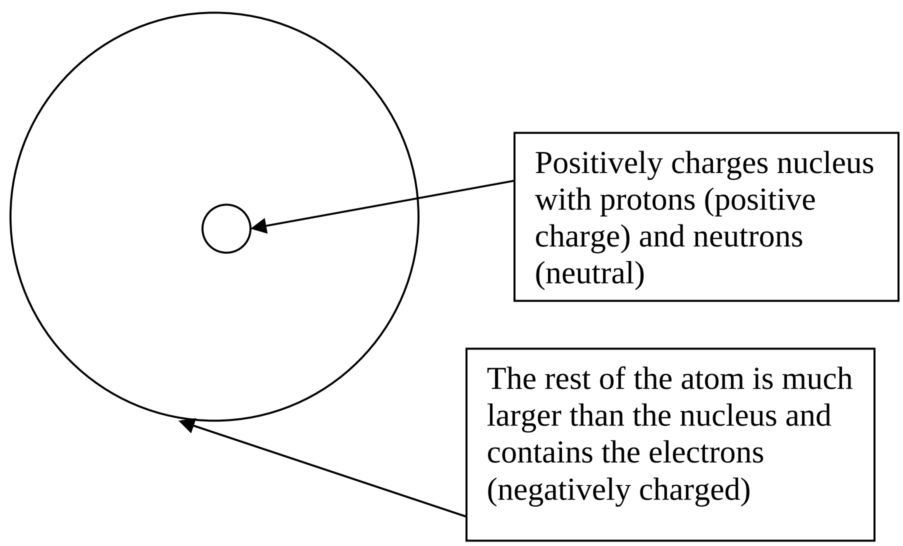
Chapter 10 reexamines the structure of the atom in more detail.

** Electron arrangement in atoms
** How electron arrangement influences properties of the elements

Chapter 10, Section 1

Rutherford's Model of the Atom

Remember from Chapter 4, Ernest Rutherford's Atomic Model.



The problem with Rutherford's Model of the atom is that it does not explain several important items:

1. How are the electrons arranged?
2. How are the electrons moving?

Rutherford guessed that the electrons revolve around the nucleus like the planets around the sun, but this was not supported scientifically.

In order to understand the structure of the atom more completely, a better understanding of some key concepts is necessary.

Chapter 10, Section 2

The first concept needed is the understanding of electromagnetic energy.

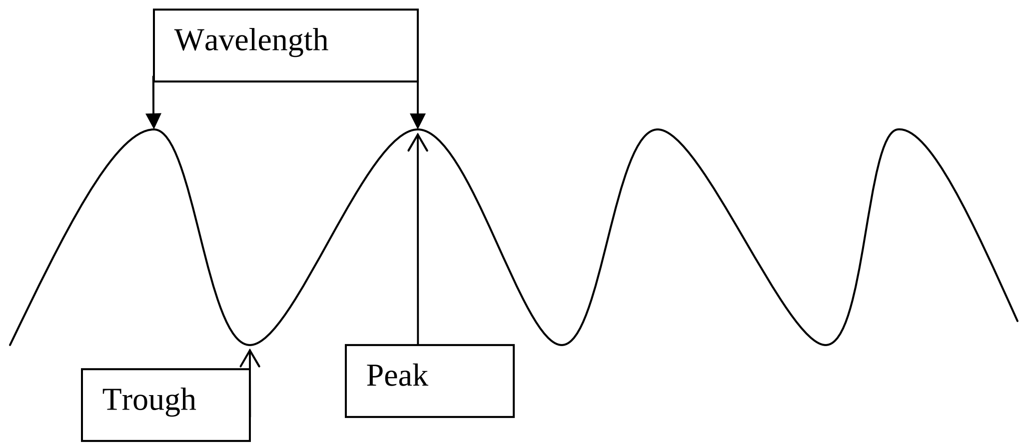
Energy being transported by light is known as ELECTROMAGNETIC ENERGY.

In order to understand electromagnetic energy, first an understanding of wave structure is necessary.

Consider the following drawing of a wave. (Picture the waves on the ocean to help)

There are three properties that are used to describe a wave.

