**Ms. Randall**
**Unit 3: Atomic Theory Unit Notes**

**Unit Objectives:**

- Understand that the modern model of the atom has evolved over a long period of time through the work of many scientists
- Discuss the evolution of the atomic model
- Relate experimental evidence to models of the atom
- Identify the subatomic particles of an atom (electron, proton, and neutron)
- Know the properties (mass, location, and charge) of subatomic particles
- Determine the number of protons, electrons, and neutrons in a neutral atom and an ion
- Calculate the mass number and average atomic weight of an atom
- Differentiate between an anion and a cation
- Identify what element the amu unit is derived from
- Distinguish between ground and excited state
- Identify and define isotopes
- Write electron configurations
- Generate Bohr diagrams
- Differentiate between kernel and valence electrons
- Draw Lewis Dot Diagrams for an element or an ion

***Focus Questions for the Unit:***

- What is the structure of an atom?
- How did the model of the atom change over time?
- How does the structure of an atom relate to chemical properties of elements?

**YOU SHOULD BE ABLE TO ANSWER THESE *IN DETAIL* BY THE END OF THE UNIT**

**Define the following vocabulary:**

| | |
|---|---|
| Allotrope | Anion |
| Atom | Atomic Mass |
| Atomic Mass unit (a.m.u.) | Atomic number |
| Bohr model | Cation |
| Compound | Electron |
| Electron Configuration | Element |
| Excited state | Ground state |
| Ion | Isotope |
| Lewis Dot Diagram | Mass number |
| Neutron | Nuclear Charge |

**Lesson 1: Chapter Diary 3 Atomic Theory**

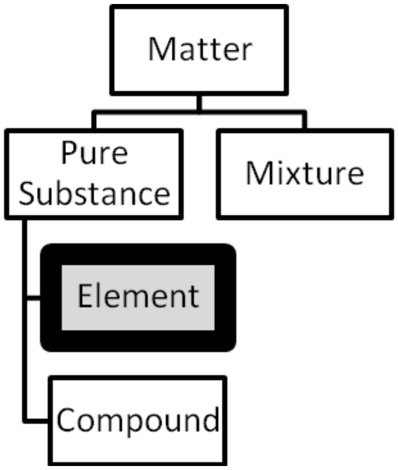
**Objective**: To summarize unit concepts

**Directions:** After reading the Atomic Theory Chapter diary 3 answer the questions in your workbook.

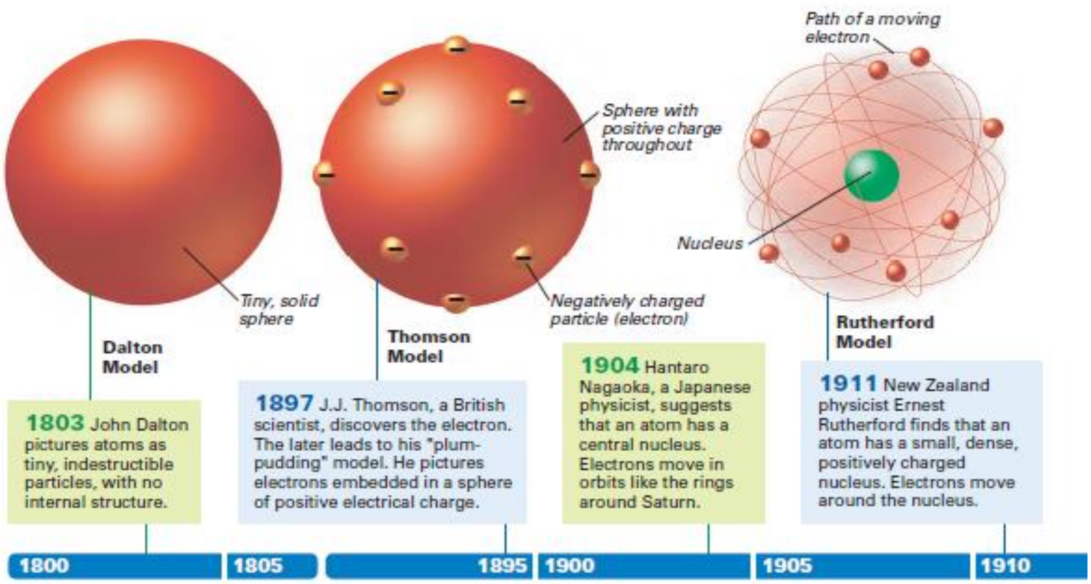
**Lesson 2: Concept of an Atom**

**Objective:** To compare and contrast the development of the models of the atom.

Often in science, we are only aware of a certain amount of information at a time. As we learn new facts, we can edit our current knowledge and develop deeper understandings of what we are studying. This is why science is often changing... we are constantly learning new things! We learned in Unit 2 that the smallest unit of matter is called an element. An **element** is a type of pure substance. Atoms are the particles that make up elements. An element can be **made of many of the same type of atoms.**




The concept of an atom is one that has evolved over time. Many different scientists have contributed to our understanding of the atom by doing various experiments to form their theory or view of the atom. Each person relied heavily on the scientist before them and they either supported or refuted the theory that existed at the time.




Atomic models are theories that describe what the physical properties of an atom are (how it looks structurally). These physical properties will help determine the chemical proper ties of an atom (how it behaves). Our current model of the atom, the electron cloud model, is based on all the ones that came before it.

**Please watch the following video: Structure of an Atom**

## The Atom – Background & History

**Atom:** smallest particle of an element that retains its properties.

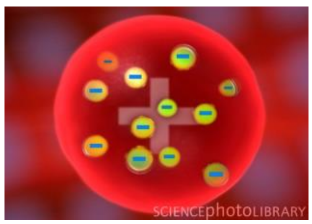
It is the **basic** building block of matter; cannot be broken down chemically

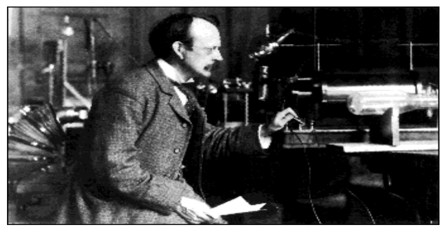
**1. Democritus =** (450 BC) Greek man who tried to answer the question about the divisibility of matter; used the term "atomos" to describe the ultimate/smallest particle of matter.

