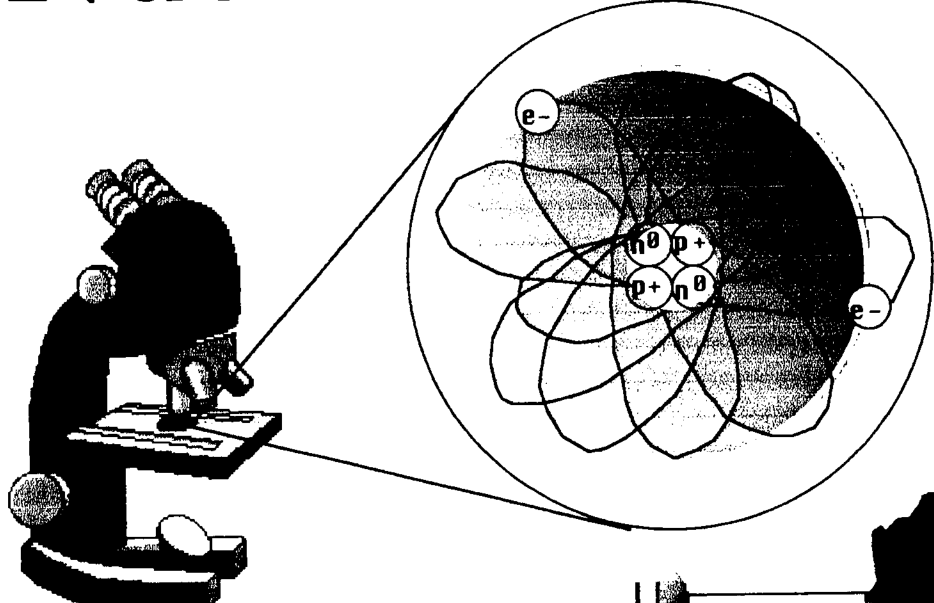
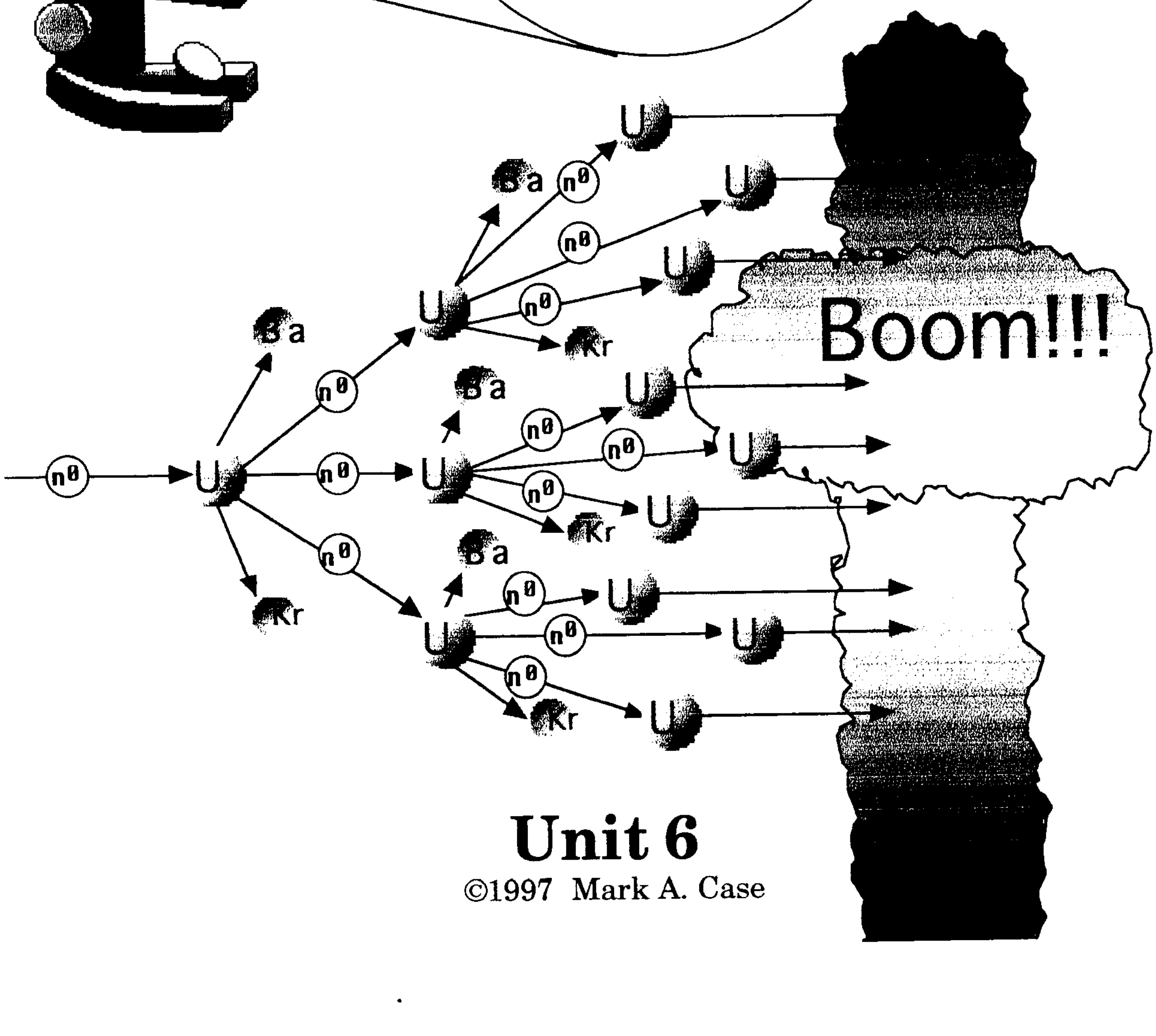
Atomic Structure & Nuclear Chemistry





Unit 6

Unit 6: The Atom...Building Block of the Universe

In order to be able to explain the differences in chemical reactivity of the elements and how they combine to form compounds, it is necessary to have an understanding of atomic structure. What do atoms look like? How are atoms changed during chemical(or nuclear) reactions? Many scientists have labored and struggled to try to answer these questions.

I. Early Atomic Theory - History

For centuries, man has tried to understand the nature of matter and describe its characteristics. As early as 400 B.C., a Greek philosopher named Democritus suggested that all matter is made up of tiny, indivisible properties, called **atoms**. To support the idea that atoms exist even though they cannot be seen, he compared them by an analogy to a single grain of sand versus an entire desert! Individually, atoms are too small to be seen, but in large numbers they become visible. In the 1600s, Isaac Newton and Robert Boyle published articles supporting the atomic nature of elements, but they offered no proof.

Antoine Lavoisier discovered the law of conservation of mass which showed that matter was neither created nor destroyed during chemical reactions. Joseph Proust noted that the compounds formed always had a definite proportion of each element by mass(the law of definite proportions). John Dalton found that the same two elements could combine in different proportions, but then they would form different compounds. For example, hydrogen and oxygen can combine in two proportions, H_2O or H_2O_2. The first compound is water and the second is hydrogen peroxide, a distinctly different substance.

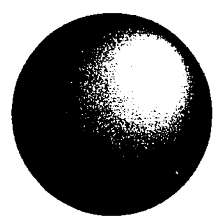
II. Dalton's Atomic Theory

In 1803, John Dalton: summarized the early experimental evidence that supported the Atomic Theory of matter. His theory consisted of 4 postulates:

1. all matter is composed of atoms
*2. all atoms of the same chemical element have similar chemical behavior and have the same mass
3. atoms of different elements have different chemical properties and masses
4. atoms unite in definite ratios to form compounds

*NOTE: We still believe this except for the statement about all atoms of the same element having the same mass. If you substitute the words average mass, then it fits our modern description!

