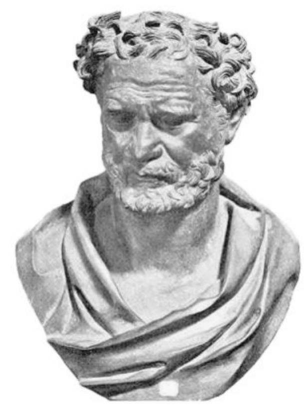
Atomic Basics

These are our atomic theory 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 modern model of the atom has evolved over a long period of time, through the work of many scientists.

Here are some of the highlights of atomic models through history.

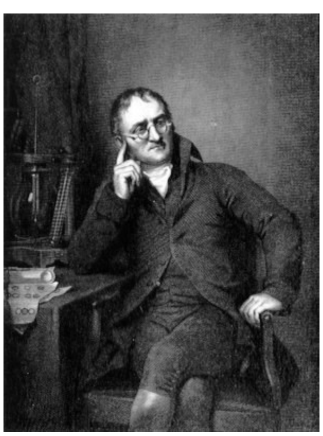


Democritus was a philosopher in ancient Greece who "thought" about things, and came up with his ideas. His idea was that all kinds of matter were unique and that you could cut it in half over and over until you reached some tiny part that could no longer be cut in half anymore.

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.

Unfortunately this meant you could have an "atom of fish" or an "atom of wood"! Still, he did come up with the name to that ultimately small particle.

John Dalton was a farmer who examined atoms, and "invented" modern chemistry. He noticed all atoms of an element has unique mass. He decided that ALL properties of elements were due to this difference in mass. That was not the reason really. He imagined atoms to be like BILLIARD Balls, which are like pool balls on a pool table, small, hard, spheres of different mass.



He published his 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

Much of Dalton's theory was correct, but we now know that atoms are made up of sub-atomic particles called neutrons, protons, and electrons. Although every element is made up of atoms that are chemically identical, they are not identical, scientists know that **isotopes** of every atom exist. (more on those later). Still, he was clearly on the right track.

JJ Thompson was the person who discovered the electron. He used a device called the cathode ray tube, and was able to find electrons.

He had no knowledge of protons or neutrons, or real atomic structure, so he imagined these electrons "stuck" into a sort of positively charged atomic stuff. He called it his "plum pudding" model of an atom.

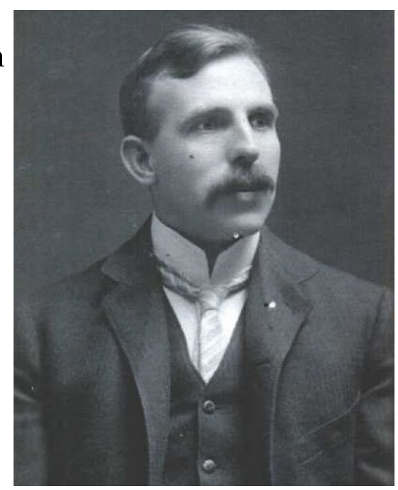
Try to imagine that electrons are the chips in a chocolate chip cookie, and the rest of the atom (the cookie part) is all positively charged, enough to cancel out the negatively charged electrons. Plum pudding is ripped up bread, with chopped up bits of plum, mixed with egg and lots of sugar, baked together and forming into a sweet loaf of bready dessert. It's heavy, but served warm with cream on top, it's quite tasty, but a silly model for an atom!

A small oops, but hey, he discovered the electron and that was a great achievement!

Ernest Rutherford (is one of my scientific heroes) furthered atomic theory along with an experiment that I love to 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 discovered that the nucleus was positive, and the atom was neutral because of the electrons flying around it.

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

Unfortunately a big problem he could not explain were how the negatively charged electrons didn't just collapse into the positive charged nucleus. You would think that they would. In addition, the electrons didn't fly off either, nor did they run out of energy. This problem was to be solved by Neils Bohr.



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:

He knew he could pound gold super thin, it's the most MALLEABLE metal of them all. He pounded it so thin, approximately 1/1000th the thickness of a sheet of paper!