**2. Dalton (1803)**

- Known as the founder of the atomic theory
- Dalton invented the word atom as the basic unit of matter which were
- considered to be indivisible
- Dalton also claimed that all atoms of a given element are identical
- He also discovered that atoms of different elements have different properties and masses
- Found that combining atoms of different elements formed compounds
- Theory referred to as the cannonball theory because it looked like a simple sphere with uniform density.

**3. J.J. Thomson (1897)**



**View the following link :** Cathode ray tube

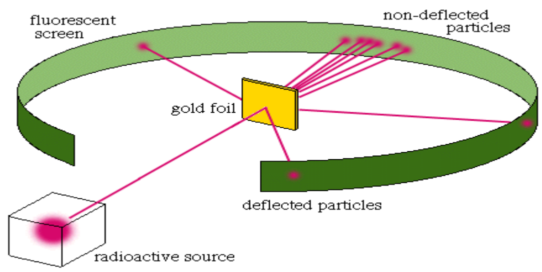
You can find this on my website!

- While using a cathode ray tube he discovered that the ray was deflected (due to a magnetic/electrical field)
- From this discovery he concluded that atoms contain small negatively charges particles called electrons
- Theory famously referred to as the plum pudding model because he visualized the electrons being embedded within the structure of the atom (just like raisin bread)
- The mass of the rest of the atom (besides the electrons) was thought to be evenly distributed and positively charged

**4. Ernst Rutherford (1909)**

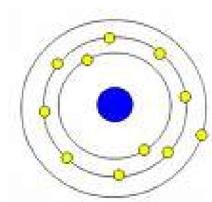
Simulation: http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/ruther14.swf

**(You can access link on my website, instead of re-typing it ☺)**




- Gold foil experiment where he bombarded a thin piece of gold foil with a positive stream of alpha particles
- Often referred to as the nuclear model
- Most alpha particles went straight through and some were deflected
- Two conclusions were therefore made:
- 
    1) most of the atom is empty space
    2) atoms have a dense positive central core called the nucleus

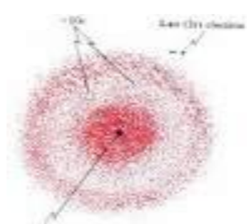
**5. Neils Bohr (1913)**



- Proposed that the atom consists of a dense nucleus with electrons found in well-defined orbits/paths
- He therefore stated that each electron orbiting the nucleus must possess a definite amount of energy to keep it in place within its orbital
- Known as the planetary model (looks much like our solar system)

**6. Wave-Mechanical/Cloud Model** (Modern, present-day model)

- Developed after the famous discovery that energy is made up of both waves and particles
- Still the same dense positively nucleus
- Electrons now have distinct amounts of energy and move in areas called orbitals or regions around the nucleus where there is a high probability of being found
- Many scientist have contributed to this theory
- Different from the Bohr diagram—now the location of the electron is based on probability within the orbital



**Lesson 3: Subatomic Particles**

**Objective: To define the parts of an atom**

***Please watch the following video:* Bill Nye Atoms**

Each subatomic particle has very unique physical properties that will in turn; affect the chemical properties of the atoms they make up. The following reading passage will help you to better understand the structure of the atom.

***Understanding Atomic Structure...***

The word nucleus of the atom is just like the nucleus of the cell that you learned about in Living Environment. One of the properties of the nucleus is that it is very dense compared to the rest of the atom. This is a qualitative observation. To put this into quantitative terms, we say that each of the subatomic particles found in the nucleus has an approximate relative mass of 1 amu or **atomic mass unit.**

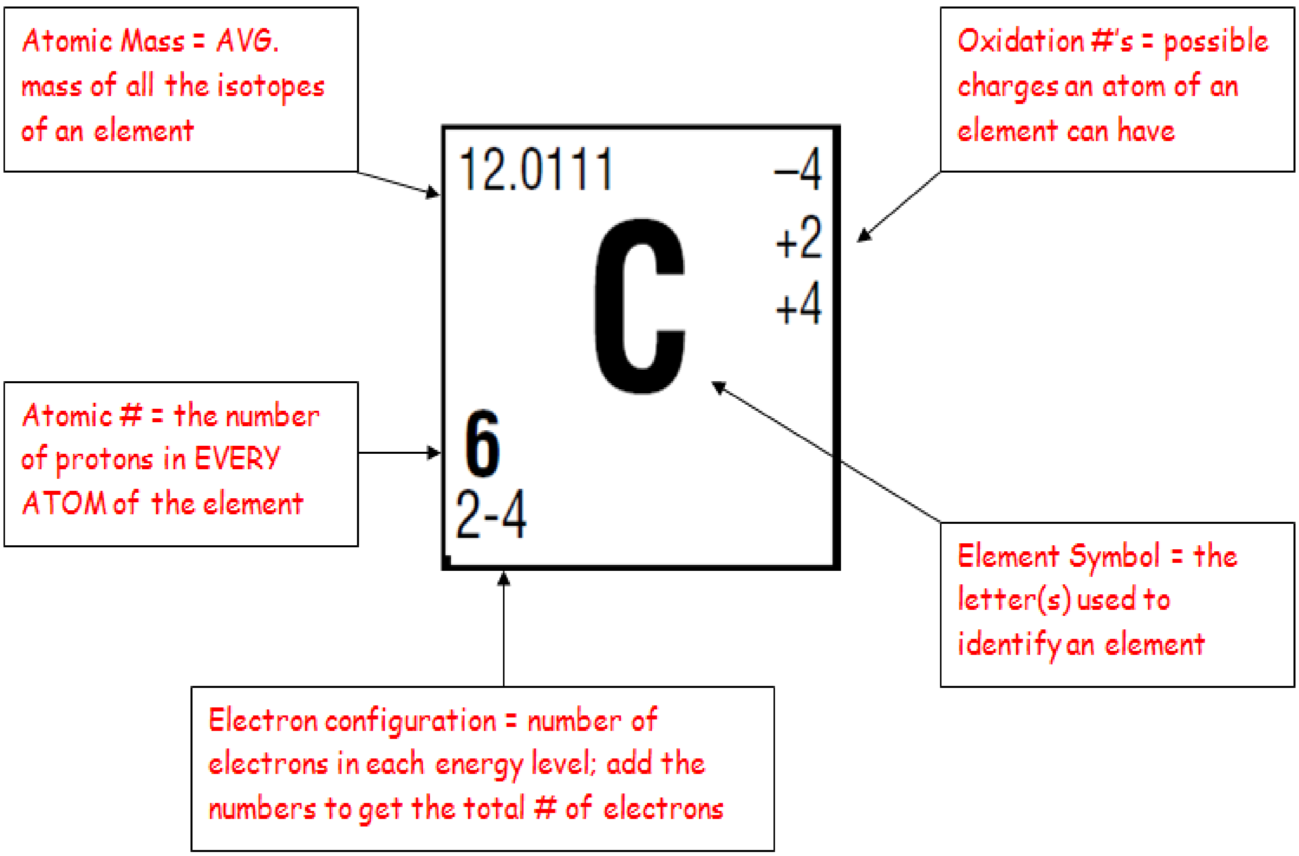
The majority of the atom's mass comes from the dense center of the atom (the nucleus). However, most of the atom is not dense; it is made of empty space. Have you ever heard of the idea that opposites attract? This principle comes from science. Positives and negatives attract one another and this attraction holds the atom together. Positively charged particles are called **protons** and given the symbol $p^+$**.** The number of protons is an atom's **atomic number**. The negatively charged subatomic particles are called **electrons** and given the symbol $e^-$. The third type of subatomic particle has no charge and so Chemists say it is neutral. (A neutral party in an argument is one who has no opinion.) We call these particles **neutrons** (sounds like neutral), and use an $n^0$. Each of the three types of subatomic particles has a different charge so that they can all co-exist in the small space that is the atom. Protons and neutrons are collectively called **nucleons** as they are found in the nucleus of an atom. Electrons surround the nucleus.