Dalton described the atom as a solid, marble-like object similar to the diagrams to the right. The atoms of different elements varied primarily in size and mass only.



oxygen atom

hydrogen atom

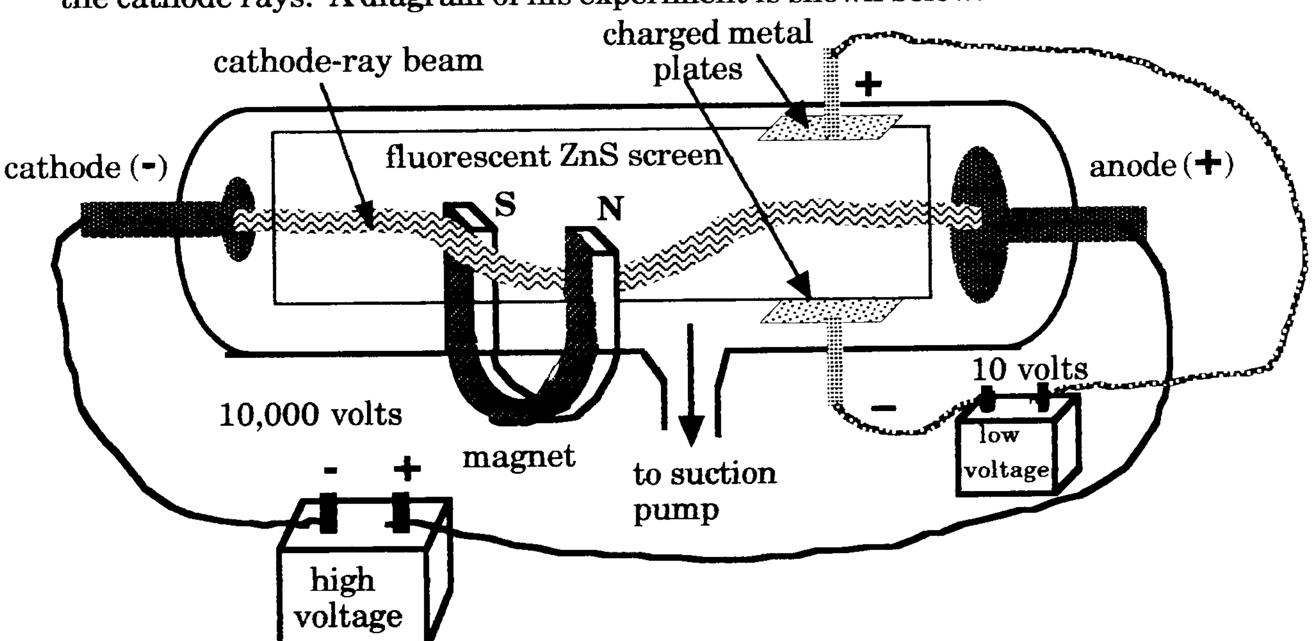
Unfortunately, we cannot see inside atoms to determine their structure so we rely on indirect measurements and observations, instead. For example, there are 2 ways to inventory the contents of a house......one method would be to go inside the house and write down all that you see. A second method would be to blow the house up, then standing outside, you write down everything that comes flying by. Atomic research has been forced to use the latter method.

Problem:
Explain the experimental evidence that supports the existence of atoms.

The classical attractive forces of gravity or magnetism could not explain the great strength with which atoms bond together in a compound. Gravity decreases as the mass of the objects decreases. As an example, the moon has only about 1/5 the gravitational force as the Earth. Since atoms are extremely small, gravitational forces between them are at a minimum. Also, most atoms have no net magnetic field around them, which eliminates this type of force as a possible explanation. The only forces strong enough to bond atoms together are **electrostatic forces**, such as those created when you rub a balloon against your head and stick it on the wall. From experiments with static electricity, it has been determined that <u>oppositely charged particles attract each other</u> and <u>like-charged particles repel each other</u>. Atoms must be capable of becoming electrically charged, but how do they do it?

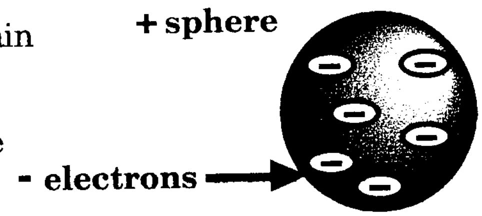
III. Early atomic particle research

A. Cathode ray tubes are glass tubes containing a gas at very low pressure. Metal electrodes are embedded in the glass, and when about 10,000 volts of electricity is connected to the electrodes, the gas in the tube glows. This is called **fluorescence**.(The picture tube in your TV is a cathode ray tube.) An English scientist, William Crookes, noted that the glowing beam originated from the negatively charged electrode(called the cathode). He assumed that the beam(known as a cathode ray) was composed of some kind of electrically charged particles. In 1895, Roentgen noticed that invisible, highly penetrating "x-rays" were also produced by the cathode ray tubes, and he made the first x-ray picture of his wife's hand. Using a cathode with holes in it and a tube filled with hydrogen gas, Goldstein(1886) observed rays traveling in the opposite direction of the cathode rays. These "canal rays", as he called them, were later found to be positively charged protons. In 1897, J.J. Thomson devised an experiment to test the nature of the cathode rays. A diagram of his experiment is shown below:

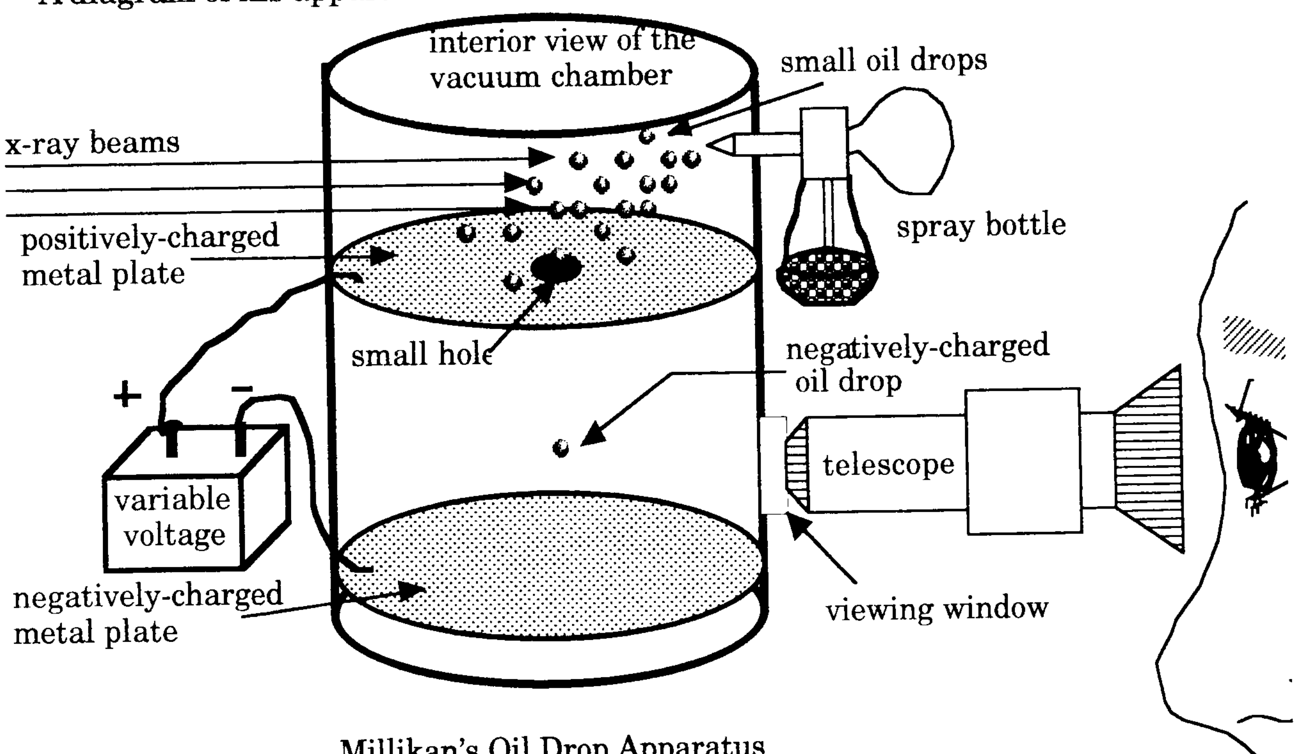


When the cathode ray beam strikes the ZnS screen, the screen glows. JJ Thomson noted that the beam could be bent using a magnet, then straightened out again using an electric field. He found that the amount of deflection of the beam was always the same, no matter what gas was used in the tube or what metals were used as electrodes. By comparing the charge/mass ratio of cathode rays, Thomson showed that the ray was made of high speed, negative particles, which he called **electrons**. Electrons must be a part of the atoms of all elements. Thomson calculated the mass of an electron to be almost 1840 times lighter than that of a hydrogen atom(a proton).

