Chapter 2: The Atom

An atom is the smallest particle of an element, having the same chemical properties as the bulk element. The first accurate theory explaining the nature of matter was ***Dalton's Atomic Theory***:

1. All matter is composed of atoms, and atoms are indivisible and indestructible. Elements are composed of identical atoms, but the atoms of each element are different and distinguishable from each other.

2. A compound is composed of atoms from at least two elements. The atoms of these elements are present in the compound in simple whole number ratios, i.e., 2 hydrogen atoms for every 1 oxygen atom in water.

3. Chemical reactions involve the rearrangement of atoms in the reacting substances to produce new substances. During this rearrangement, atoms are neither created nor destroyed; they are redistributed.

Dalton proposed his theory in the early 1800's, and since then the first postulate has been slightly modified. We now know that atoms are composed of smaller particles of matter, and that atoms of a given element can show small differences among themselves. The second and third postulates remain unchanged.

Dalton's knowledge of the structure of the atom was limited to knowing that atoms of different elements have different weights. A physical picture of Dalton's atom is a simple, hard sphere.

Structure of the atom

In 1897, British physicist Joseph John (J.J.) Thomson (1856 – 1940) discovered the ***electron***, a very small negatively charged particle. Further experimentation by Thomson and others led to the conclusion that the electron was part of the atom. Electrically neutral atoms are readily found in nature. If atoms contain negatively charged particles, there must be corresponding positively charged material. Thomson proposed a new model of the atom, consisting of a mass of positive charged material with individual electrons stuck into it. This model is called the "***plum pudding***" model, and is shown in Figure 2.1.

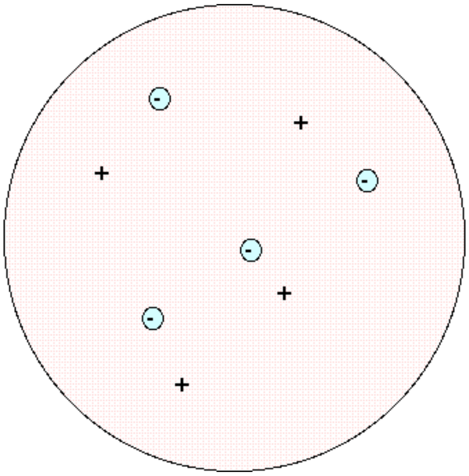

Figure 2.1. Thomson plum pudding model of the atom.

In 1911, British physicist Ernest Rutherford (1871 – 1937), working with British physicist Ernest Marsden (1889 - 1970) and German physicist Johannes "Hans" Wilhelm Geiger (1882 – 1945), performed an important experiment. They set up a gold foil target (similar to aluminum foil, but much thinner and made of gold), and bombarded this target with ***alpha particles***. (Alpha particles are one type of high-energy particle produced when certain types of radioactive materials decay.) The general experiment is illustrated in Figure 2.2.

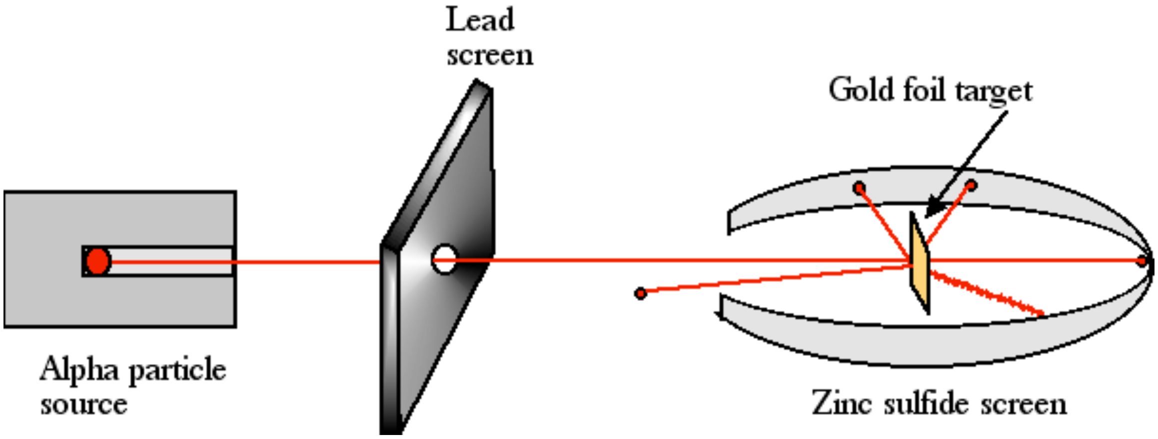

Figure 2.2. Rutherford's gold foil experiment. The alpha particles are shown as red dots, while the paths of the particles are red lines.

The zinc sulfide screen produced flashes of light when struck by the alpha particles, allowing Rutherford to determine how the alpha particles interacted with the target. Most of the particles penetrated the gold foil as if it weren't there. A relatively small percentage of alpha particles were deflected from their original path. However, about 1 alpha particle in 8000 was deflected through an enormous angle, and some of them bounced back from the gold foil target. Rutherford described this phenomenon saying, "It was almost as incredible as if you fired a fifteen inch (artillery) shell at a piece of tissue paper and it came back to hit you."

From this experiment, Rutherford proposed the ***nuclear model of the atom***: an atom contains a relatively small central nucleus containing all of the protons and neutrons. Surrounding the nucleus are the electrons (Figure 2.3). With this model, Rutherford's experimental results made sense. Most of the time, the alpha particles passed through the gold foil, pushing the relatively tiny electrons aside as they went. Sometimes the alpha particle came close to the nucleus and was deflected. Once in a while, the alpha particle "collided" with the heavy gold nucleus, and "bounced off".

