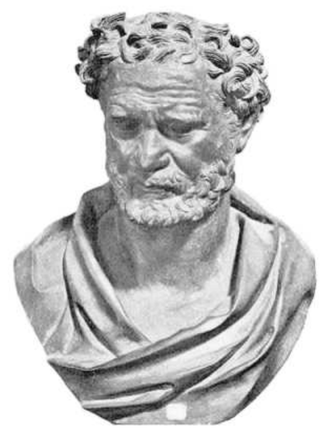
Atomic Basics

These are our most recent objectives

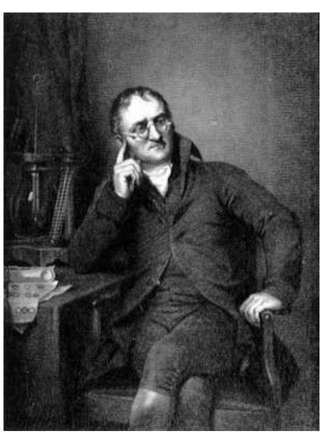
1. Models and theories of the atom
2. Subatomic particles: protons, neutrons, and electrons as well as their masses, their charges, and their locations
3. The Gold foil experiment by Ernest Rutherford
4. Determining numbers of protons, neutrons and electrons in an atom using the Periodic Table of Elements
5. Isotopes & calculating the average atomic mass as shown on the Periodic Table
6. Spectra, how and why they are produced with electron movement
7. Ground state vs. excited state for electrons, electron orbitals/energy levels

The atom has been thought of for thousands of years. Here are some of the highlights of atomic models.

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. That is what he called the "atomos", which means indivisible. Not too bad considering science was not even invented yet. His nick-name of “atom” has stuck to this day.



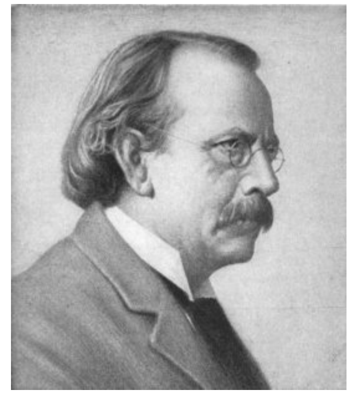
John Dalton was a farmer who meddled in science, and he used gases to examine atoms. He thought atoms were like billiard balls, small hard spheres, unique for each element. They differed only by mass, which somehow accounted for all the property differences between atoms. He published his four part Atomic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



A. all matter is made up of extremely small particles called atoms
B. elements are made up of only one kind of atom, each identical to the others in properties and mass
C. two or more atoms can combine in small, whole number ratios, to form compounds
D. in a chemical reaction, atoms are re-arranged (combined or separated) - but not destroyed

Dalton's theory is still most excellent, but we now know that atoms are in fact made up of sub-atomic particles called neutrons, protons, and electrons. Although every element is made up of atoms that are chemically identical, **isotopes** exist. (more on those later).

Still, he was clearly on the right track.



<u>JJ Thompson</u> was the person 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. A small oops, but hey, he discovered the electron and that was a great achievement!

<u>Ernest Rutherford</u> (is one of my scientific heroes) furthered atomic theory along with an experiment that I love to draw on the board and talk about: 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 protons were in the nucleus and said that this nucleus was relatively small compared to the size of the whole atom.

Although he did not understand about orbits for the electrons (or ORBITALS either), this was a most grand development in the history of the atom.

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

This problem was to be solved by a better mathematician.

Rutherford's Gold Foil experiment is one of the most important experiments in the history of science, it’s really that important. Here’s what he did:

First he pounded the heck out of some gold, so thin, maybe only 25 atoms thick. He put it in a frame to hold it upright. Next, he put some radioactive polonium into a lead box. The polonium emitted particles of radiation called alpha particles, which had a +2 charge. The box had a thin hole drilled into one side, so that the alpha particles could be aimed, only at the gold foil sheet. He surrounded the foil with a detecting screen that was coated with zinc sulfide. If an alpha particle hit the screen he could see (and count how many) little dots of light were created.

