k3 Rearrangements of the Periodic Table < weight; mass, number>

In Lavoisier's influential textbook *An Elementary Treatise on Chemistry*, published in 1789,[1] thirty-three basic substances were identified as elements. The list began with caloric and light before continuing with oxygen, nitrogen, and hydrogen. Two decades later the great Swedish chemist Jöns Jakob Berzelius divided matter into ordinary elements and compounds, plus five additional weightless and invisible substances: positive and negative electricity, magnetism, light, and caloric.
—Hans Christian von Baeyer, *Maxwell's Demon*, 1998.[2]

"The Republic has no need for savants" opined the Paris mob and the sentencing judge of Antoine Lavoisier (b. 1743, and when a student: "I am young and avid for glory," and guillotined in 1794).[3]

Elements are massy (weighable) substances that cannot be formed by mixing, or chemical union of other substances, nor can they be decomposed by ordinary types of chemical change. Naturally occurring elements (ninety-two in all) are distinguished by their weights, which are different for equal volumes of each (their atomic weights), and can be classified by their chemical behavior (alkali metal, transition metal, nonmetal, inert). The smallest elemental particle is called an *atom*. Atomic weights are scaled so that one atom of the element oxygen (Lavoisier's misnomer—Gk. *oxy* (sour) -*genes* means *acid forming*) has the exact atomic weight of 16. Alternative scalings are for carbon to be 12, or for hydrogen to be 1 (Lavoisier, Gk. *hydro-genes* means *water forming*) and so used by the pioneering first modern atomist William Prout (1785-1850) to scale the forty-or-so described elements of his day, and which John Dalton ca.1803 had guessed is the lightest. In 1911, Jöns Jakob Berzelius (1779-1848) introduced chemical symbols for elements (a capitalized first with, if needed, a lowercase second letter of their Latin names; viz: O for oxygen, Fe for iron).[4]

Dmitri Ivanovich Mendeléev
In February 1869, he had written on cards the symbols and weights of the 63 chemical elements known to him. These he played with as in *Patience*,[8] an open solitaire card game that begins with a shuffled deck dealt face-up into a row of columns followed by moving top cards or stacks of these to achieve an array ordered by suite (row) and rank (columns). His cheat was that he began play with an ordered deck but he was not certain as to the number of suites. Tired, he fell asleep. Then: "I saw in a dream a table where all the elements fell into place as required" with suites of elements of like chemical properties and ranking by atomic weight.[9]

The two criteria, atomic weight and chemical behavior, allow elements to be organized as an array called the *Periodic Table* (***Figure k3.1***) the first by Julius Lothar Meyer (1830-1895) in 1864 (published 1870) and famously by **Dimitri Ivanovich Mendeléev** (1834-1907) who boldly left spaces among the 63 elements he listed for three, now called *Gallium* (found 1875), *Scandium* (found 1879), and *Germanium* (found 1885 in Himmelsfurst mine, Freiberg, Germany)[5] with the properties he foretold.[6] Yet success was limited by the use of atomic weight. This is *not* a fundamental characteristic of an element (atomic number is, as isotopic studies would reveal). In 1913, Anton van den Broek suggested an ordering criterion for the Periodic Table be the nuclear charge, and not the atomic weight, of each atom. To test this hypothesis, Henry Mosely in the same year performed X-ray spectrographic studies of ten elements that followed each other consecutively in their atomic weights and found that "there is in the atom a fundamental quantity, which increases by regular steps as we pass from one element to the next." In 1920, Ernest Rutherford found this is the number of protons in the nucleus (his 1911 discovery of which proved atoms exist!).[10] Thus an ordinal number, called its *atomic number*, can characterize each element, and this (instead of its fractional atomic weight) now positions an element in the Periodic Table. For example, Hydrogen 1 instead of 1.008, Helium 2 instead of 4.003, ... Oxygen 8 instead of 16.000, ...).

The Periodic Table reordered according to atomic numbers instead of atomic weights, places a few elements in swapped positions (argon now precedes potassium, and tellurium comes before iodine), and this improves the chemical periodicity revealed. By then it was also known that element uranium spontaneously changes to other elements.

