

IX

Inorganic Myths in the Freshman Textbook

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1. Introduction

Roughly two years ago I was assigned to teach one of the lecture sections of the freshman chemistry course at Cincinnati. This services up to 1300 students and can have as many as four different professors teaching as many different lecture sections. Unlike the courses I had taught my entire career, I was for the first time put into a position where I had no control over the choice of textbook, the course content, or the course organization. Instead of the textbook serving to supplement otherwise original lectures by a professionally competent professor, I found that I was instead suppose to function basically as a TA by slavishly following the textbook so as to “insure uniformity” between the various lecture sections.

Although I was strongly admonished not to correct or criticize the textbook in front of the students, since our job was not to provide them with the best and most accurate chemical information, but rather to ensure that they were happy and satisfied educational customers, I nevertheless found it impossible, as a professionally trained inorganic chemist, to ignore the large number of errors and distortions present in the textbook’s somewhat perfunctory treatment of inorganic chemistry. Though I was required to bite my tongue, albeit not always successfully, and teach this questionable material to the students, I thought this morning I would at least share with you my dissatisfaction with what I have discovered.

I do not do this simply to vent but because the problems which I uncovered are not unique to the textbook used at Cincinnati but are in fact found in most freshman textbooks, as well as being widely disseminated on the internet. This is in part a result of the fact that most authors of freshman textbooks simply copy their material from other freshman textbooks and the fact that the marketing departments of most textbook publishers are careful to edit out any deviations from the accepted norm for fear that they will damage potential sales. The net result is a system which is virtually impervious to corrections or improvements in either content or organization. Thus, though I will phrase my remarks in terms of the textbook currently being used at Cincinnati, the errors under discussion are sufficiently widespread to truly qualify as common

myths. For obvious reasons, both the textbook and its author will remain anonymous.

2. The Four Primary Myths

In my opinion, the numerous errors and distortions related to inorganic chemistry which I encountered in our choice of freshman textbook are all traceable to four fundamental underlying myths that were either explicitly or implicitly accepted by the textbook author:

1. There is a discrete and discontinuous boundary separating ionic bonding from covalent bonding.
2. The only known type of localized covalent bond is the conventional 2c-2e bond used in writing Lewis dot diagrams and, as a consequence, many main-block inorganic species cannot be represented using a Lewis diagram without violating the octet rule.
3. All covalently bonded inorganic species are composed of discrete molecules.
4. The concepts and vocabulary of coordination chemistry apply only to the chemistry of the transition metals.

Rather than randomly discussing the various errors and distortions, I will instead attempt to group them under the particular myth most responsible for their persistence in the freshman textbook.

3. Ionic versus Covalent Bonds

The chapter in the textbook on chemical bonding begins in a promising enough fashion by introducing the well-known van Arkel bond-type triangle (figure 1) in which both the progressively colored shading and the book’s figure caption correctly imply that the ionic, covalent, and metallic bonding models are in fact limiting-case idealizations which progressively shade into one another through a series of intermediate bond-types (2). However, this promising start is immediately nullified by the announcement that any discussion of metallic bonding will be postponed until the chapter on solids and by the introduction, near the end of the

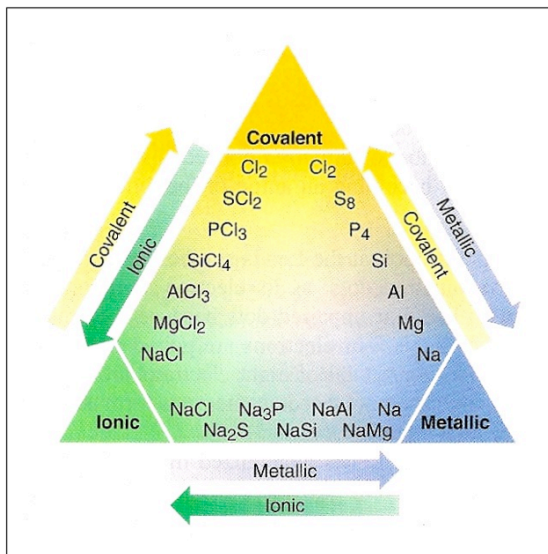


Figure 1. The van Arkel bond-type triangle given by in the textbook and intended to illustrate the progressive changes in bond type on passing between the three limiting extremes of ionic, covalent and metallic bonding.

chapter, of an outdated diagram (figure 2) taken from Pauling's *Nature of the Chemical Bond* (3). Though Pauling's original intention in this diagram was also to illustrate the progressive change from covalent to ionic bonding as a function of electronegativity difference, the textbook author arbitrarily imposes a sharp boundary on the diagram and declares that all bonds lying above this boundary are ionic and all bonds lying below it are covalent.

