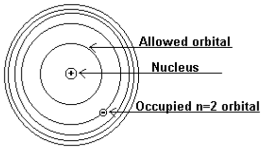
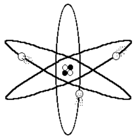
PHYS-115 (4) Kaldon Spring 2005

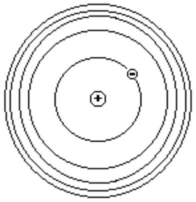
Atomic and Nuclear Physics



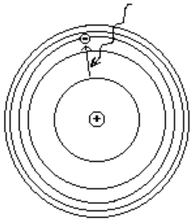
Bohr Hydrogen Atom



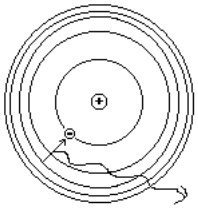
Classic Electrons Whirling Around Nucleus



Ground State

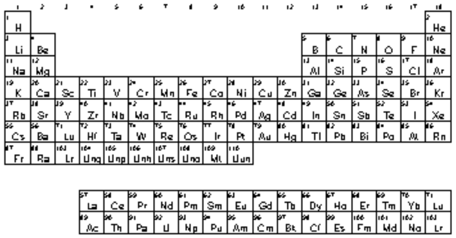


Excited State: Need to Add Energy to Ground State



Energy Released as Light When Dropping to a Lower Level

All models of the atom are fundamentally wrong – that's one way to think about the topsy-turvy world that quantum mechanics brings to physics. It's simplest to think about atoms like this: You've got a nucleus which consists of positively charged protons and neutral neutrons, surrounded by orbiting negatively charged electrons. An atom, by the way, has the same number of electrons as protons. If we have more (or fewer) electrons, we call the atom an *ion*. The *Bohr Model* of the atom makes these electron orbits be circular, and it looks kind of like a miniature solar system. Probably the classic "picture" of the atom has the electrons whirling around the nucleus in all directions. In truth, the electrons are subject to a special kind of physical law that deals more in probabilities than absolutes – we say (in the biz) that the electron is smeared around the atom in a cloud, where the most probable places we might find the electron are represented by thicker cloud wisps than elsewhere. But all of these models have to have some grounding in reality, and one of the realities of quantum mechanics is that the energy of the electron orbitals is *quantized*, or only allowed to have specific values, which depend on the type of atom and how many other electrons happen to be around at that time. If all the electrons are in the closest orbits to the nucleus that are allowed for that atom, then the electrons are said to be in the *ground state*. If energy is *added* to a ground state electron by the absorption of some energy, it is said to be in an *excited state*. If an excited state electron gives up some or all of its extra energy, which it can by emitting a *photon* (or particle) of light. A particular photon has only one color of light, and so we can look at the series of colors of light coming from a particular type of atom as a kind of fingerprint for that atom called a *spectrum*. It was these spectra that helped established what the elements were, and how we can identify them. *Helium*, for example, was first discovered in the spectrum of our Sun, as a series of lines leftover by identifying all the other spectral lines as belonging to known elements. Now, let's go back to those ground state electrons. You can't stuff all the electrons into *the* closest orbit – the Pauli Exclusion Principle states that no two electrons can exist in the same exact state at the same time. So every electron has to be different in some way. The quantum number *n* tells you how big the radius of the orbit is; so the innermost orbital is for $n = 1$. You can have two of these $n = 1$ electrons, because the are said to have *spin* in the opposite direction, *spin up* or *spin down*. These $n = 1$ electrons are in spherical orbits, or *s*-orbitals, so we can call them *1s* electrons. For $n = 2$ electrons, you can have two *2s* electrons, but you can also have electrons in orbitals that move in teardrop shaped clouds, that point in the *x*, *y* or *z* directions – these are the six *2p* orbitals. You don't have to memorize all these numbers to learn that the fact that there *are* such numbers dictated by the rules of quantum mechanics that determines the shape and order of the *Periodic Table of the Elements.* And so also all of chemistry. *Elements* are determined solely by the number of protons in the nucleus. An atom or ion that has one proton in the nucleus is the element Hydrogen, irrespective of how many (or how few) electrons are in orbit around that nucleus. Two protons in the nucleus is Helium, three protons is Lithium, etc. In the pre-Nuclear Age Periodic Table, there were 92 naturally occurring elements; Hydrogen to Uranium. (Technetium (43) isn't found in Nature, but Plutonium (94) probably is.) Since then, we have expanded the Periodic Table to include another 18+ *man-made elements*, some of which may actually occur in Nature, but in very small amounts.