1. Wavelength- distance between two consecutive peaks (or troughs)
2. Frequency- how many peaks (or troughs) pass by a point in a given time period
3. Speed- how fast a given peak is moving



Light (electromagnetic radiation) also travels in waves, just like water in the ocean.

The different types of electromagnetic radiation form the electromagnetic spectrum.

This can be found on PAGE 282 in the text as FIGURE 10.4

KNOW THE ELECTROMAGNETIC SPECTRUM.
Light can also behave as a stream of “photons”.

PHOTONS are “particles” of electromagnetic energy.

This “dual nature” of light, acting as both a wave and a stream of particles is referred to as the Wave-Particle Nature of Light.

RULE: Photon Energy and Wavelength

** The longer the wavelength, the lower the energy of its corresponding photons.

Chapter 10, Section 3

Check out the photo on page 284 in the textbook.

Atoms are able to be excited by heat from a flame.

"EXCITED" is a way to describe an atom with excess energy.

This excess energy can be released by giving of light in the form of a photon.

The energy of the photon directly correlates to the energy change of the atom

Remember wavelength dictates energy, so color can be used to determine relative energy.

Example:

A photon of red light carries less energy than a photon of blue energy

RULE: Light Color and Photon Energy

A particular color of light carries a particular amount of energy per photon.

Chapter 10, Section 4

Hydrogen's Energy Levels

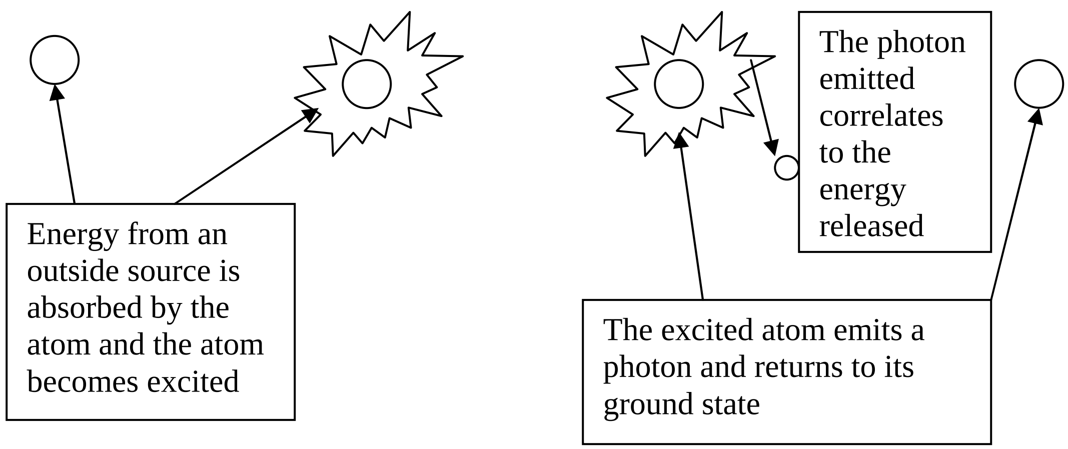
An atom with excess energy is able to release this excess energy by releasing a photon which allows the atom to move to a less excited state.

GROUND STATE is the lowest possible energy state of an atom.

Examining the photons emitted by excited hydrogen atoms allows for the study of the energy states of the hydrogen atom.

** Remember: the different wavelengths of light indicate different amounts of energy per photon!!

If a hydrogen atom is excited by absorbing energy from an outside source, it can release the excess energy and return to a less excited state by emitting a photon.



RULE: ENERGY LOSS AND PHOTONS

** The energy contained in the photon corresponds to the change in energy in the atom.

A study of very excited hydrogen atoms shows that only certain wavelengths of light can be seen.

This means that only certain photons are being emitted, which means there are specific energy changes that are occurring.

These specific energy changes indicate that the hydrogen atom has "certain discrete energy levels".

It is these specific energy levels that cause the specific colors to be seen as a hydrogen atom gives off photons to return to ground state after being in a very excited state.

See emission spectrum for hydrogen on page 286, figure 10.11.

Because there are only a few distinct colors that will be given off by hydrogen atoms, hydrogen atoms are said to have quantized energy levels.

QUANTIZED- only certain, specific values are allowed

The energy levels of all atoms are quantized.

