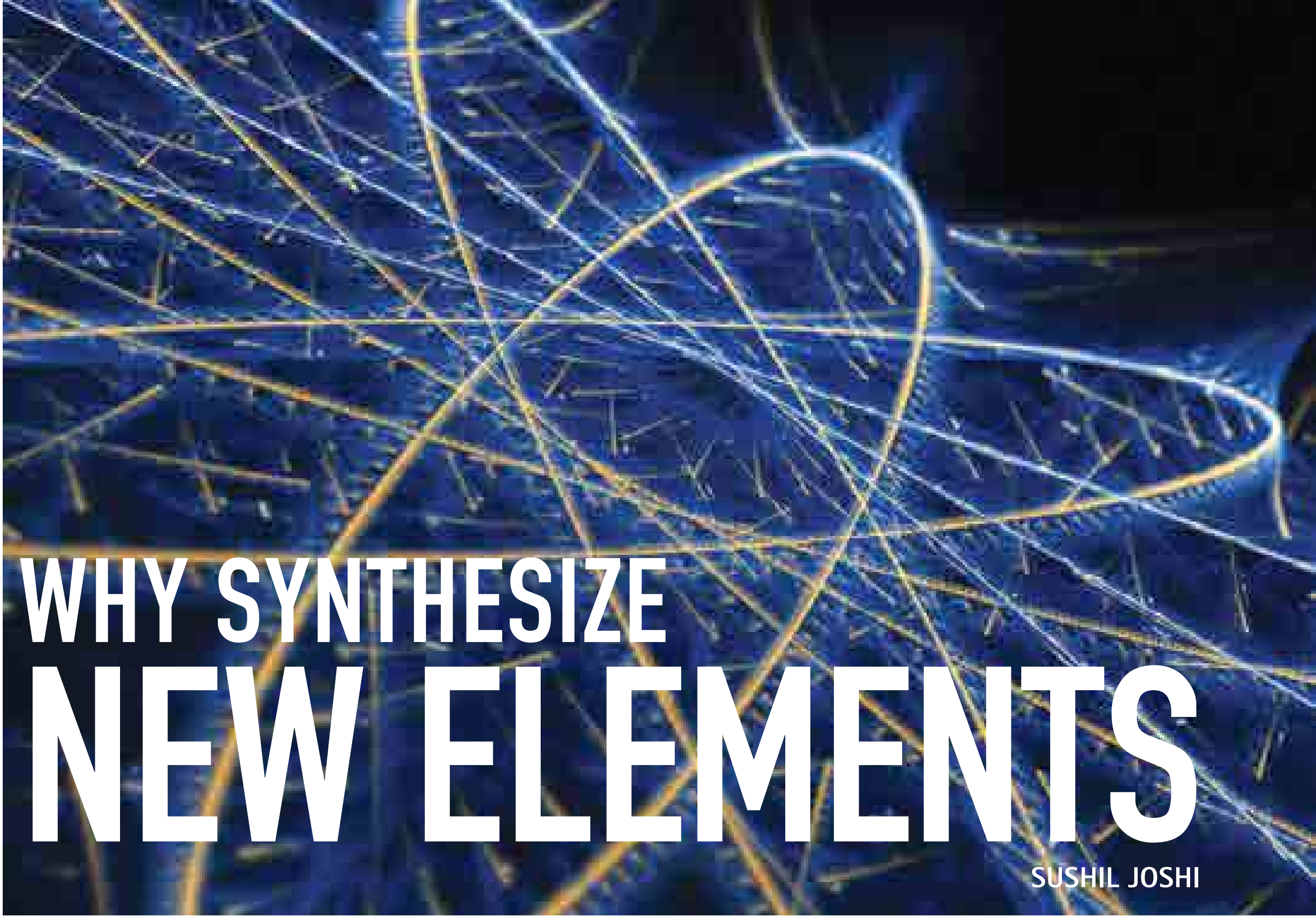

WHY SYNTHESIZE NEW ELEMENTS

SUSHIL JOSHI

We are in a race to synthesize new elements — each with more protons and neutrons than ever before. Unlike naturally occurring elements, these newly created elements seem ephemeral — ceasing to exist almost as soon as they are created. What makes them so unstable? Why do we synthesize them? What does their creation reveal to us?

In September 2009, a team of scientists from the U.S. Department of Energy's Lawrence Berkeley National Laboratory confirmed a claim made by a team of Russian scientists at the Joint Institute for Nuclear Research in Dubna in 1998 — it was possible to create element 114. The Russian team (led by Yuri Oganessian) had described element 114, now called flevorium, as being "very stable". Today, however, this element does not exist.