As it turns out, despite the implied promise, the chapter on the solid state makes no mention of inter-metallic compounds and alloys or of the progressive transitions between covalent and metallic bonding, on the one hand, or between ionic and metallic bonding,

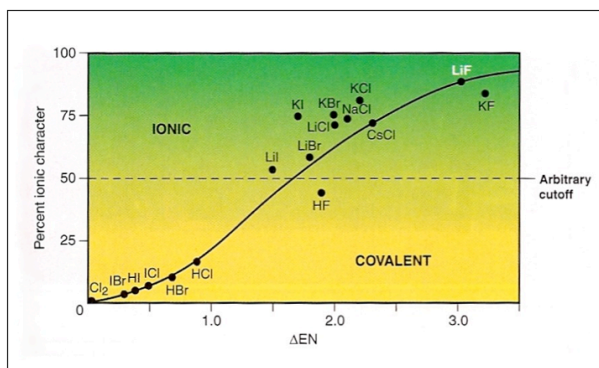


Figure 2. The textbook author's dichotomization of Pauling's original diagram illustrating the progressive transition from covalent to ionic bonding.

on the other, but rather simply discusses the unit cell packing for the pure metallic elements. In short, when push comes to shove, the author ignores the primary purpose of the van Arkel triangle by discussing only the three limiting-case vertices and the ionic-covalent transition along the left edge, which he incorrectly dichotomizes into an either/or situation. In so doing, he ignores all of the bonds lying along the other two edges and within the body of the triangle and thus 95% of all known bonds found in inorganic compounds.

Indeed, the necessity of maintaining the myth of the ionic-covalent discontinuity had already been introduced much earlier in the textbook when the author incorrectly stated that ionic inorganic compounds with variable valence are always named using the Stock oxidation-number system, whereas covalent inorganic compounds are always named using the stoichiometric prefix system. This is out and out incorrect. Since, as already stated, there is no sharp dividing line separating ionic and covalent bonding and no universally accepted criterion for assigning a borderline compound to one category or another, no one in their right mind would base their nomenclature on such an arbitrary dichotomy.

Since this discussion occurs several chapters before those dealing with atomic structure and with the ionic and covalent bonding models, the author tries to simplify the bond-type criteria by equating ionic bonds with metal-nonmetal element combinations and covalent bonds with nonmetal-nonmetal combinations. Ignoring the ambiguities of the metalloid band, this criteria is based on an either/or bulk-phase characterization of the simple substances in question. In reality not all metals are equally metallic and they may vary greatly in terms of their degrees of ductility, malleability and electrical conductivity. In any case, bond-type is not truly a function of the bulk-phase characters of the component simple substances but rather of the electronegativities of their component atoms and these vary in a progressive, rather than a discontinuous, fashion across the periodic table.

In the end, however, such subtleties are beside the point, since a moment's glance at the *IUPAC Redbook on Inorganic Nomenclature* quickly reveals that the Stock and prefix systems were never intended to be restricted to ionic and covalent compounds, respectively, but are rather two *alternative* systems for the naming of all binary inorganic compounds (4, 5). Thus Table 1 shows some of the nonmetal-nonmetal combinations or so-called "covalent" compounds used by IUPAC to illustrate the application of the Stock system and Table 2 shows some of the metal-nonmetal or so-called "ionic" combinations used by IUPAC to illustrate the application of the prefix system. Sadly, for the

Table 1. Examples from the IUPAC Redbook Illustrating the Naming of “Covalent” Compounds Using the Stock Oxidation Number System.

Formula	Name
N ₂ O	nitrogen(I) oxide
NO ₂	nitrogen(IV) oxide
CO	carbon(II) oxide
SF ₆	sulfur(VI) fluoride

last five years or more, students at Cincinnati have had certain answers to exam questions dealing with inorganic nomenclature marked incorrect, even though they were in fact correct according to IUPAC standards.

These nomenclature errors are further compounded by the fact that the textbook author, when first introducing the Stock system, incorrectly describes the Roman numerals in the names as “ionic charges” rather than as “oxidation numbers,” though in the chapter on transition elements near the end of the book he describes them as “oxidation states” when discussing the nomenclature of coordination complexes. In between, in the chapter dealing with redox reactions, he calls them “oxidation numbers” and represents them with Arabic numerals rather than Roman numerals, being careful to distinguish them from ionic charges by placing the sign in front of the numeral rather than after the

Table 2. Examples from the IUPAC Redbook Illustrating the Naming of “Ionic” Compounds Using the Stoichiometric Prefix System.