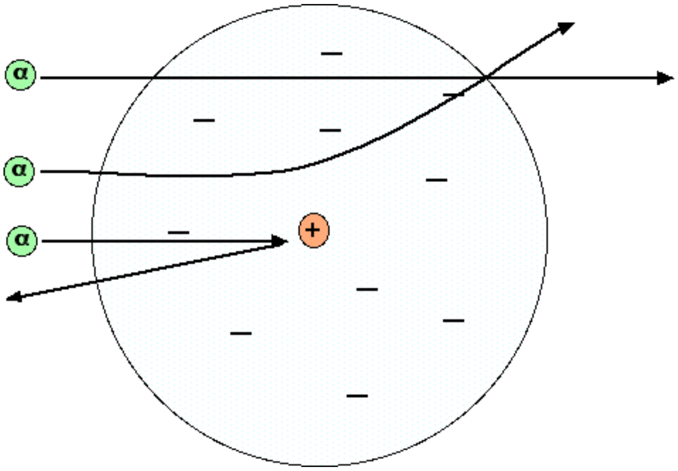

Figure 2.3. Rutherford's nuclear model of atom.

(The alpha particles didn't physically collide with the nucleus. Both the alpha particles and the gold nucleus are positively charged, and the repulsive force prevents physical contact. Nevertheless, the repulsion effect is the same as if a physical collision occurred.)

In 1912, Thomson passed a stream of ionized neon atoms through magnetic and electrical fields. A sketch of this experiment is shown in Figure 2.4. In Thomson's experiment, the magnets and electrodes were arranged perpendicular to each other, while in the sketch they are shown separately for clarity.

Thomson was surprised to see two distinct spots on his photographic plate, and concluded there were two different varieties of neon atoms. This device for deflecting charged atoms was developed into the *mass spectrometer* by one of Thomson's students, English physicist Francis William Aston (1877 – 1945) in 1919. Aston won the 1922 Nobel Prize in chemistry for his contribution to our understanding of the structure of the atom and isotopes. Thomson's experiment was the first demonstration that different forms of the same element existed.

In 1913, English chemist Frederick Soddy (1877 – 1956) observed that several "elements" appeared to occupy the same place on the periodic table. He discussed this interesting observation with a distant relative, Margaret Todd (1859 – 1918), a Scottish doctor and author. She suggested the term *isotope* (Greek, meaning "at the same place"); all isotopes of a given element are placed at the same location in the periodic table of elements.

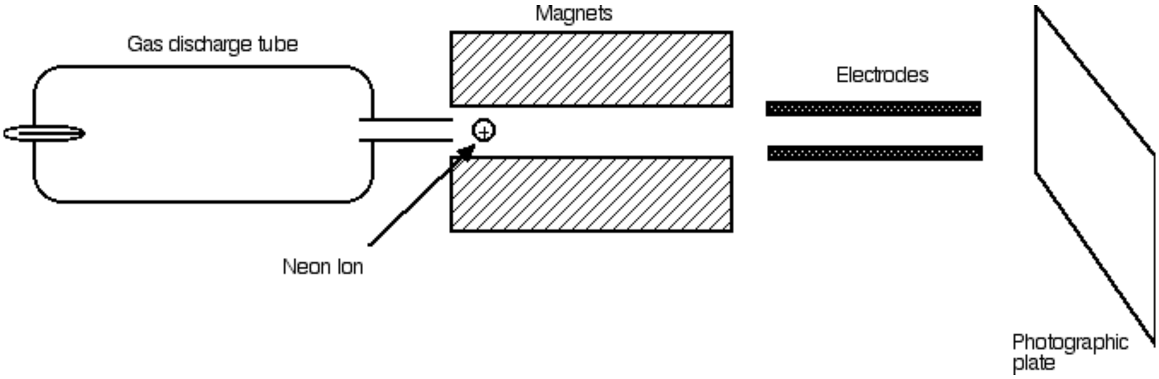

Figure 2.4. Thomson's experiment leading to the discovery of isotopes.

Of course, one important question is, "How are different isotopes of elements formed – what causes the differences?" For various reasons, physicists proposed the existence of another, uncharged particle in the atom. The three major ***subatomic particles*** are:

Electron – (symbol e^-) a very tiny subatomic particle having a negative electric charge.

Proton – (symbol p^+) a tiny subatomic particle having a positive electric charge, located in the center of the atom (nucleus). The electron and proton have equal sized, but opposite polarity, electric charges. The protons mass is ~1830 times larger than the electron.

Neutron – (symbol n) a tiny subatomic particle without an electric charge. The neutron is about the same mass as the proton, and is found in the nucleus with the proton.

Isotopes are varieties of atoms having the same number of protons, but different numbers of neutrons. We know that all elements have isotopes, either naturally occurring or artificial. There are about 2000 total identified isotopes, most of them radioactive and many of them artificial. In neon, Thomson identified two naturally occurring isotopes: one having 10 protons and 10 neutrons, and a second having 10 protons and 12 neutrons.

The existence of isotopes directly affects the values of average ***atomic weight*** shown on the periodic table. These values are normally shown at the bottom of an elements block (Figure 2.5). The atomic weight is the average weight of the common isotopes of the element, adjusted for their abundance.

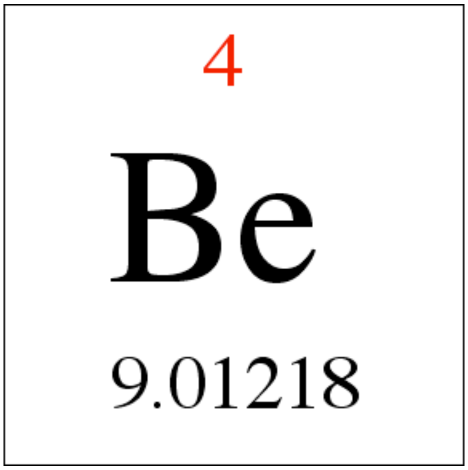

Figure 2.5. Element block for beryllium. Average atomic weight is shown at the bottom of the block.

