Structure of the Atom

Elementary Particles

Elementary particles can be thought of as the building blocks of matter. Just as matter is made of atoms and molecules, and molecules themselves are made from atoms, atoms in turn are made from even smaller particles. There are three particles which we will have to consider, although there are many more that exist. The three that concern us are protons, electrons, and neutrons.

The elementary particles are distinguish by a variety of different properties, of which we shall restrict ourselves to just their mass and their electrical charge. It takes a very large number of atoms, and an even larger number of elementary particles, to make 1 kg of mass. It follows then that their mass (and also the electrical charge) of a single elementary particle is very small. It therefore makes sense to use a scale of comparable size

- for mass, the atomic mass unit. The atomic mass unit (amu or just u for short) is equal to $1.66053886 \times 10^{-27}$ kg.

- For electrical charge, the magnitude of the charge on an electronic, which is equal to 1.602×10^{-19} C.

On this scale the mass of the proton and the mass of the neutron are very nearly equal, the neutron being slightly heavier. For all but the most careful calculations we can set each mass to 1 amu (or 1 u). However the mass of the electron is very small, only about $1/2000^{th}$ the mass of the proton.

The electrical charge of the electron and the electrical charge of the proton are equal in magnitude, but of opposite sign. By convention the proton is positively charged and the electron negatively charged. The neutron has no charge at all, hence its name. It is neutral.
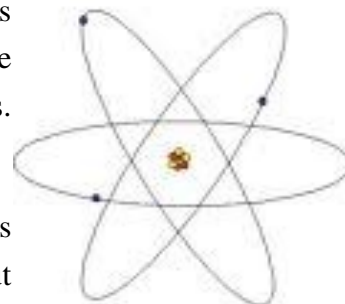
The data is summarized in the following table:

Particle	Mass (kg)	Mass (amu)	Approximate mass	Electrical charge
Electron	9.11 x 10^{-31} kg	0. 000548 amu	0 amu	-1
Proton	1.66 x 10^{-27} kg	1. 007276 amu	1 amu	1
Neutron	1.67 x 10^{-27} kg	1. 008665 amu	1 amu	0

For all but the most detailed calculations only the last two columns will be important to us.

Nucleus and electrons

A famous experiment by Ernest Rutherford showed that the atom is almost completely empty. We now know that its structure consists of a very small nucleus containing all of the protons and neutrons, surrounded by a "cloud" of electrons. This is illustrated in figure 1.

Note that this figure is not to scale. The size of an atom, as defined by its electrons, varies a little from atom to atom, but is generally speaking of the oder of 10^{-10} m. that is a line of atoms would require 10 billion to make a line about 3 feet long. The nucleus is much smaller than this, only about 10^{-15} m in diameter. That makes the nucleus only about 1/100,000th the size of an atom. In the diagram above, if drawn to scale the nucleus but be a dot too small to see.

Fig 1: Schematic diagram of an atom

For the most part, that means that on the nuclear scale the electrons are so far away that to first approximation we can ignore them. This is not true in all cases, but will give us a good conceptual introduction, with the influence of electrons being discussed at the appropriate time.

Atomic number and mass number

The atomic number (Z) of a nucleus is equal to the number of protons in its nucleus. Since atoms are normally neutral[1], the number of electrons surrounding the nucleus is normally equal to the number of protons within the nucleus. The atomic number of any given atom also tells us the chemical identity of the atom. For example if the atomic number is equal to 6 then that atom is carbon. All carbon atoms in the universe have 6 protons (and 6 electrons), but no non-carbon atom has 6 protons. The atomic number is said to define the element.

The mass number (A) is the total number of nucleons (protons plus neutrons) in the nucleus. It also follows that the number of neutrons in the nucleus is equal to the difference between the mass number and the atomic number (A-Z). Once the mass number (or equivalently the number of neutrons) is known then we have specified the isotope of the element. Since there is no definite rule for the number of neutrons in the nucleus any given element can have many different isotopes. For example, Xenon has 36 known isotopes, although only a few of them occur naturally.

