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Unraveling Isotopic Elements: A Quark-Based Investigation

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Abstract: ***Isotopes of an element share the same atomic number, indicating they possess identical proton counts, yet they differ in atomic mass due to variations in neutron numbers. While isotopes are conventionally analyzed in terms of protons and neutrons, this paper explores an alternative perspective by examining isotopes through the lens of quark composition. Quarks, the fundamental constituents of protons and neutrons, come in six varieties: up, down, charm, strange, bottom, and top. However, only up and down quarks contribute to the formation of the stable baryons protons and neutrons that constitute atomic nuclei. This study begins by delineating the relationship between isotopes and their underlying quark structures, focusing exclusively on the stable up and down quarks. By investigating the number and arrangement of these quarks within various isotopes, we aim to uncover intrinsic patterns that correlate with isotopic behavior. Utilizing the modern periodic table as a foundational framework, we assess the elements based on their atomic number, which corresponds to their proton count, and subsequently analyze how these factors influence isotopic variations. In our comprehensive evaluation, we derive relationships concerning the counts of up and down quarks in the nuclei of isotopic elements, revealing critical insights into the stability and behavior of these isotopes. The developed formulas were rigorously tested across a range of isotopes, encompassing both lighter and heavier atomic masses, thereby ensuring the robustness of our findings. Furthermore, we critically analyze these relationships to draw hypotheses regarding the interplay between quark composition and isotopic stability, contributing to a nuanced understanding of atomic structure. Ultimately, this paper emphasizes the significance of quark-based investigations in advancing our comprehension of isotopes, suggesting that a deeper insight into quark dynamics may not only illuminate the complexities of isotopic behavior but also pave the way for future research in nuclear physics and related fields.***

Keywords: isotopes, quark composition, atomic stability, nuclear physics, proton and neutron structure

1. Introduction

Elements are ultimately made up of atoms, which in turn, consist of a nucleus and individual electrons revolving around the nucleus. The quark is an essential fundamental particle in the Standard Model which attempts to explain the fundamental forces of nature (except gravity) in terms of particles, i.e., either fermions [4] or bosons. The fermion constitutes physical matter, while the exchange of the boson gives rise to forces between particles. The electron is a fermion. Fermions can be divided into two groups: quarks and leptons. Leptons are weakly-interacting particles (particles that interact using the weak nuclear force), for example, electrons. The nucleus, made up of protons and neutrons, is comparatively massive and lies at the center of the atom. Protons are made up of two up quarks and one down quark. Neutrons, which are charge less particles, have two down quarks and one up quark. (It should be noted that, as per modern science, quark-antiquark pairs may be present in the protons or neutrons. The antiquark is the antiparticle for the quark. It has similar properties but is oppositely-charged. Being equally and oppositely-charged, quark-antiquark pairs do not affect the overall charge of the protons or neutrons.)

Atoms are neutral, i.e., they have equal amounts of positive and negative charges. This is because the electrons are negatively charged, while the protons have a charge equal in magnitude to that of the electron but positive in nature. The number of protons and electrons are equal in a neutral atom. When this does not hold, an ion is formed, which can be positively-charged (cation) or negatively-charged (anion). Here, however, the discussion is regarding isotopes of neutral atoms.

The neutron is a chargeless particle because the relative charge of an up quark is positive two-thirds the charge of an electron, i.e., $\frac{2}{3}$ units; while the relative charge of a down quark is negative one-third the charge of an electron i.e., $-\frac{1}{3}$ units. The combination of two up and one down quark in a proton will have a relative charge of positive one units $[2(2/3)+(-1/3) = 4/3 – 1/3 = 3/3 = 1]$, while it will be zero for a neutron (i.e., two down and one up quark).

Isotopes of an element have atoms that have equal numbers of protons and electrons but excess neutrons, which affects the mass but not the charge. The mass of the isotope may exceed the original mass of the element by one atomic mass unit, two units or more. Most elements with lower atomic numbers (or masses) have an atomic mass approximately equal to twice the atomic number. This is because the atomic number is just the number of protons or, in case of a neutral atom, the number of electrons. The mass number is just the sum of the number of protons and the number of neutrons (electrons have almost no mass). It must be noted that the isotopes of an element are found in different quantities in nature. The abundance percentages of an isotope define what percentage of that isotope constitute all isotopes of that element in nature.