In the element block for beryllium, the ***atomic number*** is shown at the top. The atomic number is the number of protons in the nucleus of all atoms of each particular element; every atom of beryllium has 4 protons in its nucleus. The number at the bottom of the block is the average atomic weight of beryllium atoms, in **unified atomic mass units** (u) or **daltons** (Da). These units are non-SI units, but are accepted for use with the SI system.

In many cases, the average atomic weight is very close to the total ***nucleons*** (the collective name for neutrons and protons). Beryllium has 4 protons and 5 neutrons in its nucleus, for a total of 9 nucleons. Other elements exhibiting this correlation between atomic weight and total nucleons include helium (He), nitrogen (N), oxygen (O), fluorine (F), and many others.

In other cases, the average atomic weight is not close to the total nucleons. The classic example of this discrepancy is chlorine (Cl) with an average atomic weight of 35.453 Da. The existence of isotopes explains this discrepancy. Let’s imagine that chlorine has two isotopes, one with 17 protons and 18 neutrons (35 nucleons), and one with 17 protons and 19 neutrons (36 nucleons). If we have equal numbers of both kinds of isotopes, then we expect the average atomic weight to be the average weight of the two isotopes: (35 + 36)/2 = 35.5.

However, there is no requirement that isotopes must be present in equal amounts. Imagine I have 1000 chlorine atoms, selected at random. If I have 547 of the 35-nucleon isotope, and 453 of the 36-nucleon isotope, then the average weight of chlorine is:

$$\frac{(35 \times 547) + (36 \times 453)}{1000} = 35.453$$

The specific isotopes of an element, and the abundance of each isotope, determine the average atomic weight of the element.

Students are strongly cautioned **NOT TO BELIEVE WHAT YOU HAVE ASSUMED**! In the chlorine example above, I chose two isotopes, chlorine-35 and chlorine-36, because calculating the average atomic weight using these two isotopes is simple, clear, and understandable. However, the choice of these two isotopes is also wrong! Chlorine has two isotopes, but they are chlorine-35 and chlorine-37!! For every 10,000 chlorine atoms chosen at random, there will be 7735 chlorine-35 isotopes, and 2265 chlorine-37 isotopes. The average atomic weight of chlorine is:

$$\frac{(35 \times 7735) + (37 \times 2265)}{10{,}000} = 35.453$$

The evolution of the modern atomic model

Rutherford's "nuclear model" of the atom lasted for about two years. In 1912, Danish physicist Niels Henrik David Bohr (1885 – 1962) proposed a new model of the atom. This model is called the "***Bohr model***" or the "***planetary model***". Bohr was interested in the atomic spectra of hydrogen, the patterns of emission or adsorption lines produced by hydrogen absorbing or emitting radiation. He chose to work with hydrogen because it is the simplest element.

Bohr explained the absorption or emission of light of specific wavelengths by the movement of electrons between energy levels. Figure 2.6 illustrates this idea. When an electron in the $n = 1$ energy level absorbs a ***quantum*** (*pl. quanta*) of light having exactly the right amount of energy, the electron jumps from $n = 1$ to the $n = 2$ level. The electron could also absorb a different quantum of light, allowing it to jump from $n = 1$ to $n = 3$, or from $n = 1$ to $n = 4$. Only those quanta having the proper amount of energy, allowing the electron to move from one level to another, can be absorbed. The energy is given by Planck's equation, $E = hf$. (E is the energy, h is Planck's constant, and f is the frequency of light). The wavelength of light is related to the frequency by the equation $c = \lambda f$ (where c is the speed of light and λ is the wavelength), therefore $E = hc/\lambda$.

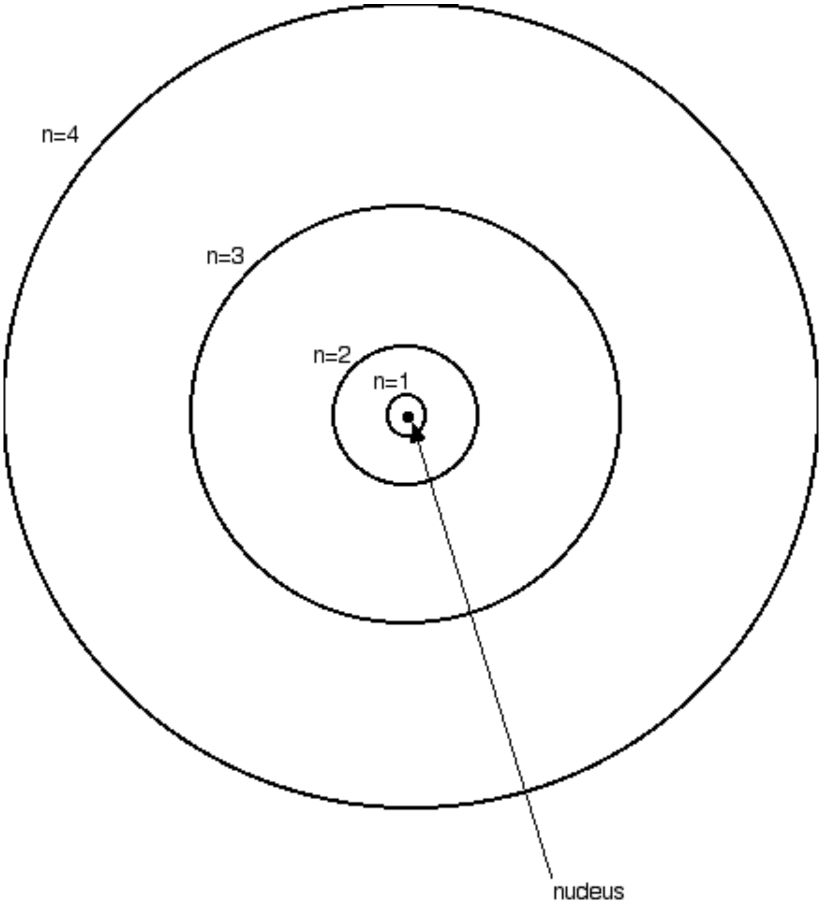

Figure 2.6. Bohr Model of Atom, showing the energy levels. Drawing is not to scale.

