IMMAGINI E STRUMENTI DIGITALI NELLA DIDATTICA DELLE SCIENZE

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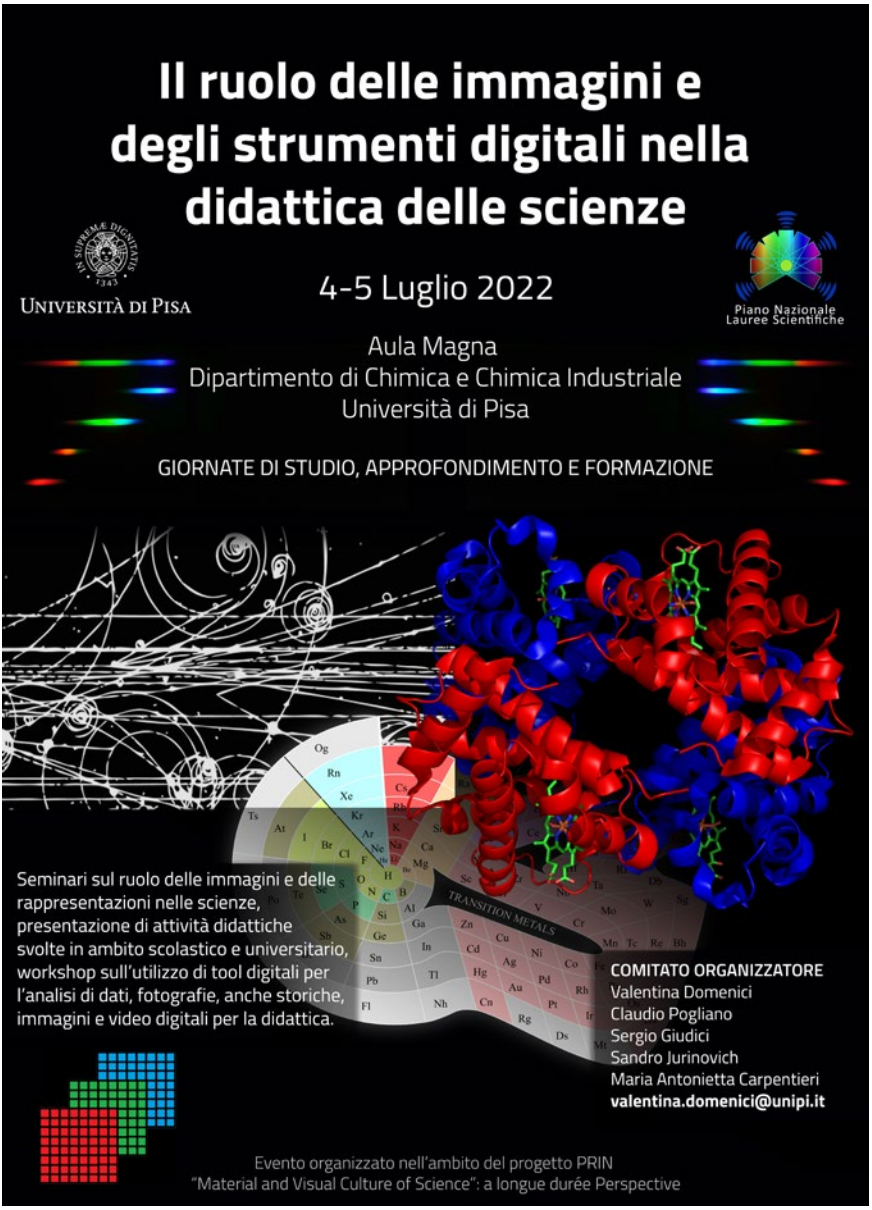

Fig. 1 - Locandina delle due giornate di studio [1].

How and why the Periodic Table should be elevated to a higher status in chemistry courses, and a very brief history of the periodic table

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1. Introduction

The most important image in chemistry and perhaps even in the whole of science is the periodic table of the elements. In this article I will be discussing the role that the periodic table plays in chemical education and a possible change to the way that the periodic table is exploited in chemistry courses. I will also be giving a very brief history of the development of the periodic table. The concluding section will consist of what may be termed remaining open questions concerning the periodic table.