Isobars [5], on the other hand, refer to elements with different atomic numbers, but the same mass numbers. The occurrence of isobars is much lower as compared to the occurrence of isotopes. The mass numbers of elements with different atomic numbers may be equivalent to one another due to their number of neutrons being the same. It must be emphasized that isotopes, especially radioisotopes [6], have wide-ranging uses. They are used in medical sciences for cancer treatment, in nuclear power plants, etc. For a better understanding of nuclear energy, and indirectly, for a better understanding of the universe, the study of subatomic particles is crucial. They can be referred to as the most fundamental unit in the universe.

2. Methods

The methods used to arrive at the results have been discussed here.

1) Ignoring quark-antiquark pairs, a proton consists of two up and one down quark. Thus, it can be represented as 2u 1d. Similarly, a neutron has 2 down and 1 up quark, i.e., 2d 1u.
2) Isotopes, generally, have more neutrons than protons. This is because, as long as the isotopes are of the same element, their atomic number, and hence, the number of protons, remains the same. Thus, although there are certain exceptions, heavier isotopes usually have more down quarks than up quarks, since neutrons contain more down quarks than up.
3) The difference between the mass number of the heavier isotope and the average atomic mass of that element, when all its isotopes are considered (if, and only if the average atomic mass has been calculated by taking into account the abundance percentages), added to the difference between the number of up and down quarks, also calculated by taking into account the abundance percentages, must always be a whole number, other than zero.
   Also, the difference between the mass number of the heavier isotope and the average atomic mass of that element, can't be zero if the element has isotopes.
   [Let A and B be the lighter and heavier isotopes, respectively, and let C be the average atomic mass of that element. If (B – C) = 0, B = C, which indicates that the given element has no isotopes. This is because B can be equal to C when, and only when, one isotope is 100% abundant and the other (say, A) is 0% abundant. In that case, C=0/100 X A + 100/100 X B = 0 + B = B.]
4) For an efficient approach to the issue, isotopic elements can be divided into three 'classes': Class 1 elements, Class 2 elements, and Class 3 elements.
5) Class 1 elements are elements with two naturally-occurring isotopes such that:
   a) The mass number of the lighter isotope is exactly twice the atomic number of the element;
   b) The difference between the mass numbers of the isotopes is exactly 1 atomic mass unit, i.e., the heavier isotope is heavier than the lighter one by 1 atomic mass unit.
6) Class 2 elements are elements with two naturally-occurring isotopes such that:
   a) The mass number of an isotope is NOT twice the atomic number of that element;
   b) The difference between the mass numbers of the isotopes is 1 atomic mass unit, i.e., the heavier isotope is heavier than the lighter one by 1 atomic mass unit.
7) Class 3 elements are elements with two naturally-occurring isotopes such that:
   a) The mass number of an isotope is not necessarily twice the atomic number of that element;
   b) The difference between the mass numbers of the isotopes is greater than 1 atomic mass unit.
8) Tiny inaccuracies might arise in the calculations due to a number of reasons, which does not necessarily question the validity of the formula. At certain points in the calculations, numbers have been rounded up to two, three or four decimal places, as convenient. For instance, 2.788692213… can be rounded up to 2.79, and so on. Also, the abundance percentages of isotopes can be rounded up similarly as convenient. However, this must satisfy that the sum of the percentage abundance of all the isotopes of that element equal 100.
9) For elements having more than 2 naturally-occurring isotopes (say, b elements, i.e. it can be a Class 1-b element, a Class 2-b element, or a Class 3-b element), any 2 isotopes can be taken into account. Taking into account their abundance percentages (when all the isotopes are considered) and the sum of abundance percentages of these 2 isotopes, their abundance percentages can be modified such that only these 2 isotopes are present, of that element. (This is just being done to test the formula for such cases, for it is not physically valid to 'eliminate' the other isotopes.) In such cases, values extremely close to a whole number, but not exactly a whole number, tend to arise.
10) Considering the abundance percentages of the isotopes and their respective masses, the average atomic mass of the element was found using known methods. The abundance percentages for all the isotopes have been determined over a long time and the working hypothesis, thus, must be that these values are correct to a reasonable degree of accuracy.
    At this point, it should be noted that many elements may have many isotopes, but few of them are stable. In this paper, only the stable isotopes are of relevance and thus, only they have been taken into account. For instance, Lithium has more than two isotopes. However, only two of them, Lithium-6 and Lithium-7, are stable.
11) After that, the difference between the mass number of the heavier isotope and the average atomic mass of that element was considered.