Notice that the +/-/0 for each symbol is in a specific location. This means something! The symbol is a way to abbreviate the subatomic particle, while still telling you something about it. This is just like your initials. If I wrote my initials as TFR, compared to TRF, compared to FTR, they would all mean something very different. This is the same for your date of birth or any other abbreviation or symbol. In Chemistry, the charge of the atom is always located in the upper right hand corner of the symbol.

Each type of atom has a specific number of protons. This is how we identify the atom. It is basically like its social security number... no two atoms can every have the same number of protons. We call these **atomic numbers**. **The word "atom" implies that the number of protons (positives) equal the number of electrons (negatives) to create a "neutral" atom.**

From now on, whenever you are asked for a charge, you should always include a positive or negative sign with it. If you do not, you are indicating that it is neutral, not positive or negative.

**We can use the reference table to help us with many of these symbols and definitions. Look at the key on the periodic table.**




**Symbol and Name-**: In compounds notation the symbols are used. When just referring to the element itself it can be written two ways:

C-12 (symbol and mass number)

**or**

12 (mass number)
C
6 (atomic number)

- **Atomic Number**- The number of protons in the element's nucleus. Always located on bottom left corner.

- **Atomic Mass-** Mass of the element's nucleus (Protons and Neutrons) Atoms of the same element always have the same number or protons. Not a whole number because the same element can have different numbers of neutrons (Isotopes) **=** sum of the protons and neutrons in an atom of an element (p + n) = mass #. Always listed in top left corner.

- **Electrons-** For a neutral atom, the number or protons and electrons are equal

- **Nuclear Charge** = charge w/in the nucleus; = to the # of protons or the atomic #
  p = nuclear charge

- **Nucleons** = any subatomic particles found w/in the nucleus; protons and neutrons

| Subatomic Particle | Charge | Relative Mass | Location | Symbol | How to Calculate |
|---|---|---|---|---|---|
| **Proton** | +1 | 1 amu | Nucleus | $p^+$ | Look at the atomic # |
| **Neutron** | 0 | 1 amu | Nucleus | $n^0$ | Mass # - Atomic # |
| **Electron** | -1 | $1/1836^{th}$ amu (negligible) | Outside nucleus | e- or e | P = e- in neutral atom |

***Lesson 4:*** **Subatomic Particles & Ions**

**Objective:** **To compare and contrast the subatomic particles of atoms and ions.**

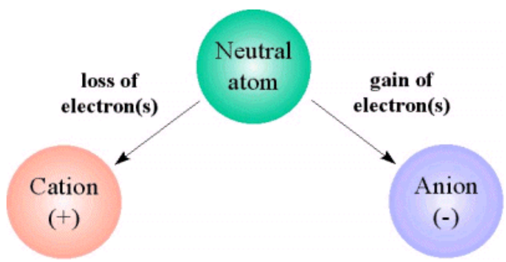
**Atoms** are neutral; they contain the same number of protons as electrons. By definition, an **ion** is an electrically charged particle produced by either removing electrons from a neutral atom to give a positive ion or adding electrons to a neutral atom to give a negative ion. When an ion is formed, the number of protons does not change.

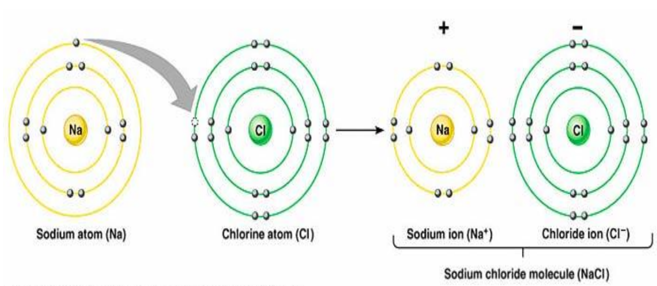
Neutral atoms can be turned into positively charged ions called **cations** by removing one or more electrons. A neutral sodium atom, for example, contains 11 protons and 11 electrons. By removing an electron from this atom we get a positively charged $Na^+$ ion that has a net charge of +1.

Atoms that gain extra electrons become negatively charged. These are called **anions**. A neutral chlorine atom, for example, contains 17 protons and 17 electrons. By adding one more electron we get a negatively charged $Cl^-$ ion with a net charge of -1.

The gain or loss of electrons by an atom to form negative or positive ions has an enormous impact on the chemical and physical properties of the atom. Sodium metal, for example, which consists of neutral sodium atoms, bursts into flame when it comes in contact with water. Neutral chlorine atoms instantly combine to form $Cl_2$ molecules, which are so reactive that entire communities are evacuated when trains carrying chlorine gas derail. Positively charged $Na^+$ and negatively charged $Cl^-$ ions are so unreactive that we can safely take them into our bodies whenever we salt our food.

