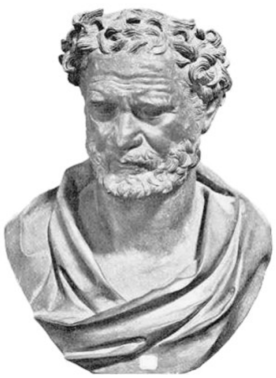
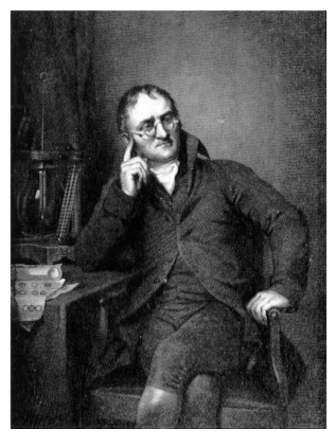
Chapter 3 Atomic Diary

History of the Atom



Democritus was a philosopher in ancient Greece who "thought" about things, and came up with his ideas. He thought that all kinds of matter were unique and that you could cut it in half over and over until you reached some tiny part that could no longer be cut in half anymore. He called this "atomos", which means indivisible. His nick-name of "atom" has stuck to this day.

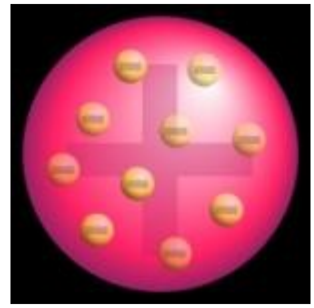
John Dalton(1803) was a farmer who meddled in science, using gases to examine atoms. He thought atoms were like billiard balls, small hard spheres, unique for each element. He stated that they differed only by mass, which somehow accounted for all the property differences between atoms. He published his four part Atomic Theory which said:



1. All matter is made up of extremely small particles called atoms
2. Elements are made up of only one kind of atom, each identical to the others in properties and mass
3. Two or more atoms can combine in small, whole number ratios to form compounds
4. In a chemical reaction atoms are re-arranged (combined or separated) - but not destroyed

Dalton's theory is still relevent, but we now know that atoms are in fact made up of sub-atomic particles called neutrons, protons, and electrons. He was clearly on the right track.

JJ Thompson(1897) was a scientist who discovered the electron. He used a device called the cathode ray tube, and was able to find the electron. He had no knowledge of protons or neutrons, or atomic structure, so he "stuck" these new found electrons into a sort of positively charged "plum pudding" model of an atom. Try to imagine that electrons are the chips in a chocolate chip cookie, and the rest of the atom (the cookie part) is all positively charged, enough to cancel out the negatively charged electrons and you will have a good idea of what he proposed.

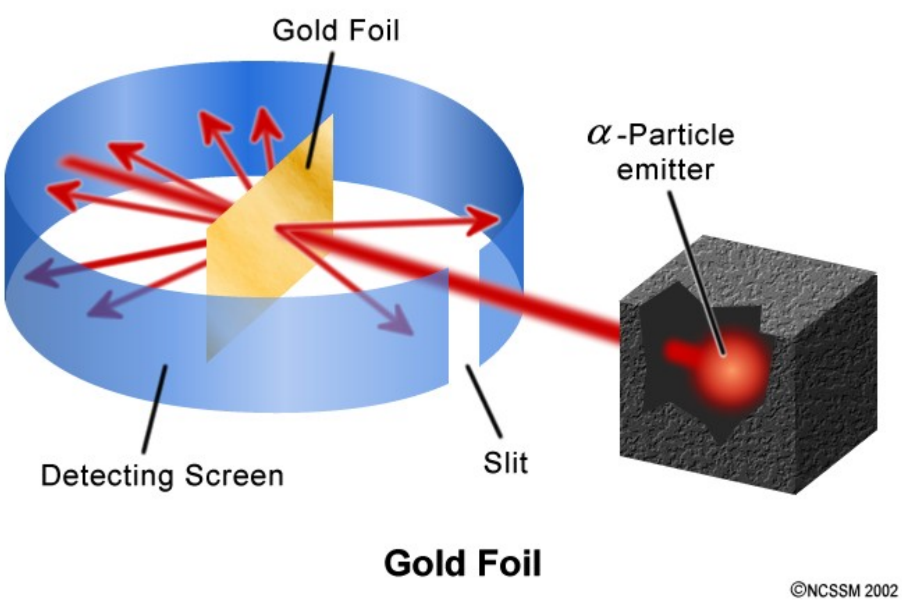


Ernest Rutherford(1897) furthered atomic theory with an experiment called the GOLD FOIL experiment. He managed to prove that the electrons were flying around at a good distance from the nucleus of the atom. He figured out that the nucleus was positively charged and that this nucleus was relatively small compared to the size of the whole atom.