Think of quantized energy levels as stair steps verses a ramp

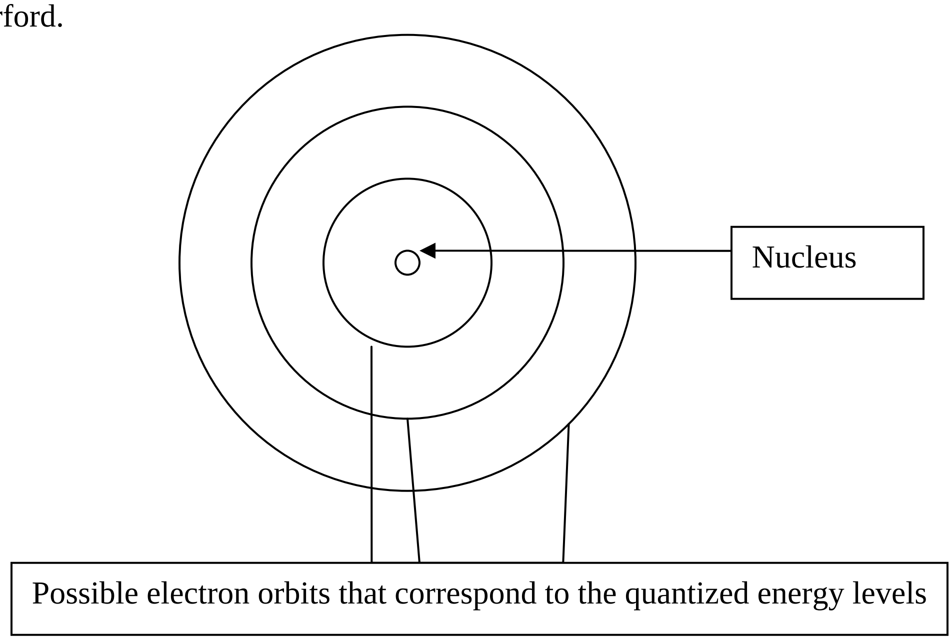
The stairs allow for specific levels, whereas the ramp allows for a continuum.

It was previously thought that the energy levels in atoms were like the ramp, that an atom could exist at any energy level. This is NOT true, the quantized energy levels, or stair step is the actual way atoms exist.

Chapter 10, Section 5

The Bohr Model of the Atom

Niels Bohr used the QUANTIZED energy level concept to modify Rutherford's model of the atom. He applied the concept of these distinct energy levels to the orbits already proposed by Rutherford.



Although initially Bohr's model seemed very correct, it only worked for hydrogen and not for other elements.

Bohr's model has since been disproved, electrons do NOT move in orbits like planets around the sun, even if the orbits correspond to specific, quantized energy levels.

Chapter 10, Section 6

The Wave Mechanical Model of the Atom

By 1920's Bohr model had been proven incorrect.

Broglie and Schrodinger suggest that electrons, like light, behave as both a wave and as a stream of particles.

This led to the Wave Mechanical Model of the Atom which was able to be applied to all elements, not just hydrogen.

The Wave Mechanical Model says that you cannot determine the exact position of an electron, nor does it allow you to predict where or how it is moving.

The Wave Mechanical Model simply predicts an area in which you are likely to find an electron.

There is a great possibility that no matter how sophisticated our science becomes, we may never be able to determine how an electron moves or exactly where it is.

Chapter 10, Section 7

The Hydrogen Orbitals

When we apply the Wave Mechanical Model to the Hydrogen atom we discover areas that we will be likely to find the electron of hydrogen.

These areas of probability are known as ORBITALS, and they represent an area for which there is a 90% probability of finding the electron.

The GROUND STATE ORBITAL of hydrogen is named 1s Orbital

Remember that hydrogen has higher energy levels for when the atom is excited. In the Wave Mechanical Model these higher energy states correspond to different shaped orbitals.

Now we need to pause and learn a bit about the different energy levels of Hydrogen (they also apply to all other elements as well)

There are four PRINCIPAL ENERGY LEVELS which are labeled with INTEGERS.

The four principal energy levels are divided into SUBLEVELS, which are labeled with the letters s, p, d, and f.