**Cations** are POSITIVELY charged ions, whereas **Anions** are NEGATIVELY charged. Because a **cation has lost an electron**, it has a smaller radius than its original atom. Because an **anion gains an electron** it has a larger radius than its original atom.



**Lesson 5** **Subatomic Particles & Isotopes**

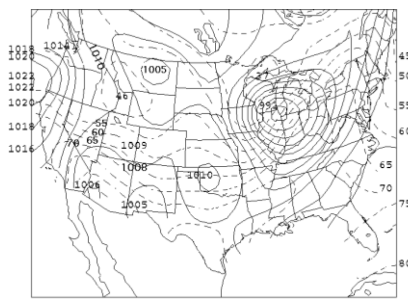
**Objective: To differentiate between isotopic forms of atoms based on number of neutrons**

**Atoms** are identified by the numbers of protons, neutrons, and electrons that they contain. Before you can understand the properties of atoms, how atoms combine to form molecules, and the properties of molecules, you must be familiar with the number of protons, neutrons, and electrons associated with atoms.

From the perspective of a chemist, the entire world is composed of atoms, and atoms are composed of protons, neutrons, and electrons. Protons and neutrons are about 2000 times heavier than an electron. A proton has a charge of +1, a neutron has no charge, and an electron has a charge of -1. The nucleus is very dense and very small compared to the entire atom.

If I take a sample of air from this room and put it into a box, there may be more than one type of hydrogen atom. Remember that in order for an atom to be hydrogen, it must have a certain number of protons (*protons determine the identity of an element*). The properties of atoms are determined by the numbers of protons, neutrons, and electrons that they contain. Atoms with the same number of protons but different number of neutrons are called **isotopes** of an element.

In Earth Science, you learned about Isobars. Isobars are lines (like the ones shown below) that represent areas of the same pressure.



00Z May 1, 1997 Sea-level pressure (solid) and surface temperature (dashed)

***"Iso-"*** is a prefix that means ***same*** "**-tope**" sounds like "type"...

The isotopic notation for an atom includes the following information: symbol of the element, the element's atomic number (Z) which specifies the number of protons in the nucleus, and the mass number (A) which indicates the number of protons plus neutrons in the nucleus. [The number of electrons in a neutral atom is equal to the number of protons in the nucleus of the atom. The mass contributed by the electrons in an atom is very small, so it is not included when calculating the mass number.]

***Atomic Symbol Notation***

| | |
|---|---|
| A<br>X<br>Z | Mass Number<br>Atomic Symbol<br>Atomic Number |

## ISOTOPES AND MASS

**Isotope =** atoms of the same element with different mass numbers but same atomic #

= same number of protons, different number of neutrons

Example: Isotopes of Carbon (C-12, C-13, & C-14)

$^{12}_{6}C$ $^{13}_{6}C$ $^{14}_{6}C$

So why does Carbon have a mass of 12.011 on the Reference Table? This is carbon's atomic mass which is the **average weighted mass** of all naturally occurring isotopes of carbon (there exist three different isotopes of carbon in the atmosphere – as seen just above)

**Atomic Mass =** the weighted average of an element's naturally occurring isotopes

$$\frac{(\%\text{abundance} \times \text{mass of isotope 1}) + (\%\text{abundance} \times \text{mass of isotope 2})}{100}$$

Carbon has two naturally occurring isotopes, carbon-12 ($^{12}C$) and carbon-13 ($^{13}C$). In nature 98.9% of carbon atoms have a mass of 12 and 1.1% of carbon atoms have a mass of 13.

So…

| Isotope | Percent abundance in nature |
|---|---|
| $^{12}C$ | 98.9% |
| $^{13}C$ | 1.1% |

Calculate the average atomic mass of carbon based on its naturally occurring isotopes and their percent abundance.

Set it up like this!

% abundance

$$\frac{(12 \times 98.9\%) + (13 \times 1.1\%)}{100}$$

mass of isotope

______________

12.011 amu

**A second example:**

Nitrogen has two naturally occurring isotopes. $^{14}N$ occurs 99.6% of the time and $^{15}N$ occurs 0.4% of the time. Calculate the average atomic mass of nitrogen.

$$\frac{(14 \times 99.6\%) + (15 \times 0.4\%)}{100} =$$

14.004 amu

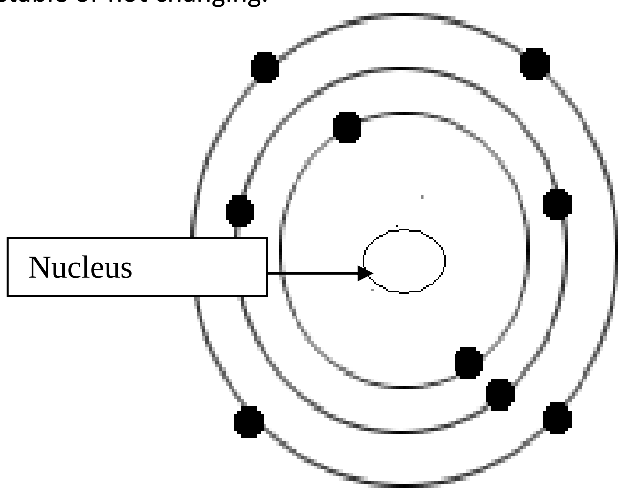
We need to make sure to differentiate between mass number and atomic mass of an atom!

| | Mass Number | Atomic Mass |
|---|---|---|
| Definition | The total number of protons and neutrons in an atom. | The weighted average of the naturally occuring isotopes of an element. |
| Explanation | Must be found for an individual atom. Can atoms ever have only half of a proton? No!!! This means that if the mass of a proton is 1 amu and the mass of a neutron is 1 amu, the mass number of an atom should never be a decimal! Decimals imply that the number was an average of multiple mass numbers. | Reported as a decimal on the periodic table. |
| Application | If a problem gives you the mass number and the number of protons (or tells you what element it is), you must find the number of neutrons by subtracting (mass number-protons) | If a problem gives you only the identity of an element, and does not give you the number of neutrons or the mass number, you can use the number given to you on the periodic table and round it to the nearest whole number. |

**Lesson 6: Bohr Diagrams**

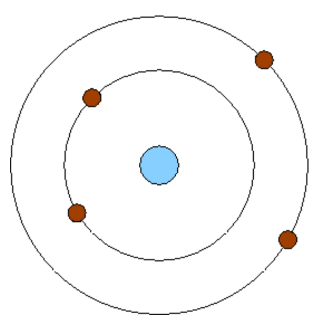
**Objective: To represent the atom using the model defined by Bohr.**

Look at the model of the atom proposed by Bohr (below). We know that the subatomic particle that determines the identity of the atom is the proton and it is located in the nucleus. Does the identity of an atom ever change in chemical reactions? NO! So do you think that the number of **protons** ever changes in a chemical reaction? NO! Chemists might say then that the number of protons is stable or not changing.

