What is an Electron? by **Mr Stephen Pilkington** on Zoom on 21st August 2020

Mr Pilkington, after a non-scientific career, but having an interest in science, set out to find out the history, discovery, character and construction of the electron. He gave us a layman's view of the electron and its place in the Atom. He avoided the use of equations, but even so gave a most illuminating talk.

The Electron is named after the ancient Greek word *elektron* for what we know as Amber; which if rubbed with fur attracts small objects. The word entered English in 16C when scientist William Gilbert coined the term 'electricus' to refer to this attraction. In 1874 the Irish physicist George Stoney proposed the electron as a "fundamental unit quantity of electricity" This has since been accepted as: $1.602176634 \times 10^{-19}$ Coulomb.

John Dalton, an English scientist, deduced his atomic theory from knowledge of how elements combined or reacted in ratios of small whole numbers (the law of multiple proportions). In the case of Tin, 100g of tin would combine with 13.5g of oxygen to form Tin Oxide, or 27g of oxygen to form Tin Dioxide – one atom of tin combining with one or two of oxygen. In 1803, a century before it could be tested, he proposed that:

- all matter was composed of very small indivisible particles - called **Atoms**;
- atoms of a given element were all identical in mass and properties;
- compounds were formed by a combination of two or more different kinds of atoms; a chemical reaction was simply a rearrangement of atoms.

The **Electron** was the first sub-atomic particle to be discovered. In 1878 Sir William Crookes, a British scientist, using an improved Geissler evacuated tube with an electrode at each end, mounted an object in it, and saw that particles from the cathode produced a sharp-edged shadow of the object at the anode end of the tube. In another tube he mounted a small off-centre paddlewheel, and noted that it was rotated by the particles. The particles were negatively charged, and had momentum/mass. J J Thomson and colleagues performed experiments that: showed cathode rays were unique particles – not waves, atoms or molecules as was believed earlier. Negatively charged particles produced by radioactive, heated and illuminated materials were uniform and universal, whatever the emission source. Later experiments using electromagnetic deflection and suspension in oil drops confirmed their now accepted mass value of $9.1093837015 \times 10^{-31}$ kg.

The Atom became a 'plum pudding' with a positive substrate containing negatively charged electrons.

The **Proton** was next to be discovered in 1909, by Ernest Rutherford. A narrow beam of positively charged α-articles (emitted from a radioactive source) was directed onto a thin sheet of gold foil. Most α-particles went straight through, but a few were scattered in other directions (a few even back towards the source). This showed that the atom was mostly empty space with a small positively charged nucleus. The charge matched that of the electrons, which would form a cloud around the nucleus.

Neutrons were not discovered until 1932 when James Chadwick observed that a beryllium foil, when exposed to bombardment by alpha particles, emitted from a Radium source, released an unknown radiation that in turn ejected protons (hydrogen ions) from a paraffin wax target. Chadwick interpreted that radiation as being composed of particles of a mass similar to protons but without electrical charge – i.e. neutrons. This explained the 'missing mass' of atoms. It also explained the existence of isotopes – atoms with the same number of protons but differing numbers of neutrons in the nucleus.

In 1913 Niels Bohr had proposed a model in which the electrons of an atom could only orbit the nucleus in a finite set of 'quantised' orbits, each associated with a discrete energy. This was based on the line spectrum of Hydrogen - electrons could jump between a pair of these orbits by absorbing or emitting a photon with energy: $E = hf$ (as proposed by Albert Einstein - h being the Planck constant, and f the frequency) and where E corresponds to the difference in energy levels associated with the orbits. This quantization – of both orbits and energy levels - explained why electron orbits were stable and gave the unique atomic spectra of each element. This model is known as 'old Quantum Mechanics'. Schrödinger, Heisenberg, and others developed this model to give the Quantum Mechanics we know today (see below).