When he shot the alpha particles at the gold foil, he was amazed with what he discovered.

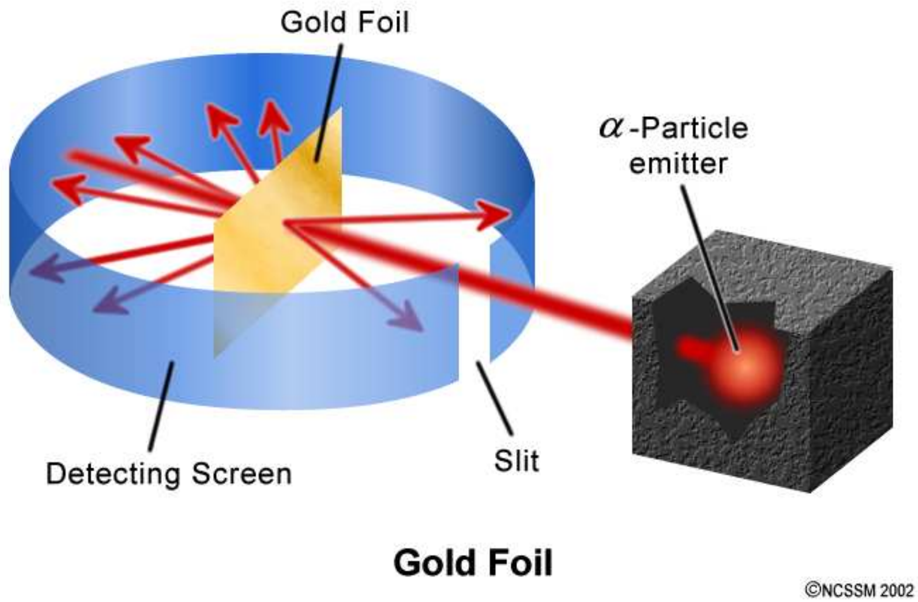
Almost all of the alpha particles zoomed through the gold, like it wasn’t there. Some alpha particles seemed to ding off at angles, and once in a while an alpha particle would bounce back towards the source! How to explain this data proved difficult.

Gold Foil Experiment set up:

You should be able to draw this, label this, list the three ideas that it gave Rutherford to describe his model of the atom, and the three problems associated with his model that he could not make sense of.

Soon: get some white paper and try it. Drawing plus 3 ideas and 3 problems.

Start soon.



Rutherford concluded the following:

1. The (gold) atom was mostly empty space -most of the alpha particles missed completely.

2. The (gold) atoms had a dense and positively charged nucleus (the positive alpha particles didn't stick, and he knew nothing of protons or neutrons so it was just positive stuff).

3. He knew the atoms were neutral and somehow contained electrons(from Thomson) so he postulated that the atoms had electrons flying around outside this nucleus, creating a large but mostly empty atom.

This was all exactly correct, but no one believed him!

Major Problems with the Rutherford model of the atom

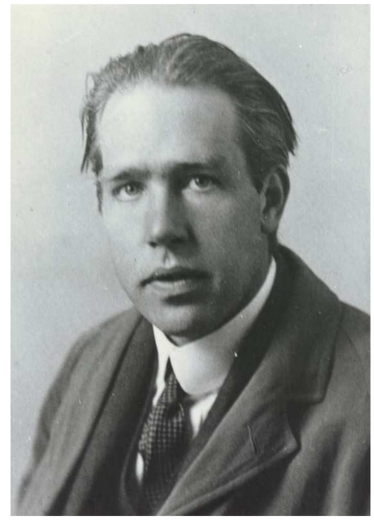
A. How could atoms be 99% not really there? Why couldn't we just walk through walls (or at least see through them)?

B. Why did these flying electrons never get tired (lose their kinetic energy) then slow down, and get sucked into the nucleus due to opposite charges?

