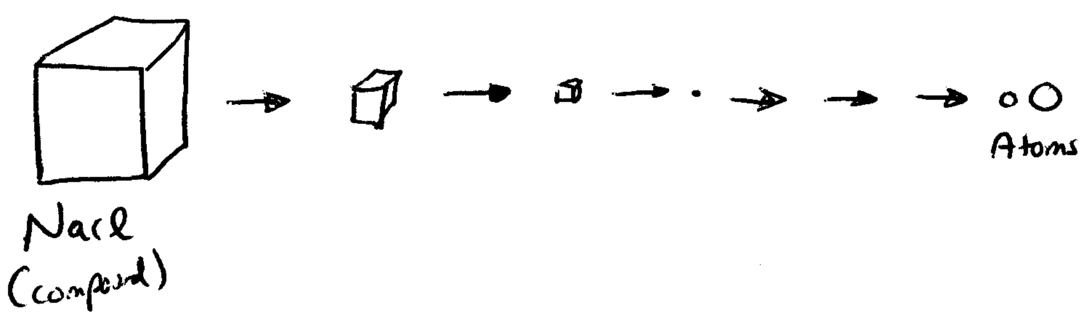
A bit of history is required before we start into the chemistry stuff. During the good old days (i.e., 100's of years ago), "chemists" were essentially magicians. They had learned some basic data about the planet around them, including basic elements. At the time, most people thought that their universe consisted of 4 basic elements:

1) Fire
2) water
3) earth
4) air

Of course some earth was better than others. Gold came from the Earth, and was highly sought after. Some chemists tried to figure out ways to convert earth into gold. They were the **alchemists**. They never really figured out how to do this (it is possible, but you need a nuclear reactor to do it).

After a while, chemists settled down and actually started to try and figure out the materials that made up the universe. They asked a simple question *what are the basic building blocks of matter?*

Consider the mineral halite (NaCl; pictured above from http://www.geoclassics.com/halite.jpg). This mineral has a cubic crystal habit. It also has 3 excellent cleavages (all at 90-degrees to one another), meaning that if you hit halite with a hammer, it will always break into smaller and smaller cubes of halite:



But eventually you come down to a single **molecule** - one sodium particle (Na) attached to one chlorine particle (Cl). Up until this century, chemists felt that these particles (**atoms** or **ions**) were the basic building blocks of all matter. Atoms were defined as *the smallest division of matter that retains characteristics of a particular "thing".* Ions were *atoms that possessed positive or negative charges* (we'll get into these shortly)
Those particular things were called **atoms**. At last count, there were just over 100 elements (although several of them were produced in labs rather than found in nature). Each has a specific chemical symbol.

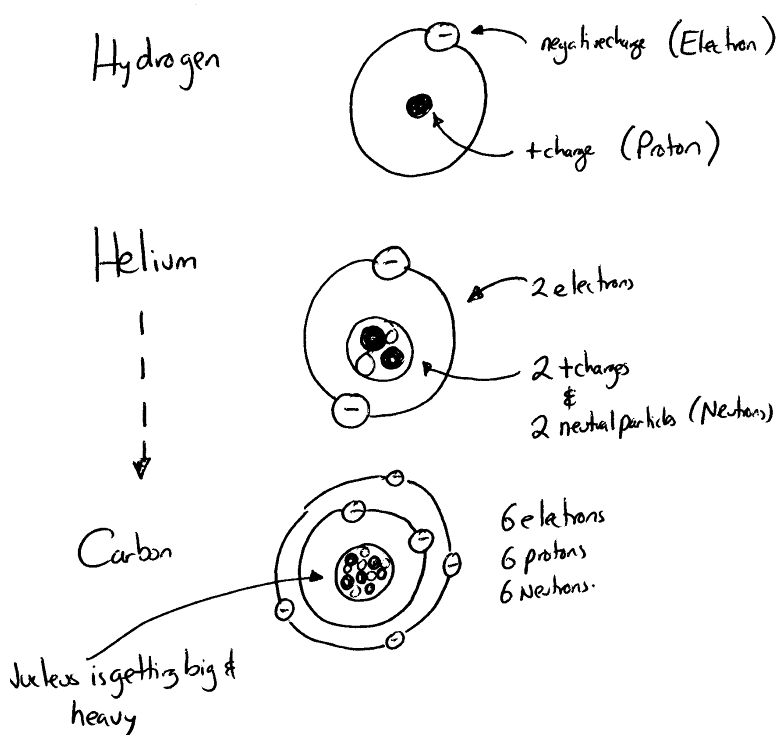
The elements can combine through various chemical reactions to for **compounds**. For example:

water is H_2O (one part hydrogen + 2 parts oxygen)

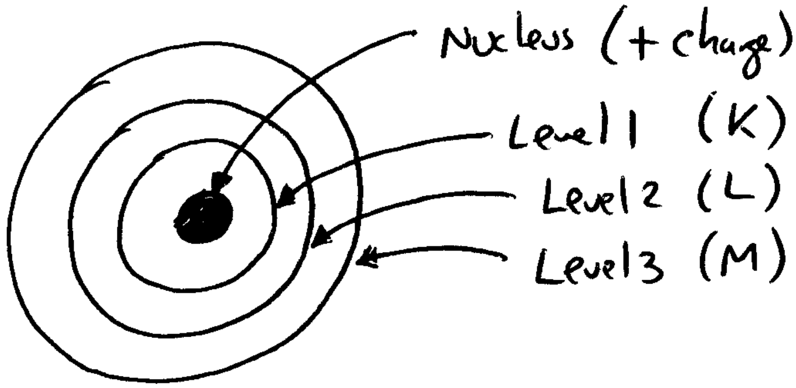
galena is PbS (one part lead - Latin is *plumbum*, + 1 part sulfur)

chalcopyrite is $CuFeS_2$ (one part copper + 1 part iron + 2 parts sulfur)

Many of the man-made elements and even some of the naturally occurring ones are unstable. They tend to break apart over time through a process called or **radioactivity**. This must mean that there is something smaller than an atom (i.e. atoms and ions are <u>not</u> the smallest particles). It has been known for well over 100 years that atoms are composed of other (sub-atomic) particles. The important ones are **protons, electrons** and **neutrons**. Each atom contains various numbers of these particles. Protons always carry a single positive charge; electrons always carry a single negative charge and neutrons are as the name implies, neutral or non charged. Each atom contains an equal number of protons and electrons; ions contain differing numbers. For example:



The protons and neutrons are held within the **nucleus** of the atom/ion. Electrons occur within clouds around the nucleus. Electrons do not really orbit the nucleus, although it does sometimes help to imagine that they do. They are assigned to different energy levels or **electron shells.** Each successive level is at a higher energy level than the previous one and each energy level contains a different maximum number of electrons:



Level (n)	**"Name"**	**# of electrons** ($2n^2$)
1	K	2
2	L	8
3	M	18
4	N	32
5	O	50

Now we need to start thinking about how atomic properties affect minerals. The more particles that you have in an atom\ion, the heavier it becomes. But, not all particles are created equal.

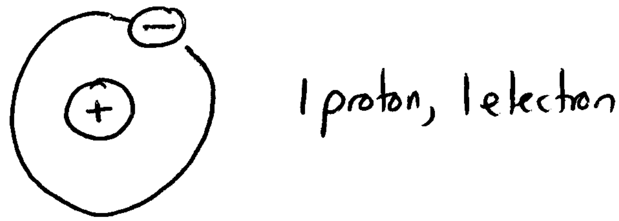
Particle	**Mass**
Proton	1.6×10^{-24} g
Neutron	1.8×10^{-24} g
Electron	8.0×10^{-28} g

The mass of an electron is about 1/2000th the mass of a proton, so electrons can largely be ignored when estimating molecular weights. Since all of an elements mass rests in its nucleus, the mass of an ion is not significantly different from the mass of an uncharged atom, but not all atoms are created equal!