Let an element have two isotopes, A and B such that B is heavier than A. Let C be the average atomic mass. Then, first (B –C) was considered.

Then, the difference between the *total* number of down and up quarks in a single atom of that element [taking into consideration their abundance percentages in nature] was considered, i.e., say (d – u).
Finally, the expression (B – C) + (d – u) was considered for Classes 1, 2 and 3 elements. This is explained clearly with examples under the "Calculations" section.

3. Postulates

1) For neutral atoms, arranged in the order of increasing atomic number, with the exception of hydrogen, the number of up as well as down quarks in the nucleus increase by 3, for every element. [There may be certain exceptions to this rule.]
2) Without considering hydrogen, the number of up quarks in an element (that is to say, in a single atom of that element) cannot exceed the number of down quarks for

the same element. [There are exceptions to this rule, but it holds in most cases. For instance, Helium-3, which is a stable isotope, has 2 protons and 1 neutron, i.e., 5 up quarks and 4 down quarks.]

3) For isotopic elements, if x up and y down represent the total number of up and down quarks in a single atom of one isotope of that element; and a up and b down represent the same for the heavier isotope, a=x+(1m) and b=y+(2m), where m is the difference between the mass numbers of the two (or more) isotopes.
4) In isotopes, the number of down quarks must always exceed the number of up quarks.
5) Radioactive elements must always have a greater number of down quarks than up quarks, i.e., for radioactive elements, the difference between the number of down and up quarks in one atom of that element (taking into account abundance percentages), (d-u) must be any positive natural number (i.e., not a fraction). For general elements, (d-u) may be zero or any positive natural number. [However, Helium-3, as stated earlier, is also a stable radioactive element. There are also other exceptions.]
6) For Class 1 elements,

   n-p=(d-u)

   where (n – p) is the difference between the total number of neutrons and protons in one atom of both the isotopes.

7) For Class 1 elements,

   $|p1–p2|+|n1–n2|=(d-u)$,

   where the Left-Hand Side of the equation represents the sum of the modulus (i.e., only the magnitude) of the difference between the number of protons and that of neutrons [without taking into account abundance percentages] for both the isotopes.

8) For Class 1 elements, B-C+d-u=1

9) For Class 2 elements, B-C+d-u=n,

where n is some whole number other than 0.
More generally, this may be written as:
For Class 2 elements, B-C+d-u=1+z,
where z = the number by which the lighter isotope’s atomic mass is more than twice the atomic number of that element.
This is explained under the “Calculations” section.
Class 3 elements as well, B-C+d-u=n,
where n is some whole number other than 0.

10) It is also evident that C-(d-u)=2Z, where Z is the atomic number of the element. From the following calculations, it would be clear that the expression after the decimal point of the average atomic mass of an element C, and the expression after the decimal point of (d-u)are the same; and their difference yields a number which is twice the atomic number of the element in question. If the element in question has an isotope whose mass is exactly twice the atomic number, then C-(d-u) would equal the mass of that isotope, usually the lighter isotope. [It should be noted that the lightest possible atomic mass of any element cannot be less than twice the atomic number, since the number of neutrons, usually, is NOT less than the number of protons. (However, there are some exceptions like protium, Helium-3 etc.)]

4. Calculations

1) The first calculation verifies the formula established under point 3 of the “Method” section.

Sulfur has 4 isotopes with mass numbers 32, 33, 34, and 36. In S-32, we have 16 neutrons(n0), 16 protons(p+) and 16 electrons(e–). As one neutron has 2 down and 1 up quarks, 16 neutrons will have (16X2)d and 16u quarks, i.e., 32d 16u. Each proton has 2 up and 1 down quark, i.e., 32u 16d for 16 protons. Thus, we have a total of 48u 48d. Here, x = 48, y = 48.