2. General and educational aspects

The periodic table has become something of a cultural icon in addition to its very well-established role as a scientific icon. It appears on all manner of household goods from coffee mugs to shower curtains to T-shirts etc. The periodic table has been a source of inspiration to visual artists such as Blair Bradshaw [1] and poets such as Paul Davern [2]. These days one can find periodic tables of virtually any kind of object or theme that one may think of. Just a few examples would include periodic tables of soccer players, philosophers, poets, jazz musicians, psychological conditions, fonts, swear-words and so on [3].

If we were to limit ourselves to just genuine periodic tables of the chemical elements, it is a remarkable fact that well over 1000 versions have been published in print form or on the Internet. They come in all shapes and sizes, including giant periodic tables on the sides of buildings, such as a famous ver-

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sion in St. Petersburg outside the institute of weights and measures where Mendeleev worked for many years [4].

As chemical educators we need to ask ourselves about how to react to such a cultural phenomenon. I believe the simple answer to this question is that chemical educators should embrace and exploit all this publicity for the periodic table, since chemistry is sadly lacking in the area of 'public relations' and popularity, as compared with modern physics and biology.

In this article I would like to take the embrace of the periodic table a good deal further, and I would suggest that the periodic table could be exploited to a greater extent by making it the focal point and main thread running through chemistry courses.

A good deal has been written by the proponents of an 'atoms first' approach to chemistry to say that the traditional introduction to chemistry courses tends to alienate students from the subject [5]. Why do we begin chemistry courses by agonizing over the difference between chemical and physical changes at a stage when students are probably not familiar with either type? Why do we put such a premium on significant figures and mole calculations which makes chemistry courses look like an exercise in mathematics?

What I wish to suggest here is that, contrary to the notion of atoms first, we should plunge directly into chemical reactions and properties by exposing students to the properties a number of elements and quickly turn to the periodic table as the central organizing principle for the discipline. For example, students can be shown the properties of some elements in group 1 of the table, such as lithium, sodium and potassium so that they begin to see their similarities. One can then turn to the halogens and show how some of their representative members display distinctive similarities. Having laid this entirely chemical foundation one can then move on to studying the periodic table and how it also groups together many other kinds of elements.

My point is that whereas the atoms first approach is correct in avoiding mole calculations and the like at the outset, it represents a blatantly reductionist approach of putting atomic structure and therefore physics *before* chemistry. This is an example of 'putting the cart before the horse' to use that well-worn phrase. As I see, it we should put chemistry and the periodic table before atomic structure. Moreover, the periodic table, and attempts to explain it, can serve as a motivation for introducing atomic structure and quantum theory into chemistry courses. In this way the 'horse' of the periodic table can lead the 'cart' of atomic structure rather than vice versa.

Of course, the periodic table already commands a large role in the chemistry curriculum in providing a list of atomic weights and unifying explanations of all manner of chemical reactivity. My point is that it should play a more central role than it does in present courses and should serve as a main thread to which instructors can return to in order to provide a more unified and intellectually satisfying account of chemistry.

Before leaving this theme, let me say that I believe there is another profound reason why the periodic table can and should be made more central. This is an idea that comes from the growth of the philosophy of chemistry in recent years and the realization that far from lacking any profound ideas, chemistry possesses two very deep philosophical ideas, namely periodicity and chemical bonding which are of course intimately related [6]. Stated otherwise, elevating the periodic table to greater heights serves the purpose of making the discipline of chemistry, which is generally regarded as being rather pedestrian and merely of practical value, into one with philosophical value to match the deep ideas in modern physics and biology.

3. A brief historical tour of the development of the periodic table

I will begin this lightning quick tour of the history of the periodic table at an arbitrary place and time, namely the work of the Manchester school-teacher John Dalton, who revived the ancient Greek notion of atomism. As is well known, Dalton went a great deal further than the philosophers by assigning relative weights to the atoms of each of the then known elements. Moreover, his atomic theory provided a good explanation for a number of chemical laws that had been observed, namely the laws of the conservation of matter (Lavoisier), of constant composition (Proust) and of multiple proportion (Dalton).