C. Finally, why did these electrons flying around so relatively far from the nucleus not just fly away?

It was due to the Nobel Prize winning work of his student Niels Bohr who saved the Rutherford model (which was mostly correct)

Niels Bohr saves his teacher Ernest Rutherford's Model of the atom

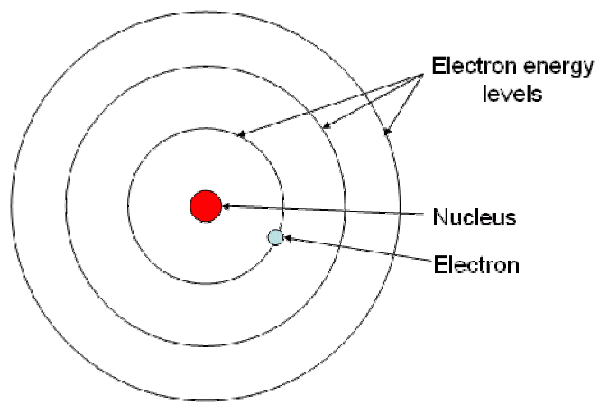


Bohr is able to prove (mathematically) that electrons do not lose energy if they can stay in precise orbits around the nucleus. Each of these orbits has an exact energy level associated with it. Each atom has unique electron orbits, and each electron has a unique energy associated with it. Finally, given precisely the right amount of energy (the quantum required) electrons could jump up to higher orbits (or higher energy levels).

His orbits are energy levels. The closer to the nucleus, the smaller the orbit, and the lower energy they are. The further an orbit is from the nucleus the higher the energy levels the electrons in them have.

Neils' math works great for the atom hydrogen with a single electron to cope with, but as soon as he works on helium with just 2 electrons, or any other atoms, the math is too hard to do. Still, his ideas stand, and the electrons are drawn into nice little planetary diagrams, with the electrons filling up these orbits from the "lowest" energy orbit, to the higher energy orbits. Below is the "Bohr Model".

The 1st orbit can fit just 2 electrons.
The 2nd orbit holds up to 8 more electrons.
The 3rd orbit can fill up with 8 electrons,
    or it can STRETCH to hold up to 18 (that's tricky).



To remember how many electrons fill up any orbital, just look at Group 18 on your Periodic Table. Those noble gases always have COMPLETE electron orbits or ENERGY LEVELS.

We no longer think of electrons in "orbits", but they are still in energy levels known as <u>orbitals</u>.

If an atom (an it's electrons) are given enough energy, the electrons can jump up to higher than normal electron orbits (orbitals now). When this happens, the electrons are not in the ground state (normal, lowest energy) but are in the excited state (higher energy than normal and not as stable).

When the electrons drop back down to the normal ground state, them EMIT the energy gained as visible light we can see, called spectra.

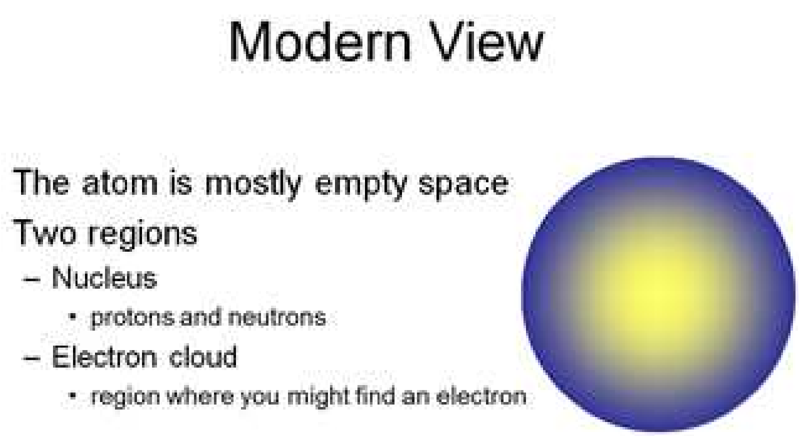
Modern Model

Finally, in the early part of the 20th century, as the math gets fancier and quantum theory becomes the rage, the atom is again reconfigured. The modern model, or the wave-mechanical model, we find the nucleus still central, neutrons are already discovered, the charges of the neutrons is 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. It seems that they are not as easy to grasp as the planetary model, but that is just how it is.



