The Discovery of Subatomic Particles

Revised Edition

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A World of Particles

How many men and women, studying the tiny boulders in a handful of sand, may have conceived of the finer and harder grains that make up all forms of matter? The explicit statement that matter is composed of indivisible particles called atoms (from the Greek ατομοσ, "uncuttable") we trace to the ancient town of Abdera, on the seacoast of Thrace. There, in the latter part of the fifth century B.C., the Greek philosophers Leucippus and Democritus taught that all matter is made up of atoms and empty space.

Abdera now lies in ruins. No word written by Leucippus has survived, and of the writings of Democritus we have only a few unhelpful fragments. However, their idea of atoms survived and was quoted endlessly in the following millennia. This idea allows one to make sense of a great many commonplace observations that would be quite puzzling if it were thought that matter was a continuum that fills the space it occupies. How better can we understand the dissolving of a piece of salt in a pot of water than by supposing that the atoms of which the salt is composed spread into the empty spaces between the atoms of the water? How better can we understand the spread of a drop of oil on the surface of water, out to a definite area and no farther, than by supposing that the film of oil spreads until it is a few atoms thick?

After the birth of modern science, the idea of atoms came to be used as a basis of quantitative theories of matter. In the seventeenth century Isaac Newton (1642–1727) attempted to account for the expansion of gases in terms of the outrush of their atoms into empty space. More influentially, in the early nineteenth century John Dalton (1766–1844) explained the fixed ratios of the weights of chemical elements that make up common compounds in terms of the relative weights of the atoms of these elements.

By the end of the nineteenth century the idea of the atom had become familiar to most scientists – familiar, but not yet universally accepted. Partly because of the heritage of Newton and Dalton, there was a disposition to use atomic theories in England. On the other hand, resistance to atomism persisted in Germany. It was not so much that the German physicists and chemists positively disbelieved in atoms. Rather, under the influence of an empiricist philosophical school centered on Ernst Mach (1836–1916) of Vienna, many of them held back from incorporating into their theories anything that – like atoms – could not be observed directly. Others, like the great theorist Ludwig Boltzmann (1844–1906), did use atomistic assumptions to build theories of phenomena such as heat, but had to suffer the disapproval of their colleagues; it is said that the opposition to Boltzmann's work by the followers of Mach contributed to Boltzmann's suicide in 1906.

All this changed in the first decades of the twentieth century. Oddly, the general acceptance of the atomic nature of matter came about through the discoveries of the constituents of the atom, the electron and atomic nucleus – discoveries that undercut the old idea that atoms are indivisible. These discoveries are the subject of this book. Before we go into the history of these discoveries, however, let us anticipate them and recall what is now understood about the constituents of the atom. This is only a brief overview; we will be going into all this in much greater detail in the subsequent chapters of this book.

Most of the mass of any atom is contained in the small, dense nucleus at its center, which carries a positive electric charge. Revolving in orbit around the nucleus are one or more electrons, which carry negative electric charge and are held in orbit by the force of electrical attraction. The typical radius of an electron's orbit is about 10^{-10} meters* (a unit of length called the angstrom), while the nucleus is much smaller, with typical diameter of about 10^{-15} meters (a unit called the fermi). The various chemical elements each consist of atoms of one specific kind, the atoms of one element differing from those of another element in the number of electrons they contain: one for hydrogen, two for helium, and so on up to 109 for meitnerium. Atoms can combine into larger aggregates – molecules – by lending, trading, or sharing their electrons; each chemical compound consists of molecules of one specific kind. Under ordinary circumstances, visible light is absorbed or emitted when the electrons in an atom or molecule are excited to orbits of higher energy or sink back to orbits of lower energy, respectively. Electrons can also be shaken loose from atoms, and by traveling through a metal wire produce an ordinary electric current.

In all these phenomena – chemical, optical, and electrical – the nucleus of the atom is essentially inert. However, the nucleus itself is a composite system with its own constituents, the particles known as protons and neutrons. The proton carries an electric charge equal and opposite to that of the electron; the neutron is electrically neutral. The proton has a mass of 1.6726×10^{-27} kilograms, the

* For a short discussion of scientific notation, see the box at the end of this chapter.

neutron's mass is a little larger (1.6750×10^{-27} kilograms), and the electron's mass is much smaller (9.1094×10^{-31} kilograms). The protons and neutrons in nuclei, like the electrons surrounding the nuclei, can be excited to states of higher energy or, if excited, can fall back to a state of lower energy, but the energies needed to excite the nuclear particles in the nucleus are typically a million times those needed to excite the electrons in the outer part of the atom.