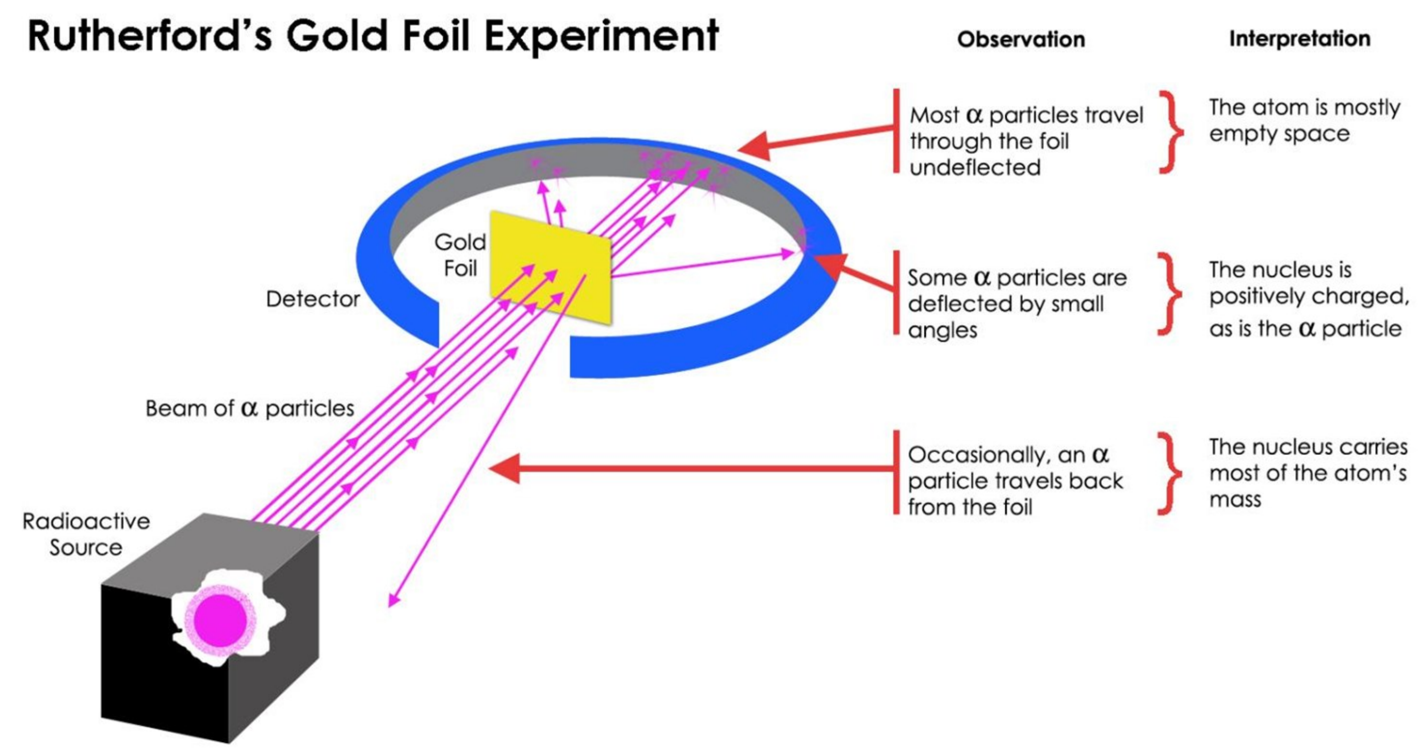
He was told, by Marie Curie, that radioactive Polonium emitted a kind of radiation called ALPHA PARTICLES, that were little and which had a charge of +2. He would "shoot" this radiation at the gold and see what it did.

He put it into a lead box (to keep the radiation from scattering) and drill a small hole into the box so that the radiation came out only through this tiny hole, and he could aim the radiation particles at the gold.

He surrounded his experiment with a zinc sulfide coated frame. It turns out whenever an alpha particle hits this zinc sulfide screen, a small flash of light can be seen (and counted).

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

Look hard at the diagram on the next page, you need to know it, be able to draw it, and explain it .