In all of this discussion we have not said anything about the neutrons in the nucleus. That's because they have nothing to do with the arrangement of the elements in the Periodic Table or in the determination of what the electrons do in Chemistry. The neutrons are there for the sake of the nucleus, and are therefore the domain of Nuclear Physics, not Chemistry. (Okay, there is a field called Nuclear Chemistry – let's not quibble.) So what are those neutrons doing anyway? They don't have an electric charge like the electrons and the protons – and that's *exactly* the point. Without the neutron, there would be only one element possible: Hydrogen. To make Helium, you need two protons. But the protons *repel* each other, and at the *tiny* distances across the nucleus (10^{-15} m or 0.000 000 000 000 001 m or 1/100,000th the size of the electron orbitals!) they repel each other really fiercely. The Force between two protons in the nucleus is about 230 N – that's the same Newtons we were using before, and it's equivalent to the weight of a 52 pound object!!! But neutrons are attracted to protons through the *strong nuclear force*, which is a force that exists only in the short distances of the nucleus, and it is really strong. So you need to add some neutrons (not too many, not too few) in order keep the nuclei of all the other elements together, but there is not just one value for *N*, the number of neutrons, to go with *Z*, the number of protons in each elements. And that means that there are *isotopes*, elements that are the same chemically, but have different nuclei, with different numbers of neutrons. So there is not just one value for $A = Z + N$, the *atomic weight*. The atomic weights that are listed in the Periodic Table are actually an average value – based on the usual ratio of isotopes of a particular element in naturally occurring samples of the elements. You can usually make a good guess of what A is for the most common isotope, by rounding the atomic weight in the Periodic Table to a whole number and subtracting the proton number, Z. For example, Gold is Au (from the Latin, aurum). It has an atomic number of 79 and an atomic weight of 196.9665. If you panned for gold in the streams of California, and accumulated *1 mole* of gold atoms (1 mole of anything is one Avogadro's Number of anything; Avogadro's Number, $N_A = 6.02 \times 10^{23}$) then it would have a mass of 196.9665 grams (which is 0.1969665 kg or, on the surface of the Earth, it would have a weight of 0.43 pounds). The most common isotope of gold should have an A = 197. Gold-197, or ${}^{197}_{79}Au$, would have Z = 79 protons and N = A - Z = 118 neutrons. For most of the naturally occurring elements in the Periodic Table, the most common isotope should be expected to be stable, and not radioactive.

79
Au
Gold
196.9665

So, what does it mean to be radioactive? A radioactive nucleus is one that emits something, either particles or energy or both. (Usually both.) This is because there is too much stuff in the nucleus and that makes it unstable. And so it has to get rid of that excess in some manner. Take excessive energy – the protons and neutrons in the nucleus can absorb and release energy just like those orbiting electrons could, and just like electrons, nuclear quantum mechanics controls which amounts of energy are allowed and which are not. And so every isotope has a *nuclear spectrum* just like it has an *electronic or atomic spectrum* – so every isotope's nucleus reveals a fingerprint of energy that identify it. But while most electrons give up energy as visible, infrared or ultraviolet light, the energies in the nucleus are much much higher, and they give off light as X-rays and gamma rays (γ-rays).