All ordinary matter is composed of atoms, which in turn consist of protons, neutrons, and electrons. However, it would be a mistake to conclude that protons, neutrons, and electrons make up the whole list of fundamental entities. The electron is just one member of a family of particles called leptons, of which some half dozen are now known. The proton and the neutron are members of a much larger family of particles called hadrons, of which hundreds are known. The special property that makes electrons, protons, and neutrons the ubiquitous ingredients of ordinary matter is their relative stability. Electrons are believed to be absolutely stable, and protons and neutrons (when bound in an atomic nucleus) live at least 10^{30} years. With a few exceptions, all other particles have very short lifetimes, and are therefore very rare in the present universe. (The only other stable particles are those that have zero or very small values of mass and charge and therefore cannot be trapped into atoms or molecules.)

The proton, the neutron, and the other hadrons are now believed to be composites themselves, made up of more elementary constituents called quarks. A proton consists of two quarks of a type known as "up" quarks and one of another kind called "down" quarks, while a neutron consists of two down quarks and an up quark. There are four other quark types, that are too unstable to be found in ordinary matter, though there are hints that certain stars may be composed of equal numbers of up quarks, down quarks, and a third kind of quark known as "strange" quarks. As far as is known, the electron and the members of the lepton family are truly elementary. But elementary or not, it is the particles that make up ordinary atoms – protons, neutrons, and electrons – that will concern us in this book.

Just as ancient Abdera symbolizes for us the birth of atomism, there is one place with which the discovery of the constituents of the atom is especially associated: It is the Cavendish Laboratory of the University of Cambridge. There, in 1897, Joseph John Thomson (1856–1940) performed the experiments on cathode rays that led him to conclude that there is a particle – the electron – that is both the carrier of electricity and a basic constituent of all atoms. It was at the Cavendish in 1895–98 that Ernest Rutherford (1871–1937) began his work on radioactivity, and to the Cavendish in 1919 that Rutherford returned, after his discovery of the atomic nucleus, to succeed Thomson as Cavendish Professor of Experimental Physics and to found what was long the preeminent center for nuclear physics. The list of constituents of the atom was completed

The exterior of the Cavendish Laboratory at Cambridge as it appeared from Maxwell's time onward. The building is now used for other university purposes, and the laboratory has been moved to more modern quarters.

at the Cavendish in 1932, when James Chadwick (1891–1974) discovered the neutron.

I first visited the Cavendish Laboratory in the spring of 1962, when as a very junior physicist I was on leave from the University of California at Berkeley for a year in London. The laboratory then still occupied its original gray stone buildings in Free School Lane, where it had stood since 1874, on land purchased by the University of Cambridge in 1786 for use as a botanical garden. I remember it as a warren of little rooms connected by an incomprehensible network of stairways and corridors. It was very different from California's great Radiation Laboratory, which looked commandingly out over the bay from its sunlit site in the Berkeley hills. The Cavendish Laboratory gave the impression that it was the scene not so much of a massive assault on the secrets of nature, but of a guerilla campaign, an effort of limited resources, in which the chief weapons were the cleverness and bravado of gifted individuals.

The Cavendish Laboratory had its origin in the report of a university committee that met in the winter of 1868–69 to consider how to make a place for experimental physics in Cambridge. It was a time of widespread enthusiasm for experimental science. A great new laboratory for experimental physics had recently been opened in Berlin, and university laboratories were being constructed at Oxford and Manchester. Cambridge had not played a leading role in experimental science, despite (or possibly because of) a tradition of excellence in mathematics that went back to the seventeenth-century Lucasian Professor of Mathematics, Sir Isaac Newton. But empiricism was now in the air, and the committee called for a new Professorship of Experimental Physics and a new building to house lectures and experiments.

It remained to find funds and a professor. The first need was quickly met. The chancellor of the university at that time was William Cavendish, seventh Duke of Devonshire and a member of the family that had earlier produced the distinguished physicist Henry Cavendish (1731–1810), who first measured the attractive force of gravity between laboratory masses. Devonshire had done spectacularly well in mathematics as an undergraduate at Cambridge, and had then gone on to do even better making money in the Lancashire steel industry. In October 1870 he wrote to the vice-chancellor of the university, offering to provide the funds required for building and apparatus – some £6,300. When the building was completed in 1874, a letter of thanks (in Latin) was presented to Devonshire, proposing to name the laboratory after the Cavendish family.

It was hoped that the first Cavendish Professor would be Sir William Thomson (1824–1907), later Lord Kelvin, the most eminent experimental physicist in Britain. However, Thomson wished to stay in Glasgow, and instead the Cavendish Professorship went to another Scot: James Clerk Maxwell (1831–1879), who at age 39 was living in retirement on his estate at Glenair.