Rutherford's Gold Foil Experiment

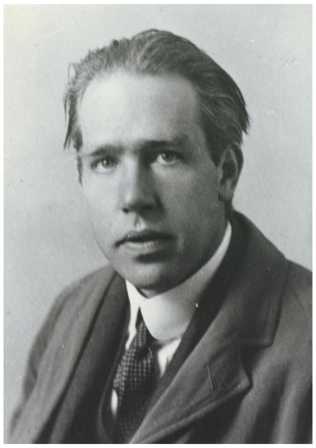
Rutherford put the radioactive atom Polonium (Po) into a big lead box. As it decayed it released alpha particles (2 protons stuck to two neutrons, +2 overall charge, mass of 4 amu) that he directed through a small hole in the lead box in order to aim them.

The particles were aimed at a near circular screen that could register these alpha particles. He placed a thin sheet of gold foil in the way of these alpha particles, and astonishingly, most of the particles seemed to travel right through the gold atoms as if it wasn't even there! Some, possibly 1 in 10,000 of the alpha particles was deflected and hit the screen at large angles, even bouncing nearly straight back to the source.



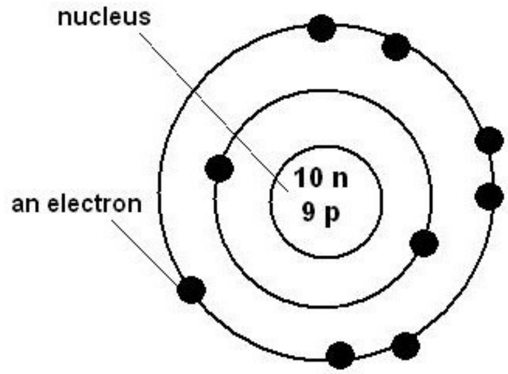
He determined that the atom was mostly empty space (most of the alpha particles missed the atoms), and that the gold atoms had a dense and positively charged nucleus (the positive alpha particles didn't stick). This was in complete disagreement with the Thompson Plum Pudding model of the atom, and a major leap forward in atomic theory.

Unfortunately a big problem he could not manage was why the negatively charged electrons didn't just collapse into the positively charged nucleus. They should. On the other hand, all the atoms exist, and they clearly don't do this. Oh well, no one has ALL the answers. This problem was to be solved by a better mathematician.



<u>Neils Bohr</u>,(1922) was a student of Rutherford. He proved with math that the electrons do not lose energy if they can stay in precise orbits around the nucleus. Each orbit has a particular amount of energy lassociated with it. He defined orbits as energy levels. The closer to the nucleus, the smaller the orbit, and the lower energy. The further an orbit is from the nucleus the more energy the electrons in them have.

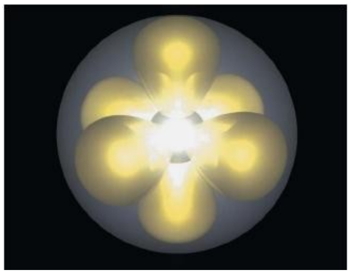
Neils math works great for the atom hydrogen, but as soon as you add even one more electron (say helium, or all the other atoms) the math falls short of proving what’s going on. Still, his ideas stand, and the electrons get drawn into nice little planetary diagrams, with the electrons filling up these orbits.



n represents one neutron
p represents one proton

The first orbit can fit just 2 electrons.
The second orbit holds up to 8 more.
The third orbit can fit up to 8 electrons as well, but sometimes it can hold up to 18.