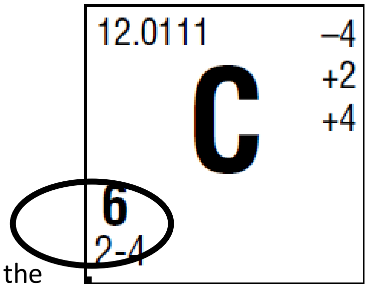



Bohr did not yet know about **neutrons**. These would not come until much later. Neutrons are also located in the nucleus of the atom. Atoms of the same element may have different numbers of neutrons. These are known as **isotopes.** However, once an atom has a specific number of neutrons, this number will not change. For example, Cl-37 cannot give one of its neutrons away to become Cl-35!

Thomson had already discovered the electron. Bohr's major contribution was the proposal that **electrons** were not spread throughout the atom as Thomson thought, but that they must exist in fixed, stable orbits around the nucleus. Because of Bohr's model of the atom, we describe the structure of an atom in terms of **ENERGY LEVELS (aka "shells" or "orbits").** Since this subatomic particle is located on the outside of the atom, it is the subatomic particle that will have the most impact on the properties of the atoms, because it is most likely to be impacted by other atoms around it. We can use the structure and locations of the electrons to help us predict these properties!



**Electron Configurations** represent the number and location of electrons around the nucleus in an atom**.** A dashed chain of numbers found in the lower left corner of an element box (see below) tells us the number of **energy levels** as well as the number of **electrons** in each level (tells us how the electrons are arranged around the nucleus)




**All electron configurations on the Periodic Table are **NEUTRAL ($p^+ = e^-$)**

| SUBSTANCE | ELECTRON CONFIGURATION |
|---|---|
| Magnesium | 2-8-2 |
| $Mg^{+2}$ | 2-8 |
| Bromine | 2-8-18-7 |
| $Br^{-1}$ | 2-8-18-8 |
| Barium | 2-8-18-18-8-2 |
| *Lead | -18-32-18-4 |

* shortcut allows you to cut out the first two orbitals to shorten the "address"

**Valence Electrons:** electrons found in the **outermost** shell or orbital; the **last** number in the electron configuration

**Kernel Electrons: inner** electrons (all non-valence electrons)

Example:

Sulfur # valence e- 6
# kernel e- 10

Nitrogen # valence e- 5
# kernel e- 2

**Principle Energy Level (n)** = electron energy levels that contain a certain number of sublevels (s, p, d, f); each sublevel contains one a set number of orbitals

Maximum # of electrons in an energy level = $\mathbf{2n^2}$ where n = quantum # (or period #)

| Principle Energy Level (n) | Maximum number of electrons ($2n^2$) |
|---|---|
| 1 | 2 |
| 2 | 8 |
| 3 | 18 |
| 4 | 32 |

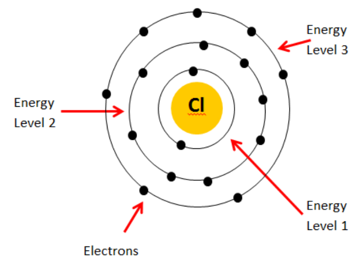
<u>**Sample question**</u>: What is the maximum number of electrons that can occupy the 3rd principal energy level in any atom? $2(3^2) = 18$

**<u>Bohr Diagrams are drawings that show the structure of the atom.</u>**

Step 1: Draw a nucleus and label it with the correct number of protons and neutrons.
Step 2: Look up the electron configuration for the element. (Use key on Periodic table.) The e-configuration tells us how many electrons are on each orbital.
Step 3: Draw the correct number of orbitals around the nucleus and label them with the correct number of electrons.

**Ex:** $^{24}_{12}Mg$


**Ex:** $^{37}_{17}Cl$




**Magnesium electron configuration 2-8-2**

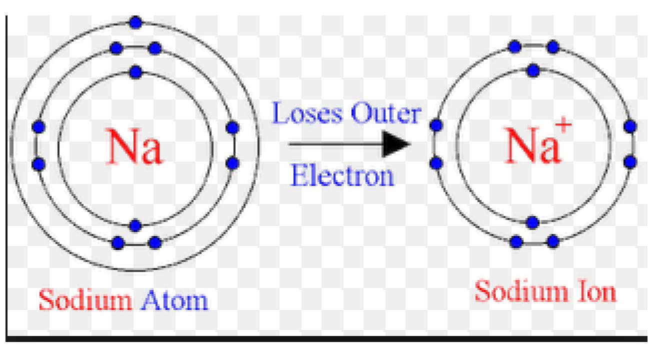
**Chlorine electron configuration 2-8-7**

**<u>Drawing Bohr Models of Ions</u>**

When drawing a Bohr model for an ion write the ion electron configuration by adding or removing electrons accordingly. Draw the Bohr model.