Physicists subsequently developed a Standard Model, where particles form two groups: Fermions (Leptons & Quarks) & Bosons (*see diagram*). The electron is a Lepton, an elementary particle with a charge of -1. Fermions are the particles of matter, while bosons are responsible for bringing about forces and mass.

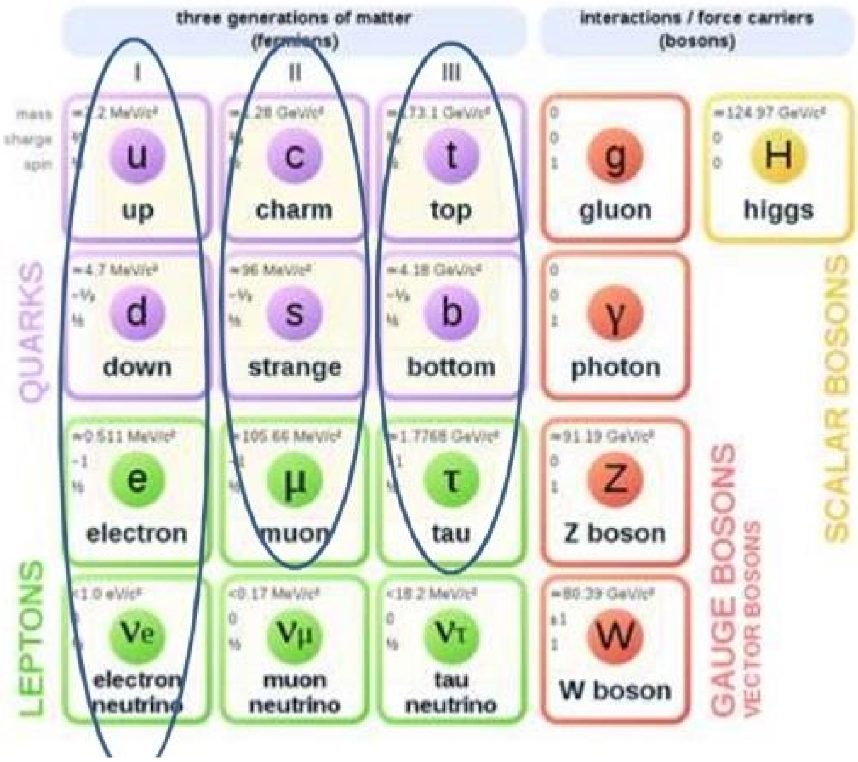
In the Standard Model neither the proton nor the neutron is a single particle: they are formed of Up (u) and Down (d) Quarks, two 'u's & a 'd' for a proton; and one 'u' & two 'd's for a neutron. Each 'u' has a charge of $+{}^{2}/_{3}$ and a 'd' of $-{}^{1}/_{3}$ giving the observed charge of protons and neutrons. They are bound together by the short-range Strong Force. (Electrons are not affected by the Strong Force.) This is why a nucleus is stable when positively charged protons would be expected to repel each other.

The Standard Model of Elementary Particles

Group I:
Only "Normal" matter is made from these particles

Group II:
Unstable – circled decay into Family l particles.

Group III:
Very unstable and circled decay into Family II and I particles.



Quantum Mechanics differs from classical physics in that Energy, Momentum, Angular Momentum and other quantities in a system are restricted to discrete values (quantisation). Particles have characteristics of both particles and waves (duality) as do objects made of many particles.

Photons - Einstein in 1905, explaining the Photo-electric effect, proposed that light consisted of photons, which could be regarded as particles but also having wave-like properties, of energy *h*f. It helped explain the well known wave phenomena of optical interference and diffraction. Einstein received the Nobel Prize for this work. Louis de Broglie applied Einstein's conclusions on photon quanta to other types matter. He proposed all particles could be treated as matter waves with a wavelength λ by the following equation:

$\lambda = h/mv$ where h is Planck's constant (6.626176×10^{-34} joule-seconds), **m** is mass, and v is velocity.