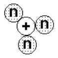
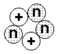
The three main types of radiation are named after the first three letters of the Greek alphabet: alpha, beta and gamma. *Alpha particles* are actually the nuclei of that really stable Helium-4 atom. It is because of alpha decay in naturally occurring radioactive materials in the ground that we have any helium on the Earth at all, and in particular, the same geological formations that tend to accumulate petroleum and natural gas, also tend to collect this helium gas. Because we are pulling out two protons from the nucleus of the unstable atom, alpha decay changes the atomic number, Z, by two and thus changes the element. Often the Z & N combination of the new element is also an unstable isotope, and so we may have a series of radioactive decays in a *decay chain*. It turns out that *beta particles* are also things that we know: they turn out to be electrons. Where do the electrons come from? Well, one of the *nucleons*, a proton or a neutron, ends up breaking into several pieces, but somehow the net charge has to remain the same, so a positive or negative electron is formed. Actually beta decay is a little complicated, because there are three main kinds of beta decay. In β^- decay, an ordinary electron is created as a neutron turns into a proton and an electron, which is kicked out of the nucleus. In β^+ decay, a proton turns into a neutron and a positive electron, or *positron*. Positrons are examples of *anti-matter*, stuff that is just like matter but opposite in sign. Positrons don't last long because they will eventually run into an electron and then we'll have *matter-anti-matter annihilation*; all the mass is turned into energy (two gamma rays) via Einstein's $E=mc^2$. In the third type of beta decay, called *electron capture*, the nucleus snatches an inner electron and a proton combines with it to form a neutron. It's kind of like doing algebra on the positron reaction, by subtracting the positron from both sides – a negative positron is a just a regular electron. In all three cases, we also change the element, because Z becomes Z ± 1. In addition, another funny particle, called a *neutrino*, or "little neutral one", is emitted. Neutrinos are very important to conservation laws in physics, but they have little impact on us or anything else. These particles, which may have a teeny mass, travel at (or extremely close to) the speed of light, and if one neutrino were to tunnel its way through something like 750 *light years* of solid lead, it would have a 50:50 chance of interacting with the lead. You've got neutrinos zipping through you all the time – they don't hurt a thing. In this course we won't sweat the differences between neutrinos and anti-neutrinos. *Gamma decay* is the only one of these that doesn't change the element. Here we start with a nucleus that is in an excited state (remember ground and excited states for electrons on the first page?), which we denote with a "star" such as ${}^{238}_{92}U^*$. The excited nucleus can give up some or all of that energy and drop back towards the ground state, the more favorable place to be, by emitting a γ-ray. Note that since the gamma ray is uncharged and isn't a nuclear particle like a proton or a neutron, that the excited nucleus doesn't change its identity, so we end up with the same isotope we started with in gamma decay. Gamma decay is sometimes combined with other decay reactions, because the new isotopes formed by alpha or beta decay may end up being in an excited nuclear state. Unstable nuclei can also eject excess neutrons.

			Radiation stopped by:	
Alpha	α	${}^{4}_{2}He^{+2}$ ion	paper	${}^{A}_{Z}X \rightarrow {}^{(A-4)}_{(Z-2)}Y + {}^{4}_{2}He$
Beta	β (β^+, β^-)	${}^{0}_{-1}e$ electron or ${}^{0}_{+1}e$ positron	foil	${}^{A}_{Z}X \rightarrow {}^{A}_{(Z+1)}Y + {}^{0}_{-1}e + {}^{0}_{0}\overline{\nu}$ Beta-minus decay ${}^{A}_{Z}X \rightarrow {}^{A}_{(Z-1)}Y + {}^{0}_{+1}e + {}^{0}_{0}\nu$ Positron (beta-plus) decay ${}^{A}_{Z}X + {}^{0}_{-1}e \rightarrow {}^{A}_{(Z-1)}Y + {}^{0}_{0}\nu$ Electron capture (e.c.)
Gamma	γ	high energy photon (a form of light)	lead	${}^{A}_{Z}X^* \rightarrow {}^{A}_{Z}X + {}^{0}_{0}\gamma$
Neutrino	ν	massless (?) bit of matter	virtually nothing, but it isn't dangerous	(see Beta reactions)
Neutron	n	"neutral proton"	paraffin (wax) and other absorbers	${}^{A}_{Z}X \rightarrow {}^{(A-1)}_{Z}Y + {}^{1}_{0}n$

So now we have the information to discuss the three basic nuclear reactions: Fission, Fusion and Matter/Anti-Matter Annihilation.