Radioactive atoms

In 1895 (the year before the discovery of the electron by Joseph John Thomson!) Antoine Henri Becquerel (1852-1980) serendipitously found uranium salts cloud photographic plates through light-shielding metal foil and that their luminescence is secondary to an invisible radiation. This property possessed by some elements of spontaneous emission of invisible particles or electromagnetic rays by the disintegration of their atomic nuclei he announced in 1896[11] and for it Marie Curie (discoverer in 1898 of polonium and radium) coined the name "radioactivity" in 1899.[12] By virtue of these unseen in radium's delicate blue glow that "looked like faint, fairy lights," Mme. Curie, née Maria Sklodowska, (b. 1867) died of leukemia in 1934.[13]

In 1902, Ernest Rutherford (unloved, later, for his put down when at Manchester of such as Greenough-style geology: "There's physics, and there's stamp collecting.") and Frederick Soddy reaped ridicule for publishing *The Cause and Nature of Radioactivity* in which they correctly, it turned out, explained how radioactive inert gas radon emanates from thorium (discovered in 1899 by Rutherford, who called it *thoron*) and from uranium (as Friedrich Dorn found in 1900).[14] Atoms of the one disintegrate into those of the other. The spontaneous transform is called *radioactive decay*.

An atom consists of a tiny nucleus occupying (like a fly in a stadium) some one-quadrillionth the volume of the atom but with almost all the atom's mass and the positive electrical charges that hold in orbit the atom's lightweight electrons each with one negative charge (1–). This Rutherford deduced (originating particle physics) in 1911 from observations made in 1909 by colleagues Hans Geiger (and his eponymous counter) and Ernest Marsden that only about 1 in 10,000 alpha particles shot through a thin sheet of gold hits anything heavy enough to cause it to bounce back and that deflections of the others reveal two positive charges (2+) on each heavy nuclei.[10]

Alpha particle bombardment of nitrogen, Rutherford found in 1919, produces oxygen and particles identical to the nucleus of a hydrogen atom. Thereby, he discovered the proton and simultaneously proved that atomic nuclei *do* transform in that one element becomes another (realizing the alchemist's dream). Indeed, particles with single positive charge and the mass of a hydrogen nucleus (a proton) can be knocked out of the nucleus of any element. Alpha particles, which are identical to the nucleus of helium atoms weigh four times as much as a proton. With atomic number two, the helium atom nucleus contains two protons (2p). The additional weight is due to two neutrons (2n). Neutrons, first described in 1932 by James Chadwick,[15] are particles with no charge. In a nucleus, they allow for a closeness of protons which, because like-charges repel, would not hold together otherwise. A neutron (of 1 up (*u*) quark and 2 down (*d*) quarks, or *udd*) that has been knocked out of a nucleus rapidly decays to a proton (*uud*)[16] and 2 leptons (an electron and an antineutrino that flees away).[17]

