

Approaching a study of the periodic table from a nature of science perspective

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Abstract

An historical approach to a study of the Periodic Table which emphasises how knowledge is constructed has great value in helping students address issues related to the nature of science. A key feature of the nature of science particularly applicable to the Periodic Law is outlined and discussed with reference to the work of Dobereiner, Newlands, Mendeleeff, and recent observations by Scerri. A teaching/learning strategy used by the author with a first year tertiary chemistry class is outlined.

Introduction

Ever since the mid-nineteenth century there have been arguments put to science curriculum developers that science education should contain not only experiences in the facts of science and their confirmation but also something of the broader issues related to how science functions as a social tool in modern society; the intellectual tools it uses and the assumptions used in constructing and justifying its knowledge base; and how its intellectual heritage has interacted with other disciplines over time to change the intellectual landscape. In 1947 J.B. Conant, president of Harvard, accomplished this by using history of science episodes in his college science courses for non-science majors. However, those of us who are involved in teaching chemistry to science majors have been so often overcome by the level of content coverage required in our chemistry courses that little space or time has been available to devote to the nature of chemistry through an exposure to the history and philosophy of chemical ideas. Arnold Arons (1983) observes that efforts to cultivate scientific literacy often flounder because of an incomprehensible stream of technical jargon not rooted in experience accessible to the student and the fact that the material is presented much too rapidly and in too great a volume. This makes it difficult to gain a sense of how concepts and theories originate, how they come to be validated and accepted, and how they connect with experience and reveal relations among seemingly disparate phenomena. Yvonne Meichtry (1999) notes that the nature of science and scientific knowledge is the area most neglected by school curricula and least understood by K-12 and university students alike. Part of the reason for this has been the emphasis given in curricula to the facts of science and how these facts can be confirmed or discovered in laboratory experiences and the fact that teachers themselves have not experienced the teaching of science from a perspective which gives due consideration to matters of scientific literacy. Nature of science issues are also regarded by teachers as too abstract and too difficult to learn (Lederman, 1999). In spite of this, curriculum bodies such as the British National Curriculum (1988), The American Association for the Advancement of Science (1993), and the New South Wales HSC Chemistry Syllabus (2000) are insisting that the nature of science be given due weight in chemistry courses.

But what is involved in a study of the nature of science and how might a study of the Periodic Table fit into the picture? Firstly, there is still significant disagreement in philosophical and educational circles about what constitutes the content and methods of science. In 1969 Herron (1969) observed the lack of a sound and precise description of the nature and structure of science and the same observation was made by Duschl (1994) twenty-five years later. Despite this difficulty McComas, Almazroa, and Clough (1998) were able to isolate fourteen consensus statements regarding the nature of science from eight international science standards documents. These are reproduced in Table 1. The authors also suggest that the fourteen statements are noncontroversial enough to be used as a basis for incorporating issues relating to the nature of science into science education curricula. Secondly, recent education reforms such as those already mentioned (NSW HSC Chemistry 2000) encourage teachers to integrate nature of science into science instruction rather than treating it as a separate topic apart from the science content. This point is emphasized by Lederman, McComas, and Matthews (1998) in their editorial for the Nature of Science issue of *Science and Education* when they say, "Educators must encourage textbook writers to move away from the treatment of the nature of science as a discrete topic and encourage them to infuse the history and philosophy of science throughout texts in meaningful and interesting ways". This paper embodies this idea by showing how one of the fourteen consensus statements of the nature of science, *scientific knowledge while durable has a tentative character*, is exemplified in a study of the Periodic Table. The focus of the paper is the lower tertiary level chemistry classroom although some adaptations could be made to make the material applicable to the senior high school.

The chemistry context leading to the organisation of the elements-a brief summary

The concept of element as a building block of matter had been postulated in early Greek philosophy and by the fifth century BC matter was considered to consist of combinations of only four elements; earth, air, fire, and water. This idea persisted well into the time of alchemical science in the middle ages and was only really challenged from about the sixteenth and seventeenth centuries. The idea that an element was a substance that could not be broken down further by chemical means became

Table 1 A consensus view of the nature of science objectives extracted from eight international science Standards Documents