<u>Modern Model</u>: Finally, in the early part of the 20th century, as the math got fancier and quantum theory became the rage, the atom was again reconfigured. The modern model, or the wave-mechanical model, describes the nucleus as still central. Neutrons have a charge of zero. Protons are still positive and exactly balanced by the negative charges of the electrons flying about. All atoms are still neutral, but the electrons no longer follow in neat little circles like in Bohr's time. Now these electrons are moving about in a sort of statistical cloud, a zone, called an ORBITAL. Orbits were constant radial paths, but now the radius of any electron is a bit fuzzy, and more difficult to pin down. Electrons can act like particles, and sometimes they act as waves of energy. You can never determine both the speed of an electron and its location at the same time - that’s called the Pauli Exclusion Principle.



ORBITALS are more like "rooms" that electrons live in, and they have remarkably complex shapes. Electrons can only have certain amounts of energy, in precise "quanta" amounts. For our class, it's all about the orbitals, don’t sweat the shapes of orbitals or the fancy equations.

<u>Sub-Atomic Particles</u>

Atoms are made of protons, neutrons, and electrons.

Protons have a +1 charge, a mass of 1 amu, they are in the nucleus only.

Table O
Symbols Used in Nuclear Chemistry

Name	Notation	Symbol
alpha particle	${}^{4}_{2}\mathrm{He}$ or ${}^{4}_{2}\alpha$	α
beta particle (electron)	${}^{0}_{-1}\mathrm{e}$ or ${}^{0}_{-1}\beta$	β^-
gamma radiation	${}^{0}_{0}\gamma$	γ
neutron	${}^{1}_{0}\mathrm{n}$	n
proton	${}^{1}_{1}\mathrm{H}$ or ${}^{1}_{1}\mathrm{p}$	p
positron	${}^{0}_{+1}\mathrm{e}$ or ${}^{0}_{+1}\beta$	β^+

Neutrons have no charge, a mass of 1 amu, and they are in the nucleus only as well.

Electrons have a -1 charge and fly around the nucleus in specific orbitals, or energy levels. In our high school chemistry class the electron mass is SO SMALL, we consider it to be zero, but it is in fact about 9.1066×10^{-28} grams (which is, 0.00000000000000000000000000091066 grams) which we all realize is pretty close to nothing anyway or 1/1836 of an amu.

Determining numbers of p^+, n° and e^- in an atom using our periodic tables.

All atoms are listed in ascending atomic number. The atomic numbers equal the number of protons an atom. Since ALL ATOMS ARE NEUTRAL, the number of electrons must equal the number of protons found in the nucleus. The negative charges balance out the positive charges in a 1:1 ratio.

The atomic mass of an atom is the mass of the nucleus, the protons and the neutrons only (remember in our class the electrons are of no mass). So for this concept, we round off the atomic mass number on the periodic table to the nearest whole number (more on atomic masses below). The atomic total mass is the protons plus the neutrons. If you know the mass of an atom, and can subtract off the atomic number, or number of protons, the left over mass is made up of only neutrons. This method will work for all atoms.

47.88

Ti

22

2-8-10-2

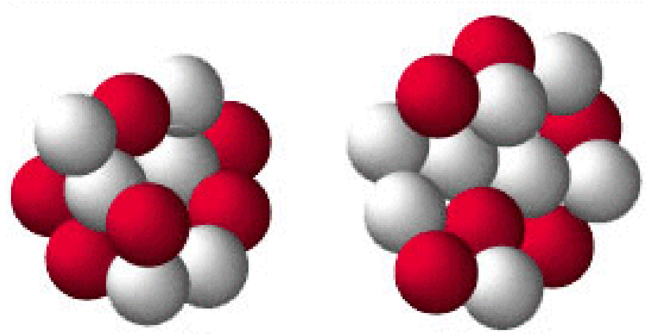
Titanium is shown at right. It’s mass is rounded to 48 amu. 48 is the total number of titanium’s protons plus neutrons. It does not tell us about how many of each, but we will figure out that now.

Titanium’s atomic number of 22 tells us that it has 22 protons as well as 22 electrons (neutral atom).

$$\begin{array}{rl} 48 & = \text{protons plus neutrons} \\ -\ \underline{22} & = \text{protons (from atomic number)} \\ 26 & = \text{neutrons} \end{array}$$

<u>Isotopes</u>

Not all carbon atoms are exactly the same. John Dalton said they were but he was incorrect. It turns out that all carbon atoms are CHEMICALLY IDENTICAL because they have the same numbers of protons and electrons, but they have DIFFERENT numbers of neutrons. Neutrons do not affect the chemistry of the atoms, just the masses.