Major Problems with the Rutherford model of the atom

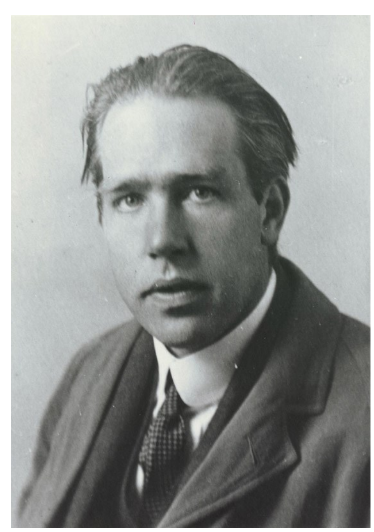
A. How could atoms be 99% not really there? Why couldn't we just walk through walls (or at least see through them)?

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?

Although Rutherford was correct in his interpretation of what atoms looked like, it was due to the Nobel Prize winning work of his student Niels Bohr who saved the Rutherford model.

Niels Bohr and the Planetary Model of the Atom



Bohr was a physicist (explaining the lack of a mustache!). He wanted to study the structure of the atom. He knew it wasn't a billiard ball, or a plum pudding dessert. He understood the golf foil experiment, and was smart enough to do math to "explain" Rutherford's results.

Normal people can't always make sense of the math, it is counter intuitive.

Atoms could be 99% "not there" and still be there. Electrons could fly around the nucleus and never lose energy, never fly off, nor collapse into the positively charged nucleus. He could do the math the "prove" it, even if you couldn't understand the math, he could.

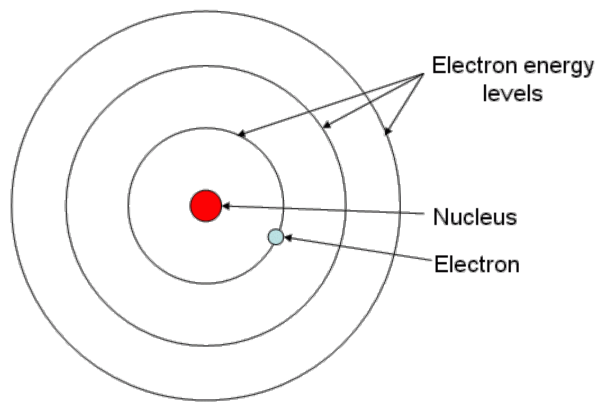
Using the atom hydrogen, which has just one proton and one electron, he was able prove mathematically that electrons do not lose energy if they can stay in precise orbits around the nucleus. Each of these orbits is also an energy level. with it. That meant the electrons could fly around their nuclei forever. Further, each electron lived in specific energy levels or orbits (like planets orbiting the Sun), and each orbit had unique energy levels to it.

If an electron gained a specific amount of energy (a quantum of energy) the electron could "jump up" to a higher than normal orbit for a while, which he called the EXCITED STATE. This was unstable, and the excited electrons would then "jump back to the lower energy level, the normal or GROUND STATE. To return, they'd give off that unique quantum of energy, which we can see as visible light. Since each atom (or compound) requires unique energy to become excited, the electrons return that quantum of energy as visible light, each substance produces a unique visible light which we call spectra, and which we will see in lab soon.

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, even for him, 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 ONLY COMPLETE electron orbits or ENERGY LEVELS.

The Modern Model or the Wave-Mecahanical 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.

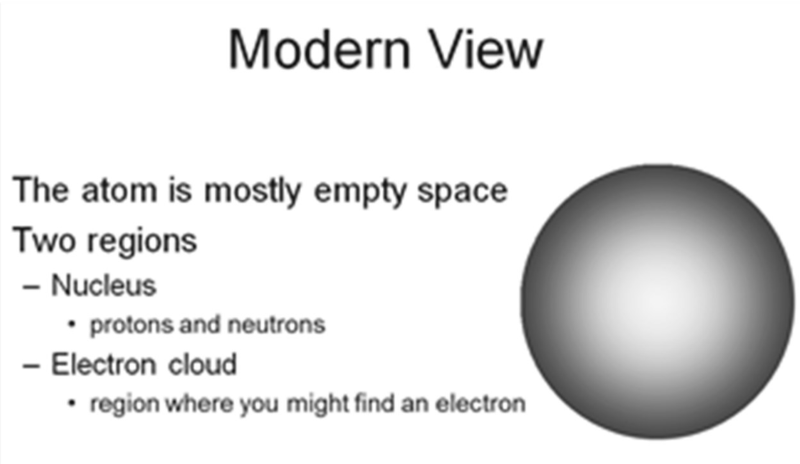
Protons are still positive and exactly balanced by the negative charges of the electrons flying about.