Thomson modified Dalton's model of the atom to explain its electrical nature. He thought the atom was a positively-charged sphere with electrons embedded inside(plum pudding model). The total positive charge MUST equal the total negative charge for the atom.



B. Robert Millikan's Oil Drop experiment(1909) was used to measure the charge and mass of a single electron. A fine spray of oil droplets was injected into an enclosed, vacuum chamber. X-ray beams from a cathode ray tube caused some of the oil droplets to become electrically charged. When a negatively-charged droplet passed through the small hole in the upper metal plate, Millikan could adjust the electrical charges on the plates so that the drop would be suspended in front of the telescope. A diagram of his apparatus is below:



Millikan's Oil Drop Apparatus

Knowing that the charge on the plates just offsets the gravitational force on the droplet, Millikan calculated the charge of an electron to be equal to -1.6×10^{-19} coulombs(a coulomb is the SI unit for electrical charge) with a mass of 9.11×10^{-28} grams. The significance of these values was shown by Thomson, who determined the electrical charge of hydrogen ions in cathode ray tubes to be $+1.6 \times 10^{-19}$ coulombs with a mass of 1.67×10^{-24} grams. Thomson believed the hydrogen ion was a single positively-charged particle, called the **proton**, which has an equal but opposite charge to that of an electron. An electron can be represented by the symbol e^-, and a proton by either p or H^+. During chemical reactions, atoms gain or lose electrons to form electrically charged ions. Oppositely charged ions then attract to form compounds that are electrically neutral.

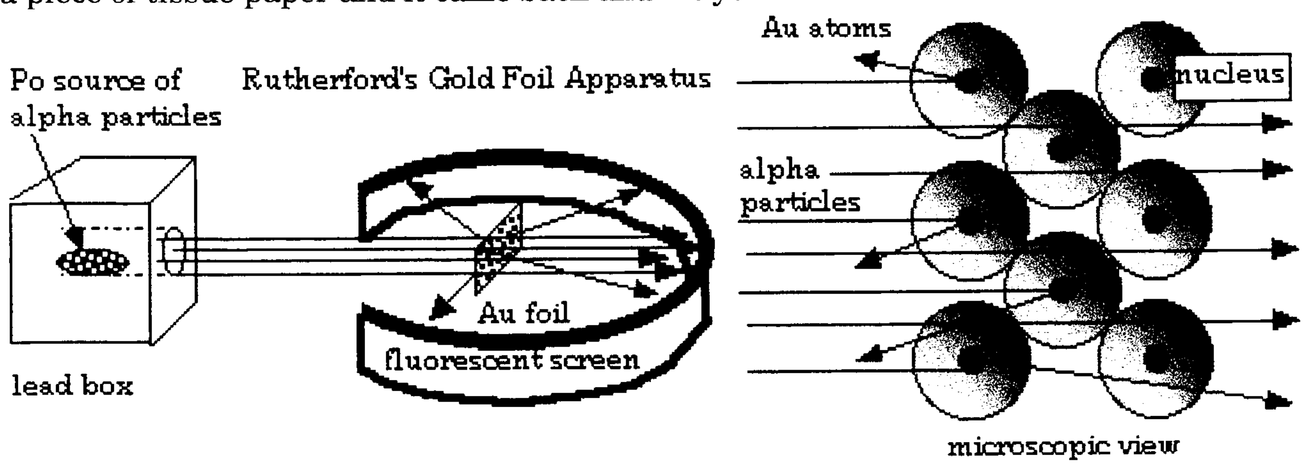
Problem:
Describe the electrical forces that hold atoms together in a compound?

IV. The Discovery of Radioactivity: Historical Glimpses of Nuclear Science

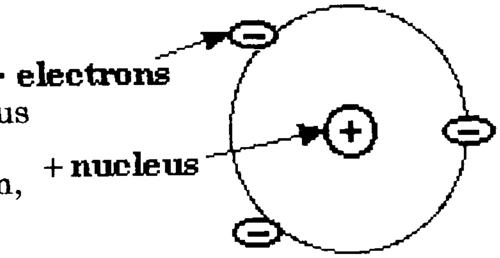
In 1896, Henri Becquerel accidentally discovered that uranium ore gave off some kind of rays which had the ability to expose a photographic plate covered by black paper. These rays were later identified as products of **radioactivity**, in which certain atoms spontaneously break down to emit particles and very penetrating rays(similar to x-rays). The study of radioactivity became the object of intense scientific curiosity. In 1898, Marie and Pierre Curie discovered the radioactive elements Po and Ra(they isolated only 0.1g from tons(!!!) of pitchblende ore). From 1896-1905, Crookes, Becquerel, Rutherford, Soddy, Dorn, et al discovered that radioactivity is produced when atoms of one element are changed into the atoms of a different element(called **transmutation** of atoms) during a nuclear change. Radiation effects chemical substances and can also cause biological damage! There are three types of radioactive emissions that can occur:

1. **alpha particles** - high speed helium ions that have low penetrating ability, but cause the most damage once inside the body; travel 3-4 inches before being stopped by molecules in the air.
2. **beta particles** - high speed electrons emitted by the atom that have moderate penetrating ability; less damaging than alpha particles due to their extremely small mass; travel 12-15 inches before being stopped by molecules in the air..
3. **gamma rays** - bursts of pure energy similar to x-rays, but with greater penetrating ability; can pass right through your body with minimal damage, except in large amounts. Cosmic gamma rays can travel millions of miles before being absorbed by molecules in the air or soil.

In 1911, Ernest Rutherford developed an experiment to test Thomson's atomic model. He directed a beam of alpha particles from radioactive polonium at a thin sheet of gold. A circular fluorescent screen was placed around the foil. This screen produced a faint flash of light at the points where alpha particles would strike it. Rutherford thought the alpha particles would pass directly through the gold foil, but to his surprise, a few bounced back. Later he described this event "as if you fired a 15-inch(artillery) shell at a piece of tissue paper and it came back and hit you!"



Rutherford concluded that most of the mass of the atom is located in a small, dense, positively-charged center, called the **nucleus.** Electrons orbit the nucleus at the outer edges of the atom. The diameter of the nucleus is about 1/10,000 of the diameter of the atom, which means the atom is mostly empty space.



By studying the x-rays produced by cathode ray tubes with different metal anodes, Henry Moseley(1913) concluded that all the atoms of a particular element have the same number of protons. The number of protons is known as the **atomic number**(represented by the symbol Z), and it determines the identity of the element. For atoms, the number of protons = number of electrons. For ions, the size of the electrical charge = number of protons - number of electrons.

In 1932, James Chadwick identified another particle in the nucleus that had about <u>the same mass as a proton but no electrical charge</u>, which he called a **neutron**. These particles were produced during a nuclear reaction between beryllium atoms and alpha particles(helium ions) that formed carbon atoms and the stray neutrons. Most of the mass of an atom is due to the total number of protons and neutrons in the nucleus, which is called the **mass number**(symbolized as A). The number of neutrons in the atoms of a particular element can vary, giving them slightly different masses. The atomic mass(sometimes called atomic weight) value on the periodic table is actually a weighted average of the masses of all of the atoms of an element. These different atoms of the same element are called **isotopes**, which can be identified by the element's name and mass number.

For example, three isotopes of hydrogen exist. All of the atoms have one proton and one electron, but most hydrogen atoms have no neutrons, some atoms have one neutron and a very few atoms have two neutrons. They are identified as hydrogen-1, hydrogen-2 and hydrogen-3, respectively.