The short-lived element 114 is not an exception. A number of super-heavy elements invented or created (refer **Box 1**) in the last couple of decades don't seem to be very stable. In contrast, 92 elements, the heaviest of which (uranium) has 92 protons, occur naturally. This has led to the hypothesis that elements with more than 92 protons may not be stable in nature. What is it that makes one element stable, and another unstable? And why invest effort in creating new elements, if they are literally *kshan-bhangur* (transient)?

The effect of proton number

Every element is identified by its atomic number, which is the number of protons in its nucleus. For example, flevorium has an atomic number of 114 — the number of protons in its nucleus. For each atom of this element to be electrically neutral, it would require the same number of electrons. Another property of an element is its atomic weight, defined as the sum total of protons and neutrons in its nucleus. Irrespective of where the element is found, its atoms carry the same number of protons, but may have a variable number of neutrons. For example, the element hydrogen has just one proton in its nucleus. Atoms of its most abundant form have no neutrons, making their atomic number and atomic weight equal. However, some hydrogen atoms have one or two neutrons, with atomic weights of 2 (deuterium) or 3 (tritium) respectively, although their atomic number remains the same (refer **Fig. 1**). Atoms with the same atomic

Box 1. Synthesis of new elements

In theory, all it takes to create new elements is to bombard atoms of two elements with appropriate atomic numbers at each other at very high speeds. Some of these will fuse and gift you with new elements. The atomic numbers of the new elements will be a sum of those of the individual elements with which you start the process.

In reality, this is easier said than done. To achieve these high speeds (~10% of the speed of light) of collision, atoms are accelerated in a cyclotron. When two atoms collide, the repulsive force between their nuclei tends to throw them apart. Consequently, most atoms in such energetic collisions are shattered; but a few may fuse to give rise to a new element.

The presence and nature of new nuclei is inferred from piecing together data from a series of disintegrations and a careful analysis of the products of their decay. For example, new elements tend to be much bigger in size, move far more slowly (~2% of the speed of light), and respond differently to a magnetic field.

number but differing atomic weights are called isotopes of that element.

Most elements have two or more isotopes. In fact, the existence of isotopes seems more a rule than an exception. While isotopes can vary in their physical properties, their chemical properties are identical. For example, charcoal has a mixture of the three isotopes of carbon — all of which have the same atomic number (= 6), but atomic weights of 12, 13 and 14 respectively (refer **Fig. 2**). On combustion, all three isotopes burn with equal ease and are found in the same proportions in the resulting carbon dioxide as in charcoal. Naturally, all isotopes of an element are placed in the same house in the periodic table. However, the isotopes of an element can vary with respect to their individual stabilities. For example, atoms of carbon-12 and carbon-13 are quite stable, while atoms of carbon-14 are relatively unstable and break down spontaneously. Thus, if you were to keep 10g of a carbon-14 sample for 6000 years (called the half-life of carbon-14), you would find only 5 g left at the end of this period, with the rest having disintegrated into nitrogen and escaped into air. Why are some isotopes of an element stable while others are not?

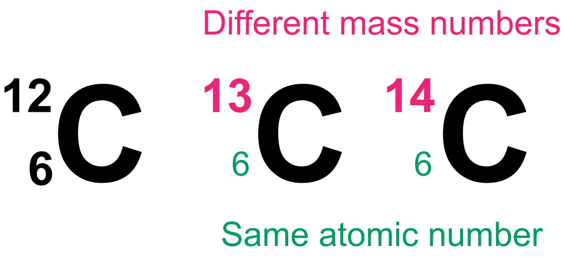

Fig. 2. The three isotopes of carbon.

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One explanation attributes this difference in the stability of isotopes to the number of protons in an atom. While protons and neutrons are packed into the nucleus of an atom, electrons revolve around it. Since the mass of an electron is negligible, all the mass of an atom is packed into its nucleus. However, the volume of the nucleus is negligible compared to that of the atom. In fact, if the size of an atom were like that of a football field, the nucleus is likely to occupy the volume of a tennis ball. Since protons are positively charged and neutrons are electrically neutral, any two protons in the nucleus will repel each other while neutrons will not interact with each other or with protons. With an increase in atomic number, the number of protons packed into the tiny volume of the nucleus increases. So does the repulsive force acting between them. Consequently, the stability of the atom is likely to decrease. Indeed, elements with an atomic number of 20 and less are observed to be relatively stable; whereas atoms of elements with an atomic number greater than that show increasing instability. The stability of helium, however, presents an interesting exception. If stability is based solely on the amount of repulsion between the protons in its nucleus, this element with its two protons should have been unstable.