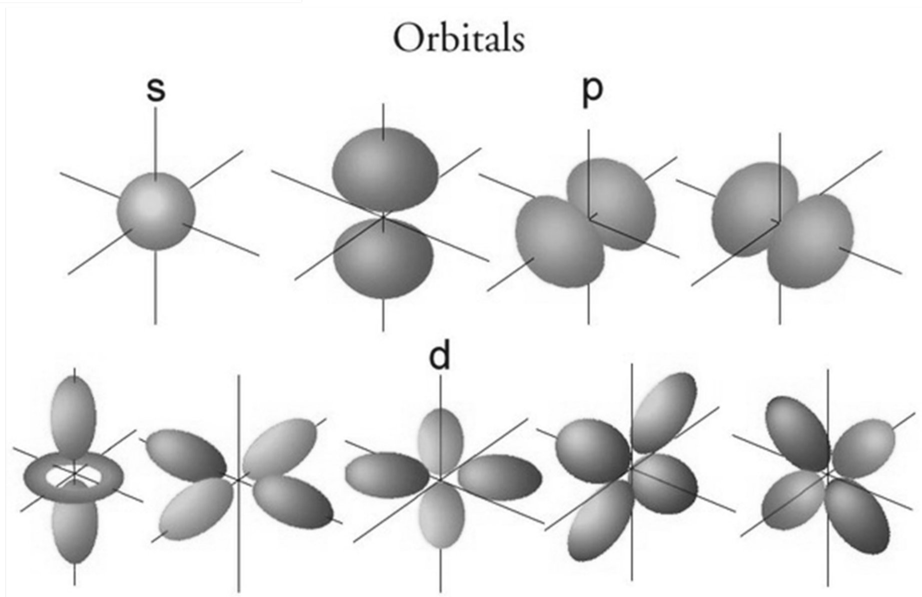
All atoms are still neutral, but the electrons no longer follow in neat little circles like in Bohr's time. Now these electrons are moving about in a sort of statistical cloud, a zone, called an ORBITAL.

Orbits were simple 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. They are actually pretty funky and weird. 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 3 biggies: protons, neutrons & 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 as pretty close to nothing in high school. They are 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 = the mass of the nucleus, or the protons plus 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 (and 22 e^-)

48 = protons plus neutrons
<u>- 22</u> = minus the protons (atomic number)
26 = neutrons

47.88
Ti
22
2-8-10-2

63.546
Cu
29
2-8-18-1

Copper….

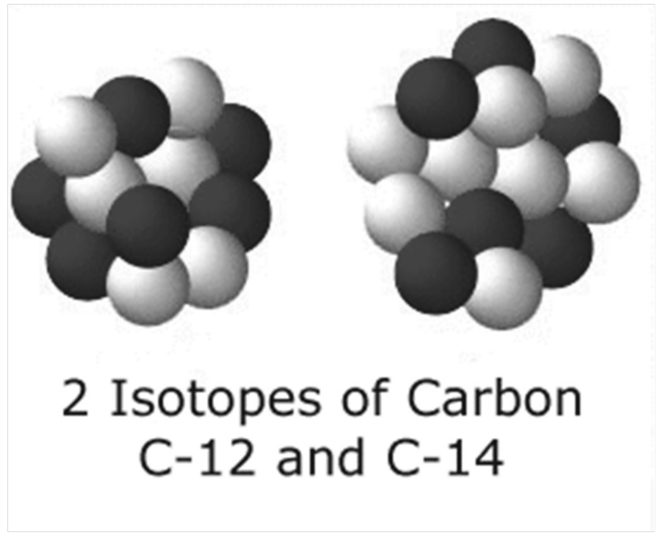
Mass = 64 (The protons plus the neutrons)
Minus <u>- 29</u> (the atomic number or # protons)

35 = (the number of neutrons)

Tungsten - W mass 183.64 amu. Atomic number 74. 184 - 74 = 110 neutrons 74 protons & 74 electrons

Iron - Fe mass 55.845 amu. Atomic number 26. 56 - 26 = 30 neutrons 26 protons & 26 electrons

Isotopes



Not all atoms of one element are identical, John Dalton said they were but he was incorrect.