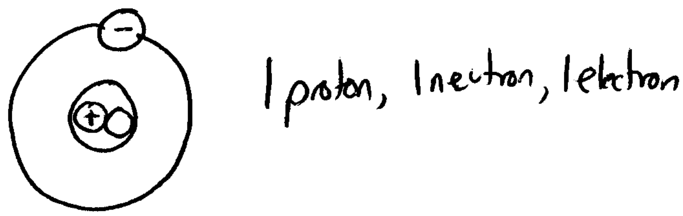
You have already been told that pure elements are composed of atoms that contain equal numbers of protons (+) and electrons (-). The number of neutrons is not as fixed. Take Hydrogen for example. Most hydrogen atoms contains 1 electron and 1 proton. A much smaller number contain 1 proton, 1 electron and 1 neutron or 1 proton, 1 electron and 2 neutrons.

None of these hydrogens are charges (all of them contain 1 electron and 1 proton), but they all have different masses. Consider that a neutron and a proton weigh about the same, so the addition of a neutron <u>doubles</u> the mass of the hydrogen atom and 2 neutrons triples its mass. As you will discover shortly, it is the number of protons and electrons that dictate an elements properties. All a neutron does is change the mass of an element. So hydrogen with 1 neutron or 2 neutrons is more or less the same element, just heavier. The same element with different numbers of neutrons are called **isotopes.** Hydrogen has 3 isotopes. Usually you indicate different isotopes by placing a number (atomic mass) adjacent to the chemical symbol (e.g., H^1, H^2, H^3, U^{235}, U^{238} etc), but for hydrogen, the isotopes are given specific names:

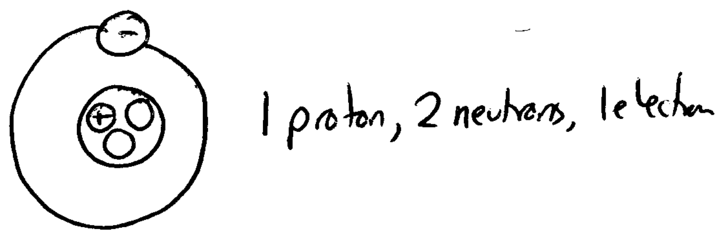
Hydrogen (H):



Deuterium (D):



Tritium (T):



Deuterium is used to make heavy water (D_2O), and tritium is radioactive (e.g., T_2O is deadly)

In the late 1960's, scientists started dissecting protons, neutrons and other sub-atomic particles. What we found was that these formerly smallest particles are actually composed of even smaller particles called **quarks** (3 quarks for each proton\neutron). Imaginative names were applied to many of them (flavor, charm, red, green etc).

B) The Periodic Table of the Elements

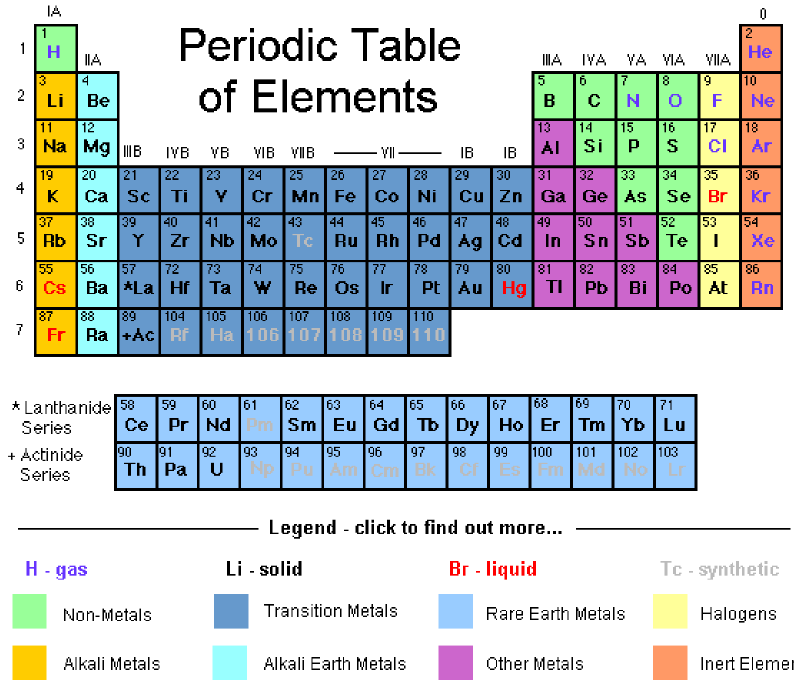
This is to chemists what the geological time scale is to geologists. The table (see hand out issued in class) organizes all of the elements according to similarities in properties.