When an electron is located in an energy shell higher than it normally occupies, the electron is in an "***excited state***". Eventually, the electron loses energy, and returns to its lowest possible level (called the ***ground state***). How does the electron lose energy? It emits a discrete quantum of energy, having a specific frequency and therefore wavelength. The energy contained in the light photon is the same as the energy difference between two orbitals.

Bohr's model provided a clear and reasonably accurate explanation of the atomic spectra of hydrogen based on the quantization of energy. While the Bohr model is commonly called the *planetary model*, this description is unfortunate because it leaves the impression that negatively charged particles (electrons) are circling the nucleus.

This is impossible. Any object with a curved or circular path must accelerate, even if it has a constant speed. From various experiments and theoretical treatments, it was known that an accelerating electric charge emits radiation. If an electron emits radiation, it must lose energy. When an electron circling a nucleus loses energy, it moves closer to the nucleus. A stable orbit rapidly becomes a spiral, and the electron strikes the nucleus in about 10^{-16} seconds (for a hydrogen atom). When this happens, the opposite charges on the electron and proton cancel out, and a neutron is formed. A literal planetary model of the atom cannot explain why atoms exist!

Bohr knew all this, and was unable to explain precisely how the electrons traveled around the nucleus in his model. The Bohr model was largely an "after the fact" (*ad hoc*) theory, developed to explain the results of spectroscopic experiments. There are serious deficiencies in the model. At best, the Bohr model represents the first application of the principles of quantization of energy to a classical mechanics treatment of the atom. This was an important first step in our understanding of the atom, but it was only the first step.

In his 1924 doctoral thesis, a French physicist and nobleman introduced a new theory of ***electron waves***. The physicist was Louis-Victor-Pierre-Raymond, 7th duc de Broglie (pronounced "broy"), although he is more commonly known as Louis de Broglie (1892 – 1987). His theory of electron waves culminated in the de Broglie hypothesis – any moving object has an associated wave, and this hypothesis led to the development of ***wave mechanics***.

Mathematically, the de Broglie wavelength of a moving object is calculated from:

$$\lambda = \frac{h}{mv}$$

Where *h* is Planck's constant, *m* is the mass of the moving object, and *v* is the velocity.

Physicists generally viewed this hypothesis with a great deal of skepticism. In 1927 two scientists working at the Bell Telephone Laboratories in Murray Hill, New Jersey, fired slow moving electrons at a crystallized nickel target and observed diffraction patterns similar to those shown by X-rays. The scientists were Clinton Joseph Davisson (1881 – 1958), and Lester Halbert Germer (1896 – 1971), both American physicists. The Davisson-Germer experiment provided experimental verification of de Broglie's hypothesis. In 1937, Davisson shared the Nobel Prize in physics with British physicist George Paget Thomson (1892 – 1975), who had independently demonstrated the wave properties of electrons. Incidentally, George Thomson was the son of J.J. Thomson, who had earlier won the Nobel Prize for discovering the electron.

The de Broglie wavelength of an ordinary object moving at ordinary speed is extremely short. Consider a baseball thrown by a major league baseball pitcher. An official baseball weighs between 5 and 5 1/4 ounces (141.75 – 148.83 grams). Let's imagine our ball weighs 145 grams. A professional pitcher can throw a fastball at about 90 mph (about 40 meters/second). What is the de Broglie wavelength of the ball? Substituting these values into the de Broglie equation, and using a value of 6.63×10^{-34} joule-sec for Planck's constant, the resulting wavelength:

$$\lambda = \frac{6.63 \times 10^{-34} \, joule - \sec}{0.145 kg \times 40 \frac{m}{s}}$$

is 1.14×10^{-34} meters. By comparison, a wavelength of visible light is from 4 to 8×10^{-7} meters.

On the other hand, if we look at the wavelength of an electron traveling at 90 mph, we get a result of 1.8×10^{-5} m, within the infrared region of the spectrum. This is because the mass of the electron is extremely small (9.11×10^{-31} kg) compared with the mass of a baseball.

De Broglie argued that each electron orbiting a nucleus was actually a ***standing wave***, also known as a stationary wave. This is a special kind of wave that remains in one particular location. If a standing electron wave has ends that match up, then the wave is stable; if the ends don't match, then the wave isn't stable and destroys itself. This requirement limits the size of electron waves to fixed values. Figure 2.7 shows standing waves circling an atoms nucleus.

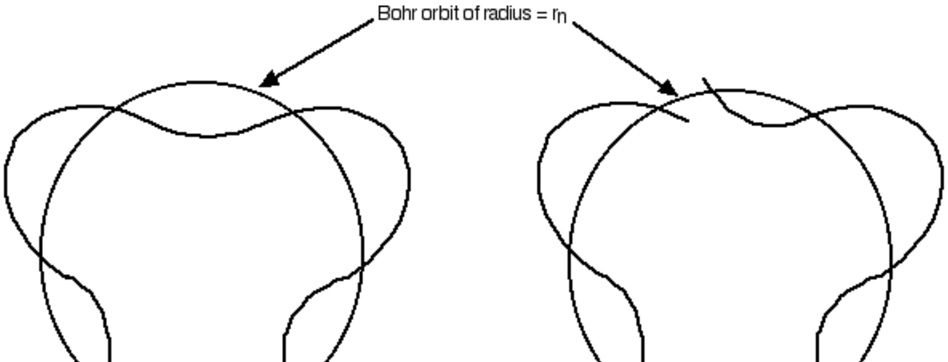

Figure 2.7. Standing waves circling an atoms nucleus.