Principal Energy Level #1 has only one sublevel: s

Principal Energy Level #2 has two sublevels: s and p

Principal Energy Level #3 has three sublevels: s, p and d

Principal Energy Level #4 has four sublevels: s, p, d, and f

SUBLEVEL s has a spherical shape

SUBLEVEL p is further subdivided into three “two-lobed” orbitals, each surrounding one of the three dimensional axis, x,y,z and can be written as p_x, p_y, p_z.

SUBLEVEL d is further divided into five orbitals (do not worry about memorizing the shapes of these)

SUBLEVEL f is further divided into seven orbitals (again, do not worry about memorizing the shapes of these)

One important note is that the larger the number of the principal energy level, the further the average distance of the electron from the nucleus of the atom. This means that the relative size of a 1s orbital is smaller than that of a 2s orbital and so on.

Another note is that as new more complex atoms are created (with more electrons) it is possible to add principal energy levels for example you could have a principal energy level of 5 which would have five sublevels. The intended naming scheme for sublevels would have g be the next sublevel after f and would presumably have 9 orbitals.

Chapter10, Section 8

The major triumph of the Wave Mechanical Model of the Atom was that it was the first model that was able to be accurately applied to all elements.

This model even allows us to explain the arrangement of the periodic table and the properties of the elements.

Remember that an atom has as many electrons as it does protons (to give a net charge of zero to the overall atom).

All atoms other than hydrogen have more than one electron.

When we are dealing with multiple electrons there are some basic rules:

* No more than two electrons can occupy an orbital at the same time.

*If there are two electrons in the same orbital they will have opposite spins.

*Spin Up is designated by an upward facing arrow. $\uparrow$

*Spin Down is designated by a downward facing arrow. $\downarrow$

This leads us to an important principle in chemistry:

THE PAULI EXCLUSION PRINCIPLE: an atomic orbital can hold a maximum of two electrons, and those two electrons must have opposite spin.

A Summary of the Wave Mechanical Model of the Atom

1. Atoms have energy levels called PRINCIPAL ENERGY LEVELS which are numbered with integers (1,2,3,4,…)
2. The energy of the level increases at the number of the Principal Energy Level increases.
3. Each Principal Energy Level contains one or more sublevels (the number of sublevels is equal to the number of the Principal Energy Level)
4. The sublevels are indicated by letters (s, p, d and f)
5. Each sublevel contains one or more orbitals (s has 1 orbital, p has 3 orbitals, d has 5 orbitals and f has 7 orbitals)
6. Each orbital can hold from 0-2 electrons. If more than one electron occupies the orbital, then the electrons have opposite spin.

Chapter 10, Section 9

Electron Arrangement- First 18 Elements

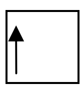
Electrons prefer to be as close to the nucleus as possible (+ charge attracts – charge)

Only two electrons are able to be very close to the nucleus,1s orbital, because only two electrons can occupy an orbital at one time. The rest of the electrons must fill in the other orbitals.

There are two ways of describing the electron arrangement of an atom.

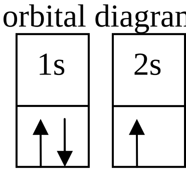
ELECTRON CONFIGURATIONS- these are short hand notations for the energy levels and sublevels. (Example: hydrogen's electron configuration is $1s^1$, indicating that there is one electron in the 1s orbital.

ORBITAL DIAGRAMS- use boxes to represent the orbital thus giving a more visual approach to electron arrangement description. (Example: hydrogen's orbital diagram looks like this: 1s indicating one electron in the 1s orbital.



The electron configuration for Lithium is $1s^2$ $2s^1$, indicating that there are two electrons in the 1s orbital and one electron in the 2s orbital.

The orbital diagram for Lithium is



*** Notice the opposite spin on the two electrons in the 1s orbital.

The orbitals fill according to their energy level, the lower energy (closer to the nucleus) levels fill first.

Electron configurations can be abbreviated using "noble gas configurations" to skip over some repetitive material. Example: The configuration of Sodium is written as [Ne] $3s^1$ instead of 1s2 $2s^2 2p^6 3s^1$.

VALENCE ELECTRONS: electrons in the outermost (highest) principal energy level.

CORE ELECTRONS: electrons in the non outer most principal energy level.