- Scientific knowledge while durable, has a tentative character
- Scientific knowledge relies heavily, but not entirely, on observation, experimental evidence, rational arguments, and skepticism
- There is no one way to do science (therefore, there is no universal step-by-step scientific method),
- Science is an attempt to explain natural phenomena
- Laws and theories serve different roles in science, therefore students should note that theories do not become laws even with additional evidence
- People from all cultures contribute to science
- New knowledge must be reported clearly and openly
- Scientists require accurate record keeping, peer review and replicability
- Observations are theory-laden
- Scientists are creative
- The history of science reveals both an evolutionary and revolutionary character
- Science is part of social and cultural traditions
- Science and technology impact each other
- Scientific ideas are affected by their social and historical milieu

prominent from the time of Lavoisier in the eighteenth century and it was recognized that on this basis there were many more elements than just four as previously thought. By the beginning of the nineteenth century, then, one of the major problems that concerned chemists was the determination of the elemental composition of pure compounds. Richter (1762-1807) used quantitative methods to determine the reacting ratios of substances and introduced the term, stoichiometry, for such a process. In 1799 the French chemist Proust (1754-1826) observed that copper carbonate had the same elemental composition regardless of how it was synthesized. This was a practical proof of the law of constant proportion. Berthollet (1748-1822), however, believed that such a law was an accidental outcome of the experiment and that the elemental composition of a compound could continuously vary depending on the conditions of synthesis. However, for simple compounds the laws of constant and multiple proportions received strong experimental support due to the excellent research of the Swedish chemist Berzelius (1779-1848). In addition the laws were consistent with Dalton's atomic theory first proposed in 1803. According to Dalton the dense spheres of atoms of a given element were all the same size but of a different size to those of a different element and numerical values for relative atomic weights based on a mass unit of one for the hydrogen atom could be assigned to the atoms of different elements if one assumed, for example, that a compound of two elements consisted of one atom of each element. Dalton and Berzelius, however, never distinguished between the concepts of atom and molecule. This was left to Avogadro to propose in 1811 but the notion of a molecule consisting of a combination of like atoms was never accepted until the 1860 Karlsruhe conference. Thus Gay-Lussac's law of combining volumes of 1809 combined with Avogadro's

hypothesis for the existence of molecules enabled atomic weights such as that for oxygen to be corrected.

By the middle of the nineteenth century more elements had been discovered through standard chemical methods and the new spectroscopic method of Bunsen (1811-1899). These elements were considered to consist of atoms or combinations of atoms whose relative weight could be compared to that of hydrogen. Classifying elements in terms of the relative weights of their atoms became a key research area of nineteenth century chemistry. The careful research of Stas (1813-1891) and Berzelius (1779-1848) in determining the atomic weights of the elements became crucial for systematizing the chemistry of the elements on the basis of atomic weight, a program of research which was to occupy chemists in the latter half of the nineteenth century. It cannot be underestimated how important the tedious research associated with the chemical determination of atomic weights was to the development of the Periodic Law.

Early ideas on the organization of the elements

The process of knowledge development in any field, and the Periodic Law is no exception, is one that builds upon the challenges of the past and the present because the challenges act as important catalysts for further knowledge generation. In 1817 thirty-five elements were known and chemists were active in assigning atomic weights to these elements based on Dalton's Atomic Theory. Chemists were also interested in whether there was any connection between an element's atomic weight and its chemical properties. Dobereiner observed that elements with similar chemical properties seemed to occur in TRIADS. Two examples of this grouping were lithium, sodium, and potassium as a metal triad and chlorine, bromine, and iodine as a nonmetal triad. Dobereiner observed that the atomic weight of the middle member of the triad was the average of the other two atomic weights and the properties of the middle element lay between the properties of the other two. For example, sodium's atomic weight (23) is the average of lithium's (7) and potassium's (39); bromine's atomic weight (80) is the average of chlorine's (35.5) and iodine's (127). The melting point of sodium (98 celsius) lies between the melting points of lithium (180 celsius) and potassium (63 celsius) and the melting point of bromine (-7 celsius) lies between the melting points of chlorine (-101 celsius) and iodine (114 celsius). These facts are still true in the modern Periodic Table but the concept of TRIAD has been superseded. By 1864 sixty-three elements had been discovered and the TRIAD law became increasingly inadequate for explaining the properties of a larger number of elements. In 1865 John Newlands observed that elements with similar chemical properties seemed to occur in OCTAVES if the elements were arranged in the order of their atomic weights. The numbers for analogous elements generally differ by 7 or some multiple of 7 as shown in Figure 1.