De Broglie's work provided an explanation for the quantized orbits, but avoids the problem of an accelerating charged object radiating energy. The wave is stationary, and since it is not moving it's not accelerating, and no energy is emitted.

In Figure 2.7, please don't make the mistake of believing that the electron is following the oscillating wave path. Instead, the oscillating wave path represents the electron wave and has been superimposed on the supposed circular orbit of the electron "particle".

In January 1926 an Austrian physicist, Erwin Rudolf Josef Alexander Schrödinger (1887 – 1961), published a paper on wave mechanics, presenting an equation now called the ***Schrödinger equation***. This equation is a version of the ***law of conservation of energy*** applied to electrons in an atom. Many forms of the Schrödinger equation are extremely complex, but a simplified version is:

$$\left(E_k + E_p\right)\Psi = E\Psi$$

Where E_k, E_p, and E represent the kinetic, potential, and total energy respectively, and Ψ is the ***wave function***. The wave function describes the amplitude of an electron wave as a function of time and position. While physicists weren't sure how to interpret Ψ, they eventually concluded that Ψ^2 (the square of the wave function), multiplied by r^2 (the square of the Bohr radius), represented the probability that the electron in a hydrogen atom would be located at a distance, r, from the nucleus. This interpretation led to the development of the "***electron cloud***" model of the atom. A sketch of the electron cloud model for the hydrogen atom is shown in Figure 2.8.

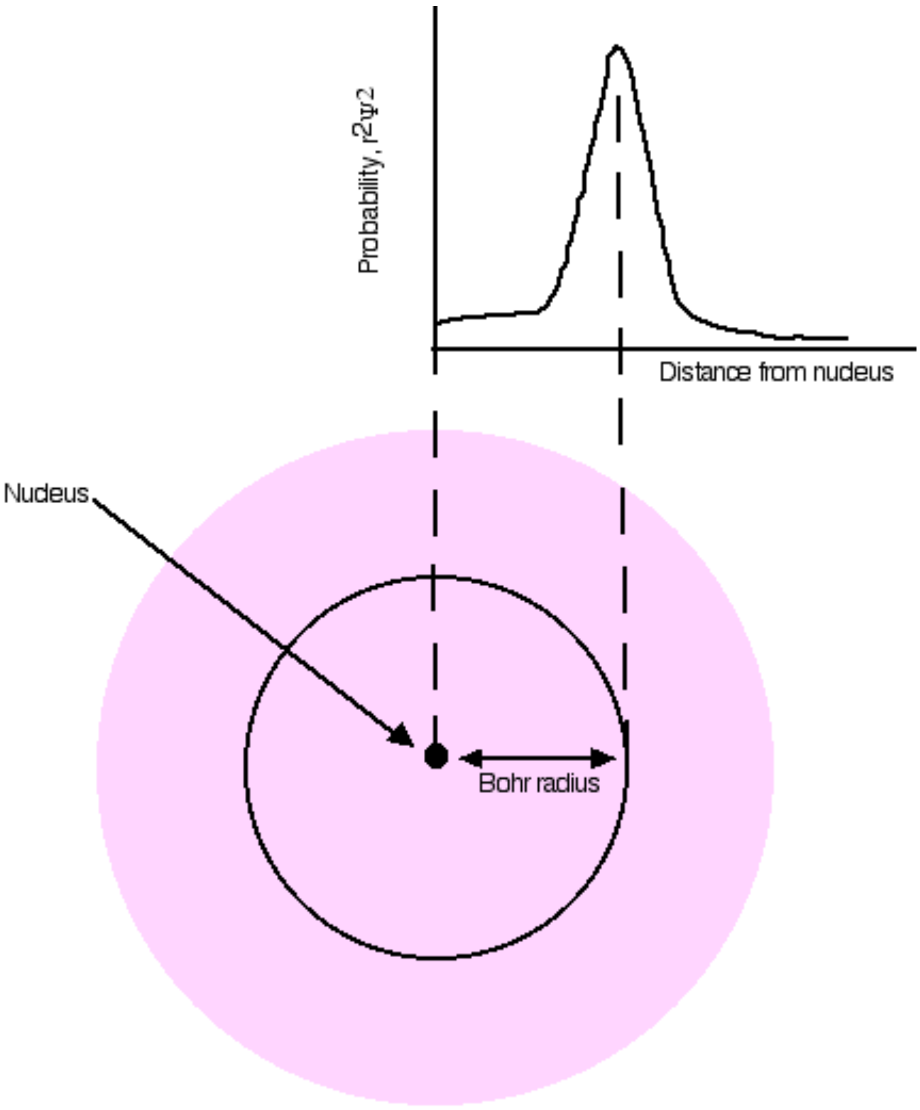

Figure 2.8. Electron cloud model of the atom.

In this model, the Bohr radius represents the most probable location of the electron. However, other locations are permitted, but with diminishing levels of probability. The probability of finding the electron at other radii leads to the idea of a diffuse "cloud" of electron probabilities, shown as the colored region. Schrödinger's equation is considered the most accurate and generally applicable description of electrons around an atom, and in 1933 Schrödinger won the Nobel Prize in physics for "the discovery of new productive forms of atomic theory."

The next person to make a significant impact on our understanding of the atom was German physicist Werner Karl Heisenberg (1901 – 1976). Heisenberg was one of Bohr's students in 1922, and in 1927 Heisenberg stated his famous "***uncertainty principle***". Mathematically, this principle is:

$$m(\Delta v)(\Delta x) \approx h$$

Where m is the mass, Δv is the uncertainty in measuring the velocity, Δx is the uncertainty in measuring the position, and h is Planck's constant.