2 Isotopes of Carbon
C-12 and C-14

The number of neutrons can vary, and there is not a clear connection between number of neutrons and other the subatomic particles (p^+ and n°).

On the right are the nuclei of the 2 isotopes of carbon. Both have 6 red protons. On the left there are six white neutrons, on the right there are eight white neutrons. Both would have six electrons. Both react identically in a chemical reaction, but have different masses (12 amu vs. 14 amu).

Isotopes and Average atomic masses

The atomic masses on the Periodic Table are mostly decimal measures while we know each atom has a whole number of protons and neutrons making up this mass. The reason for this decimal is because isotopes exist in nature and they need to be taken account of. Atoms of the same element can have different numbers of neutrons; the different possible versions of each element are called isotopes. For example, the most common isotope of hydrogen has no neutrons at all; there's also a hydrogen isotope called deuterium, with one neutron, and another, tritium, with two neutrons. The atomic masses on the Periodic Table are the weighted averages of all the masses of each isotope multiplied by its proportion of the whole. That is where the decimals come from.

Hydrogen

Deuterium

Tritium

How to Calculate an Average Atomic Weight

To do these problems you need some information: the exact atomic weight for each naturally-occurring stable isotope and its percent abundance.

Example: carbon

mass number	exact weight	percent abundance
12	12.000000	98.90
13	13.003355	1.10

To calculate the average atomic weight, each exact atomic weight is multiplied by its percent abundance (expressed as a decimal). Then, add the results together and round off to an appropriate number of significant figures.

This is the solution for carbon:

$$(12.000000)\ (0.9890) + (13.003355)\ (0.0110) = 12.011 \text{ amu}$$

Ground vs. Excited state for e

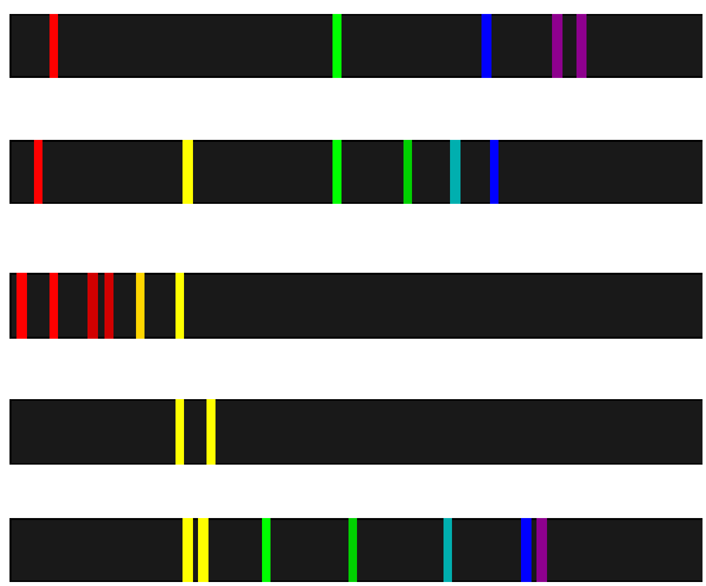
Although Neils Bohr put the electrons into orbits, which was wrong, he did put them into energy levels. These levels are exact, and the electrons tend to stay in the lowest energy levels possible. That means that the orbitals (his orbits) fill up from the inside to the out, and don't fill in the outer orbitals until the inner ones are completely filled first. That is all true. They are configured in the ground state, and all the electron configurations on your reference tables are the ground state (lowest energy state).

As electrons absorb energy they become excited and move to higher energy levels. As the electrons fall back to lower energy levels they release the energy they absorbed in set amounts called quanta. The energy that is released as electrons fall from higher to lower energy levels has a characteristic wavelength and frequency that corresponds to a particular type of electromagnetic radiation. For example, when electrons fall from a higher energy level down to the 2nd energy level, the wavelength and frequency of the energy produced correspond to that of visible light. Electrons of atoms can be excited in various ways including heat, electricity and friction.