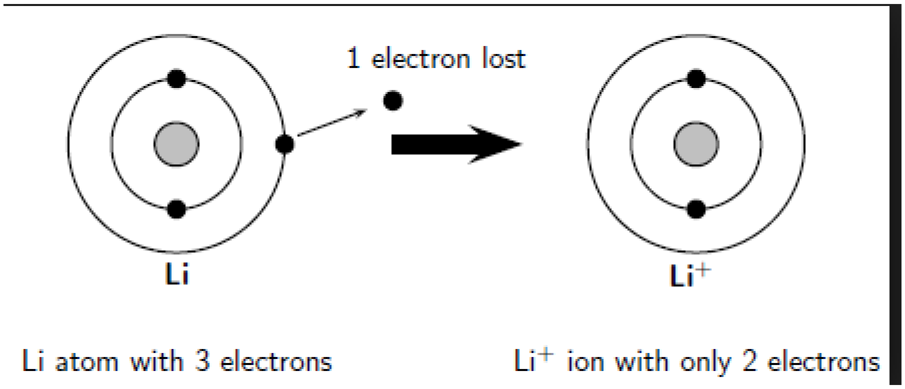
**Sodium atom configuration 2-8-1**

**Sodium cation configuration 2-8**




**Lithium atom configuraiton 2-4**

**Lithium anion configuration 2**


1 electron lost
Li
Li+
Li atom with 3 electrons
Li+ ion with only 2 electrons

**<u>Lesson 7: Electrons and Energy</u>**

**<u>Objective:</u> To relate energy state to electron location**

The modern or "quantum mechanical" model of the atom describes energy levels as being further subdivided into sublevels and **orbitals**. The word "orbital" is one you need to know, and simply refers to a region of space around the nucleus where there is a high probability of finding an electron. The more electrons an atom has, the more orbitals it will use. Each orbital contains a maximum of 2 electrons.

Each electron in an atom has a particular amount of energy depending on how close they are to the nucleus. Electrons in orbitals that are **closer to the nucleus** are more energetically **stable and are at a lower energy**. When all the electrons in an atom are in their lowest possible energy levels, the atom is said to be in the "**ground state**." The ground state electron configuration is the one displayed on the periodic table below each element.

When an atom interacts with an outside energy source such as heat, light or electricity, the electrons can be affected. The most accessible and therefore most likely to be affected electrons are the valence electrons. If the electron can **absorb** just the right amount of **energy** to pop it to a **higher energy level**, it will indeed do this. The atom is now described as being in the "**excited state**." An atom in the excited state is unstable and will immediately **emit the energy** *in the form of light* as the electron(s) **return to the ground state**. The excited state configuration of an atom has the **same number of electrons** as the ground state. For example, oxygen in the ground state is 2-6, but in the excited state might be 2-5-1, or 1-7, or 1-6-1. In all cases the total number of electrons is that of an atom of oxygen… that is 8 electrons.

**<u>Ground State</u> =** electrons in **Lowest energy** configuration possible (configuration found on PERIODIC TABLE)

→ ground state electron configuration for Li is 2-1

**<u>Excited State</u> =** electrons are found in a **Higher energy** configuration. (ANY configuration NOT found on Periodic table)

→ excited state electron configuration for Li could be 1-2, 1-1-1

**So how can you tell if you have a ground state or excited state configuration?**

**1) Shell Configuration (Principal Energy Levels)**

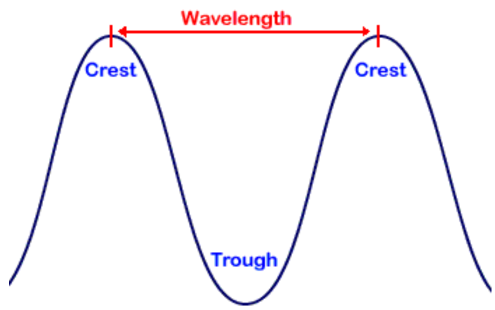
Add up all of the electrons in the configuration. Use the total (which equals the atomic number) to identify the element that the configuration belongs to. If the basic configuration matches the configuration on the Periodic Table, then it is in the ground state. If the configuration does not match that element's configuration, then it is in the excited state.

| Given Shell Configuration | Add Up The Electrons | Which Element Is It? | What is the Shell Configuration on the Periodic Table For this Element? | Does it match your given configuration? | Ground state or excited state? |
|---|---|---|---|---|---|
| 2-8-8-3 | 21 | Sc | 2-8-8-3 | YES | Ground |
| 2-7-2 | 11 | Na | 2-8-1 | NO | Excited |
| 2-8-3-1 | 14 | Si | 2-8-4 | NO | Excited |
| 2-8-15-2 | 27 | Co | 2-8-15-2 | YES | Ground |

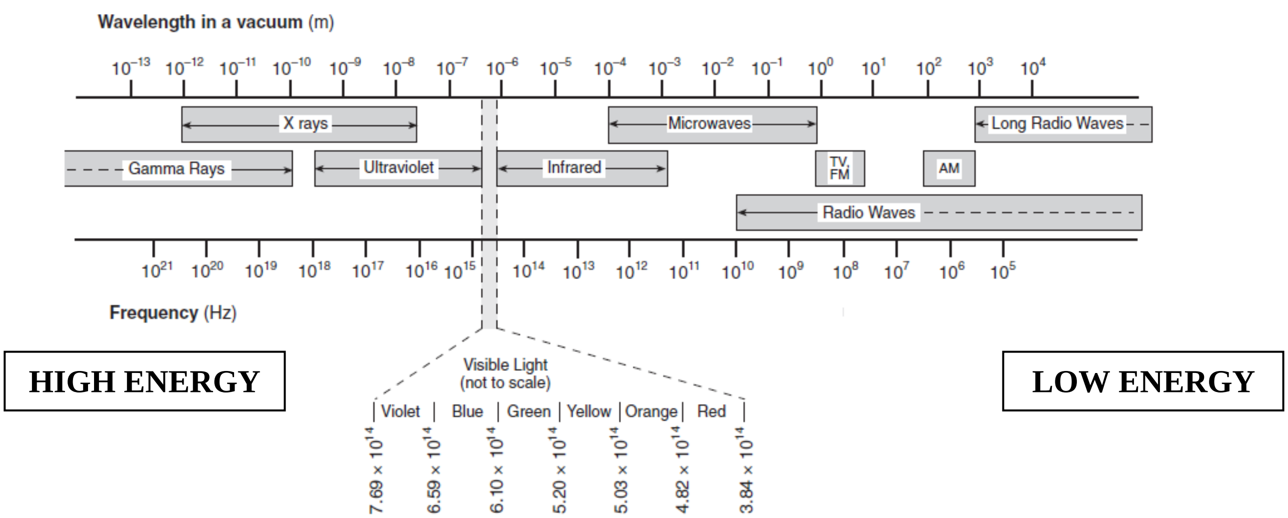
## Bright Line Spectrum

We need to consider the nature of light, or what is called the **electromagnetic spectrum**.

The electromagnetic spectrum is a continuum of all electromagnetic waves arranged according to frequency and wavelength. The sun, earth, and other bodies radiate electromagnetic energy of varying wavelengths. Electromagnetic energy passes through space at the speed of light in the form of sinusoidal waves. The wavelength is the distance from wave crest to wave crest (see figure below).