Four important pieces of information can be instantly gleamed form the table. For example, refer to sodium

(1) Name of element
(2) Chemical Symbol (Latin term is natrium)
(3) **atomic number** (number of protons in nucleus)
(4) **atomic weight**

The atomic weight isn't really an actual weight, but is in fact a ration between the element in question (say sodium) and carbon. It uses the **atomic mass** of carbon.

1 (atomic weight) =1/12 carbon mass



The rows of the periodic table are arranged to display the elements in terms of increasing atomic number and mass (and weight). There are 8 main columns (labeled I through VIII): Note: the version of the periodic table handed out in class is more detailed than the one in your textbook. This one uses IA, IIA, IIIA….IB, IIB, IIIB…VIIB, VIII. Note that there are two series that overlap near the middle portion of the table. If this were Chemistry 131, I'd have to tell you why this was done, and you'd have to know why for the test. As this is Geology 111, we can ignore it.

The main thing to remember about the periodic table is that it organizes the elements according to their properties. Each column contains elements that display similar properties and their properties are largely controlled by the number of electrons that occur in the *outermost shell of the atom.*

Column #	**Common Name**	**# of electrons in outermost shell**	**# of electrons needed to be gained or lost in order to fill outermost shell**
IA	**Alkali Metals**	1	1
IIA	**Alkaline Earths**	2	2
IIIB		3	3
IVB		4	4, 4
VB		5	3, 5
VIB		6	2
VIIB	**Halogens**	7	1
VII	**Inert Gases**	8	0

Elements are at their most stable when the outer most shell is completely filled. For the purposed of our simplified chemistry lesson, we will assume that the elements need 8 electrons to fill an out shell (except of course for H and He which only need 2). For Na (column IA) to get "stabilized", it needs to ditch an electron. Mg (column IIA) would need to ditch 2 and Boron (column IIIB) would have to lose 3. In contrast, Chlorine (column VIIB) would need to *gain* one electron and Oxygen (column VIB) would need to *gain* two.
Every gain or loss of electron(s) induces 2 important changes:

(1) the charge balance is lost. Na becomes Na^{+}; Mg becomes Mg^{2+}, B becomes B^{3+}, O becomes O^{2-} and Cl becomes Cl^{-}. They are no longer atoms. Now they are **ions.** Ions are charged atoms. There are two broad types: **cations** are positively charged ions and **anions** are negatively charged particles.

(2) the size of the ion changes. A cation has less electrons than protons so every electron now feels a bit more "pull" than it did before the other electron was lost. Consequently, cations are somewhat *smaller* than the uncharged atom. Anions are larger for the very same reason.

Chemical reactions are driven by the behavior of the different elements undergoing the reaction. For example, if Na metal is added to Cl gas, a vigorous reaction takes place according to the following:

$$Na + Cl \rightarrow Na^{+}Cl^{-} + \text{a lot of heat}$$

So what you get is the mineral halite. The reaction is strongly **exothermic** because the transfer of the electron releases energy from *both* Na and Cl. The Na^{+} and Cl^{-} ions are strongly attached or **bonded** together. Other chemical compounds are not so well bonded. Bonding will be among the topics that we discuss next time.

Important terms/concepts from today's lecture

(Google any terms that you are not familiar with)

Alchemist
molecule
atoms
ions
elements
compounds
radioactive decay/radioactive decay
protons
electrons
neutrons
electron shells
nucleus
isotopes
atomic numbers,
atomic mass,
atomic weight
ions
cations
anions

Useful Websites

Periodic Table: http://www.cs.ubc.ca/cgi-bin/nph-pertab/periodic-table
Mineral collectors information page: http://collectminerals.about.com/