Similarly, it is 49u 50d for S-33, i.e. (x+1m)u (y+2m)d = (48+1)u (48+2)d. It is of importance, here, that m, i.e., the difference between the mass numbers, is 33 – 32 = 1.

It is 50u 52d for S-34.

For S-36, when S-34 and S-36 are considered, m = 2. Thus, au bd is (50+2)u (52+4)d = 52u 56d. One might choose to consider S-32 and S-36, in which case, m = 4. Thus, au bd is (48+4)u (48+8)d = 52u 56d.

This can be verified for other isotopes as well.

2) The following calculations verify the formula presented under point 8 of the “Postulates” section.

Lithium has atomic number 3 and 2 isotopes of mass numbers 6 (i.e., 3 X 2) and 7. [7 – 6 = 1] Thus, Lithium can be considered a Class 1 element, as per the convention presented under the “Methods” section.

The average atomic mass, taking into account abundance percentages is 6.925 units.

Lithium-6 accounts for approximately 7.5% of all the Lithium available, while 92.5% of it is Lithium-7. Thus, the average atomic mass:

$$7.5 \times 6/100+92.5 \times 7/100=0.45+6.475=6.925.$$

Here, (B – C) = 7 – 6.925 = 0.075……….(i)

Li-6	Li-7
Neutrons: 6 – 3 = 3	Neutrons: 7 – 3 = 4 (atomic mass is always 3)
Protons: 3	Protons: 3
6d 3u	8d 4u
+ 6u 3d	+ 6u 3d
= 9u 9d	= 10u 11d

Considering the fact that 7.5% Lithium is Li-6 and the rest is Li-7, we get:
7.5 X 9/100u7.5 X 9/100d and 92.5 X 10/100u92.5 X 11/100d
=0.675u 0.675d and 9.25u 10.175d
=9.925u 10.85d (adding up the total number of up and down quarks obtained in the previous step.)

Thus, d-u=10.85-9.925=0.925…………(ii)

From (i) and (ii), B-C+d-u=0.075+0.925=1.
Boron (atomic number 5) has 2 isotopes of mass numbers 10 (i.e., 5 X2)[20%] and 11[80%].

Proceeding in a similar way, it is observed that the average atomic mass of Boron is 10.8 units.
Thus, (B – C) = 11 – 10.8 = 0.2

For Boron-10, we get 15u 15d; and for Boron-11, we get 16u 17d. Considering the abundance percentages, we get 3u 3d and 12.8u 13.6d. Taking the sum of the up and down quarks, the final result is 16.6d 15.8u.
(d – u) = 16.6 – 15.8 = 0.8
Thus, it is clear that B-C+d-u=0.2+0.8=1.

For Carbon (atomic number 6), there are 2 isotopes, C-12 [99%] and C-13 [1%].
C=99X12/100+ 1X13/100= 11.88 + 0.13 = 12.01.
(B-C)=13-12.01=0.99.

The number of up and down quarks, before considering the abundance percentages, is 18u 18d for C-12 and 19u 20d for C-13. This can easily be predicted from previous formulae.

Taking into account the abundance percentages, the result is 17.82u 17.82d and 0.19u 0.20d. On adding the up and down quarks, the final result is 18.02d and 18.01u. Thus, (d-u)=18.02-18.01=0.01.

B-C+d-u=0.99+0.01=1.

For Nitrogen (atomic number 6), there are 2 isotopes, N-14 [99.6%] and N-15 [0.4%].
C=99.6X14/100+ 0.4X15/100= 13.944 + 0.06 = 14.004.
(B-C) = 15-14.004=0.996.

The number of up and down quarks, before considering the abundance percentages, is 21u 21d for N-14 and 22u 23d for N-15.

Taking into account the abundance percentages, the result is 20.916u 20.916d and 0.088u 0.092d. On adding the up and down quarks, the final result is 21.008d and 21.004u. Thus, (d-u)=21.008-21.004=0.004.

B- C + d – u = 0.996+0.004=1.

From the above example, it is evident that (B – C) and (d – u) are always the abundance percentages of both the isotopes divided by 100, at least for this class of elements. For instance, in the case of Nitrogen, the abundance percentages of N-14 and N-15, respectively, are 99.6% and 0.4%. And (B-C)=99.6/100= 0.996. And (d-u)=0.4100= 0.004.