But why should there be any uncertainty in measuring the velocity or position of an object? The simplest answer is "Whenever you measure a value, you change the value you are measuring."

A common, real world example of how measuring a property changes the value of the property is the measurement of temperature. Consider a cup of hot water. We want to measure the water temperature. How do we do it? The simplest method is to put a thermometer into the cup, wait a few seconds for the mercury or alcohol to reach its highest value on the calibrated scale, and read the corresponding temperature. There is a problem with this method; when we put the thermometer into the water, the thermometer is probably not at the same temperature as the hot water. Since there is a temperature difference between the hot water and the colder thermometer, heat energy spontaneously flows from the hot water to the colder thermometer. As heat energy flows, the temperature of the hot water decreases, and the temperature of the thermometer increases. The temperature we measure is the temperature of the hot water and the thermometer, but it is not the original temperature of the hot water. In measuring the temperature of the water, we have changed the temperature.

On the macroscopic, real world scale of measurements, the magnitude of changes caused by making measurements is generally insignificant and can be safely ignored. On the atomic level, these changes are large and cannot be ignored. The Heisenberg uncertainty principle sets a fundamental limit on how accurately we can measure the velocity and position of an electron.

Working together, Bohr and Heisenberg developed the ***Copenhagen interpretation*** of quantum mechanics in 1927. This interpretation is significantly different from our ordinary, everyday experience and is worth some discussion.

Throughout history, scientists and philosophers have generally viewed the world as being ***deterministic***. Newton believed that if you knew the physical properties of an object, such as mass, velocity, location, and direction of travel, then you could predict infinitely far into the future (or infinitely far into the past) the position, velocity, and location of the object.

The Copenhagen interpretation introduced the idea of uncertainty at the subatomic scale. This interpretation of quantum mechanics says that measurement results are fundamentally non-deterministic. Physics depends directly upon the results of measurements, and therefore questions such as "Where was this particle before I measured its position?" or "Where will this particle be at some future time?"

are meaningless. The act of making a measurement randomly selects one of many possibilities allowed by the wave function, and therefore is essentially a "snapshot" at one particular instant in time.

While most physicists accept the Copenhagen interpretation of quantum mechanics, many physicists object to the non-deterministic nature of the theory. Albert Einstein expressed his objections to this interpretation by asking "Do you really think that the Moon isn't there if you're not looking at it?" Random chance and probability play a significant role in this non-deterministic theory, and Einstein responded to this idea by saying, "I, at any rate, am convinced that He (God) does not throw dice." Bohr answered, "Einstein, don't tell God what to do".

In the 1920s and 1930s, Heisenberg collaborated with Austrian theoretical physicist Wolfgang Ernst Pauli (1900 – 1958) and with British theoretical physicist Paul Adrian Maurice Dirac (1902 – 1984), further developing quantum mechanics. Pauli later developed the ***Pauli exclusion principle***, which is extremely important in modern chemistry, describing how electrons arrange themselves around the nucleus. Dirac shared the 1933 Nobel Prize in physics, and developed the Dirac equation, ultimately predicting the existence of ***antimatter***.

In 1932 Heisenberg won the Nobel Prize in physics for his development of quantum mechanics. During World War II, Heisenberg was part of the German effort to develop nuclear power and nuclear weapons. How willingly he participated in this effort is still the subject of intense historical research and debate.

I have always thought it extremely funny that, in about 123 years, our model of the atom progressed from the small, hard sphere of Dalton to the much more sophisticated small fuzzy sphere of Schrödinger. The story of our evolving understanding of the nature of the atom underscores an important idea in science: not all progress is "linear". Sometimes, you make progress by returning to your starting point. In science, it is not always the destination that is important, but rather the journey and what you learn along the way.

Periodic Table of the Elements

The periodic table of the elements is a graphical arrangement of the known elements arranged by increasing atomic number. The periodic table is probably the most important single tool in chemistry, allowing chemists to predict the chemical formulae of compounds and to distinguish between ionic and nonionic compounds. Knowledge of the underlying patterns affecting the arrangement of the elements allows chemists to make reasonable predictions of a variety of chemical and physical properties.

While there are many versions of the periodic table, the most commonly encountered form has the shape shown in Figure 2.9. A modern periodic table is available as a separate .pdf document (PeriodicTable.pdf). All elements have ***valence electrons***, electrons on the outside of the atom responsible for chemical bonding, occupying specific types of electron orbitals. The elements are grouped together based on having the same number of valence electrons in similar orbitals. The exact nature and shape of these orbitals is not important at this time.