Organizing the elements according to triplets or octaves, known musical terms, was reminiscent of a past belief in the harmony of the universe expressed as the *music of the spheres* and of the numerical beliefs of the Pythagorean

brotherhood. In the minds of the nineteenth century scientific establishment these ideas could not be given serious consideration because of their association with mysticism. A more scientific account of the organization of the elements was to await the work of Mendeleeff who outlined his Periodic Law to the Russian Chemical Society in 1869. Cassebaum and Kauffman (1971) have drawn attention to the fact that other chemists such as Lothar Meyer (1830-1895) and William Odling (1829-1921) could lay equal claim to having discovered the law but simultaneous and/or independent discovery is not our emphasis here. We will confine ourselves to the more widely recognized work of Mendeleeff and the durable and tentative character of the law as he described it.

H 1	F 8	Cl 15	Co/Ni 22
Li 2	Na 9	K 16	Cu 23
G 3	Mg 10	Ca 17	Zn 25
Bo 4	Al 11	Cr 19	Y 24
C 5	Si 12	Ti 18	In 26
N 6	P 13	Mn 20	As 27
O 7	S 14	Fe 21	Se 28

Figure 1: Newland's (1865) arrangement of the elements to illustrate the Octave Principle.

Mendeleeff and Periodicity

In 1869 Mendeleeff organized the elements according to their atomic weight and known chemical properties as shown in Figure 2.

			Ti=50	Zr=90	?=180
			V=51	Nb=94	Ta=182
			Cr=52	Mo=96	W=186
			Mn=55	Rh=104.4	Pt=197.4
			Fe=56	Ru=104.4	Ir=198
			Ni=Co=59	Pd=106.6	Os=199
H=1			Cu=63.4	Ag=108	Hg=200
	Be=9.4	Mg=24	Zn=65.2	Cd=112	
	B=11	Al=27.4	?=68	U=116	Au=197?
	C=12	Si=28	?=70	Sn=118	
	N=14	P=31	As=75	Sb=122	Bi=210?
	O=16	S=32	Se=79.4	Te=128	
	F=19	Cl=35.5	Br=80	I=127	
Li=7	Na=23	K=39	Rb=85.4	Cs=133	Tl=204
		Ca=40	Sr=87.6	Ba=137	Pb=207
		?=45	Ce=92		
		?Er=56	La=94		
		?Yt=60	Di=95		
		?In=75.6	Th=118		

Figure 2: Mendeleeff's (1889) arrangement of the elements in 1869.

As the magnitude of the atomic weight increases there is a periodic change in chemical properties, apart from a small number of anomalies. Even though the modern periodic table is based on atomic number the periodicity with atomic

weight still holds true for a large part of the periodic table. In fact, the order based on atomic weights is the same as the order based on atomic numbers apart from seven pairs: Ar/K; Co/Ni; Te/I; Th/Pa; U/Np; Pu/Am; Lr/Rf. This is a case of an important chemical law being originally based on a different premise (atomic weight) to that currently accepted (atomic number) but the trend according to atomic weight was sufficiently close to the trend according to atomic number to enable the basic features of the law and the elements to be established. Consider, for example, the identification of the two elements between Zn and As in Figure 2.

How did Mendeleeff know that there should be two elements between zinc and arsenic and how was he able to predict the properties of these new elements? Mendeleeff observed that a gap in atomic weights of the order of 1 to 3 units was typical for consecutive elements and that therefore a gap of 10 units suggested at least two elements missing in the sequence. Consider the possible properties of the analogue of silicon, ekasilicon (one element after silicon), as Mendeleeff called it. Mendeleeff (1897) suggested that the properties of ekasilicon, later to be called germanium, could be estimated from the properties of silicon, tin, zinc, and arsenic; that is, from elements above, below, and to the sides of it in Figure 2. The atomic weight of ekasilicon was estimated from summing the atomic weights of the four elements mentioned and dividing by four to give a kind of average. An atomic weight of around 70 was thus estimated. A similar procedure was performed with the specific gravities giving about 5.5 for the result. The properties of carbon, silicon, and tin suggested that ekasilicon (Es) would form the higher oxide, EsO_2 , the lower oxide, EsO , and halides of general formula, EsX_4 . The boiling points of SiCl_4 (57 celsius) and SnCl_4 (114 celsius) suggested that EsCl_4 should be a volatile liquid with a boiling point less than 100 celsius. When germanium was later discovered and its tetrachloride produced it was found to be a volatile liquid with a boiling point of about 90 celsius so the predictions proved to be correct.