3) The following calculation verifies the formula presented under point 9 of the “Results” section. The element Vanadium, with atomic number 23 but 2 isotopes with mass numbers 50 (i.e. ≠ 23 X 2)[0.25%] and 51 (i.e. ≠ 23 X 2)[99.75%], (B – C) = 51 – 50.9975 = 0.0025; while (d – u) = 78.995 – 73.9975 = 4.9975.

Thus, (B – C) + (d – u) = 0.0025 + 4.9975 = 5, which is a whole number other than 0. It must be stressed that in the above example, the result would have been 1 if the atomic mass of one isotope (i.e. lighter isotope) was 23 X 2 = 46. But the mass number of the lighter isotope is 46 + 4 = 50. Thus, we could have easily predicted that (B – C) + (d – u) would be 1 + 4 = 5.

4) The following calculations verify the formula presented under point 10 of the “Postulates” section. For Chlorine (atomic number 17), there are 2 isotopes Cl-35 (i.e. ≠17 X 2) [75%] and Cl-37 [25%]. The average atomic mass is 35.5 units.

Here, (B – C) = 37 – 35.5 = 1.5
(d – u) comes out to be 54 – 52.5 = 1.5
Thus, (B – C) + (d – u) = 1.5 + 1.5 = 3, which is a whole number.

For Argon (atomic number 18), there are 3 isotopes: Argon-36 [0.3%], Argon-38 [0.1%] and Argon-40 [99.6%].
If only Argon-36 and -38 are considered, a modified abundance percentage of Argon-36 would be:
0.3/0.3+0.1X 100 = 75%

This is because the abundance percentage of Argon-36, when all the isotopes are considered, is 0.3%. Similarly, the abundance percentage of Argon-38, when all the isotopes are considered, is 0.1% (the numerator). When only these isotopes are considered, the total abundance becomes 0.3 + 0.1 (which thus becomes the denominator) and the percentage abundance for Argon-36, would then be, 75%, as shown above. This result can be achieved by multiplying the fraction by 100.

Similarly, in this case, an abundance percentage for Argon-38 would be 25% (=100-75).

Here, the average atomic mass is 36.5 units.
Thus, (B – C) = 38 – 36.5 = 1.5
And, in this case, (d – u) = 55 – 54.5 = 0.5
Thus, (B – C) + (d – u) = 1.5 + 0.5 = 2, which is a whole number.

Similarly, for Argon-36 and Argon-40, we get,
(B – U) = 40 – 39.984 = 0.016
And, (d – u) = 61.968 – 57.984 = 3.984
Thus, (B – C) + (d – u) = 0.016 + 3.984 = 4, which is a whole number.

It is interesting, at this point, to consider the case of Potassium (atomic number 19), which has 3 isotopes of mass numbers 39[93.26%], 40[0.01%], and 41[6.73%] (approximately). Proceeding as in the case of Argon, Potassium-39 and -40 can be considered. The abundance percentages will, respectively for K-39 and K-40, be 99.99% and 0.01%. Using these values, it is observed that the average atomic mass, when ONLY these 2 isotopes of Potassium are considered, is 39.0001 units.
(B – C) = 40 – 39.0001 = 0.9999
(d – u) = 59.0002 – 58.0001 = 1.0001
Hence, (B – C) + (d – u) = 0.9999 + 1.0001 = 2, which is a whole number.

By now, a clear understanding of the process that is being used to obtain the values for (B – C) and (d – u) is attainable. Thus, only these values for the following elements have been provided.

For K-39 and K-41, after modifying the abundance percentages, (B – C) + (d – u) = 1.8654 + 1.1346 = 3, which is a whole number.

For K-40 and K-41, after modifying the abundance percentages, (B – C) + (d – u) = 0.0015 + 2.9985 = 3, which is a whole number.