It turns out that all atoms of one element are CHEMICALLY IDENTICAL because they have the same numbers of protons and electrons, but they have DIFFERENT numbers of neutrons.

He didn't know about protons, neutrons or electrons.

Neutrons do not affect the chemistry of the atoms, just the mass.

The number of neutrons can vary, and each isotope of an element has a different number of neutrons, and therefore a different mass. They all bond and react identically, but the neutron number is different.

At right are the nuclei of the 2 isotopes of carbon. Carbon-12 isotope has 6 dark colored protons and six white neutrons (6 + 6 = 12 amu)

The Carbon-14 isotope on the right has 6 dark colored protons and EIGHT white neutrons (6 + 8 = 14 amu)

Both have six electrons. They are chemically the same, but they 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.

Out of the 118 types of elements, there are almost 1500 different kinds of isotopes. Every atom has at least 2 isotopes, some have up to 15 or more isotopes. Why do these exist? No one really knows, but scientists like to measure everything, so we will learn this too.

For example, CARBON has 3 naturally occurring isotopes. They are C-12 which is the common and regular carbon. Also Carbon-14 which is radioactive. There is also some Carbon-13 which is also stable but less common. Each has a whole number mass in high school chem, 12 amu, 13 amu, or 14 amu. It's the PROPORTIONS of them that are measured carefully, to many decimal places, that yields the decimals on the periodic table masses.

To calculate the average weighted atomic mass, on the Periodic table, scientists multiply the mass by it's proportion, for each isotope, then sum the masses to get the weighted average of the atom's mass.

Isotope	Mass	Approx. Proportion	math	Partial mass
C-12	12 amu	98.80%	(12 amu)(.9880) =	11.854 amu
C-13	13 amu	1.100 %	(13amu)(.01100) =	0.1430 amu
C-14	14 amu	0.100 %	(14amu)(.00100) =	0.01400 amu
All		100.0 %	Avg. Wt. At. Mass =	12.011 amu

For an “unknown” element with just two isotopes, calculate the average weighted atomic mass. When you are given “better” measurements, as you are here, use them instead of the casual masses of the isotope symbols.

Isotope	Mass	Approx. Proportion	math	Partial mass
X - 75	74.86 amu	89.35 %	(74.86 amu)(.8935) =	66.89 amu
X - 77	77.06 amu	10.65 %	(77.06 amu)(.1065) =	8.210 amu
All		100.0 %	Avg. Wt. At. Mass = 75.10 amu	

It makes perfect sense that when we would “round” the average weighted atomic mass to 75 amu, that the most common isotope of this element is X - 75. That always happens on the periodic table.

The most common isotope for mercury is 201 amu (mass = 200.59 amu)
And the most common isotope for helium is 4 amu (mass = 4.00260 amu)

Electron Configurations

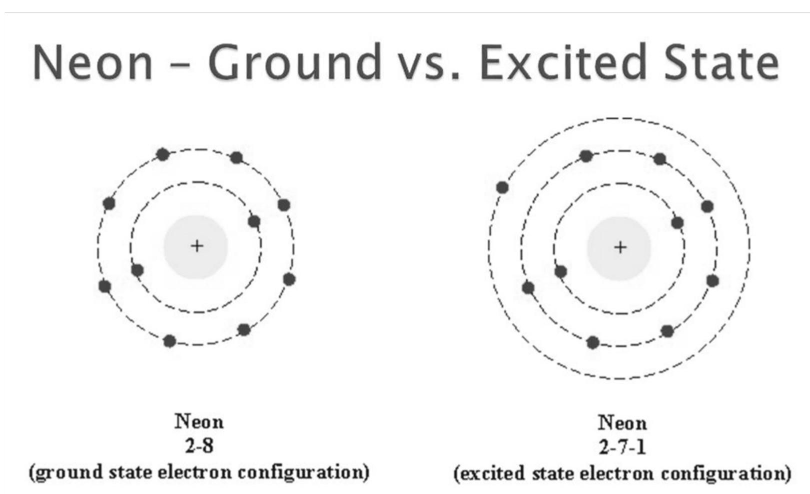
On the periodic table the electron configurations provided are always in the ground state, or lowest energy state. There are plenty of sub-orbitals that we don’t learn about, but this simple (up to 7 orbital system is pretty simple in our class. Electrons always fill up the atoms from lowest energy levels (that means lowest orbitals) first, then fill higher orbitals.