When the atoms/electrons gain this energy, the electrons get excited and are now NOT in the lowest energy levels or configurations that they usually are in.
For example, neon usually has a 2-8 electron configuration in the ground state.

An excited state for neon could be 2-7-1.

The excited state is unavoidable with the extra energy, but a bit unstable. So when the energy is given back to the universe the electron can go back to the ground state.



This energy gain, due to electricity or heat, is then emitted, and some of it is given off as visible light energy. This light creates colored flames in a flame test.

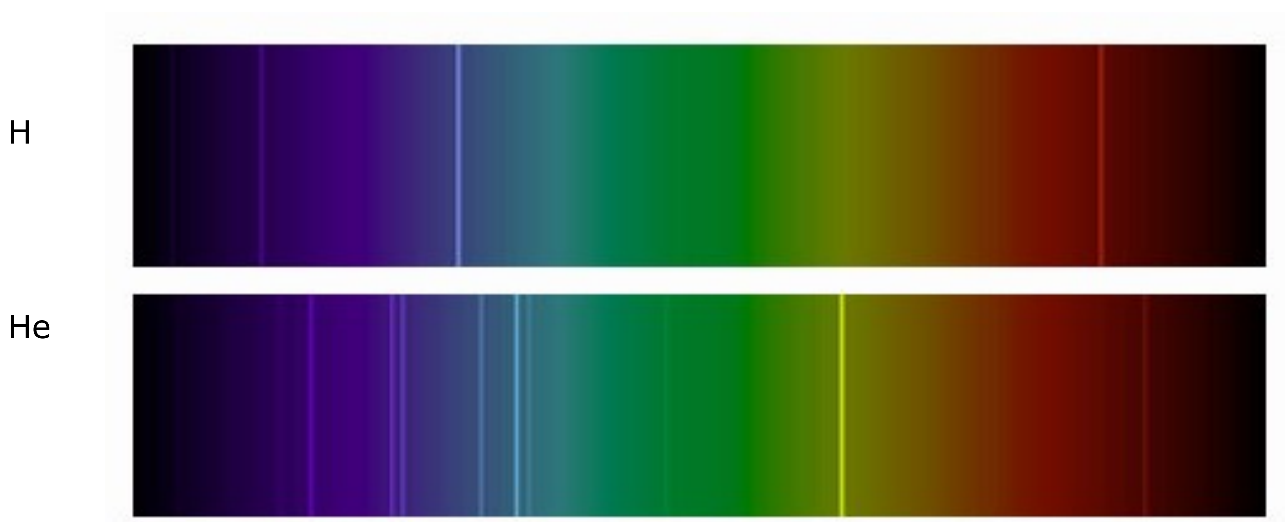
This "color" is really a mixture of colors that our eyes register as one color. This mixture can be broken apart with special glasses or lenses called refractive lenses.