The periodicity of properties based on atomic weight rather than atomic number was a good enough law to form the basis whereby some atomic weights could be corrected. Consider the case of indium which has a question mark next to it in Figure 2. The equivalent weight based on hydrogen was known to be 37.7. If indium was monovalent then one would expect its atomic weight to be 37.7, but there was no space for an element of this atomic weight in the periodic grid. If indium was divalent its atomic weight would be 75.4, close to the value shown in Figure 2. This also presents a problem because there is no space between arsenic and selenium in the grid. Trivalency would yield an atomic weight of 113.1, putting it between cadmium and tin in the boron/aluminium group. Such a space exists on the periodic grid and this position is now confirmed in that indium's density lies between that of cadmium and tin, and the nature of its oxide is consistent with the formula, In_2O_3 , with its properties intermediate between those of CdO and SnO_2 .

Comparing Mendeleeff's table of 1869 (Figure 2) with a modern periodic table reveals similarities and differences. The obvious differences are: the *s* and *p* block elements in their groups in Figure 2 are arranged horizontally rather than vertically; the *3d*, *4d*, and *5d* transition elements are arranged vertically in Figure 2 rather than horizontally; no noble gases appear in Figure 2 because they had not yet been discovered; the lanthanide and actinide elements are rather incomplete in Figure 2; and there are obvious errors in some atomic weights and the group placement of the element such as in the case of uranium. Despite these differences, however, the principle of the periodicity of chemical properties has not changed even though the basis for periodicity (atomic weight to atomic number) has changed. The alkali metals, the halogens, the oxygen group, and the carbon group, for example, are clearly identified in Mendeleeff's table. Mendeleeff's later tables show some refinement over the 1869 table, of course, but the principle of the periodicity of chemical properties remains.

Mendeleeff probably had good reason to be confident in the role of atomic weights as the basis of periodicity given the fact that in the modern table only seven pairs of elements arranged according to atomic number do not sequence similarly when arranged according to atomic weight. In fact, Mendeleeff was so confident in the role of atomic weights that, on occasion, he was led to accept the results of experiments that were later shown to be somewhat inaccurate. The paradox is that on some occasions his firm belief in atomic weights paid off, as in the case of indium previously discussed, but on other occasions it did not pay off. In the case of tellurium and iodine shown in Figure 2, for example, Mendeleeff knew that iodine's chemical properties put it into the halogen group even though its atomic weight (127) was less than the previous element, tellurium (128). Mendeleeff was certain that tellurium's atomic weight was too high and that on refinement would be shown to be somewhere between 123 and 126. When the chemist, Brauner, reported an atomic weight for tellurium close to 125, Mendeleeff was ready to support this result even though it was later shown to be incorrect.

Early and recent reflections on the completeness of the Periodic Law

Is modern chemistry's view of the basis of the periodicity of chemical properties now complete? When Professor W.A. Tilden delivered the Mendeleeff Memorial lecture to the Chemical Society in 1909, he challenged his audience with a most pertinent question. "Can it be truly said that the elements arranged in the order of their atomic weights show without exception periodic changes of properties?" (Tilden, 1909, p. 2094). Tilden implies that there were anomalies in the system that could not be explained because towards the end of his address he summarizes his conclusion by saying, "...the periodic law... is destined to be absorbed into a more comprehensive scheme by which obscurities and anomalies will be cleared away, the true relations of all the elements to one another revealed, and doubts as to the

doctrine of evolution resolved in one sense or the other. But as with Atomic Theory itself, there is no reason to doubt that the essential features of the periodic scheme will be clearly distinguished through all time, and in association with it the name of Mendeleeff will be forever preserved among the Fathers or Founders of Chemistry" (Tilden, 1909, p. 2105). Even since the modern version of the Periodic Table has been established on the basis of atomic number and electron structure, Tilden's statement and question are still pertinent as echoed in the writings of Eric Scerri (1997, 1998). Tilden's question could be repeated this way. Can it be truly said that the elements arranged in the order of their atomic numbers and in groups according to the pattern of their valence electron structure show without exception periodic changes of properties? The fact is, of course, as pointed out by Scerri (1997), that there are exceptions to periodicity based on electron structure. Take, for example, the ten valence electrons in the nickel, palladium, and platinum group. The valence structures are $3d^8 4s^2$, $4d^{10} 5s^0$, and $5d^9 6s^1$ respectively. Scerri (1997, p. 533) comments in this respect, "Each shows a different outer-shell configuration, yet they are grouped together because of their marked chemical similarities. If it were the case that possession of a particular configuration is a sufficient condition for membership to a particular group, possession of a certain configuration would ensure that the atoms of those elements would fall into a particular group. Yet the elements, helium, beryllium, and magnesium - all of which share the property of having two outer-shell electrons - do not fall into the same group". Scerri (1997) suggests that a better basis for periodicity probably lies in the approximate quantum mechanical calculations of nonvisualizable properties of the atom such as total energy. Reflecting on the historical development of the Periodic Law, Scerri (1997, p. 553) notes that, "Over 125 years ago Mendeleev, probably the leading discoverer of the periodic system, refused to adopt a realistic view of the system and emphasized only its classifying aspects. Exactly 100 years ago, the electron, the first subatomic component, was discovered and pointed the way for a swing back toward a realistic account of atomic physics. In due course, this led to the equally realistic electron-shell approach to the periodic system. About 30 years later, the reality of electron shells and orbitals had evaporated into the formalism of quantum mechanics, leaving behind just the mathematical utility of superimposed expressions of electronic configurations".