The true mass of an isotope is very nearly equal to the total number of nucleons (protons plus neutrons) in the nucleus. Referring back to the table above, the masses of the proton and the neutron are very nearly equal to 1 amu, whereas the mass of the electron is so very small that their contribution to the total mass of the atom is negligibly small. For example, the isotope of chlorine ^{35}Cl (17 protons and 18 neutrons) actually has a mass of 34.9689 amu, very close to but not exactly equal to the total number of nucleons.

For shorthand notation, the atomic number if commonly written as a preceding subscript to the chemical symbol for the element. Thus $_{33}As$ would be the notation for the element arsenic, with 33 protons in its nucleus. To specify the isotope we would add the mass number as a preceding superscript to the chemical symbol. Thus $^{75}_{33}As$ would be the notation for one of the isotopes of the element arsenic, the one with 33 protons in the nucleus, 75 nucleons in total, and so 75-33 = 42 neutrons in the nucleus.

Since the atomic number has a one-to-one relationship with the chemical identity of the atom, specifying the atomic number is rather superfluous, and can be omitted without loss of information. Thus $^{75}_{33}As$ could equally well be written as ^{75}As. Knowing that the element is arsenic would allow you to conclude that there are 33 protons in its nucleus.

Remember

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1   If the number of electrons is not equal to the number of protons, then the atom will have a net charge, and in that case is usually referred to as an ion.

- Cl is an element (chlorine)

- ^{35}Cl is an isotope

Binding energy

It seems obvious that if you assemble an atom from its constituent protons, neutrons, and electrons, then the mass of the atom that you produce should equal the total mass of the elementary particles. However, accurate measurements show that this is not true.

Let us look at (for example) ^{65}Cu. It has 29 protons, 36 neutrons, and 29 electrons. The total mass of the elementary particles is therefore

- protons: mass of 29 protons = 29 * 1.007276 = 29.21100 amu

- neutrons: mass of 36 neutrons = 36 * 1.008665 = 36.31194 amu

- electrons: mass of 29 electrons = 29 * 0.000548 = 0.015892 amu

- Total mass of elementary particles = 65.53884 amu

However, if we look up the mass of a ^{65}Cu atom, it is only 64.9278 amu[2]. We have to conclude that in assembling the atom about 0.611 amu of mass has been lost. The obvious question is, what happened to it?

To answer this question we have to look at the structure of the nucleus. It is comprised of protons and neutrons, the first of which carry a positive charge. Like (that is, of same sign) charges (in this case two positive charges) experience a force of repulsion which becomes very strong when their separation becomes small. (Bear in mind that the size of the nucleus is only about 10^{-15} m.) Without any other influence the protons in the nucleus would fly apart. Something must be responsible for stabilizing the nucleus, a nuclear force whose is confined to the dimensions of the nucleus. This nuclear force needs an origin, the energy to do its task, and that is where the mass went, into proving the energy to stabilize the nucleus. The lost mass is referred to as the mas defect, and the energy that is created to stabilize the nucleus is known as the binding energy.

The binding energy is not the same for all nuclei. Generally speaking it increases with increasing number of nucleons. After all, the more there are of them, the harder it is to keep them all together. For that reason the data is usually presented in terms of the binding energy per nucleon (see figure 2). The curve that shows a steady increase for the

2 http://environmentalchemistry.com/yogi/periodic/Cu-pg2.html#Nuclides

elements hydrogen through iron, after which it slowly decreases again. However, the curve is not a smooth one, there are fluctuation of the data about the average, which has consequences when considering which nuclei are radioactivity and which are not.