A shorthand method to represent the composition of any atom or ion is called **isotope notation**, as shown below:

mass number → 23 ← electrical charge 1+

$$^{23}_{11}\mathrm{Na}^{1+}$$

atomic number → 11

This notation indicates that the given element is a sodium ion with 11 protons(equal to the atomic number). The number of neutrons in this isotope = mass number - atomic number = 23 - 11 or 12 neutrons. Since there is a 1+ electrical charge, this is an ion of sodium, not a neutral atom. The number of protons - number of electrons will equal the electric charge. Therefore, 11p - e^- = 1+, and there are 10 electrons present. One electron had to be lost by the atom during a chemical reaction. The notation for a sodium-23 atom would be similar to that of the ion, except the superscript position for the electrical charge would be left blank to show no charge is present.

mass number → 23 ← electrical charge, NONE

$$^{23}_{11}\mathrm{Na}$$

atomic number → 11

Problems:

Complete the following chart, using a periodic table for reference:

isotope name	isotope notation	atomic number	mass number	# of proton	# of neutron	# of electron	electric charge
calcium atom					20		
	$^{15}_{7}N$						
sulfide ion			36				
	$^{54}_{26}Fe^{3+}$						
				53	74	54	
			65			28	2+

V. Atomic Mass of an Element

In 1919, Aston's mass spectrometer was used to identify the different isotopes of an element and their relative abundance. The atomic mass can then be calculated by determining a weighted average for the isotopes present. The gaseous element sample is first ionized by an electric field and then propelled into a chamber containing positively-charged plates. The isotopes in the beam are bent from their paths as they pass through the chamber. Lightweight particles are deflected the most, heavier particles the least. A sample of ionized hydrogen is shown in the mass spectrometer below:

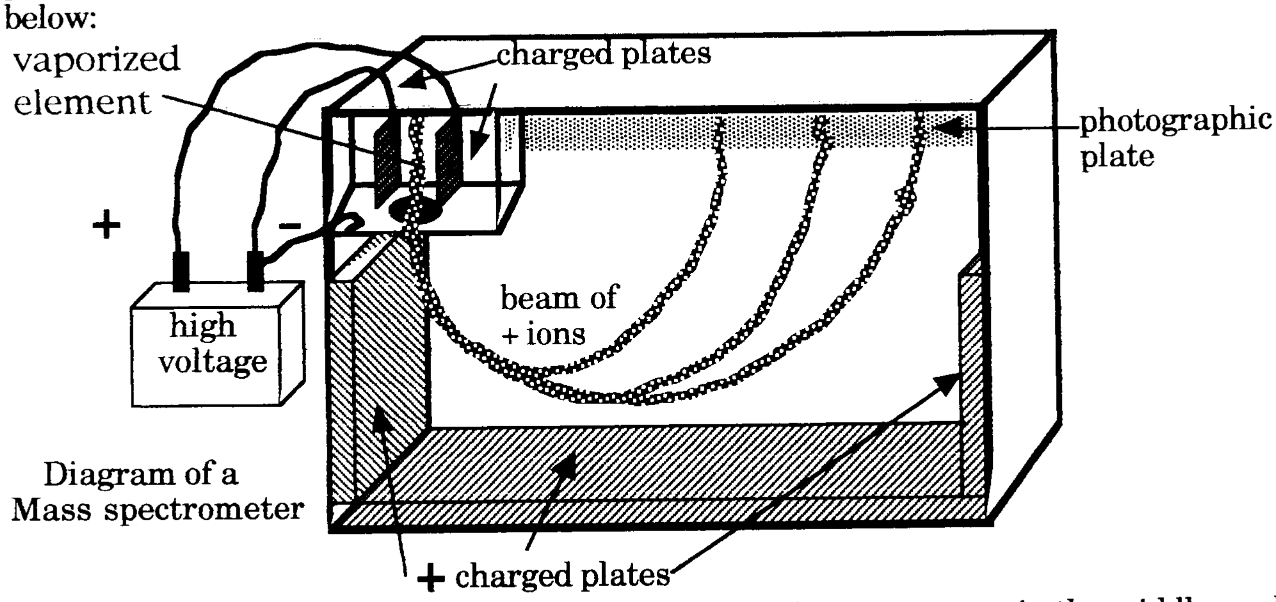


Diagram of a Mass spectrometer

Hydrogen-1 atoms would be those on the left side, hydrogen-2 atoms in the middle, and hydrogen-3 atoms on the right side.

Sample Problem:

Data from the previous experiment shows that hydrogen-1 is 99.85% of the sample with a mass of 1.0078 units, hydrogen-2 is 0.15% of the sample with a mass of 2.0165 units, and hydrogen-3 is 0.0012% of the sample with a mass of 3.0252 units.
Calculate the weighted average by multiplying the percent composition by the relative mass for each isotope, then add these products together. This is the atomic mass for hydrogen.

hydrogen-1: 99.85% x 1.0078 units = 1.00629 units
hydrogen-2: 0.15% x 2.0165 units = 0.00302 units
hydrogen-3: <u>0.0012% x 3.0252 units = 0.000036 units</u>
= 1.009346 units, rounded to 1.0 unit(usually expressed as g/mole)

Problems:

1. Neon has 2 isotopes. Ne-20 (with an abundance of 90.000%) has an exact mass of 19.981 u and Ne-22 (with an abundance of 10.000%) has an exact mass of 21.970 u. Calculate the average atomic mass of Neon (which we could compare to the mass on a periodic chart)

ANS: 20.180 u

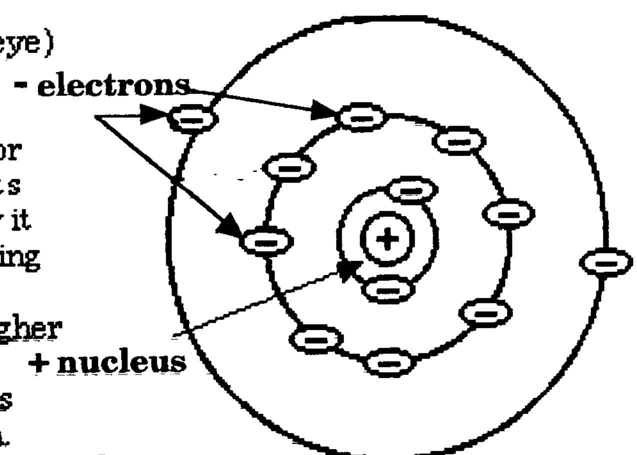
2. What is the average atomic mass of hafnium if, out of every 100 atoms, 5 have mass of 176, 19 have mass 177, 27 have mass 178, 14 have mass 179, and 35 have mass 180?

ANS: 179 u

VI. The Energetic Electrons

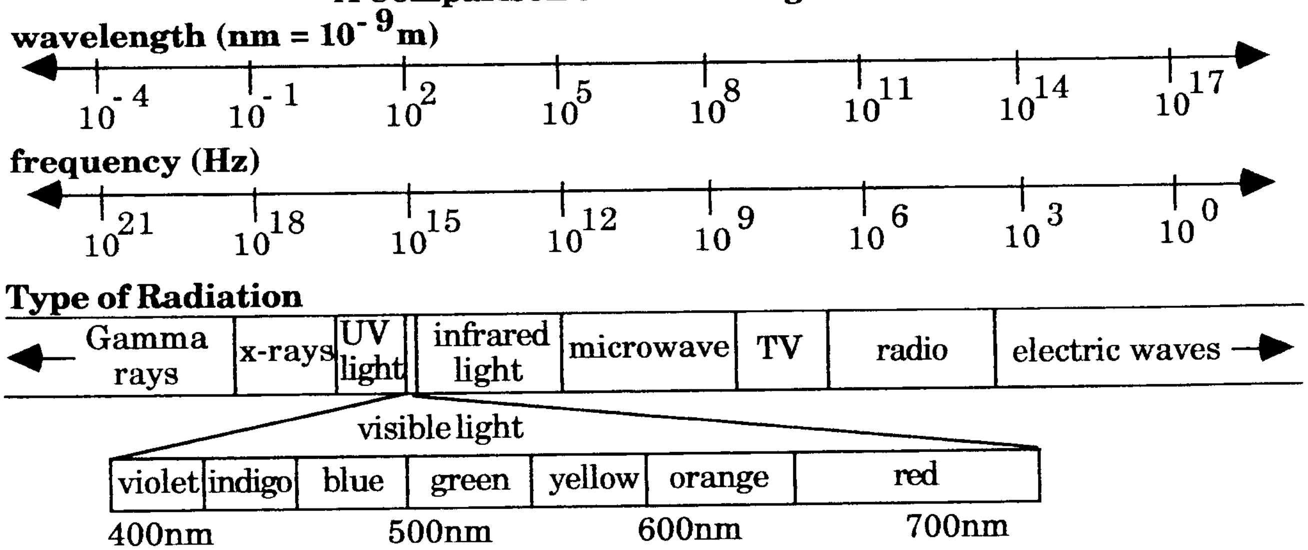
If oppositely charged particles attract each other, then what would keep the electrons from collapsing into the nucleus? Centrifugal force, similar to that which keeps the planets in orbit around the sun, was first suggested. However, experiments showed that charged moving particles (like electrons) would lose energy. This means they should slowly spiral down into the nucleus and collapse the atom, but this doesn't happen. In 1913, Neils Bohr created a new branch of atomic physics to try to resolve this problem.