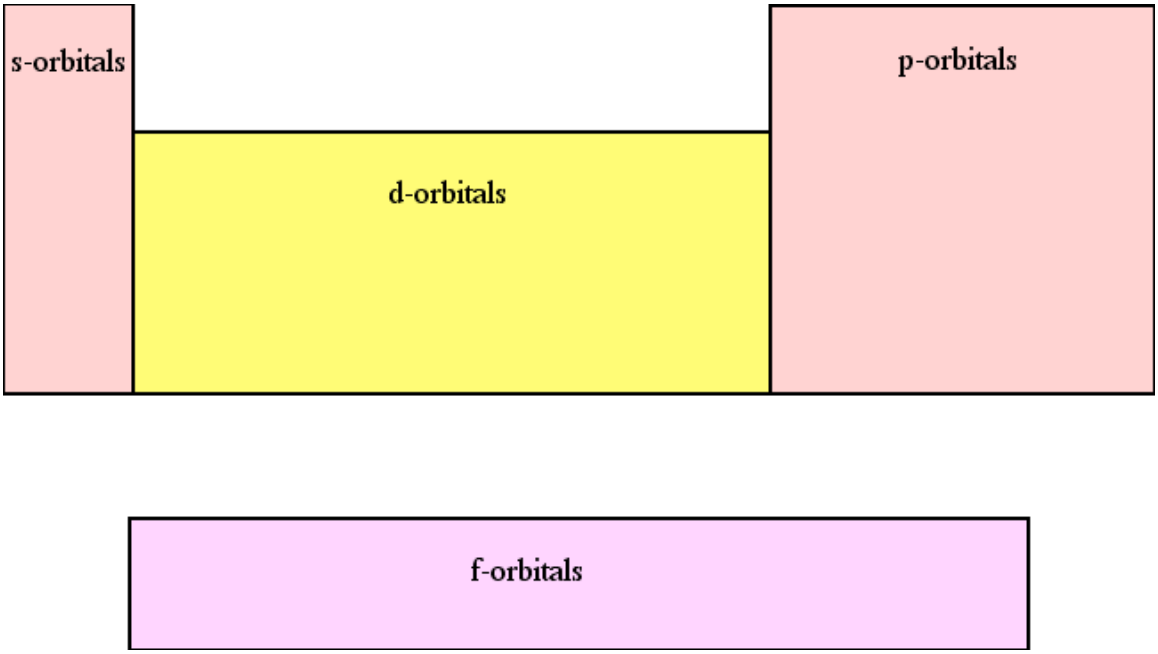

Figure 2.9. Common version of the periodic table.

Each horizontal row of the periodic table is called a ***period***. Each column on the periodic table is called a "***group***". The groups are designated using numbers or a combination of numbers (Arabic or Roman) and letters. At one time, each group was given its own distinctive name. Currently only four groups retain these names and these groups will be indicated when appropriate.

The block labeled "s-orbitals" is composed of two groups called 1A and 2A (Group 1 and 2). All elements in 1A ("Alkali metals") have one electron in their valence shells; all elements in 2A ("Alkaline earths") have two electrons in their valence shells.

The block labeled "p-orbitals" is composed of six groups called 3A through 8A (Group 13 through 18). All elements in 3A have three valence electrons; all elements in 4A have four valence electrons, and so on. Group 7A is named "Halogens", while Group 8A is named "Noble gases". Groups 1A through 8A (Groups 1, 2, 13-18) are collectively known as "representative elements". The group number, using the number-letter format, indicates the number of valence electrons present.

The block labeled "d-orbitals" has a more complicated pattern than that shown by the representative elements. There are ten groups in this block, and they are labeled (from left to right) 3B, 4B, 5B, 6B, 7B, 8B, 1B, and 2B (Groups 3 through 12). Group 8B is composed of three adjacent columns (Groups 8, 9, and 10). The reason for this arrangement is not important for now. Elements in this block are collectively known as "transition metals".

The block labeled "f-orbitals" contains fourteen groups, but these groups are generally not individually numbered. The top row of this block is called the "Lanthanide series" while the bottom row is called the "Actinide series".

As of September 2012, there are 114 known, confirmed elements. In some schools, students have been compelled to memorize the location of the elements on the periodic table, their atomic numbers, average atomic weights, and other information found on the periodic table. This exercise is senseless – chemists write this information down on the periodic table so we don't have to memorize it! It makes as much sense to memorize the periodic table, as it makes to memorize the telephone book.

What is necessary is to learn the names of the elements and their corresponding symbols. Knowing the names of the elements is important when naming compounds. Consider the compound NaCl; it is completely unacceptable to name this compound "N-A-C-L", pronouncing each letter. This compound is sodium chloride; the name of the compound is derived from the names of the elements composing it. (Organic compounds composed predominantly of carbon, hydrogen, oxygen, nitrogen and a few other elements have their own rules of nomenclature. We won't learn these rules in this course.)

Get a set of index cards (3" x 5" will work perfectly); on one side, write the atomic symbol. On the other side, write the element name (spelling is important!). Regularly practice identifying the name of the element from the symbol, and the symbol of the element from the name. This is "brute force" memorization, but it is the only way to master this topic. As a consolation, you can limit your memorization to elements 1-57, and 72-86, if you so desire.

Mendeleyev, Meyer, and the Periodic Table

In 1869, Russian chemist Dmitri Ivanovich Mendeleyev (1834 – 1907) developed the "modern" periodic table of the elements. Several other scientists had searched for some underlying order to the chemical properties of the elements, with varying degrees of success. However, Mendeleyev is credited with discovering the underlying order, and he expressed this order in his periodic law,

> ***"The properties of the elements are a periodic function of increasing atomic weight."***

At the time, nothing was known about protons, neutrons, or electrons.

A German chemist, Julius Lothar Meyer (1830 – 1895) independently developed a similar periodic table, and published the first version of his periodic table in 1864, with a subsequent revision in 1869. Most chemists acknowledge the contributions of Meyer to the development of the periodic table. Mendeleyev's table was considered more complete and more useful, so he is generally given credit for the development of the periodic table.

During Mendeleyev's lifetime, an entirely new group of elements was added to the periodic table, representing the *noble gases*. These elements are helium, neon, argon, krypton, xenon, and radon. Of these elements, only helium was known when Mendeleyev developed his periodic table. The others were discovered between 1894 and 1900. It was a further success of Mendeleyev's work that this entire group of elements obeyed the periodic law.

There are several myths concerning Mendeleyev's development of the periodic table. It is true that he arranged the elements by increasing atomic weight, but he wasn't the first to organize the elements by atomic weight. As a practical matter, there was no other choice but to arrange the elements by increasing atomic weight, since the proton was unknown. Once Mendeleyev had his list of elements, he looked for chemical similarities, and if the chemical properties were at odds with the atomic weights, he reversed the order of elements to reflect the chemical properties.