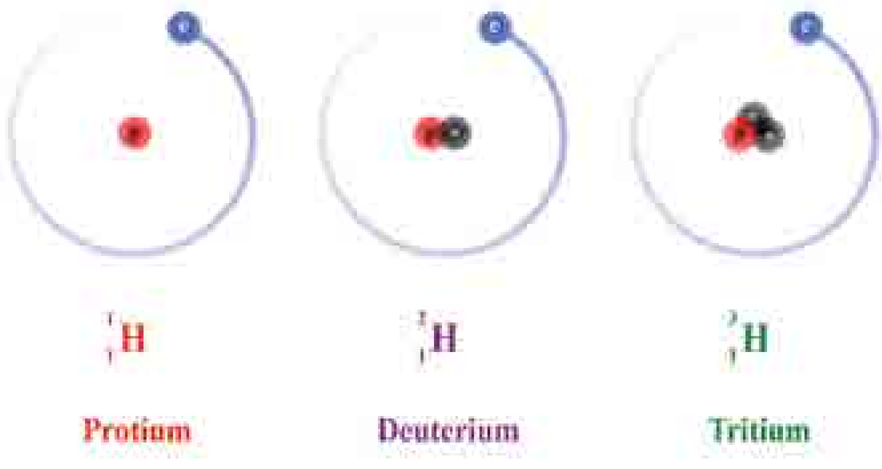

Fig. 1. The three isotopes of hydrogen.

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The strength of forces that hold the nucleus together

Fortunately for the universe, nuclear interactions are not shaped solely by repulsion between protons. According to the laws of physics, subatomic interactions are shaped by another force – the strong force. This is an attractive force, but acts only across the short distances between individual protons and neutrons within the nucleus. Thus, at any given point of time, there are two forces acting between the protons in a nucleus – an electrical repulsive force and an attractive strong force. If an atom has a small number of protons and a small nuclear volume, the strong force between the protons overcomes their electrical repulsion, and the atom remains stable. An increase in the number of protons increases both the electrical (repulsive) force and the strong (attractive) force between them. Since distances between protons in the larger nucleus will also increase, the attractive force will get weaker, and the atom will show a higher probability of disintegrating. Does this mean that all elements with higher atomic weights are unstable?

This is where neutrons seem to play an important role. Being situated amidst the protons in a nucleus, neutrons can shield protons from the repulsive force acting between them. Neutrons are also capable of exerting the strong force, thereby increasing the total attractive force in the nucleus and adding to its stability (refer **Fig. 3**). This implies that an element with an atomic number over 20 will be stable if it has more neutrons than protons in its nucleus. In reality, however, all heavy (atomic numbers > 20) and super-heavy (atomic numbers > 100) elements have more neutrons than protons (N:P ratio > 1). And, the most abundant isotopes of light elements (or elements with an atomic number less than 20) have an equal number of neutrons and protons in their nuclei (N:P ratio = 1). This has led to the hypothesis that beyond a certain atomic number, any increase in nucleons (protons and neutrons) in an atom will increase its instability. In other words, heavier atoms are, in general, more unstable. Broadly speaking, this turns out to be true.

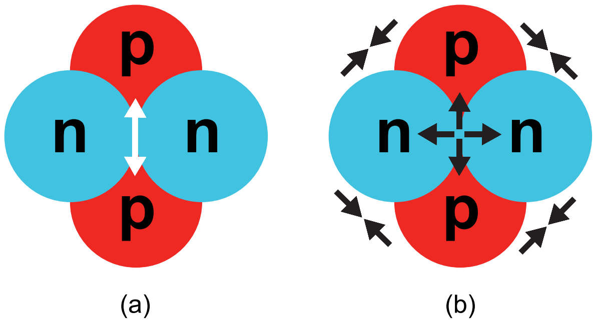

Fig. 3. The stability of an atom is determined by two opposing forces in the nucleus of an atom. (a) An electric repulsion pushes the protons in a nucleus apart. **(b)** An attractive strong force between the protons and neutrons holds the nucleus together.