Bohr proposed the first solar system(bull's eye) model of the atom. The electrons orbit the nucleus with a fixed amount of energy at definite distances in what are called **shells** or **energy levels**. The farther the electron gets from the nucleus, the more potential energy it has. If an atom becomes <u>excited</u> by absorbing energy from a flame or electric current, the electrons jump away from the nucleus to <u>higher</u> energy levels. This excess energy will be released as colors of light when the electrons drop back down to lower energy levels again.
He calculated that each characteristic energy level change in a hydrogen atom produced a specific color of light.



Each element produces very specific lines of color when they are excited by an energy source. These **line spectra** can be used as a fingerprint for the atoms of a particular element. The light released is only one form of **electromagnetic radiation(EMR)** given off when the excited atoms lose their excess energy and return to their lowest energy state, called the **ground state**. Other forms of EMR released include radio, microwave, infrared, ultraviolet, x-rays and gamma rays. All electromagnetic radiation travels at the speed of light, which is 3.00×10^8 m/s or 186,000 miles/s. The radiation moves as "waves" of energy with a specific **frequency**(number of waves per second) and **wavelength**(distance between the peaks of two successive waves). Mathematically, the speed of light is equal to the wavelength x the frequency of the radiation. The symbolized equation is: $c = \lambda f$, where **c** is speed of light, λ (lambda) is wavelength in m, **f** is frequency in hertz. As the wavelength of EMR increases, the frequency decreases, as shown in the chart below:

A Comparison of Electromagnetic Radiation

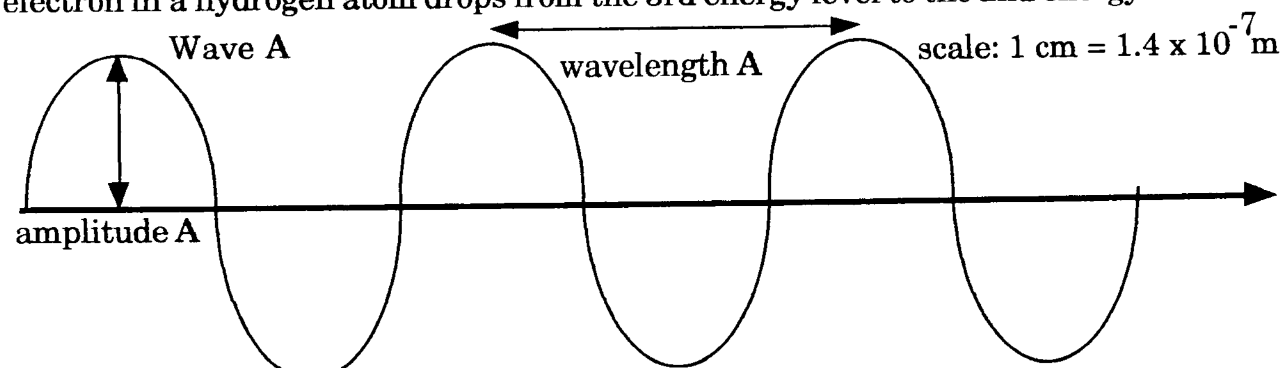


Max Planck suggested that when an electron dropped to a lower energy level, the radiant energy was given off in little packets, called **photons**. The energy of the photon was directly related to the frequency of the light by the equation,

energy of photon = Planck's constant x frequency.

The symbolized equation is: $\mathbf{E = hf}$, where **E** is the energy in J/photon, **h** has a constant value of 6.63×10^{-34} J/Hz•photon, and **f** is frequency in hertz.

Sample Problem:
The diagram below represents a segment of the red light that is produced when an electron in a hydrogen atom drops from the 3rd energy level to the 2nd energy level.



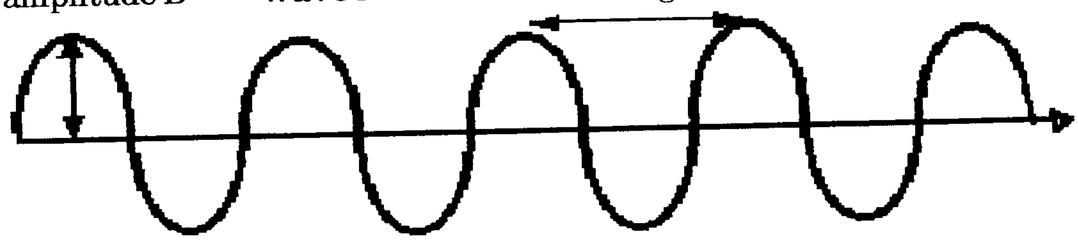
The amplitude is the maximum displacement of the wave from the base(middle) line, which measures about 2.0 cm (scaled = 2.8×10^{-7}m). The wavelength of 4.8 cm (scaled = 6.7×10^{-7}m) is the distance from the crest of 1 wave to the crest of another. Since frequency x wavelength = the speed of light, the frequency = 4.5×10^{14} waves/s or 4.5×10^{14} hertz(Hz). The calculation used to determine the frequency is:

frequency = (3.0×10^{8} m/s) ÷ (6.7×10^{-7} m/wave).

Using Planck's equation, the energy of this photon is equal to:

(4.5×10^{14} Hz/s) x (6.63×10^{-34} J/Hz•photon) = 3.0×10^{-19} J/s•photon.

Problem:
The diagram below represents a segment of the violet light that is produced when an electron in a hydrogen atom drops from the 6th energy level to the 2nd energy level. Compare the amplitude, frequency, wavelength, speed and energy of this photon compared to the photon of red light.



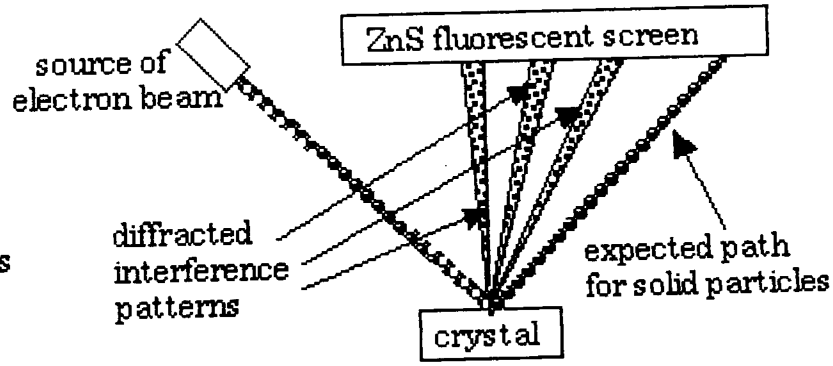
ANS: amplitude = 1.4 cm
wavelength = 2.9 cm
scaled = 4.1×10^{-7} m
frequency = 7.3×10^{14} Hz/s
speed = 3.0×10^{8} m/s
energy = 4.8×10^{-19} J/photon

In 1923, Louis de Broglie suggested that particles, such as electrons, also have a wave-like motion. By relating Einstein's equation for matter and energy($E = mc^2$) and Planck's equation for energy and frequency($E = hf$), de Broglie predicted that the wavelength of a particle is inversely proportional to the product of its mass x velocity.