Once atomic weights became available other chemists began to see relationships among their values. One such relationship was discovered as early as 1827 by the German chemist Döbereiner who found that there were groups of three elements such as lithium, sodium and potassium which formed a triad. In such a trio or triad of elements one of them had an atomic weight that was approximately the average of the weights of the other two members. In addition, that same element showed intermediate chemical behavior. This was perhaps the first time that a mathematical relationship had been found between the properties of *different* elements. Furthermore, and in retrospect, we now realize that this was essentially the discovery of chemical periodicity although it was not fully realized at the time.

As frequently happens in science, the idea of triads was taken a little too literally by some researchers who attempted to force all the elements into such groups even if the chemical facts did not warrant doing so [7]. One of the chemists who perhaps took the idea of triads too far was Peter Kremers [7]. This, and the fact that the triad relationship is only approximate if one uses atomic weights, led to triads falling into disrepute. For example, Mendeleev had strong views of the subject of triads and claimed that they had misled many chemists into becoming too engrossed in numerology.

By an interesting twist of fate, it turns out that all the previously recognized atomic weight triads, in addition to many new ones, are seen to be exact if one uses atomic numbers instead of atomic weights. This outcome is far from surprising given that atomic numbers deal with whole numbers from the outset, whereas atomic weights are subject to the vagaries of the isotopic abundances of any given element. Triads of elements may be said to have made a 'come-back' to use a phrase from the philosopher of science Imre Lakatos, who first discussed this aspect of scientific theories and concepts [8].

The second philosophical idea that contributed to the development of the periodic table was initiated by William Prout, a Scottish medical doctor who was practicing in London. Prout considered the atomic weights of the elements that were available at the time and noticed the rather obvious fact that most elements had weights that were an integral multiple of that of the lightest element of all, namely hydrogen. He then went a step further and proposed that the atoms of all the elements were literally composites of atoms of hydrogen and that all matter was essentially made of the same 'stuff'. The Prout hypothesis, as it became known, was put to the test by some of the leading chemists of the time including Staas and Berzelius and was essentially found to be incorrect. In a similar manner to triads, Prout's hypothesis fell into disrepute, and yet it too seems to have made a come-back in a modified sense. If we consider just the protons in the nucleus of any atom then of course the atoms of all other atoms truly are composites of the one proton that occurs in the hydrogen nucleus. From the completely different domain of astrophysics, we also learn that the elements were literally formed initially by the fusion of two hydrogen atoms to form helium and then, at least in principle, the fusion of two helium nuclei to form one of beryllium, and so on. Stated otherwise the elements really are all made of the same substance [9].

Continuing on a historical trajectory we come to the important year of 1860 during which an international conference was held in Karlsruhe, Germany, at which Cannizzaro presented a set of more accurate and more consistent atomic weights of all the elements. This development opened the path for the discovery of chemical periodicity in the modern sense of the term. Far from being the discovery of one or two individuals, as is frequently portrayed in textbook and popular accounts [10], chemical periodicity was independently discovered by at least six individuals, over a period of seven years from 1862 to 1869, the latter date being usually taken as *the* discovery by Mendeleev [11].

The main reason why Mendeleev is usually given major credit for this discovery lies with the fact that he made predictions of the existence of then unknown elements, which were found over a period of 15 years. I am referring to Mendeleev's prediction of three elements which were subsequently named gallium, germanium and scandium by their discoverers [12].

However, there is a long-standing debate in the history and philosophy of science concerning the relative merit of genuine prediction as opposed to retrodiction (accommodation) of data [13].

According to some authors theories and scientific concepts deserve more credit if they can successfully accommodate already known data. If this is indeed the case it would imply that Mendeleev's successful predictions should not be lauded to quite the extent that they generally have been [14]. Furthermore, of the 16 new elements that Mendeleev predicted, only eight of them were actually found, which is surely not so impressive a record of prediction.

Returning to my selective time-line takes us to the year 1894 when the element argon was discovered in London by Ramsey and Rayleigh. This element posed a major challenge for the periodic table because it had an atomic weight of 40 which meant that it could not be accommodated into the table which already had an element, calcium, with the same atomic weight. Mendeleev's reaction was initially one of skepticism and a speculation that the new element might be triatomic nitrogen [15]. The following few years saw an intensification of this challenge to the periodic table given the additional discoveries of the elements helium, neon, krypton and xenon, also by Ramsey and Rayleigh [16]. A little-known fact is that the Danish chemist Julius Thomsen had predicted all of the noble gases, up to and including an element that is now known as oganesson [16]. The problem of how to accommodate these elements into the periodic table was solved by Ramsey who simply introduced a new column at the right-hand edge of the table, a move that was hailed as a triumph for the periodic table by Mendeleev.

