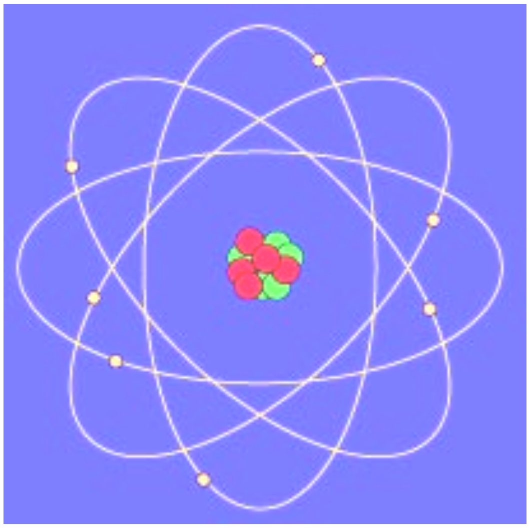
THE ATOM

All matter is made up of **atoms.** An atom is made up of two major parts: First, there is a cloud of **electrons** (very light, negatively charged particles filling most of the volume), and second, there is a central **nucleus,** which is made up of a heavy ball of **protons** and **neutrons**. The proton has a positive electric charge equal and opposite to the electron, while the neutron has almost the same mass as the proton but no electric charge.

The **atomic number** of an atom is the number of protons or electrons in the atom (the number of protons equals the number of electrons, for an electrically neutral atom). This number determines the *type of element* it is; each type of element has a different name and behaves in a unique way chemically. The different elements may have familiar names like hydrogen, oxygen, carbon, iron ..... but many have ones that are not so familiar (rhenium, praseodymium, technetium ....).

The **atomic weight** of an atom is its total number of protons and neutrons combined. You can change the number of neutrons and still have the same element -- the number of neutrons doesn't affect the chemical properties of the atom -- but it does of course affect its mass or atomic weight. If you have atoms of the same element but different numbers of neutrons, they are called different **isotopes**.

EXAMPLE: An atom of carbon normally has 6 electrons, 6 protons, and 6 neutrons, giving it an atomic number of 6 and an atomic weight of 12 (thus, "Carbon-12"). However, you can add or take away a neutron or two and still have carbon, but a different isotope of carbon. (For example, "Carbon-14" has 6 electrons, 6 protons, and 8 neutrons.)

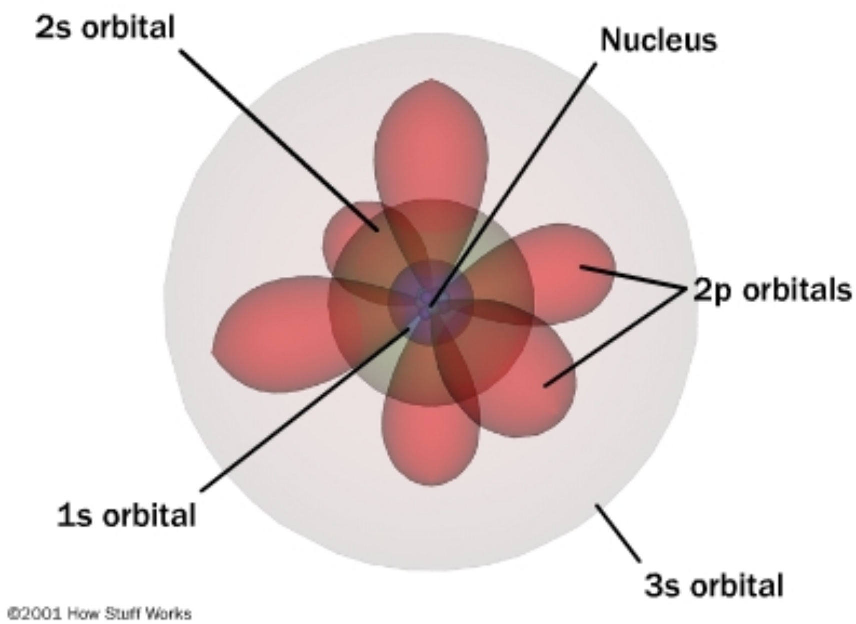


Most people have a mental image of the atom something like the first illustration above: the electrons are like little billiard balls in neat little orbits, and the nucleus is a tight ball of protons and neutrons. However, there are *two major things wrong with this picture:*

(1) The nucleus in reality is *very much tinier* in size than the picture leads you to believe: it is only about one hundred-thousandth of the atom's size. That is, if an atom were the size of a football field, the nucleus would be the size of a grain of sand at the center of the field! Yet, the nucleus contains 99.9% of all the atom's mass. The electrons fill up most of the volume of the atom but have almost none of the mass.