Mathematically, the expression is written as: $\lambda = h / mv$

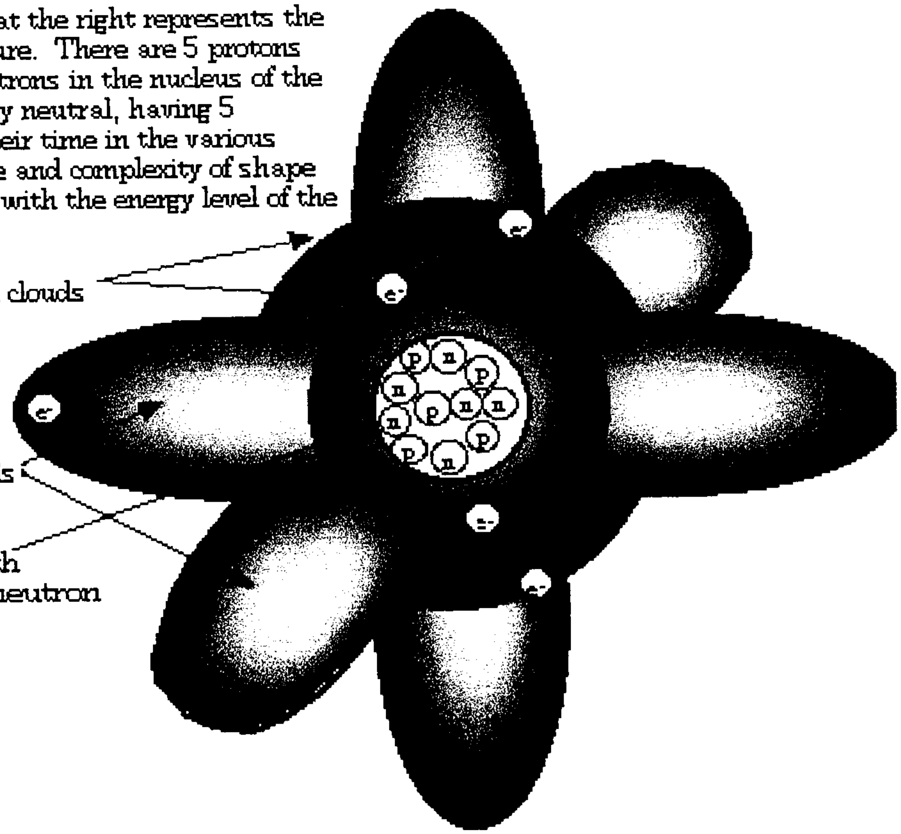
Electrons have a **wave-particle duality**, exhibiting properties of both waves and particles.

Davisson and Germer were the first to demonstrate the electron's wave nature in 1927 by bouncing an electron beam off of a crystal. Instead of the particles simply bouncing off like billiard balls, the beam was diffracted(bent) and created interference patterns similar to a beam of light.



Since the electrons in an atom have different energy levels(Bohr's model) and a wave-like motion around the nucleus, they also must have different frequencies, wavelengths and amplitudes. Werner **Heisenberg's uncertainty principle** pointed out that it is impossible to know both the exact position and momentum of an electron at any given time. Erwin Schodinger treated the electron as a wave to develop an equation which could predict its most probable positions. By solving Schrodinger's equation, the position of an electron can be described as an **electron cloud**. The size and shape of the electron clouds will vary, depending on the energy of the electrons that occupy that particular space. This is the present theory of atomic structure.

The model of a boron-11 atom at the right represents the present theory of atomic structure. There are 5 protons (atomic number = 5) and 6 neutrons in the nucleus of the isotope. The atom is electrically neutral, having 5 electrons that spend most of their time in the various electron cloud regions. The size and complexity of shape for the electron cloud increases with the energy level of the electron. (not drawn to scale)



Problem:
Explain what happens to an atom's electrons when it absorbs energy from a flame or electric current and then releases the energy as colored lines of visible light(forming a line spectra).

VII. Subnuclear particles and Antimatter

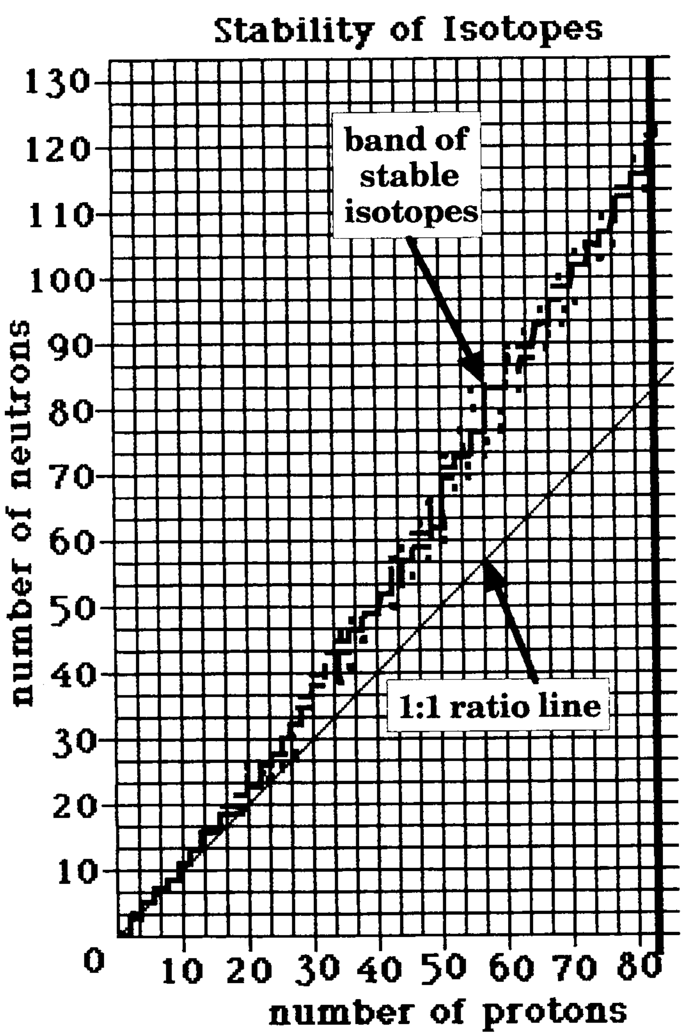
The idea that all atoms were made up of only protons, neutrons and electrons suggested a great simplicity in the structure of matter. In 1932, Lawrence and Livingston build the first cyclotron for smashing high energy particles and radiation into a target nucleus. As scientists began studying the results of these experiments, hundreds of new subatomic particles were discovered. These particles are unstable and do not exist in ordinary matter, but they are common in nuclear reactions, such as those occurring in the stars. Physicists have grouped the particles into two families:

1. <u>leptons</u> (lightweight particles) are truly elemental particles, which include the electron, neutrinos(have an infinitely small mass and no electrical charge), and mesons.
2. <u>hadrons</u> (heavier particles) are composed of smaller particles called quarks. They are subject to the strong forces that hold the nucleus together. There are 6 different quarks, three have a +2/3 charge and three have a -1/3 charge. Protons and neutrons are each made up of 3 quarks.

For every particle of matter, there is an antiparticle with an opposite charge. When matter and antimatter collide, they annihilate each other and release large amounts of energy. The search for gaining a clear understanding of atomic structure continues. The nuclear age was born in 1939 when Otto Hahn, Fritz Strassman and Lise Meitner first split the uranium atom. A year later, Glenn Seaborg and team produced the first transuranium elements, Np and Pu, by bombarding uranium with slow moving neutrons. In 1942, Fermi and his team created the first controlled nuclear chain reaction (atomic pile) under the University of Chicago stadium. From 1940-1945, Oppenheimer and team worked on the Manhattan Project to produce uncontrolled nuclear chain reactions and the atomic bomb used in Japan. The first practical use for nuclear power occurred in 1951 when the first breeder reactor in Idaho produced electric power.