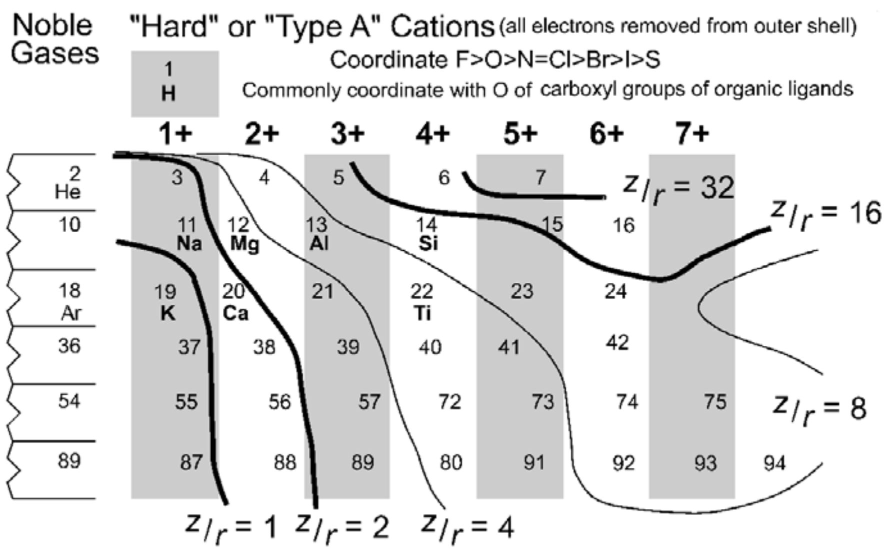
Note: Chemists originally referred to weights of substances but this is not very precise as weight is dependent on the force of gravity (and so will vary from place to place), whereas mass is the measure of the massiness of a substance at rest. Careful mass-spectrometer measurements showed that individual atoms of each element exist with quite different rest masses (the first shown so by Francis William Aston in 1920 was stable neon with two isotopes of atomic masses 20 and 22—but not quite it was later shown). Isotopic varieties of an element have different numbers of neutrons in the nucleus of each. The number of neutrons in the nucleus does not affect the ordinary chemical nature of an element. That is, the atomic number of an element is not changed by the number of neutrons in its nucleus. However, the stability, or radioactivity, of a nucleus is affected (no clear-cut rule exists) by the number of neutrons it contains.

The chemically pure substances listed in the Periodic Table are now referred to as *elements* when their chemical behavior is considered, and as *isotopes* when to the count of protons (atomic number) in the nucleus of an element's atom is added the number of neutrons. So, an element's atoms all have the same number of protons and all are isotopes that, owing to neutrons, can differ in atomic masses.

Isotope geology is the study of isotope equilibrium abundances that environmental conditions determine in chemical compounds.[18] ❑

Figure **k3.1** [19] Theme and part of *An Earth Scientist's Periodic Table of the Elements and Their Ions* by L. Bruce Railsback.[20]

The table is of chemical entities arranged by charge (*Z*), ionic radius (*r*), and charge density (*Z*/*r*) contours. The sample of it here shows, left to right: 1) noble gases, 2) hard or type A cations (hard cations with no outer-shell-electrons bond strongly to F^- and O^{2-} but not to S^{2-}); and (not shown): 3) intermediate to soft or type B cations (have at least some outer-shell cations, bond strongly to S^{2-} and the larger halides, Br^- and I^-), 4) elemental (uncharged) forms, 5) anions, and 6) the noble gases again. Because different natural conditions cause most elements to assume different charges, many elements (e.g., P and U) thus appear twice, a few appear three times (e.g., V, Fe, C, and N), and a few appear four times (most notably S, as S^{2-}, S^0, S^{4+}, and S^{6+}).



Uranium, atomic number 92, is the heaviest of the naturally occurring elements. Transuranic elements to element 103 (its atomic number) occur fleetingly in nature. Isotopes of these, and higher numbered ones, twenty three in all so far, made, some in abundance by physicists operating particle accelerators and nuclear reactors, have, most of them, extremely short half-lives.[21] However, element 114 does last a fair time, and isotopes of elements 112 and 108 in their decay chains have been observed to last 15 minutes and 17 minutes, respectively, before disintegrating.[22] Americium (element 95), is used in smoke detectors, and plutonium (element 94), is used in fission atomic bombs.

All matter is hypothetically of 12 things: 6 quarks and 6 leptons. Nucleons (protons and neutrons) are comminglings of 3 quarks. An electron is a single lepton.[23]

Figure **k4.1** [1] Nuclides graphed - number of neutrons (horizontal) - number of protons, Z, (vertical). Plotted are isotopes: stable (gray), long lived radioactive (white) as are Potassium-40, Rubidium 87, Samarium 147, Thorium 232, and Uranium 234, 235, 238 useful for rock and mineral dating, and short lived radioactive (black) as is Carbon-14. None of the short lived can have existed from the beginning of geologic time and so have limited application in geology for dating ages of archeological objects and determining rates of on going processes. Isotopes in addition to those shown may yet be identified.