Light is a particular type of electromagnetic radiation that can be seen and sensed by the human eye, but this energy exists at a wide range of wavelengths. The micron is the basic unit for measuring the wavelength of electromagnetic waves. The spectrum of waves is divided into sections based on wavelength. The shortest waves are gamma rays, which have wavelengths of 10e-6 microns or less. The longest waves are radio waves, which have wavelengths of many kilometers. The range of visible consists of the narrow portion of the spectrum, from 0.4 microns (blue) to 0.7 microns (red).



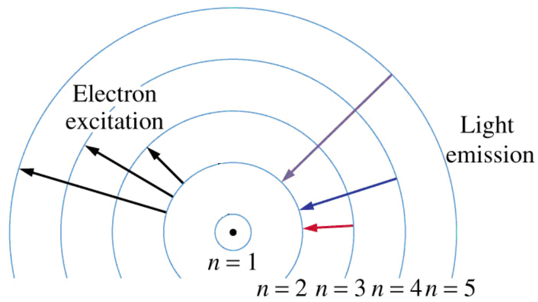

**Ground → excited** **energy is absorbed**

**Ground ← Excited** **energy is released (in the form of light energy)**

When atoms **absorb energy** their electrons will shift to **a higher energy level** or an **excited state**. **The greater the distance from the nucleus, the greater the energy of the electron**.

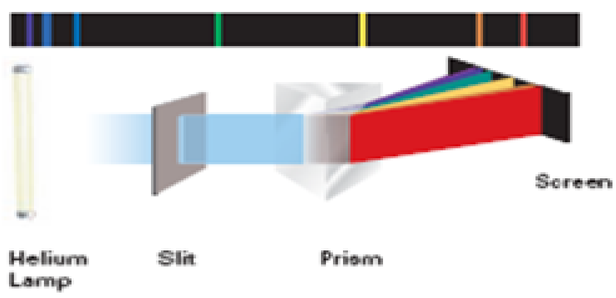
This is a very unstable and temporary condition so the electrons will fall back down or drop to a lower state energy level (or the ground state.) When they **fall back down from the excited state to the ground state** they **release energy** in the **form of light**.

On the left of the diagram, electrons are shown being excited from the $2^{nd}$ energy level (n = 2) to the $3^{rd}$, $4^{th}$ and $5^{th}$ energy levels. As they return to their original energy level (to the ground state), they emit their extra energy in the form of a particular frequency (color) of light.




The specific set of frequencies of light emitted by atoms as they return to the ground state produces a **bright-line spectrum**. Every element has its own unique spectrum, and could be thought of as the "fingerprint" of that element. A spectrum of samples of unknown composition can be observed in order to identify the elements present. Much of what we know about the structure and history of the universe has been figured out by observing the spectra of stars.

This diagram shows an example of how a spectrum is taken and what it would look like. It may be helpful to look at the COLOR PICTURE on p 141 of your Prentice Hall textbook.




**Another Resource:**

- ☐ http://www.visionlearning.com/library/module_viewer.php?mid=51&l=&c3=
  **(scroll down to "Bohr Atom: Quantum Behavior of Hydrogen" animation)**

**Lesson 8: Lewis Dot Diagrams:**

**Objective: To represent the atom using the model defined by Lewis**

The electrons are the subatomic particle that is most important for our study of Chemistry because it is on the outside of the atom and therefore it is the particle that is most likely to be affected by other atoms around it. Due to this, Chemist's use another kind of diagram to help keep track of an atoms electrons. It is called a Lewis Dot Diagram. **Lewis Dot Diagrams** show only the outer most energy level of electrons. This level is called the ***valence electrons.*** To find the number of valence electrons, you start by writing down the electron configuration. The last number in the configuration is the number of valence electrons!

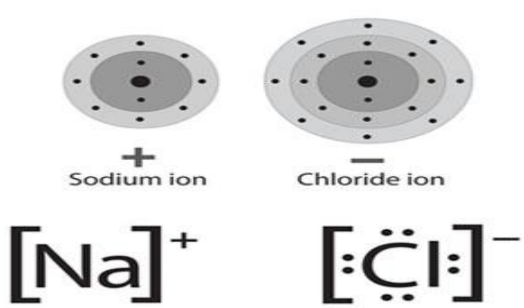
1. Write the element's symbol

2. Retrieve electron configuration from Periodic Table. The last number in the configuration is the NUMBER OF VALENCE ELECTRONS

1. Arrange the valence electrons (DOTS) around the symbol using the following rules:
    - Only two electrons maximum per side of the symbol (therefore no more than 8 total surrounding symbol – 8 is great!)
    - Always "pair" the first two
    - If you have more than 2 valence electrons, deal them one at a time to the other three sides until you run out

| | | |
|---|---|---|
| | 1 2 | |
| 8<br>5 | X | 3<br>6 |
| | 4 7 | |

2. If you are working with an **ion** you must adjust the valence electrons (add or subtract electrons) in the configuration before constructing your Dot Diagram – be sure to draw your final diagram with the initial charge on the ion. An **anion** will always have brackets without any electrons drawn inside and its charge outside on the top right. A **cation** will always have all eight electrons drawn inside the brackets with its charge outside on the top right.



**<u>Lesson 9:</u> Article Analysis**

**<u>Objective:</u> To relate the nature of science to historical influences.**

**<u>Reading:</u> *Niels Bohr and Ernst Rutherford: The Life of a Scientist.***

Scientists are people too and can lead very interesting lives. Just as our childhood and environment mold us into who we are, the life experiences of a scientist mold their character and ability to achieve greatness. Read the attached two passages about two scientists that were instrumental in defining the structure of the atom. Answer the following questions on a separate piece of paper to be collected upon completion of your reading.

1. What type of education did both scientists receive in their youth and how did it influence their careers? Do they have anything in common?

2. The work of one scientist can very much influence the work of another scientist. How does this statement apply to Bohr and Rutherford? Support your answer with examples from the passage.

3. Both scientists lived and worked during Hitler's reign. How did this influence their ability to complete their work? Support your answer with examples from the passage.

4. If Rutherford were still alive, what do you think his reaction would be to the creation of the atomic bomb and Bohr's role in its creation?