Formula	Name
Fe ₃ O ₄	triiron tetraoxide
U ₃ O ₈	triuranium octaoxide
MnO ₂	manganese dioxide
FeCl ₂	iron dichloride
FeCl ₃	iron trichloride
TlI ₃	thallium triiodide
Cr ₂₃ C ₆	tricosachromium hexacarbide

numeral – thus +2 represents an oxidation number whereas 2+ represents an ionic charge. Finally, in the chapter on Lewis diagrams he uses the exact same Arabic symbolism to represent formal charges rather than oxidation numbers.

While it is true that oxidation numbers for simple ions are numerically identical to their true ionic charges, the converse is not universally true. Rather oxidation numbers represent hypothetical or idealized ionic charges which correspond to true ionic charges in only a few limiting cases. Not only is it allowable to assign oxidation numbers to covalent species, where they have only an artificial significance – as amply illustrated by the author’s treatment of redox reactions – they are also purely artificial for many of the author’s so-called metal-nonmetal ionic compounds as well. Thus it is doubtful whether a Ti⁴⁺ ion really exists in titanium(IV) chloride or a Cr⁶⁺ ion in chromium(VI) oxide. Of course the student is never informed why it is legitimate to assign oxidation numbers to covalent compounds when balancing redox equations but supposedly verboten when naming them.

Table 3. Examples from the IUPAC Redbook Illustrating the Assignment and Proper Representation of Oxidation Numbers.

Species	Oxidation Number
[Co(NH ₃) ₆] ³⁺	III
[CoCl ₄] ²⁻	II
[MnO ₄] ⁻	VII
Fe(CO) ₄ ²⁻	-II

As illustrated in Tables 3 and 4, whether in writing Stock nomenclature or in assigning oxidation numbers in redox reactions, oxidation numbers, according to IUPAC, are always represented using Roman numerals and never using Arabic numerals. The use of Arabic numerals with an inverted sign convention should be reserved instead for formal charges so as to properly distinguish them from true ionic charges or “charge numbers,” as IUPAC now prefers to call them.

And if any more reason is needed, calling the Roman numerals used in Stock nomenclature “ionic charges” has the potential for confusing this system with yet a third naming system for inorganic compounds also officially endorsed by IUPAC and known as the Ewens-Bassett or charge-number system in

Table 4. Examples from the IUPAC Redbook Illustrating the Proper Representation of Oxidation Numbers within Inorganic Formulas.

Formula	Name
$\text{Os}^0(\text{CO})_5$	pentacarbonylosmium
$\text{K}[\text{Mn}^{\text{VII}}\text{O}_4]$	potassium tetraoxomanganate(VII)
$\text{Pb}^{\text{II}}_2\text{Pb}^{\text{IV}}\text{O}_4$	trilead tetraoxide

which iron(II) chloride becomes iron(2+) chloride and the number in parentheses is indeed intended to literally represent the ionic charge on iron.

4. 2c-2e bonds versus the Octet Rule

The debate over whether the 2c-2e covalent bond or the octet rule is the more universal bonding principle has a long history dating back to the 1920s and the work of G. N. Lewis and Irving Langmuir (6). Since the 1980s the evidence has pretty much supported the conclusion that the octet rule represents a rigorous upper bonding limit for all main-block elements and that, instead, it is necessary to supplement the 2c-2e bonds of conventional Lewis diagrams with a variety of other localized multicentered bonding components (7).

Thus the late 1940s and early 1950s saw the introduction of the concepts of back-bonding for such electron-deficient species as BF_3 and localized multicentered 3c-2e bonds for such electron-deficient species as the boron hydrides. Though both of these concepts have been standard features of the inorganic textbook since the 1960s and both are now over 60 years old, our freshman textbook appears to be blissfully unaware of them. Absolutely no mention is made of 3c-2e bonds or of the boron hydrides and the author literally states that BF_3 must of necessity violate the octet rule because the presence of back-bonding would result in unreasonable formal charges (8).

The same is true of so-called electron-rich or hypervalent compounds, such as PCl_5 and SF_6 . Once again the author insists that these must be formulated using 2c-2e bonds and must therefore violate the octet rule via octet expansion, even though Rundle and Pimentel had shown how to formulate them in accord with the octet rule using multicentered 3c-4e bonds as early as the late 1940s and early 1950s and these results have since been repeatedly confirmed using advanced quantum mechanical calculations. Indeed, the author goes even further by also invoking octet expan-

sion for such electron-precise species as SO_4^{2-} and PO_4^{3-} , though once again quantum mechanical calculations have shown that octet expansion is unimportant for these species (9). As will be commented on later, the dogged persistence of these circa-1930 bonding concepts in the freshman text appears to be an unintended by-product of their approach to the VSEPR model for the prediction of molecular geometries.