Later Erwin Schrödinger and Werner Heisenberg independently developed modern quantum mechanics. This model, unlike the Bohr model, does not define the exact path of an electron but rather predicts the probability of finding an electron in a particular location. The atom can be portrayed as a nucleus surrounded by a cloud of electrons. This model introduced the concept of sub-energy levels and also described the electron as a wave with an associated *wavefunction* denoted as Ψ (psi). At points where the square of the wavefunction (actually mathematically its complex conjugate) Ψ^2 is most dense the probability of finding the electron is greatest; and conversely, the electron is less likely to be in a less dense area of the cloud.

In principle any problem in quantum mechanics can be solved by writing down the Schrödinger equation for the system under consideration and then solving that equation for Ψ. This then delivers quantisation in three key areas of electron behaviour in an atom: **n**: the principal quantum number (associated with a definite particular energy level); **l**: orbital angular momentum; **ml**: magnetic moment. Ψ^2 the probability density, then also defines the shapes and sizes of electron orbitals.

Spin – Particles including electrons have a property called spin, which is analogous to a classical object spinning about an axis (e.g. the Earth rotating). It is therefore a form of angular momentum, not to be confused with orbital angular momentum which is analogous to the Moon orbiting the Earth. Spin was introduced into quantum mechanics by Wolfgang Pauli as an add-on necessary to complete the quantum picture of the electron in an atom – Schrödinger's formulation did not account for spin. Later Paul Dirac showed that electron spin arose as a natural consequence of quantum behaviour and integrated it into the quantum model so that it no longer appeared as an add-on.

Electrons within the Atom - Within atoms, electrons are organized into Shells and Orbitals with the number of orbitals in a shell given by the square of the principle quantum number. Each shell is associated with an energy so that all the electrons in the shell have the same energy whichever orbital they occupy. Early/simple models (Bohr) show electrons orbiting the nucleus in a near-circular orbit. Actual behaviour is far more complex: some orbitals resemble spheres; others have dumb bell or other shapes. Technically, an electron can be found anywhere within the atom, but will mostly be in the region where Ψ^2 for the orbital is greatest.

The principal quantum numbers are labelled as integers n: 1, 2, 3, 4 … ; the higher the number the greater the energy required to occupy that level. Each principal energy level can contain up to $2n^2$ electrons, many in higher energy levels occupying sublevels (s, p, d & f), each with the same energy:

Principal Quantum No.	Sublevel	No. of Orbitals in Sublevel	Total possible No. of Electrons	
1	s	1	2	
2	s	1	2	) 8
	p	3	6	)
3	s	1	2	)
	p	3	6	) 18
	d	5	10	)
4	s	1	2	)
	p	3	6	) 32
	d	5	10	)
	f	7	14	)

The description of the atom provided by quantum mechanics explains why the elements can be arranged in a periodic table where elements in a column have similar chemical properties.

The Periodic Table of Elements

Column number = Groups containing the same number of electrons in their Outer Shell (*exception*: Helium)
Row number = Highest Principal Electron Energy levels for that element

Group →	1	2	3		4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																		2 He
2	3 Li	4 Be												5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg												13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc		22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y		40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 La	*	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 Ac	**	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og
				*	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	
				**	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr	

Atomic Number = number of protons and electrons.
Atoms with full electron shells/ levels are stable (e.g. Inert Noble gases).
Atoms with vacancies in their outer electron (or Valence) shell are reactive.
Atoms with same number of electrons in their outer shells/levels have similar properties.

Quantum Field Theory - Mr Pilkington concluded with a description of the complex theory designed to account for how bosons are responsible for mediating the four possible interactions (forces) between particles, and for mass (the Higgs). A field is a space with a value associated with each point in the space, for example the temperature in a room has a value everywhere and therefore forms a (scalar) temperature field. The quantity at each point need not be a scalar but could be a vector quantity such as magnetic field, giving a vector field. A field theory describes how a field changes in response to some disturbance, while a quantum field theory does the same for a field where the values at each point are discrete.

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