Maxwell is generally regarded as the greatest physicist between Newton and Einstein, but it is odd to think of him as a professor of experimental physics. Although he did experimental work of some distinction on color perception (with his wife as colleague) and on electrical resistance, his greatness rests almost entirely on his theoretical work. Above all, it was Maxwell who completed the equations that describe the phenomena of electricity and magnetism, and then used these equations to predict the existence of electromagnetic waves, thereby explaining the nature of light. Maxwell's melding together of the theories of electricity and magnetism into a unified theory of electromagnetism has been a paradigm for the efforts of theoretical physicists ever since. Although Maxwell's work lent great prestige to the Cavendish Professorship, the Cavendish Laboratory did not develop into a leading center of experimental physics during his tenure. For instance, the existence of electromagnetic waves was demonstrated experimentally not at the Cavendish Laboratory, but rather at Karlsruhe, by the German experimentalist Heinrich Hertz (1857–1894).

After Maxwell's death in 1879, the Cavendish Professorship was again offered to William Thomson, who again declined. This time the professorship went to John William Strutt (1842–1919), the third Lord Rayleigh. Rayleigh, who was gifted both as a theorist (though not on Maxwell's level) and as an experimentalist, worked on a huge variety of physical problems. Even today, when one is confronted with a problem in hydrodynamics or optics, a good place to start in looking for a solution is in his collected works. Under Rayleigh's leadership the Cavendish Laboratory remained small, most of its research consisting of Rayleigh's own work, but important improvements were made. New apparatus was purchased, instruction was reorganized, a workshop was opened, and beginning in 1882 women were admitted on the same terms as men. In 1884 Rayleigh resigned the Cavendish Professorship, and shortly thereafter he accepted the less demanding position of professor at the Royal Institution in London.

Once again, the Cavendish Professorship was offered to William Thomson (now Lord Kelvin), and once again Kelvin decided to stay in Glasgow. The obvious next choice was between R. T. Glazebrook and W. N. Shaw, who did most of the work of preparing apparatus for lectures and experiments. To almost everyone's surprise, the professorship went instead to a young man of mostly mathematical talent – J. J. Thomson. Although it is not clear if there were any good reasons then for this decision, it was the right choice. Following Rayleigh's advice, Thomson began his epoch-making experimental work on cathode rays. What is more, under his direction the Cavendish Laboratory came alive. Flocks of talented experimentalists came to work there, including in 1895 a young New Zealander, Ernest Rutherford. The stage was now set for the discovery of the constituents of the atom.

Scientific or Exponential Notation

Atoms and subatomic particles are very small, and there are a large number of them in any ordinary piece of matter. We can get nowhere in speaking of them without using the convenient "scientific" or "exponential" notation for very large and very small numbers. This notation employs powers of ten: 10^1 is just 10; 10^2 is the product of two tens, or 100; and so on. Also, 10^{-1} is the reciprocal of 10^1, or 0.1; 10^{-2} is the reciprocal of 10^2, or 0.01, and so on. (That is, 10^n is a one with n zeroes, and 10^{-n} is a decimal point followed by $n-1$ zeroes and a one.) Here is a list of some powers of 10, with their American names and the prefixes that are used to denote them.

Power of 10	American name	Prefix
10^1	ten	deka
10^2	hundred	hecto
10^3	thousand	kilo
10^6	million	mega
10^9	billion	giga
10^{12}	trillion	tera
10^{-1}	tenth	deci
10^{-2}	hundredth	centi
10^{-3}	thousandth	milli
10^{-6}	millionth	micro
10^{-9}	billionth	nano
10^{-12}	trillionth	pico
10^{-15}	quadrillionth	femto

(I say American here because for the British a billion is 10^{12}, and 10^9 is a milliard.) For instance, 10^3 grams is one kilogram; 10^{-2} meter is one centimeter; and 10^{-3} ampere is one milliampere. The great thing about this scientific notation is not just that it saves writing words like "quadrillionth," but that it makes arithmetic so easy. If we want to multiply 10^{23} times 10^5, we are taking the product of 23 tens times the product of 5 tens, or 28 tens in all; so the answer is 10^{28}. Likewise if we want to multiply 10^{23} by 10^{-19} (or divide 10^{23} by 10^{19}), then we are dividing the product of 23 tens by the product of 19 tens; so the answer is 10^4.

This is the general rule: in multiplying powers of 10 add the powers; in dividing subtract them. According to this rule ten to any power divided by ten to the same power would be ten to the zeroth power; so 10^0 is taken to be 1. To deal with numbers that are not simply a power of 10, we can always write them as a number between 1 and 10 times a power of 10: thus 186,324 is 1.86324×10^5, and 0.0005495 is 5.495×10^{-4}. In multiplying or dividing such numbers, we multiply or divide the numbers that accompany the powers of ten and combine the powers of ten as before; thus 1.86324×10^5 times 5.495×10^{-4} is 1.86324 times 5.495, or 10.238, times $10^5 \times 10^{-4}$, or 10^1, which we could also write as 1.0238×10^2. These days scientific notation is used pretty widely, from the pages of *Scientific American* to the electronic calculators that one can buy for under \$20 (2×10^1 dollars), so I will use it freely in this book.