VIII. Nuclear Chemistry

Different isotopes have different nuclear properties but chemically behave the same. Chemical reactions are dependent on oxidation numbers, nuclear reactions are not. Transmutation(change of one element into another) for nuclear reactions must have a conservation of mass numbers and atomic numbers, and there are comparatively larger energy yields for nuclear reactions. Unstable nuclei will undergo characteristic nuclear changes and emit radioactivity in order to become more stable.



The dots on the graph at the left represent the combination of protons and neutrons which form stable nuclei. Notice that the lighter elements are very stable with a 1:1 neutron to proton ratio, but as the size of the nucleus increases, the atoms need more neutrons in order to maintain stability. Scientists have also discovered that an even number of both protons *and* neutrons produces a more stable nucleus. Any unstable nuclei will undergo some type of nuclear change in order to achieve a better balance of protons and neutrons.

All atoms with an atomic number greater than 82 are simply too large to be stable, and they typically emit alpha particles.

Isotopes that are found on the left side of the band of stability have too many neutrons, and emit beta radiation to become more stable.

Isotopes to the right of the band with an atomic number ≤ 82 have too many protons. They will emit positrons or undergo electron capture to become more stable.

A. Types of radioactivity and accompanying reactions

Natural Nuclear Decay:

1. Alpha radiation consists of ${}^{4}_{2}He^{2+}$ with a net loss of two protons and of two neutrons from the source element. All elements with more than 82 protons are unstable due to their large size and will emit alpha radiation. These elements will undergo a series of nuclear changes until a stable nuclei is produced, which will normally be an isotope of lead.

   Example: $${}^{238}_{92}U \longrightarrow {}^{234}_{90}Th + {}^{4}_{2}He$$

2. Beta radiation consists of ${}^{0}_{1-}e$ (an electron) with a net loss of a neutron and gain of a proton by the source element. A neutron is converted into a proton and an electron by unstable isotopes that have too many neutrons.

   Example: $${}^{14}_{6}C \longrightarrow {}^{14}_{7}N + {}^{0}_{1-}e$$

3. Gamma radiation, γ, consists of high energy photons, with no change in atomic number nor in mass number (usually accompanies alpha and beta radiation)

   Example: $${}^{60}_{27}Co \longrightarrow {}^{60}_{27}Co + {}^{0}_{0}\gamma$$

Synthetic Nuclear Decay:

When target atoms are bombarded by particles or radiation, unstable nuclei can form, then quickly decay to try to become more stable. Normally, these unstable nuclei are not found in nature, and the changes are referred to as artificial transmutations.

1. Positron emission consists of ${}^{0}_{1+}e$ (a type of antimatter which is an electron with a positive charge) produced by unstable isotopes that have too many protons. A proton is converted into a positron and a neutron .

   Example: $${}^{25}_{13}Al \longrightarrow {}^{25}_{12}Mg + {}^{0}_{1+}e$$

2. Electron capture occurs when an orbital electron is pulled into the nucleus and combines with a proton. This results in the emission of x-rays with a net loss of a proton and a gain of a neutron.

   Example: $${}^{18}_{9}F + {}^{0}_{1-}e \longrightarrow {}^{18}_{8}O$$

3. Spontaneous emission occurs when unstable isotopes formed during nuclear bombardment emit protons and neutrons in an attempt to become more stable.

   Example: $${}^{14}_{7}N + {}^{4}_{2}He \longrightarrow {}^{17}_{8}O + {}^{1}_{1}H$$

Example: $^{12}_{6}C + ^{244}_{96}Cm \longrightarrow ^{254}_{102}No + 2\,^{1}_{0}n$

4. Fission can occur when slow moving neutrons are used to bombard a heavy element to begin the process. The large nuclei splits into two smaller elements and a few extra neutrons are produced. NOTE: Hypothetically, if one "shoots" a neutron at an atom, it might (a)pass directly through (b)artificially transmutate (c)be captured if it is a slow (thermal) neutron.
U-238 + a slow neutron --> transuranium elements (as Pu for the breeder reactor)
U-235 + a slow neutron --> fission (as in the next example)
Example: $^{235}_{92}U + ^{1}_{0}n \longrightarrow ^{90}_{37}Rb + ^{144}_{55}Cs + 2\,^{1}_{0}n$

5. Fusion involves "fusing" or combining two small nuclei into one larger element. Fusion requires extreme temperatures and pressures to occur and is common only in the stars.
Example: $2\,^{2}_{1}H \longrightarrow ^{4}_{2}He$

Problems:

Balance the following nuclear reactions and state what type of nuclear change occurs:

1) $^{218}_{84}Po \longrightarrow$ ____ $+ ^{4}_{2}He$

2) $^{3}_{1}H + ^{1}_{1}H \longrightarrow$ ____

3) $^{242}_{95}Am \longrightarrow$ ____ $+ ^{0}_{1-}e$

4) $^{235}_{92}U + ^{1}_{0}n \longrightarrow ^{142}_{57}La + 3\,^{1}_{0}n +$ ____

5) $^{13}_{7}N \longrightarrow ^{0}_{1+}e +$ ____

6) $^{0}_{1-}e + ^{30}_{15}P \longrightarrow$ ____

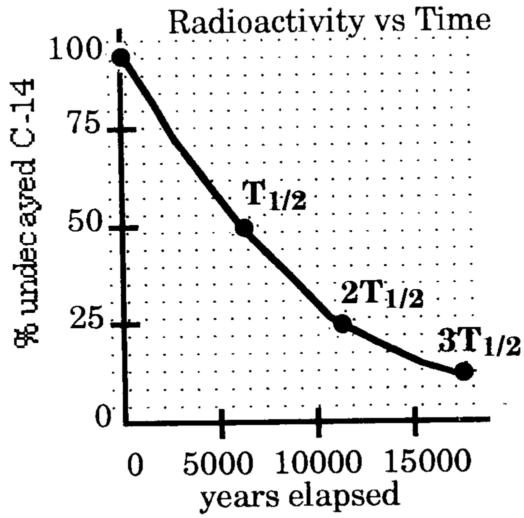
B. Half-life and Radioactive Dating

How quickly the nuclei of a particular isotope decays is an indication of its stability. **Half-life** is the time it takes for one half of the nuclei in a radioactive sample to decay. The shorter the half-life time period is, the more unstable the nuclei. For example, the most common isotope of natural uranium has a half-life of 4.5 billion years. Therefore, the uranium will be around for a long time, slowly releasing its radioactivity as part of the background radiation. One of uranium's decay by-products(called a radioactive daughter) is the gas, radon-222, which has a half-life of only 3.8 days. This highly radioactive gas seeps into the basements of homes and buildings, creating a health hazard.

The intensity of radiation can be measured by a device called a Geiger counter that records each atomic disintigration as an electronic impulse. For every half-life time period that passes, the amount of recorded radioactivity from a sample will decrease by one-half. Scientists use this fact to determine the age of rocks and organic materials due to their "natural clocks". All living plants and organisms contain the same percentage of radioactive carbon-14. Once the organism dies, it does not absorb any new carbon-14, so its "clock" starts ticking.

The decay rate for C-14 is shown in the chart below and by the graph to the right:

% undecayed carbon-14	# half-life cycles	elapsed time (in years)
100	0	0
50	1	5730
25	2	11460
12.5	3	17190
6.25	4	22920
3.125	5	28650
1.562	6	34380
0.781	7	40110
0.391	8	45840
0.195	9	51570



After a sample undergoes 10 half-life cycles, its radiation levels will be no greater than that of ordinary background radiation, and it is no longer considered to be hazardous.

Sample Problem:
How old is an ancient Egyptian mummy if a sample of the cloth body wrap has 25% the radioactivity of a living tree sample?

ANS: Using the chart above, a sample containing carbon-14 will have only 25% of its original activity after two half-life cycles, which is equivalent to 11460 years old.

Problem 1:
A thyroid patient is administered 15.0 mg of iodine-131 during a medical treatment. If the half-life for I-131 is 8 days, how much of the isotope will remain in the body after 4 weeks(28 days)?

ANS: 1.41 mg

Problem 2:
Using a mass spectrometer, it is found that the mass of barium-131 in a sample used by a local hospital's nuclear medicine department has decreased from 3.20 mg to 0.10 mg in a period of 60 days. What is the half-life of this isotope?