The turn of the 20th century saw three major discoveries that were to have a profound influence on chemistry, physics and the eventual explanation for the periodic table. In 1895 Roentgen discovered X-rays, in 1896 Becquerel discovered radioactivity and in 1897 Thomson discovered the electron [17]. These discoveries were rapidly followed by further work on radioactivity by the Curies and on the structure of the atom by Rutherford among many others.

Between the years 1913 and 1914 Moseley working in Manchester, and later Oxford, discovered that atomic number provided a better ordering principle for the elements in the periodic table than atomic weight did. The mysterious question of pair reversals such as the cases of tellurium and iodine, as well as cobalt and nickel, thus received a natural and independent explanation, which did not depend just on using their chemical properties in order to determine where they should be placed.

The year 1913 also saw the publication of Bohr's model of the atom and his explanation that the elements that belonged to any particular group shared the same number of outer electrons. This development was a direct outcome of the growth of the old quantum theory of Planck and Einstein and its application by Bohr to the nature of the hydrogen atom.

Between the years 1924 and 1926 quantum theory was placed on a more rigorous and axiomatic foundation and became known as quantum mechanics [18]. In this improved theory one solves the Schrödinger equation for the hydrogen atom to obtain a number of solutions, each of which is characterized by three quantum numbers. To this set of three values, one adds a fourth quantum number as first introduced by Pauli. Then by means of four quantum numbers to each electron one can rigorously predict that successive electron shells should contain 2, 8, 18 and 32 electrons. The fact that these values coincide with the possible length of periods in the table is no coincidence. It is rather a major triumph when it comes to the explanation or reduction of the periodic table to quantum mechanics. Nevertheless, it is not a complete explanation for the form of the periodic table because the electron shells do not fill sequentially [19].

The order of filling of electron orbitals occurs in a diagonal rather than a horizontal fashion as can be seen by consulting the Madelung diagram below (Figure 1). The accompanying Madelung rule dictates that the order of occupation of orbitals follows the order of increasing values of the n + ℓ quantum numbers for any given orbital.

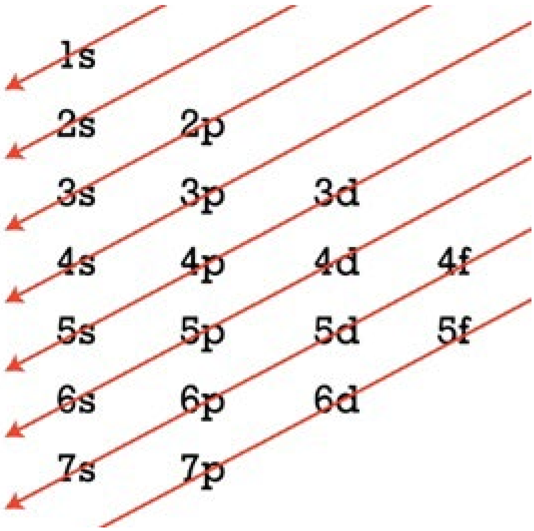

Fig. 1 - The Madelung diagram showing the order of occupation of atomic orbitals starting with 1s, 2s, 2p, 3s etc.

4. Remaining open questions about the periodic table

A more severe test of whether quantum mechanics truly explains the periodic table is to demand that this simple rule should also be derived from first principles. Unfortunately for those who believe that the periodic table has been fully explained by quantum mechanics, this feat has yet to be achieved, many decades after the rule was first introduced [20].

Another important and contemporary issue concerning the periodic table has been the consideration of relativistic effects. It has become well known that as atomic number increases, and as atoms become heavier, the electrons begin to travel at relativistic speeds, which requires the application of the special theory of relativity in addition to quantum mechanics. Several kinds of relativistic effects result in some profound modification to the macroscopic properties of the elements concerned. For example, the distinctive color of the metal gold is now explained as being a result of relativistic effects [21]. The very next element in the periodic table after gold (79) is mercury (80). The fact that this metal is unique in its being a liquid at room temperature is also explained by the application of special relativity [22].