**Niels Bohr**
**1885 - 1962**

Niels Bohr was born and educated in Copenhagen, Denmark. He lived, worked, and died there, too. But his mark on science and history was worldwide. His professional work and personal convictions were part of the larger stories of the century.

At the University of Copenhagen, he studied physics and played soccer (though not as well as his brother, who helped the 1908 Danish soccer team win an Olympic silver medal). After receiving his doctorate in 1911, Bohr traveled to England on a study grant and worked under J.J. Thomson, who had discovered the electron 15 years earlier.

Bohr began to work on the problem of the atom's structure. Ernest Rutherford had recently suggested the atom had a miniature, dense nucleus surrounded by a cloud of nearly weightless electrons. There were a few problems with the model, however. For example, according to classical physics, the electrons orbiting the nucleus should lose energy until they spiral down into the center, collapsing the atom. Bohr proposed adding to the model the new idea of quanta put forth by Max Planck in 1901. That way, electrons existed at set levels of energy, that is, at fixed distances from the nucleus. If the atom absorbed energy, the electron jumped to a level further from the nucleus; if it radiated energy, it fell to a level closer to the nucleus. His model was a huge leap forward in making theory fit the experimental evidence that other physicists had found over the years. A few inaccuracies remained to be ironed out by others over the next few years, but his essential idea was proved correct. He received the Nobel Prize for this work in 1922, and it's what he's most famous for. But he was only 37 at the time, and he didn't stop there. Among other things, he put forth the theory of the nucleus as a liquid drop, and the idea of "complementarity" -- that things may have a dual nature (as the electron is both particle and wave) but we can only experience one aspect at a time.

In 1912 Bohr married Margrethe Nørlund. They had six sons, one of whom, Aage, followed his father into physics -- and into the ranks of Nobel Prize-winners. Bohr returned to Denmark as a professor at the University of Copenhagen, and in 1920 founded the Institute for Theoretical Physics -- sponsored by the Carlsberg brewery! Bohr remained director of the institute for the rest of his life, except for his absence during World War II. Bohr's personal warmth, good humor ("Never express yourself more clearly than you can think," he once said), and hospitality combined with world events to make Copenhagen a refuge for many of the century's greatest physicists.

After Hitler took power in Germany, Bohr was deeply concerned for his colleagues there, and offered a place for many escaping Jewish scientists to live and work. He later donated his gold Nobel medal to the Finnish war effort. In 1939 Bohr visited the United States with the news from Lise Meitner (who had escaped German-occupied Austria) that German scientists were working on splitting the atom. This spurred the United States to launch the Manhattan Project to develop the atomic bomb. Shortly after Bohr's return home, the German army occupied Denmark. Three years later Bohr's family fled to Sweden in a fishing boat. Then Bohr and his son Aage left Sweden traveling in the empty bomb rack of a British military plane. They ultimately went to the United States, where both joined the government's team of physicists working on atomic bomb at Los Alamos. Bohr had qualms about the consequences of the bomb. He angered Winston Churchill by wanting to share information with the Soviet Union and supporting postwar arms control. Bohr went on to organize the Atoms for Peace Conference in Geneva in 1955.

In addition to his major contributions to theoretical physics, Bohr was an excellent administrator. The institute he headed is now named for him, and he helped found CERN, Europe's great particle accelerator and research station. He died at home in 1962, following a stroke.

"An expert is a man who has made all the mistakes which can be made, in a very narrow field."

# PEOPLE AND DISCOVERIES

**Ernest Rutherford**
**1871 - 1937**

Ernest Rutherford's family emigrated from England to New Zealand before he was born. They ran a successful farm near Nelson, where Ernest was born. One of 12 children, he liked the hard work and open air of farming, but was a good student and won a university scholarship. After college, he won another scholarship to study at Cambridge University in England -- a turning point in his life. There he met J.J. Thomson (who would soon discover the electron), and Thomson encouraged him to study recently-discovered x-rays.

This was the start of a long, productive, and influential career in atomic physics. Rutherford eventually coined the terms for some of the most basic principles in the field: alpha, beta, and gamma rays, the proton, the neutron, half-life, and daughter atoms. Several of the century's giants in physics studied under him, including Niels Bohr, James Chadwick, and Robert Oppenheimer.

Early on he found that all known radioactive elements emit two kinds of radiation: positively and negatively charged, or alpha and beta. He showed that every radioactive element decreases in radioactivity over a unique and regular time, or half-life, ultimately becoming stable. In 1901 and 1902 he worked with Frederick Soddy to prove that atoms of one radioactive element would spontaneously turn into another, by expelling a piece of the atom at high velocity. Many scientists of the day scorned the idea as alchemy. They stuck with the age-old belief that the atom is indivisible and unchangeable. But by 1904 Rutherford's publications and achievements gained recognition. He was an extremely energetic researcher: in the span of seven years, he published 80 papers.

In 1907 he went to the University of Manchester and with Hans Geiger (of the Geiger counter) set up a center to study radiation. In 1909 he began experiments that were to change the face of physics. He discovered the atomic nucleus and developed a model of the atom that was similar to the solar system. Like planets, electrons orbited a central, sun-like nucleus. Acceptance of this model grew after it was modified with quantum theory by Niels Bohr. For his work with radiation and the atomic nucleus, Rutherford received the 1908 Nobel Prize in chemistry. He was slightly put out, since he was a physicist and felt a bit superior to chemistry! In 1914 Rutherford was knighted.

During World War I, he left his research to help the British Admiralty with problems of submarine detection, but was soon back in the lab. He managed to produce the disintegration of a non-radioactive atom, dislodging a single particle. The particle had a positive charge, so it must have come from nucleus: he called this new particle a proton. With this experiment, he was the first human to create a "nuclear reaction," though a weak one. In 1919 he took over as director of the Cavendish Laboratory. His warm, outgoing personality made him an outstanding mentor to researchers attracted there by his scientific achievements.

He took on more supervision and less direct research as years went by. In 1931 he was made the first Baron Rutherford of Nelson, allowing him to join the House of Lords. He was fiercely anti-Nazi, and in 1933 he served as president of the Academic Assistance Council, established to help German refugees. He would not personally help chemist Fritz Haber, however, who had been instrumental in creating chemical weapons in World War I. Rutherford died two years before the discovery of atomic fission.

"All science is either physics or stamp collecting."

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