** Only Valence Electrons are important for bonding and reaction purposes.

Chapter 10, Section 10

Just as in the English language there are exceptions to the rules of electron configurations.

Chromium and Copper are transition metals which have irregular electron configurations for reasons beyond the scope of this introductory class.

Chromium would be expected to have the configuration of [Ar] $4s^2\ 3d^4$, but instead Chromium's electron configuration is [Ar] $4s^1\ 3d^5$.

Copper would be expected to have the configuration of [Ar] $4s^2\ 3d^9$, but instead Copper's electron configuration is [Ar] $4s^1\ 3d^{10}$.

Reasons for this odd behavior will be covered in Chemistry II, but for now you will just have to memorize these two exceptions.

Do not try to memorize all the electron configurations, just these two exceptions. The periodic table can be used to help determine electron configurations for all rule following elements.

There are a few rules that will help you work with the elements having more than 18 electrons.

1. If an atom has enough elements to be using a principal energy level that contains a d sublevel then the s sublevel of the next higher energy level will fill first, then the d sublevel.
2. After lanthanum a group of 14 elements, the lanthanide series, fills the 4f orbitals.
3. After actinium a group of 14 elements, the actinide series, fills the 5f orbitals.
4. Except for helium, the group numbers indicate the total number of electrons in the s and p sublevels of the highest energy level. These are the valence electrons and influence the bonding and chemical properties of the elements.
5. d sublevel electrons are never in the highest energy level of the atom, so they do NOT count as valance electrons.

**** NOTE: You are NOT responsible for being able to produce the electron configurations for the elements in the lanthanide and actinide series, just for knowing which orbitals are filling in each series.

Elements in Groups 1,2,3,4,5,6,7 and 8 are referred to as MAIN GROUP or REPRESENTATIVE ELEMENTS.

Chapter 10, Section 11

Trends in Atomic Properties in the Periodic Table

Remember:

Metals are located on the left side of the periodic table and have similar observable properties.

*Luster
*Malleability
*Conductivity

Non Metals are located on the right side of the periodic table and have similar observable properties.

*Dull
*Brittle
*Insulator

Chemically, metals want to loose electrons and become positive ions. Non Metals want to gain electrons and become negative ions.

When a metal and a non metal meet, a transfer of electrons takes place. The metal will give up electrons and the non metal will gain those electrons. This creates a positive ion and a negative ion. These two opposing charges are attracted to each other and form an ionic bond.

Remember:

The further an electron is from the nucleus the less control the nucleus has over the electron.

This means the farther from the nucleus (the higher the energy level) the more easily an electron is to give up.

If we apply this knowledge to the Group 1 (alkali metals) cesium looses an electron more easily than rubidium which looses an electron more easily than potassium which looses an electron more easily than sodium which looses an electron more easily than lithium.

As we proceed down a group the valence electrons are further from the nucleus, because they are in a higher energy level with each progression down the group.

The easier it is to remove an electron the less energy it takes to remove the electron.

The energy required to remove an electron is known as IONIZATION ENERGY. The energy required to create an ion (remove electrons) of the atom.

If we look at ionization energy from the perspective of the periodic table we can establish trends. We already recognize that as we move down a group ionization energy decreases (less energy required to remove an electron because the electron is farther away from the control of the nucleus)

Non metals want to gain electrons, not loose electrons so they have very high ionization energies. This means that as we move across a period on the periodic table ionization energy increases.

Combining these two concepts gives the PERIODIC TREND FOR IONIZTION ENERGY: Ionization energy decreased down a group and increases from left to right across a period.

It may be easier to remember the trend with both directions increasing or decreasing:
Decreases down and to the left
Increases up and to the right

Which method you use to learn the trend for Ionization Energy is your option.

There is another periodic trend that you should be familiar with:

Atomic Size

The periodic trend for Atomic Size is that it increases down a group and decreases as you move from left to right across a period.

Decreases up and to the right
Increases down and to the left

It may be obvious that the atomic size would increase down a group as we are increasing principal energy level and that indicates further distance from the nucleus.

The reason for decreasing size across a period may not be so obvious.

All elements in a given period are filling the same principal energy level. This means the orbitals are the same size. As we move across the period there are more protons (less shielding) and this leads to smaller total atomic size.