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The existence of 'islands of stability'

While the instability of elements tends to increase with increasing atomic number, this does not happen as a continuum. The sudden appearance of atoms that are relatively stable (or more stable than is expected from their atomic numbers alone) punctuates it.

Box 2. Electron configurations

Electrons of an atom are arranged in concentric shells around its nucleus, with certain arrangements being more stable than others. This distribution of electrons is expressed by the mathematical rule: 2 (n^2), where n is the serial number of the shell. Thus, the first shell can have 2 [= 2 (1^2)] electrons, the second can accommodate 8 [= 2 (2^2)] electrons, the third can have 18 [= 2 (3^2)] electrons and so on. This rule has one proviso, referred to as the famous rule of eight – the last or outermost shell of an atom cannot contain more than 8 electrons. When the last shell is completely filled, as in the case of noble gases, the atom is stable. If the outermost shell has less than 8 electrons, the atom tends to react with other atoms to either complete or empty (give up this shell altogether) its outermost shell. The only exceptions to this rule are hydrogen and helium – both have one shell that can accommodate a maximum two electrons. With only one electron in this shell, hydrogen is very reactive. On the other hand, with a completely filled shell (having two electrons), helium is inert.

This creates a zigzag trend of increasing instability (sea of instability) with some **islands of stability.** In other words, certain atoms are relatively more stable than predicted on the basis of their atomic number. The atomic numbers of these atoms, dubbed as magical atomic numbers, have been calculated to be 2, 8, 20, 28, 50, 82, and potentially 126. Of these, elements with atomic numbers as high as 92 exist naturally. Scientists are attempting to create elements with atomic numbers as high as 118 in laboratories in the hope of testing these theoretical predictions.

The existence of islands of stability implies that nucleons may not be randomly packed in the nucleus; they may exist in some definite arrangements. That certain electron configurations tend to be more stable than others lends support to this idea (refer **Box 2**). This has led to the hypothesis that certain arrangements of nucleons may contribute to the unexpected degree of stability that some elements exhibit. While several models of nuclear arrangement have been proposed, attempts to express the supposed magical atomic numbers in some kind of mathematical formula are yet to succeed.

Some interesting observations

There have been other attempts to explain the stability of atoms. For example, it has been observed that elements with even atomic numbers (2, 22, 76 etc.) show greater stability and have more stable isotopes than those with odd atomic numbers (refer **Table 1**).

Number of stable isotopes of elements with even atomic number	170
Number of stable isotopes of elements with odd atomic number	63

Table 1. Elements with even atomic numbers have a larger number of stable isotopes.

Analysis shows that the stability of these isotopes may also depend on the number of neutrons (refer **Table 2**). An atom seems to achieve maximum stability when it has even numbers of protons as well as neutrons. If either of these numbers is odd, it compromises the stability of the atom. While the calculations that explain these patterns are quite complicated, scientists have used them to speculate that an element with an atomic number of 114 would be relatively stable. On synthesis, this element did indeed turn out to be relatively (to its atomic number) 'stable'.

Number of protons	Number of neutrons	Number of stable isotopes
Even	Even	163
Even	Odd	53
Odd	Even	50

Table 2. Elements where both protons and neutrons are even-numbered tend to have a larger number of stable isotopes.

To conclude

While the synthesis of super-heavy elements may seem pointless from a utilitarian view of science, it helps refine our understanding of factors responsible for stability. It may be pertinent to note that often, we tend to measure the stability of an element in terms of its ability to survive for years, decades or even centuries. Element 114 is not stable with respect to such time scales. It is not even as stable as predicted and, therefore, not a hypothetical island of stability (refer **Table 3**). However, given that most elements break down the instant they are created, element 114 is stable in the sense that it survived long enough for us to know that it exists. Its synthesis has strengthened our quest to answer questions like — do islands of stability really exist? Are there any limits to the periodic table? The next stability island is predicted to occur at an atomic number of 126. Can this element be created in laboratory? Will its synthesis bring us any closer to these answers? We'll have to wait and see.

Name	Symbol	Atomic number	Most stable isotope	Half-life of its most stable isotope
Nihonium	Nh	113	^{286}Nh	9.5 s
Flevorium	Fl	114	^{289}Fl	1.9 s
Muscovium	Mc	115	^{290}Mc	650 ms
Livermorium	Lv	116	^{293}Lv	57 ms
Tennessine	Ts	117	^{294}Ts	51 ms
Oganesson	Og	118	^{294}Og	0.69 ms

Table 3. The relative stability of some newly synthesized elements.

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