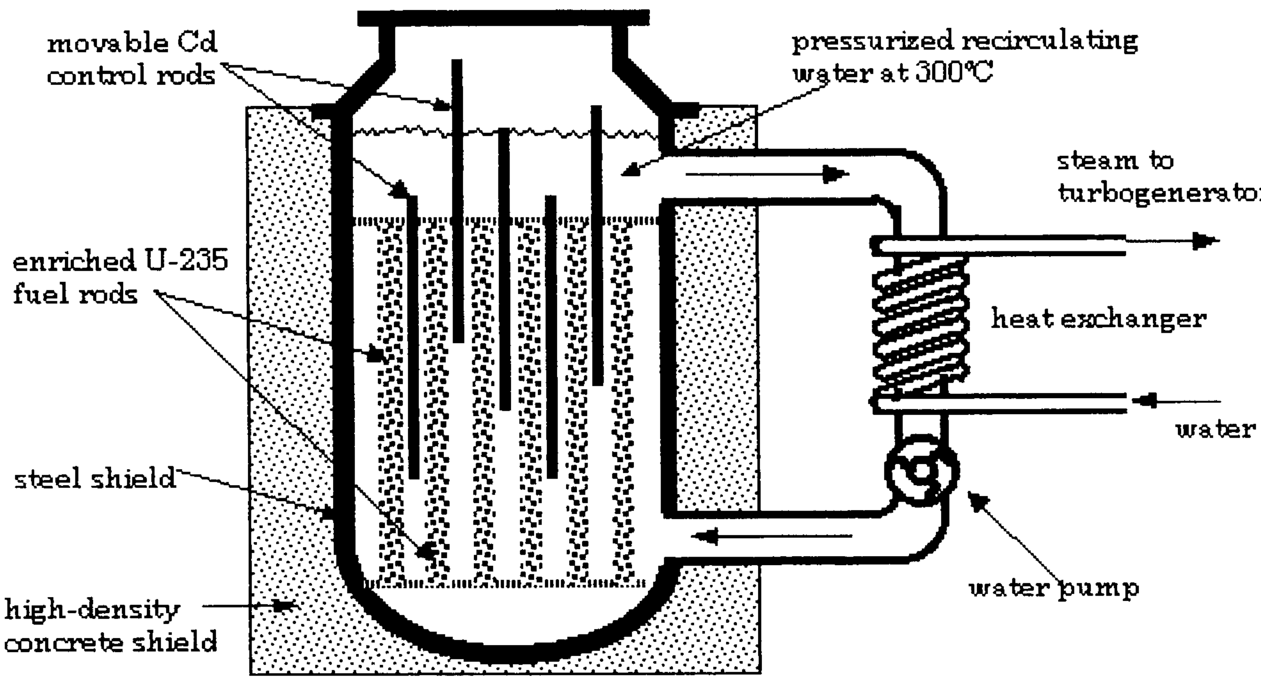
ANS: 12 days

Problem 3:
Technitium-99m emits gamma radiation that can be detected during brain scans. Its half-life is only 6.0 hours. How much time would the doctors have to complete all of their studies on a patient who is given a 20.0 mg dose of Tc-99m before the radiation levels drop down to that of background radiaton?

ANS: 60 hours

C. Nuclear Power: Fission vs. Fusion

Interior Structure of a Nuclear Fission Reactor



Extra neutrons produced by the fission of one U-235 atom can hit other atoms to produce a self-propagating chain reaction. The fuel rods contain the enriched uranium-235 that has been concentrated to obtain a minimum amount needed to sustain the chain reaction, called the critical mass. The neutrons are slowed down by moderators, such as water or graphite, so that they can be readily absorbed by the U-235. The rate of reaction is regulated by control rods, usually made of cadmium, that can be raised or lowered into the reactor to absorb excess neutrons and prevent them from causing further fission. Steel and high density concrete shielding is used to contain the radiation. The heat from the reaction warms the pressurized water to 300°C without boiling. This hot water is continually pumped through the reactor to an outside heat exchanger, which converts external water into steam to drive electric generators. The major problem with fission reactors is the disposal of the radioactive waste contained in the used fuel rods. These rods need to be encased in glass and buried deep underground in salt deposits so that they will not leach into the water table. They will continue to be radioactive for the next 10 million years!

Nuclear accidents are rare occurrences, but they can happen. The 1979 accident at Three Mile Island was the result of an improperly adjusted valve to control the flow of the water inside the reactor. As a result, the core overheated and vaporized the water, venting it as radioactive steam. Scientists were worried that the core would get so hot that it would melt right through the containment building and burn a hole into the ground("The China Syndrome"- burn a hole all the way to China). Fortunately, everything except the steam was contained inside the radiation shields. The Russian accident at Chernobyl in 1986 was a different kind of disaster. Since their fuel rods contain a lower grade of uranium than in U.S. reactors, graphite rods must be used to moderate the neutrons instead of water. While Russian scientists were doing several experimental tests on the reactor simultaneously, the core rapidly overheated. Instead of the water boiling off and stopping the reaction, the graphite rods remained in position to speed up the reaction until the reactor exploded. Since Russian reactors do not have any radiation containment buildings or safeguards, radioactive by-products, such as strontium-90 and cesium-137, were spewed into the atmosphere and the glowing nuclear core ended up on the ground outside of the reactor. Workers and fireman, most of whom have since died from radiation exposure, quickly tried to clean up the waste and the entire damaged reactor was entombed in concrete. The area is still an unsafe place to live because of the high radioactivity.

Controlled nuclear fusion, like that which occurs in the sun, is a potential energy source for the future. Enough deuterium and tritium(hydrogen-2 and hydrogen-3) used as the fuel for this reaction can be extracted from sea water to meet all of our future energy needs. Also, the helium produced by the reaction is not radioactive. Unfortunately, temperatures in excess of 20,000,000°C are required to initiate the reaction. At this temperature, the hydrogen exists as a Jell-o-like plasma, consisting of only the nuclei smashed together without any electrons. A hydrogen bomb achieves this temperature by using a fission atomic bomb to start the reaction, but obviously the side effects are not very desirable as an energy source. Experimental tokamak reactors have been built to contain the hydrogen plasma in a doughnut-shaped vessel using electromagnets. Lasers heat the plasma to the needed temperatures, but it requires as much electricity as that used by a city of 500,000 people to begin the reaction. Only recently have scientist been able to reach the <u>break-even point</u>, where they are producing slightly more energy than they are consuming to start the reaction. Improvements in the technology are still needed for fusion to become a practical energy source.

Problem:
Explain the advantages and disadvantages of producing electrical energy using fission versus fusion reactors.

Unit 6 Objectives

Having studied the unit notes and done the problems, you should be able to:

1. State a brief history of the development of the theory of atomic structure.

2. Identify the important persons associated with the development of the atomic theory and the discovery of subatomic particles and their properties.

3. Explain the electrical nature of the atom.

4. Differentiate among e^-, p, and n.

5. Define atomic number and mass number and the relationship to isotopes.

6. Given atomic number, mass number, and ionic charge calculate n, p, and e^- and vice-versa

7. Differentiate among the three forms of radiation from naturally radioactive nuclei.

8. Describe the structure of an atom's nucleus.

9. Explain how a mass spectrometer can be used to determine the relative masses and abundances of isotopes.

10. Calculate average atomic mass, given % of isotopes and their masses.

11. Describe the position of the electrons in an atom based on their different energy levels and what happens when an atom absorbs or releases energy..

12. Be able to calculate the frequency, wavelength or energy of a photon of EMR.

13. Explain the wave-particle duality of electrons and define "electron cloud".

14. Describe the differences between chemical and nuclear changes.

15. Compare the penetrating power and biological damage of the naturally occurring types of radiation.

16. Write balanced equations to represent nuclear reactions.

17. Define and differentiate between fission and fusion.

18. Define half-life, and solve simple problems using this concept.

19. Compare the pros and cons of nuclear power using fission versus fusion reactors.

20. Explain why the 1986 nuclear accident at the Russian reactor in Chernobyl was much more deadly than the 1979 accident at the Three Mile Island power plant in the U.S.