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.

ORBITALS are more like "zones" that electrons live in, and they have remarkably complex shapes. Atoms cannot just have energy, rather they 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, that is not for us.

The Three Sub-Atomic Particles:

As far as the sub-atomic particles, we need to know about the three biggies: protons, neutrons, and electrons. Protons have a +1 charge, they have a mass of 1 amu, they are in the nucleus only.

Neutrons have no charge, they too have a mass of 1 amu, and they are in the nucleus only as well. Electrons have a -1 charge, and they 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 accept is pretty close to nothing anyway, only about 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 and also the number of electrons in 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 total mass is the protons plus the neutrons. If you know this mass, 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. YOU NEED TO GRASP the concepts below concerning average atomic masses.

Titanium is shown at right. It's mass is rounded to 48 amu. That is 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 that out now.

Titanium's atomic number of 22 tells us that it has 22 protons as well as 22 electrons.

48	=	protons plus neutrons
- 22	=	protons (from atomic number)
26	=	neutrons

47.88

Ti

22

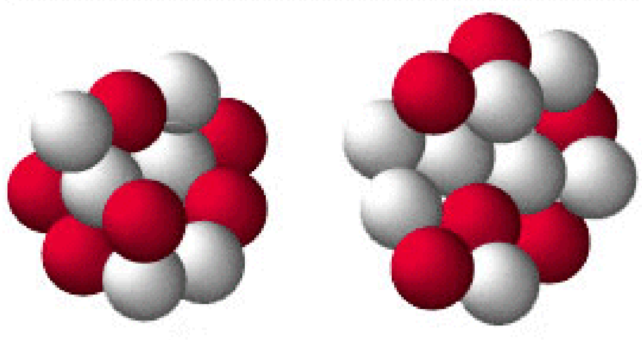
2-8-10-2

Isotopes

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.

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



2 Isotopes of Carbon
C-12 and C-14

At right are the nuclei of the 2 isotopes of carbon. Both have 6 red protons. At left there are six white neutrons, at right there are eight white neutrons. Both would have six electrons. Both react identically in a chemical way, 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. All carbon, for example, is made up of 15 different isotopes, all of various proportion. We have seen only the two most common on the previous page. All of the "kinds" or isotopes are in certain proportions.

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.

Ground vs. Excited state for e^-, then, Bright Line Emission Spectra

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).

Now, if you had some atoms and could give them some energy, say by shocking them with electricity, or heating them up in a fire, they would be able to gain some of that energy. The electrons would gain energy and have enough to "jump" up to higher energy orbitals. They would get excited, and find themselves in what's called the "excited state". We'll do this in lab, with electricity and with fire. It makes some nice colors too!

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.

Normally that would not happen, the electrons tend to lowest energy levels. 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.

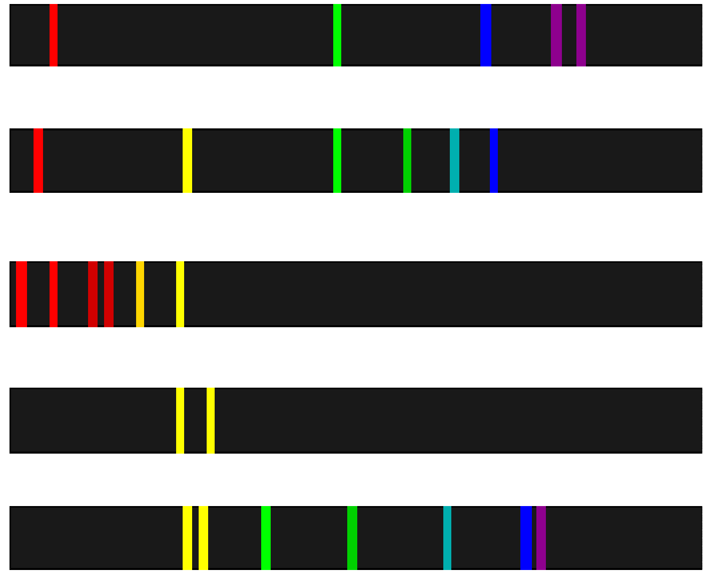
This energy gain, due to electricity or heat, is then emitted, and some of it is given off as visible light energy. This light creates colored flames in a flame test. It comes out as a orange color in a neon lamp.