The fact that the curve peaks at the element iron has important consequences in the production of the elements in the universe. Initially, only hydrogen and helium were created in appreciable quantities by the Big Bang. All the other elements had to be produced. For elements up to iron this is relatively straightforward, they were created in the interiors of long dead stars, and then distribu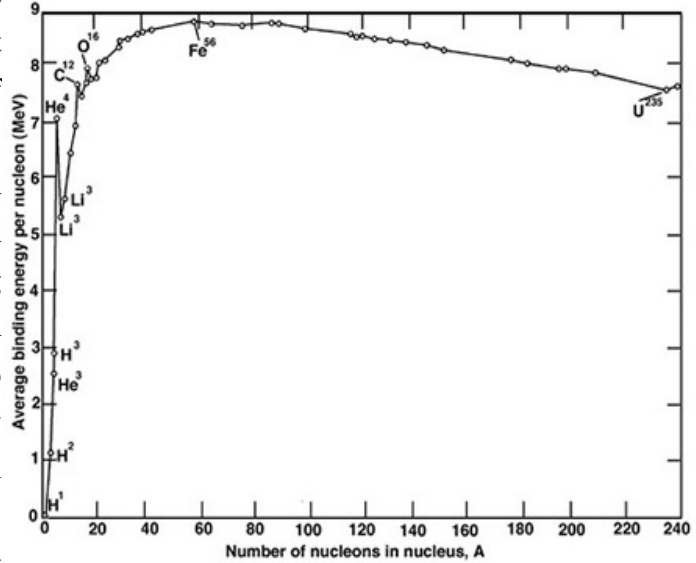ted throughout the universe when those stars exploded. This is still going on today. However elements heavier than iron cannot be created in this way, their creation needs even more extreme conditions, and these condition are only encountered in the explosion of the most massive stars, stars with much more mass than our own Sun. An explosion of a star of such large mass is termed a supernova.

Fig 2: Binding energy per nucleon as a function of the number of nucleons

Einstein's formula

We have seen qualitatively in the previous section that there is a loss of mass required to create the energy to hold the nucleus together against electrostatic repulsion of the protons. The amount of energy that is created had originally been calculated by Albert Einstein in his treatment of relativity (the relative motion of particles as seen from different vantage points). The result is a famous formula which almost everyone has seen

$$E = m\ c^2$$

in which E is the energy equivalent corresponding to a mass m, and c is the speed of light.

If we convert the energy from Joules to MeV, and the mass from kg to amu, then the above equation can be written as

$$E = m * 931.4 \text{ MeV/amu}$$

For example, if the mass loss is 0.0155 amu, then the energy that is produced is equal to 0.0155 * 931.4 = 14.4 MeV.

Note that we have talked here about losing mass and creating energy, because that is the calculation we needed in the discussion above for stabilizing the nuclei, and which we will need later when discussing radioactivity and nuclear reactions. However, it is possible to go the other way also, to destroy energy and create mass. This is known as pair production (it always produces a particle and its own antiparticle), but we do not need to look at this further.

Nuclear Energy vs Chemical Energy

Although both involving atoms, nuclear and chemical processes have different origins and different energies

- Chemical processes involve the electrons of the atoms. Chemical bonds arise from the electrical energy derived from the interaction of the electrons with both the electrons and nuclei of adjacent atoms. Since the electrons are relatively far from the nucleus (about 10^{-10} m) these interactions are not particularly strong, and typical energies of these interactions are in the electron-volt (eV) range. For example, the ionization potential of helium (the largest ionization potential of all atoms) is 24.8 eV and the molecular bond in atmospheric nitrogen has a strength (dissociation energy) of about 9.8 eV.

- Nuclear processes involve only the particles inside the nucleus, that is the protons and neutrons. The electrons play no role. The energy comes from both the electrostatic energy between protons, and the binding energy required to hold the nuclei together. Since the distances involved are much smaller (about 10^{-15} m) these interactions are much stronger than the chemical interactions, and typical nuclear energies are in the millions of electron-volt (MeV) range.

The difference in magnitude of the these interactions shows up in the amount of fuel that is required for a power station. Those that use fossil fuels (coal, oil, natural gas, etc) are taking advantage of chemical processes to generate the electrical energy. They require enormous amounts of fuel to maintain electrical energy generation. A typical coal fired power plant uses about 10,000 metric tons of coal *each day*[3], coal which must be

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3   http://en.citizendium.org/wiki/Conventional_coal-fired_power_plant

brought in (usually by train) and whose waste has to be shipped back out again, all on a continual basis.

By contrast nuclear power stations rely on the energy of the nucleus, and that is of the order of one million times larger per reaction, much less fuel is required. All the nuclear reactors in the USA require about 20,000 metric tons of nuclear fuel each *year*[4]. That works out to about only about one half of a metric ton per day per reactor. This is so small that reactors are only refueled once every few months to once every couple of years.

4   http://www.world-nuclear.org/info/reactors.html