But the deeper story appears to be more complicated in that only some elements with high atomic numbers seem to suffer from modifications to their chemical properties. For example, the element rutherfordium (104) falls into group 4 of the periodic table, lying below zirconium and hafnium. However, chemical experiments conducted on this element suggest that it differs considerably from the elements above it and behaves more like the element plutonium, which is some distance away in the periodic table. On the other hand, the even heavier element copernicium which les in group 12 of the table does behave similarly to the elements above it, namely cadmium and mercury [23]. For example, this is true of the sublimation enthalpies of these elements which fall on an almost perfectly straight-line graph [23]. It is therefore not clear whether relativistic effect do provide a major challenge to chemical periodicity or not, although the majority view seems to be that they do [24].

Turning to another current and controversial question, there is some debate as to whether there exists an optimal or most fundamental form of the periodic table. Perhaps the main candidate for such a role is the left-step periodic table that was first introduced by a French engineer, Charles Janet, in the 1930s [25]. This distinctively shaped table is obtained by placing helium into group 2 of the periodic table and by locating the entire s-block of the table at the right-hand edge as shown in Figure 2.

On one hand this table makes perfect sense in terms of electronic configurations, since helium has two electrons and therefore seems to be analogous to the group 2 atoms, all of which have two outer-electrons. On the other hand, in purely chemical terms such a relocation of helium appears unjustified since this element would seem to be a noble gas, and perhaps

																														H	He	2
																														Li	Be	2
																								B	C	N	O	F	Ne	Na	Mg	8
																								Al	Si	P	S	Cl	Ar	K	Ca	8
														Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr	Rb	Sr	18
														Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe	Cs	Ba	18
La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn	Fr	Ra	32
Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Nh	Fl	Mc	Lv	Ts	Og	119	120	32

Fig. 2 - Left-step periodic table, as first proposed by C. Janet.

the most noble of the noble gases [26]. What we see here is a good example of the tension that exists between concentrating on physical or chemical properties when it comes to the placement of elements in the periodic table.

Among the other virtues of the left-step table is the fact that all period lengths, without exception, are repeated, whereas the first period length of just two elements is not repeated in the conventional table. Another advantage would seem to be that the left-step table more faithfully reflects the Madelung rule for the occupation of atomic orbitals that was discussed above, in that each period corresponds to a particular value of the $n + \ell$ quantum numbers.

The final contemporary issue that will be briefly discussed concerns group 3 of the periodic table. It is a remarkable fact that some textbooks display a periodic table in which the elements of group 3 consist of scandium, yttrium, lanthanum and actinium, whereas other books consider group 3 to consist of scandium, yttrium, lutetium and lawrencium. Various authors have attempted to resolve this issue in recent years and there appears to be little by way of consensus. One such attempt has been by myself, in which I have argued that group 3 should consist of Sc, Y, Lu and Lr on the basis of two simple requirements. First, I propose that the elements should be displayed on a 32-column periodic table rather than on the more usually seen 18-column format. The 32-column or long form table has the advantage of incorporating all elements into the main body of the table rather than banishing the f-block elements to a disconnected footnote. If one follows this suggestion and if one makes the further requirement that all the elements should follow each other in order of increasing atomic number, then it follows that group 3 should consist of Sc, Y, Lu and Lr [27].

I find it rather unfortunate that what is sometimes said to be the official IUPAC periodic table fails to take sides on this issue, and simply leaves empty the two spaces below the elements scandium and yttrium (Figure 3).

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
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
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
55 Cs	56 Ba		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
87 Fr	88 Ra		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

57 La	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
89 Ac	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

Fig. 3 - This format is sometimes designated as the official IUPAC periodic table.

Finally, I hope that this quick survey of the periodic table, and some associated controversial issues, might serve to stimulate chemical educators into elevating the periodic table to a higher level than it already commands within the teaching of chemistry.

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[15] Since nitrogen has an atomic weight of 14, tri-atomic nitrogen would have a weight of 42 which is not so different from the measured 40 for argon. In addition, argon was extracted from a sample of nitrogen gas, this making the speculation a little more plausible still.

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[18] This terminology is a little unfortunate since it would seem to reinforce the incorrect view that a theory is a vague concept that eventually matures into something other than a theory. Quantum mechanics is of course as much of a

theory as the old quantum theory when one takes account of the more scientific sense of the word.

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