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 is 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. These patterns at right show some spectra.

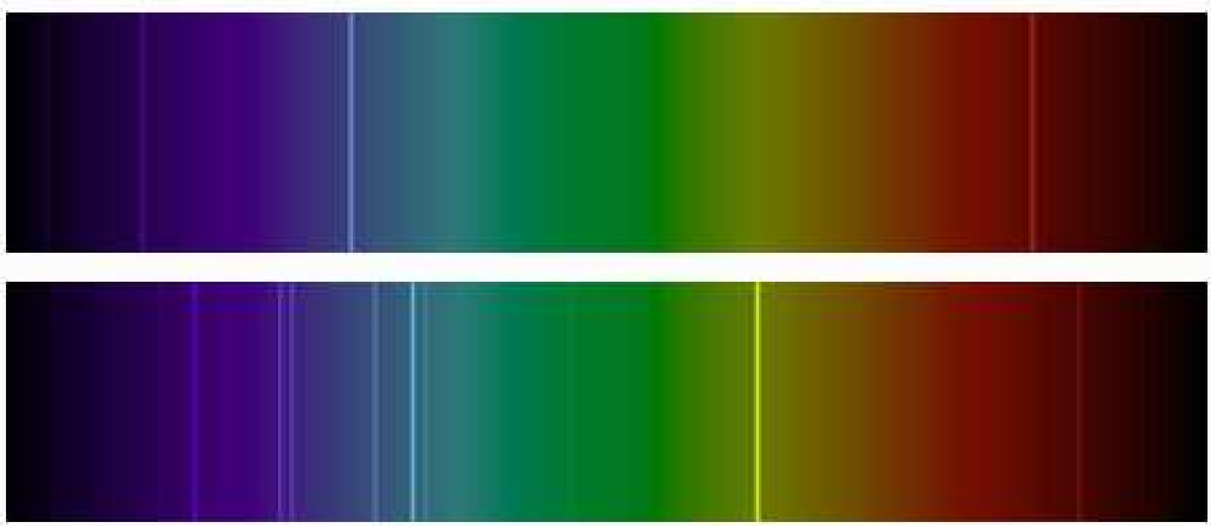
Uses for this technique include determining what elements and compounds are found on distant planets and stars. If you were to photograph the light from a star through a telescope, the many mixed up elements and compounds would produce a virtual smear of colored lines as all the spectra would be mixed 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.