Fission	Fusion	Annihilation
An excited nucleus, after absorbing a neutron, can split into fragments	Two nuclei are forced together and form a single nucleus	Two nuclei, of matter and anti-matter, collide and convert all their mass into energy ($E = mc^2$)
${}^{A}_{Z}X + {}^{1}_{0}n \rightarrow {}^{(A+1)}_{Z}X^* \rightarrow {}^{a}_{z}A + {}^{(A-a+1)}_{(Z-z)}B + Energy$	${}^{A}_{Z}X + {}^{a}_{z}Y \rightarrow {}^{(A+a)}_{(Z+z)}Z + Energy + Etc.$	${}^{A}_{Z}X + {}^{-A}_{-Z}\overline{X} \rightarrow Energy$

*** Fusion processes may result in the emission of neutrinos as well, because of internal conversions of *p* & *n*'s.

		# of Protons	# of Neutrons	Comments
	Hydrogen-1 $^{1}_{1}H$	1	0	The most common element in the Universe
	Neutron $^{1}_{0}n$	0	1	By itself, the neutron is unstable and will break up into a proton and an electron 
	Helium-2 $^{2}_{2}He$	2	0	Without a neutron, these two protons won't form a nucleus
	Hydrogen-2 (Deuterium) $^{2}_{1}H$ or $^{2}_{1}D$	1	1	Heavy hydrogen is about twice as massive as regular hydrogen; it is stable, but not nearly so common
	Hydrogen-3 (Tritium) $^{3}_{1}H$ or $^{3}_{1}T$	1	2	Super heavy hydrogen is unstable (radioactive). Rare.
	Hydrogen-4	1	3	Three neutrons with only one proton – this just falls apart. Doesn't exist.
	Helium-4 $^{4}_{2}He$	2	2	Very strongly held together nucleus. Helium likes to be Helium-4.
	Helium-3 $^{3}_{2}He$	2	1	Another stable isotope of helium, but not as common.
	Helium-5	2	3	Helium-4 is so favorable, that Helium-5 would rather shed a neutron and become Helium-4.

Mass Energy

The game we play in nuclear reactions is one where we move energy and mass around. Deuterium (heavy hydrogen or Hydrogen-2) has a nucleus that consists of a proton and a neutron. So why do the p and the n stick together? Because it is *energetically feasible*, to use the words of nuclear physics. Every particle and every nucleus has a mass. Currently we define the *amu* or atomic mass unit as 1/12th of the mass of a Carbon-12 nucleus. (Note: If you check the Periodic Table you'll see that Carbon is listed as having an atomic weight of 12.011. This is not a contradiction, because naturally occurring carbon is going to have both Carbon-12 and Carbon-13, etc.) $^{12}_{6}C$ has 6 protons and 6 neutrons, and a mass of 12.000 amu. But each proton by itself has a mass of 1.007825 amu and each neutron has a mass of 1.008665 amu. So, is the whole greater than the sum of its parts? Well, as you can see, 6 p's and 6 n's have more mass than that of the Carbon-12 nucleus. Since you can't just "hide" mass, forming the Carbon-12 nucleus must have some freed that excess mass as the *binding energy* for the nucleus, or the energy that holds it together.

6 protons × 1.007825 amu/proton = 6.046950 amu

6 neutrons × 1.008665 amu/neutron = 6.051990 amu

Total mass of 6 p + 6 n = 12.098940 amu

Carbon-12 (made of 6 p and 6 n) = 12.000000 amu

Excess mass of pieces = 0.098940 amu

The conversion from mass to energy for our purposes is that 1 amu converts to 931 MeV. What's an MeV you say? Well, an *electron volt* or eV is an amount of energy equal to 1.602×10^{-19} J. It's a tiny amount of energy. 1 MeV is 1,000,000 eV or 1.602×10^{-13} J. So the formation of one Carbon-12 nucleus from bare protons and neutrons frees up 0.098940 amu × 931MeV/amu = 92.0142 MeV = 1.47407e-11 J. This is the energy released by forming a Carbon-12 from scratch, and it is also the energy you'd have to give a Carbon-12 nucleus to get it to completely fall apart into 6 protons and 6 neutrons. The nuclear reaction can be written as: $6(^{1}_{1}H) + 6(^{1}_{0}n) \rightarrow ^{12}_{6}C$ + Energy.

$^{1}_{0}n$	neutron	1.008665 amu
$^{1}_{1}H$	proton	1.007825 amu
$^{2}_{1}H$	deuteron	2.014102 amu
$^{3}_{1}H$	triton	3.016050 amu
$^{4}_{2}He$	helium	4.002603 amu

`Find the binding energy of Carbon-12 if built out of Helium-4's.`