In college you will learn the more complex (not harder) orbital and sub orbital system, and this will all make even more sense to you.

In the first orbital there only 2 electrons can fit. In the second orbital up to 8 electrons fit. The rest of the orbitals are weird. They can “fill up” with 8 electrons, or stretch out and fill with more electrons. How will you ever remember all of this? Look at group 18 on the periodic table, the noble gases. They have ONLY full orbitals. So you can see (example) argon has 3 full orbitals. 2 in the first, 8 in the second and 8 more in the third orbital. But krypton has a different pattern. 2 in the first orbital, 8 in the second, 18 (!) in the third orbital and finally 8 in the fourth orbital.

The third orbital can be “full up” with 8 electrons, or with 18 electrons!

All the atoms on the periodic table have the electrons in the ground state.
They are all in their lowest energy state. This is normal. But atoms (and electrons) can gain energy by electricity, or heat, or even by radiation, and the electrons can gain enough energy to “jump up” to higher than normal electron configurations. This is called being in the excited state.



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

Carbon's electron configuration in the ground state is 2-4. A possible excited state is 2-3-1 where one electron in the second orbital gains sufficient energy go move up to the new third orbital.

Sodium's electron configuration in the ground state is 2-8-1. A possible excited state is 2-7-2.

Argon's electron configuration in the ground state is 2-8-8. A possible excited state is 2-8-7-1 OR 2-7-9.

When electrons are in this excited state they are holding onto some extra energy. It takes a unique amount of energy for any atom to get excited, they each absorb their own "quanta" amount of energy.

The excited state is unstable, and the electrons would "rather" revert back to the ground state. They need to emit this exact quanta of energy do return to the ground state. This unique amount of energy can be seen by your eyes as visible light.

Energy can be absorbed as electrical, heat, or even radiation (not in our class though), but it is emitted as visible light which we call spectra.

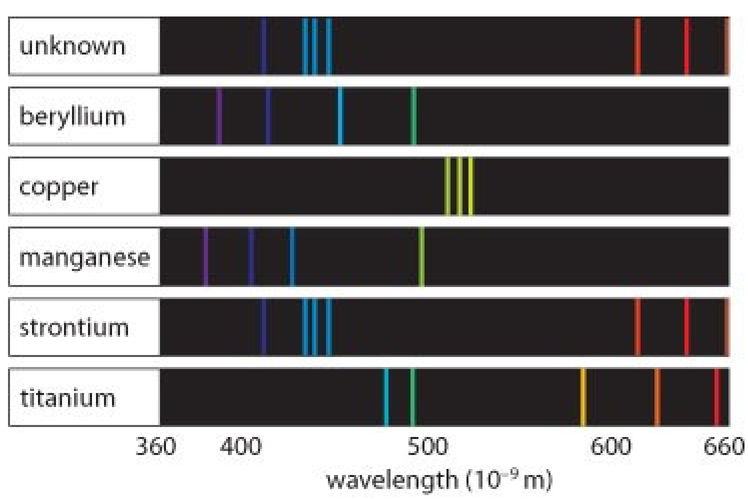
In lab we will see that this energy gain, due to electricity or heat, is then emitted, and it is given off as visible light energy. This light creates colored flames in a flame test. It comes out as a single color of orange color in a neon lamp.

This "color" flame, or color lamp light 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.