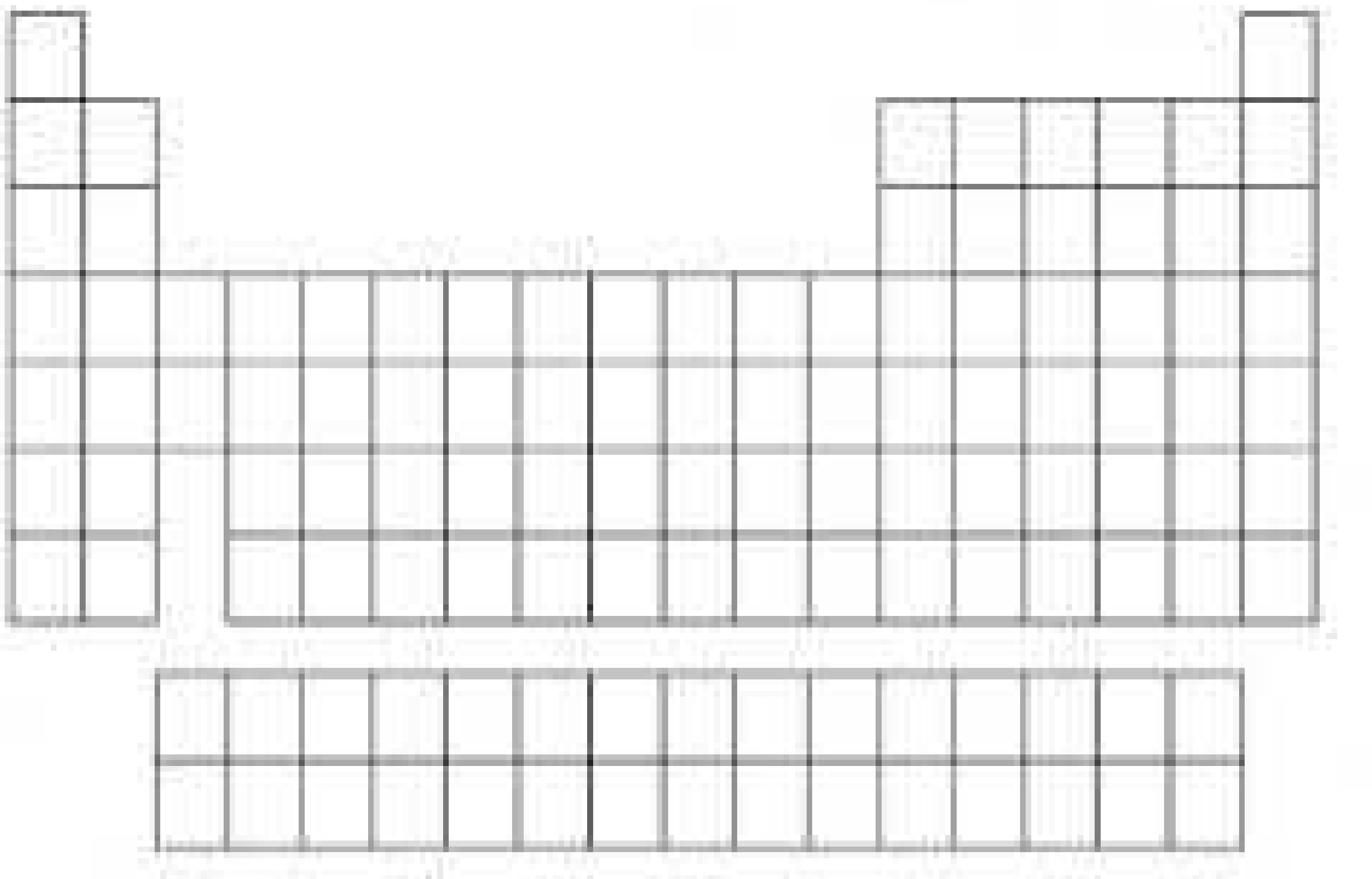


Below are spectra for H and He. Similar but clearly different.

H

He





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 are the alkali metals.
Group 2 are called the alkaline earth metals.
Groups 3-12 + the triangle of metals from Al to Tl to Po, are the transitional metals.

Group 18 are the noble gases, which are nearly inert.
Group 17 are the halogens.
The inner transitional metals are the two rows at the very bottom that fit into Group 3 under Y.
The staircase divides metals on the left, nonmetals on right (hydrogen is the exception)

9 atoms touch the stairs, but only 7 are metalloids. Al + Po are the “dog food” exception (ask!)

A metalloid is a metal with some nonmetallic properties (odd), or a nonmetal with some metallic properties (also odd). They sometimes are called semi-metals, but not in our class.

The 7 Metalloids are: B, Si, Ge, As, Sb, Te, and At.
On your table all elements that touch the darkened stairs (except Al, Po) are metalloids.