Sulfur (atomic number 16) has 4 isotopes of mass numbers 32 [95.02%], 33[0.75%], 34[4.21%], and 36[0.02%].
For S-32 and S-33, (B – C) + (d – u) = 0.9922+0.0078= 1, which is a whole number.
For S-32 and S-34, (B – C) + (d – u) =1.9152+0.0848= 2, which is a whole number.
For S-32 and S-36, (B – C) + (d – u) = 3.9992+0.0008= 4, which is a whole number.
For S-33 and S-34, (B – C) + (d – u) = 0.1512+1.8488= 2, which is a whole number.
For S-33 and S-36, (B – C) + (d – u) = 2.922+1.078= 4, which is a whole number.
For S-34 and S-36, (B – C) + (d – u) = 1.9906+2.0094= 4, which is a whole number.
Copper (atomic number 29) has 2 isotopes: Cu-63 [69%] and Cu-65 [31%]. For Copper, (B – C) + (d – u) = (65 – 63.62) + (98.24 – 92.62) = 1.38 + 5.62 = 7, which is a whole number.

Finally, the case of a much heavier element, Uranium (atomic number 92) and its 2 main isotopes, U-235 [0.7%] and U-238 [99.3%] can be considered.

In this case, (B – C) = 238 – 237.979 = 0.021.

Uranium-235	Uranium-238
Neutrons: 235 – 92 = 143	Neutrons: 238 – 92 = 146
Protons: 92	Protons: 92
(143 X 2)d 143u	(146 X 2)d 146u
(92 X 2)u 92d	(92 X 2)u 92d
327u 378d	330u 384d

= 0.7 X 327/100u 0.7 X 378/100d and 99.3 X 330/100u 99.3 X 384/100d
= 2.289u 2.646d and 330u 384d
= 329.979u 383.958d
(d – u) = 383.958 – 329.979 = 53.979

Hence, (B – C) + (d – u) = 0.021 + 53.979 = 54, which is a whole number.

From point 9 under the "Postulates" section, we have seen that: For Class 2 elements, (B-C)+(d-u)=1+z,

where z = the number by which the lighter isotope's atomic mass is more than twice the atomic number of that element.

If the case of Uranium is considered more closely, it is evident that B-c+d-u equals 1 + z + {(B – A) – 1}

Thus, the formula obtained in the case of Vanadium was only a special case. In the case of Vanadium, (B – A) = difference between the mass numbers of the 2 isotopes = 51 – 50 = 1. Thus, {(B – A) – 1} is 1 – 1 = 0. In the cases where the isotopes differ by a mass of 1 unit, B-c+d-u=1+z would work.

But the general formula to directly obtain the result of (B – C) + (d – u), thus, for 2 isotopes of mass numbers A and B, A < B, with average atomic mass, C, and z = (A – 2Z), where Z = the atomic number of that element is:

$$B-C+d-u=1+z+B-A-1=z+B-A$$

It may also be written in this form:

$$\mathbf{(d-u)=z+(C-A)}$$

This formula directly predicts the value of (d-u) if z, C and A are known, and can be of importance in certain cases.

This formula is far more general than anything achieved so far. The above formula can be verified for ALL the cases considered, and, if possible, even beyond.

For instance, in the case of Lithium, (d-u) was 0.925. The average atomic mass C was 6.925. A, or the mass of the lighter isotope is clearly 6. Atomic number of Lithium, Z = 3. Thus, z = A – 2Z = 6 – 2(3) = 6 – 6 = 0.

$$z+(C-A)=0+(6.925-6)=0.925=(d-u).$$

For elements with the mass number of one isotope precisely equal to twice the atomic number, the formula collapses to (B – A).

Even for elements such as sulfur, in which cases different combinations of 2 isotopes have been considered, the formula works after modifying the abundance percentages.

5. Discussion

The existence of isotopes (more fundamentally, the fact that the number of down quarks must always equal or exceed the number of up quarks in an atom) seems to indicate that nature has a strong preference for down quarks over up, while both these quarks are stable enough. It may be reasoned that in a parallel reality, the number of up quarks must exceed the number of down quarks. Also, it is suggestive that the magnitude of the (relative) charge of a down quark (–1/3) is less than the magnitude of the (relative) charge of an up quark (+2/3). However, the fact that stable isotopes seem to have more down quarks than up quarks can also be understood this way: the neutrons and protons have a strong mutual attraction, but the protons also have an electrostatic repulsion because they are positively-charged and like charges repel while unlike charges attract. To account for the many protons confined in the nucleus, more neutrons are required to ensure the stability of the nucleus. If the number of protons is large, then this repulsion also becomes large and, in such cases, having more neutrons aids in balancing the repulsion. In particular, as we go to higher atomic numbers, the neutron/proton ratio becomes larger. In fact, when this ratio is greater than 1.5

(approximately) for an element, then that element is radioactive.