Other scientists had made similar accommodations. The classic example is that of tellurium (atomic weight 127.60) and iodine (atomic weight 126.9045). Iodine should strictly precede tellurium, but its chemical properties place it under bromine. Mendeleyev did not make this change because "chemical properties override atomic weight". Mendeleyev believed the key to organizing the elements was increasing atomic weight. He also knew the difficulties in determining atomic weights and believed the atomic weights of iodine and tellurium were incorrect; once correct values of atomic weights for these two elements were determined they would naturally occupy the proper place on the periodic table.

When Mendeleyev encountered a position on his table that the next element did not fit, he skipped the position, leaving a hole, and predicted the existence of a new, undiscovered element. This was not unique to Mendeleyev; others had predicted the existence of undiscovered elements based on their versions of the periodic table. Another reason Mendeleyev is given priority over Meyer is because Mendeleyev predicted new elements: Meyer didn't predict new elements to any significant extent.

Many chemistry teachers, and chemistry textbooks, claim that the accuracy of Mendeleyev's predictions led to the wide acceptance of his periodic table. However, in the Davy Medal citation, awarded to Mendeleyev and Meyer for the work developing the periodic table, the accommodation of physical and chemical properties of the elements are given as reasons for awarding the medal. The prediction of new elements isn't mentioned. In fact, Mendeleyev predicted the existence of several new elements that have never been discovered and can never exist. His record of correct predictions is 10 out of 20 (50%), not particularly impressive.

Harry Moseley and Atomic Number

In 1913, a young English physicist named Henry Gwyn Jeffreys Moseley (1887 – 1915) conducted X-ray experiments, dramatically establishing the fundamental nature of the atom. In these experiments, Moseley bombarded various elements with high-energy electrons, producing X-rays. These X-rays were diffracted through crystals, allowing the wavelengths and frequencies of the X-rays to be determined. Moseley discovered a relationship between the frequency of the main X-ray emission line produced, and a simple whole number designated as "Z". This relationship is called Moseley's law, and takes the form:

$$k_1(Z + k_2) = \sqrt{f}$$

Where k_1 and k_2 are constants that depend upon the specific type of X-ray emission line. The whole number, Z, corresponds exactly to the number of protons in the nucleus. Moseley's work provided an experimental justification of the concept of atomic number.

Prior to 1913, the atomic number of an element was purely a formality. Elements were listed by increasing average atomic weight, and the atomic number designated the position of the element on this list. At one time, the atomic number of lithium was 2. When helium was discovered, lithium was the third element on the new list, and its atomic number became 3. Moseley's work demonstrated a more significant, non-arbitrary, meaning for the atomic number. For a series of elements, listed by increasing atomic weight, the value of Z increased in discrete, single unit values.

The only exceptions occurred when there was a hole in the periodic table, representing an undiscovered element. When hafnium was discovered in 1923, it fit between lutetium (discovered in 1907) and tantalum (1802). Similarly, technetium (1937) fits between molybdenum (1781) and ruthenium (1808), and rhenium (1923) fits between tungsten (1783) and osmium (1803). The Z values for the series lutetium-hafnium-tantalum are 71, 72, and 73. The Z values for the other two series follow the same pattern.

Moseley's work provided a way to count protons in the nuclei of atoms, eliminating the possibility of discovering new elements between adjacent elements. There can't be an element between carbon and nitrogen on the periodic table for the simple reason that carbon has six protons in its nucleus, and nitrogen has seven protons. The idea of a fractional proton is meaningless. Moseley's work provided the basis for the only change in Mendeleyev's periodic law in over 140 years:

> ***"The properties of the elements are a periodic function of increasing atomic number."***

Many scientists believe Moseley would have won a Nobel Prize for his work, because of its importance in explaining the fundamental structure of the atom. But Moseley did not win a Nobel Prize, and he never will. In 1914, the First World War started, and Moseley joined the British Army, becoming a signals officer. On August 10, 1915, at the battle of Gallipoli, Moseley was shot and killed by a Turkish sniper.

The Nobel Prize is only awarded to the living.

Homework: Start learning the names and symbols of the elements. The table (back cover) gives the proper symbol and correct name.

Vocabulary. The following terms are defined and explained in the text. Make sure that you are familiar with the meanings of the terms as used in chemistry. Understand that you may have been given incomplete or mistaken meanings for these terms in earlier courses. The meanings given in the text are correct and proper.

Electron
Plum pudding model
Nuclear model
Alpha particles
Proton
Neutron
Atomic number
Unified atomic mass units (Daltons)
Nucleons
Bohr model
Isotopes
Atomic weight
Electron cloud model
Valence electrons
Group
Period

1. Use the periodic table and the list of element names/chemical symbols, and fill in the missing information in the chart below.

Name	Symbol	Atomic #	At. Wt.	Group #
Hydrogen	H	1	1.0079	IA
		17		
Iron				
	Hg			
		82		
	Br			
Vanadium				
	Rn			
		20		
Fluorine				

2. Write a statement of the modern periodic law.

Answers:

1.

Name	Symbol	Atomic #	At. Wt.	Group #
Hydrogen	H	1	1.0079	IA
Chlorine	**Cl**	17	**35.453**	**VII A**
Iron	**Fe**	**26**	**55.847**	**VIII B**
Mercury	Hg	**80**	**200.59**	**II B**
Lead	**Pb**	82	**207.2**	**IV A**
Bromine	Br	**35**	**79.904**	**VII A**
Vanadium	**V**	**23**	**50.9415**	**V B**
Radon	Rn	**86**	**(222)**	**VIII A**
Calcium	**Ca**	20	**40.078**	**II A**
Fluorine	**F**	**9**	**18.998**	**VII A**

2. **The properties of the elements are a periodic function of increasing atomic number.**