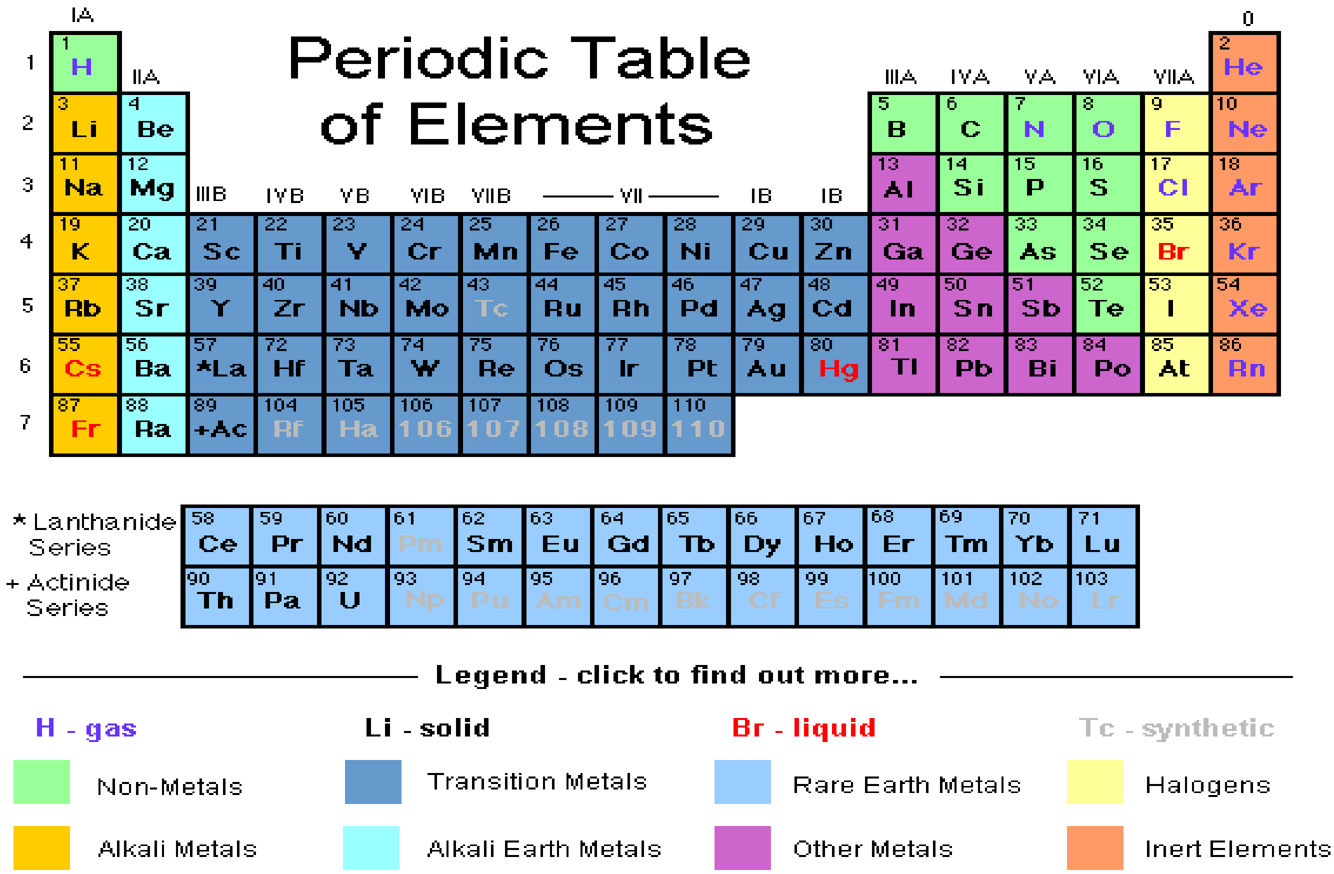
If you break up the mixture of light into component colors with a refractive lens, you see a unique pattern of bright color lines. These lines are the exact wavelengths of energy that are being emitted (looks like one color to the eyes) and can be measured. Each pattern, or spectra, is unique to the atom or molecule due to particular electron movements. The colors are due to the electrons moving from an excited state to the ground state.

Each element or compound has unique spectra that can be broken into a color pattern that is unique for that particular
substance.

Uses for this technique include determining what elements and compounds are found on distant planets and stars. If you were to photograph the light from a star through a telescope, the many mixed up elements and compounds would produce a virtual smear of colored lines as all the spectra would be mixed up together. By comparing known spectra, one atom and one molecule at a time, if the spectra line up with the star spectra lines, you know that a particular compound or element exists on that star.

Below are spectra for H and He. Similar but clearly different.





The Periodic Table

The periodic table is organized into groups which go up and down, and are labeled differently here than in your reference table. This is the "old fashioned" style. We'll stick to the groups 1-18, but this color coded chart is important and still nearly correct.

Group 1 (IA here) are the alkali metals.
Group 2 (IIA) are called the alkaline earth metals.
Groups 3-12 are the transitional metals.

Group 18 (0 here) are the noble gases, not all of which are inert.
Group 17 (VIIA) are the halogens.
The greens are the non-metals and include hydrogen at top left.
The powder blues (called rare earth here) are the inner transitional metals.

Along the "staircase" between the other metals in purple and the non-metals, include 7 elements called the metalloids. These seven have properties that cross over, metals with some non-metallic property, or non-metals with some metallic properties.

Metalloids are: B, Si, Ge, As, Sb, Te, and At.