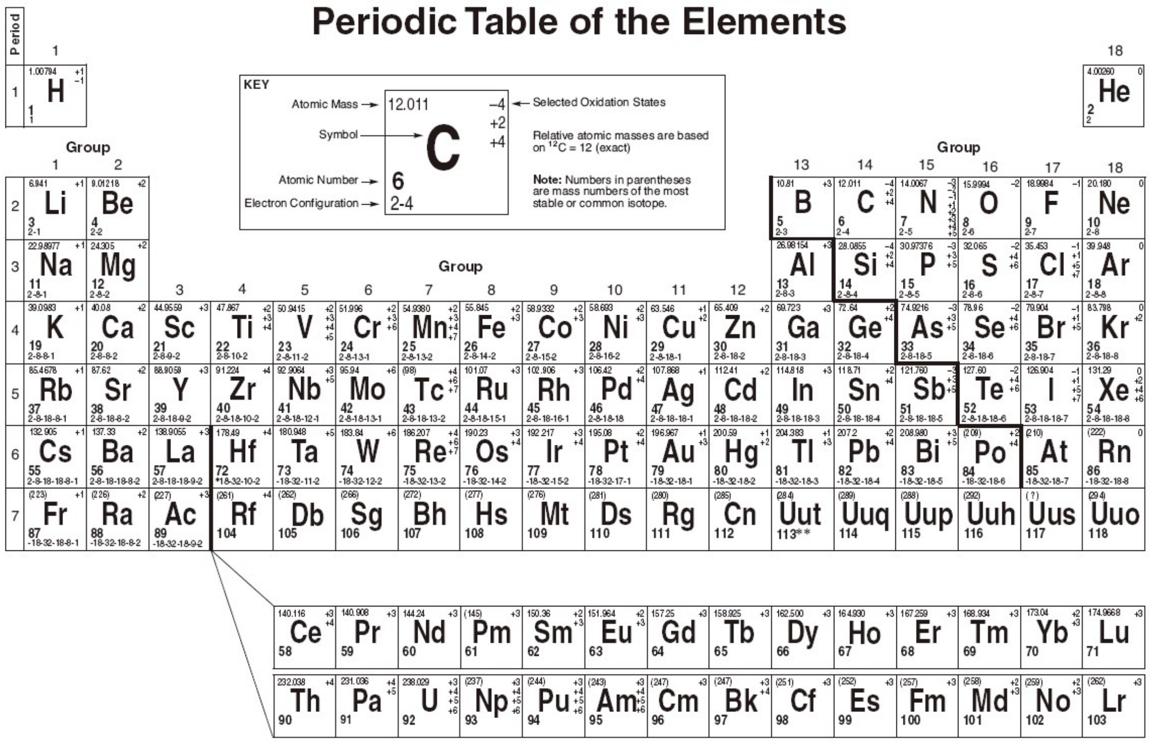
Each element or compound has unique spectra that can be broken into a color pattern that is unique for that particular substance.
These patterns at right show some spectra.

The top is the unknown spectra.
Below we are comparing four known spectra to it.
What is the unknown substance (which of the known spectra matches this unknown?)



It's strontium, of course.

Uses for this technique include determining what elements and compounds are found on distant planets and stars. Or discovering any unknown substance at hand, by comparing the unknown to known spectra.



*denotes the presence of (2-8-) for elements 72 and above

**The systematic names and symbols for elements of atomic numbers 113 and above will be used until the approval of trivial names by IUPAC.

Source: *CRC Handbook of Chemistry and Physics*, 91st ed., 2010–2011, CRC Press

The Periodic Table
The periodic table is organized into groups which go up and down, and need to be labeled 1-18.
The periods run left to right, and need to be labeled 1-7.

Group 1 are the alkali metals. Lithium to Francium.
Group 2 are called the alkaline earth metals. Beryllium to Radium.
Groups 3-12 and the "triangle" of metals from Al to Tl to Po, are the Transitional metals.
Group 17 are the halogens. Fluorine to Iodine (not At)
Group 18 are the noble gases, which are nearly inert. Helium to Radon (not Uuo)
The inner transitional metals are the two rows at the very bottom that fit into Group 3, under Yttrium. They include both La and Ac. In group 3, period 6 fit La to Lu. In group 3 period 7 fit Ac to Lr).

The staircase divides metals on the left, nonmetals on top right (hydrogen is the exception)

9 atoms touch the stairs, but only 7 are metalloids, they are B, Si, Ge, As, Sb, Te, and At.
Al + Po are the "dog food" exception. Aluminum and polonium are metals,
but they do touch the staircase.
Al and Po remind your teacher of AlPo Dog Food!

A metalloid is a metal with some nonmetallic properties, or a nonmetal with
some metallic properties. They sometimes are called semi-metals.

All metals are to the left and below the stairs, including the inner transitional metals

All nonmetals are to top right corner, to the right of the stairs.

Only hydrogen is "misplaced", it is a nonmetal, but it is in group 1.
We'll see why another day.