Also, it should be suggestive that the total relative charge of an up antiquark (i.e., an up quark with the opposite charge of the same magnitude, i.e., in this case, a relative charge of –2/3) and a down quark is:
$-2/3+-1/3= -2/3-1/3=-3/3= -1$, i.e., the relative charge of an electron.

It can be hypothesized that the formation of electrons and quarks, with the evolution of the universe, are linked to a greater extent. However, it must be noted that electrons and quarks are fundamentally different. This is because electrons are leptons, while quarks, according to the Standard Model, form an entirely different group of fermions. As fermions (and bosons) are the most fundamental particles that have been discovered till now, it is unlikely that electrons and quarks might have common characteristics.

Isobars can be assumed to be the result of pure accidents, due, mainly, to the fact that the mass number is, clearly, a less fundamental quantity than the atomic number. (Interestingly, the Mendeleev periodic table ultimately was superseded by the modern periodic table as the former was based on the atomic mass, while the latter was based on a much more fundamental property, atomic number. This is because the atomic number remains constant for all elements while an element may, due to isotopes, have different mass numbers.)

The terms proton and neutron are undeniably most appropriate when discussing isotopic elements. No particular work has been done, so far, specifically addressing isotopic elements in terms of the number of up and down quarks. Isotopes are usually studied under chemistry, and the relevance of isotopes with quarks is not direct. However, the results and calculations suggest that certain relations can be established for isotopic elements, in terms of the number of up/down quarks, the atomic number of that element, etc. Also, keeping in mind that values such as abundance percentages have been considered and there have been round-ups during the calculations, a few inaccuracies can be ignored. It is highly probable, and as of now, must be taken as a working hypothesis, that initially there were (and are) more down quarks or more up antiquarks, in the universe. It may be hypothesized that protons that were present (or formed) near the boundary of the universe, during the early time of the evolution of the universe, were affected and may be attracted by some external electric and/or magnetic field from other universes (multiverse [7]). [The multiverse refers to the idea that parallel/alternate universes exist alongside this universe, where different possibilities, which are not fulfilled in this universe, might take place.] Neutrons, being electrically-neutral, on the other hand, can be comfortably assumed to not be affected by such fields. As a consequence, neutrons are more abundant. However, there are many contradictions at this point. For instance, as per the latest knowledge, most of the universe is hydrogen, and in the case of the regular hydrogen, i.e., protium, the rule of more-down-quarks-than-up, does not hold. Therefore, making statements regarding the overall composition of the universe in terms of up and down quarks and/or neutrons and protons may be futile.

Though there are no direct implications of the results obtained, the equations determined can be used to get the results of quantities such as the number of up/down quarks, etc., when some other information, such as the atomic mass of the element, etc., are known. These results can be used to further investigate the greater philosophical implications of the model, and whether the quark model can produce useful and practical results, for isotopic elements. Further investigations might reveal greater connections between quarks and the isotopic nature of elements, which could be manipulated to, even, produce useful isotopes and even predictions on isotopes of the elements of the Periodic Table. However, the fact remains that most of the naturally-occurring and artificial isotopes of elements have already been determined. This is one of the major limitations of this paper.

6. Conclusion

It has been verified that the mass of the lighter isotope subtracted from the average atomic mass of an element (by considering abundance percentages and by considering only two isotopes at a time), added to the number by which the lighter isotope's atomic mass is more than twice the atomic number of the element, equals the total number of up quarks subtracted from the total number of down quarks, of both the isotopes (by taking in account the abundance percentages).

Concluding, it should be noted that it is in itself obvious why the number of neutrons should be greater than the number of protons for most elements (to compensate for the forces of repulsion due to so many like-charged protons confined in a small nucleus, and keep the nucleus stable). Also, although this is an attempt to rephrase the physics of isotopes in terms of quarks, the proton-neutron language may be even more appropriate for the physics of isotopes. The main point of this paper was to determine certain relations regarding isotopic elements in terms of up and down quarks. Though the relations seem to hold, they may have minimal practical applications, due to the fact that the properties of the isotopes of most of the 118 elements are already known.

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