What can be said about this analysis? In terms of chemical education the use of pictureable models will, I think, retain an important place in chemistry but we must help our students understand that these models are limited in what they can explain as evidenced by the anomalies mentioned. Our understanding of the Periodic Table must revolve around two separate points. Firstly, the principle upon which the elements are placed in a group of the Periodic Table relates to their chemical properties. This has been a durable feature of the Periodic Law since its inception. Secondly, the fundamental nature of matter upon which this classification takes place is still open to question, but

will probably lie in the nonpictureable entities of quantum mechanics. This aspect illustrates the tentative character of the law. The pictureable entities such as electron shells are still useful in most cases but are limited. This is an important aspect of the nature of science for students to grasp because the most common picture of science advanced in our classes is one of a system closed to any further discoveries rather than as a system still open to further enquiry. It turns out that the Periodic Law is ideal for conveying this aspect of the nature of science. But how might one incorporate these ideas in the teaching/learning strategies pertinent to a chemistry class? The next section illustrates one strategy I have used with a first year tertiary general chemistry class.

A teaching/learning strategy for incorporating the Periodic Law and the Nature of Science in the chemistry curriculum

Science teaching and learning has often been criticised for its dependence on algorithmic problem-solving and rigid laboratory tasks which do not allow for creative thinking and writing skills (Stenhouse, 1985). To rectify this situation I decided to set my first year chemistry class (19 year-olds mainly) an historical essay of 1500 words which would address how scientific knowledge was generated in the case of the Periodic Law and how the development of the Periodic Law illustrates the durable and tentative character of science. The students were asked to read the articles by Wynn and Wiggins (1997), Scerri (1997 and 1998), and Mendeleeff (undated) and to build their essay around a discussion of the following questions:

1. What are the differences between Mendeleeff's version/s of the table and the current version and give some reasons for the differences?
2. What anomalies exist in Mendeleeff's system?
3. How did Mendeleeff identify the element 'ekaboron' and the elements estimated by Mendeleeff to have atomic weights around 68 and 70?
4. What is the meaning of the term 'periodic' and illustrate using a particular property like melting point?
5. What do you consider to be the 'durable' character of the law? Give justifications.
6. What do you consider to be the 'tentative' character of the law? Give justifications.
7. How does the development of the Periodic Law illustrate how scientific knowledge is constructed?

Students were free to use other references in addition to those quoted and the website.

My experience is that students find writing essays on the nature of science and responding to open-ended questions rather difficult. For example, the task of identifying the durable and tentative aspects of the law proves much more challenging to the students than solving an equation for a single unknown. The fact that there is not just one correct answer in an exercise like this is an extremely valuable experience for a student who characteristically views science in dualistic terms, that is, answers to questions in science are either correct or incorrect. But the experience of writing is a creative experience and with determination and effort students find the task ultimately rewarding. The

task is more easily accomplished when nature of science questions relate to specific content. The success of the exercise also depends on the quality of the reading materials. Those listed here are quite suitable for first year tertiary chemistry classes. The Periodic Law is one science area where suitable reading materials are available. This is not the case with many other scientific ideas.

Conclusion

While controversy still surrounds how one might define science and its methods, there appears to be a good case for drawing upon a consensus model of the nature of science for educational purposes and letting the development of scientific knowledge in a content area like the Periodic Law illustrate important components of this consensus model. A writing task, guided by some deliberate comments and questions, gives the student an opportunity to develop their communication skills while learning how scientific knowledge develops. This kind of teaching/learning strategy broadens the educational experience of science students and encourages them to think outside what is often regarded as the rigid boundary of a science education experience.

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