(2) The electrons don't really behave like little, hard "balls" orbiting around the center. While they are inside an atom, they should be thought of as more like puffy "clouds" which fill the space and have vague boundaries, something like the second illustration below. The darker colored regions show places where an electron has the highest probability of being found at any one time, but you can't make any

precise statements about just where it will be. (This was one of the chief discoveries of the new physics of quantum mechanics in the 1920's and 1930's.)



In other words, an electron can act like a vague cotton-candy cloud *inside* an atom, but acts like a tiny, well-defined particle when it is all by itself *outside* the atom. Don't let this worry you. It is a well established feature of all subatomic particles, which do not have to behave according to our common-sense notions.

History:

The concept that all materials were made up of tiny invisible particles dates back at least as far as ancient Greece, to the philosophers Leucippus and Democritus (ca. 600-500 BC). Their idea was that every substance was formed by putting together different combinations of these particles ("atomos" in Greek means "indivisible", or "uncuttable"). However, they had no experimental basis for such ideas; they argued strictly on philosophical grounds. A contrasting picture was the Aristotelian philosophy that everything was made up of different combinations of four basic "elements" (earth, air, fire, water). This was historically a more successful idea, and the idea of "atoms" fell by the wayside for many centuries.

The concept of atoms was brought back in the late 1700's, after the Newtonian revolution and the invention of modern science was well underway. The new ideas about atomic properties originated with chemistry, not physics, and started in 1808 with John Dalton, one of the pioneers of chemistry. He proposed that there was a small number of basically different kinds of atoms which represented the smallest units of the elements then known (iron, sulfur, oxygen, carbon, and so forth). The idea worked well to explain how chemical compounds formed. It was developed further in the mid-1800's by the Italian, Avogadro, who measured the relative weights (masses) of the most common elements, and by the Russian, Mendeleev, who worked out the central organizing principle of all of chemistry: the **Periodic Table**. (See our mini-tutorial "The Elements" for more.) Still, however, in the 1800's the idea of atoms remained a useful conceptual and theoretical tool rather than a hard experimental reality; no one had ever seen an atom or had any truly direct laboratory evidence for their existence.

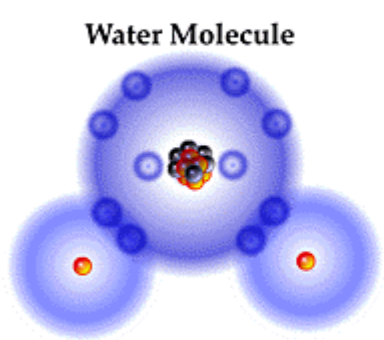
In the late 1800's, the story of uncovering the structure of the atom passes from chemistry to a series of key experiments in physics which are among the highlights of 20^{th}-century science. In 1897, J.J.Thomson found that electrons (the carriers of electric current) could be removed from within atoms, and thus that the inside of an atom must have a mixture of positively and negatively charged objects. (Paradoxically, this meant that the very term "atom", meaning "indivisible", was the wrong choice of word!) Not long after that, in 1909 Rutherford discovered that the atom had a tiny, incredibly dense central nucleus which held all the positive charge and most of the mass. In later experiments, Rutherford and Chadwick found that the nucleus itself consisted of two distinct types of massive particles, protons and neutrons. The basic structure of the atom was now known and their reality was now firmly established.

The later history of atomic structure is closely associated with the step-by-step discovery of quantum mechanics and particle physics: the strange and unexpected behavior of things in the submicroscopic world. For more on this part of the story, see our next primer “Quarks and Leptons”.

Atoms in a Crowd

The common objects and materials around us are usually made up, not of single atoms, but of ***molecules.*** A molecule is a combination of two or more atoms locked tightly together in a single new entity. When the atoms get close together, their outer electrons act as if they are shared among the atoms and bond everything together. The formation or breakup of molecules is what is called a **chemical reaction.**

A molecule of water has three atoms: two hydrogen, one oxygen, labelled H_2O and looking something like this:



The little blue dots indicate the positions of the electrons within the molecule (eight from the oxygen atom, two from the two hydrogens).

The air around us is made of molecular nitrogen (N_2) and molecular oxygen (O_2, which we need to breathe).

But molecules can be assembled from much larger numbers of atoms in infinite variety: the famous DNA molecule which is the basis of the genetic code has thousands of atoms strung together in a double helix shape.

When Good Atoms Go Bad

We need to introduce one other term here: the ***ion***. The type of atom (i.e., the type of element it is) is defined by the number of protons it has in its nucleus. Normally, the number of orbiting electrons is the same, and the whole atom is electrically neutral. However, electrons can be removed entirely from the atom, or extra ones can be added. When this happens, we say that the atom has been ***ionized***, and it has a net electric charge which is either positive (if there are fewer electrons than protons) or negative (if there are more electrons than protons).

Ions can behave rather violently in a chemical sense. The atom would “prefer” to be electrically neutral. So if it is short one or two electrons, it will eagerly grab any spare electrons that happen to be nearby.