5. Molecular versus Nonmolecular Solids

If one classifies solids based on their known structures – that is, solely on a distinction between which components are chemically bonded and which are merely interacting via weak intermolecular attractions, one obtains the result in figure 3. Whereas the solid-state structures of most known organic species fall into the class of discrete molecules, over 95% of known inorganic species fall into the classes of infinite chain, infinite layer, or infinite framework polymers.

Though this sort of classification is the basis of such major reference works as A. F. Wells' classic *Structural Inorganic Chemistry* (10), our textbook author is apparently unaware of it and in the chapter on the solid state he substitutes instead, as shown in Table 5, a classification of solids based on bond type rather than structure type. From a purely structural point of view both the atomic and molecular classes in this table belong together since both consist of finite neutral units interacting with one another solely by means of weak intermolecular attractions, their only difference being that the noble gases form monoatomic, rather than polyatomic molecules. Likewise, from a purely structural point of view the remaining three classes, though interacting via a variety of different bonding

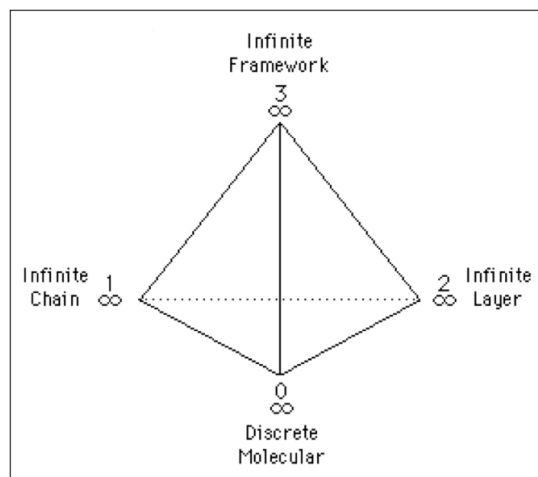


Figure 3. A structural classification of simple substances and compounds.

Table 5. The Textbook's Classification of Solids Based on Bond Type Rather than Structure Type.

Class	Examples
Atomic	Noble Gases
Molecular	O ₂ , C ₄ H ₁₀ , Cl ₂ , P ₄
Ionic	NaCl, CaF ₂ , MgO
Metallic	Na, Zn, Fe
Covalent Network	Diamond, SiO ₂

mechanisms, all belong together as examples of infinitely polymerized framework structures. Missing from the table are the equally important classes of infinitely polymerized layer and chain structures.

So what is the problem with using bond type instead of structure type to classify solids? Aside from the fact, emphasized earlier, that there is no sharp dividing line separating the metallic, covalent and ionic bonding extremes, the major difficulty is that such a classification falls apart the moment it is extended beyond binary compounds. Thus any ternary species containing a finite complex ion, such as K₂(SO₄) or [Co(NH₃)₆]Cl₃, etc. must of necessity contain a mixture of both ionic and polar covalent bonds. Indeed, if one looks at a sufficiently broad spectrum of examples, the system breaks down even for the case of binary compounds. Thus binaries containing poly-anions, such as Cs(I₃) or Na₂(S₂), contain a mixture of both ionic and purely covalent bonds, whereas species, such as Hg₂Cl₂ and NiS, contain a mixture of polar covalent and metallic bonds.

It is also telling that this table is one of the few places the author explicitly mentions co-called covalent network structures (11). Elsewhere he repeatedly fails to mention the nonmolecular character of many covalent inorganic solids and so, by implication, leaves the student with the impression that they are all composed of discrete finite molecules. Thus, in the chapter on VSEPR, he assigns BeCl₂ a discrete linear triatomic molecular structure and treats it as a covalently bonded species which violates the octet rule. This is certainly true of its gas-phase structure, but at normal temperatures and pressures, BeCl₂ is not a gas. Rather it is a crystalline solid having an infinite-chain structure. Each Be has a coordination number of 4 and it certainly obeys the octet rule. Indeed, using the textbook's earlier criteria for nomenclature, this compound would have to be classified as ionic rather than as co-

valent. Even if one ignores this inconsistency and properly classifies this compound as having polar covalent bonding, the author, using his criterion for denying the presence of back-bonding in BF₃, would have to declare that the solid-state structure of BeCl₂ is impossible on the basis of unreasonable formal charges.

A similar failure to mention infinitely polymerized chain and layer structures occurs in the chapter on the chemistry of the main-block elements, where the molecular structures of the S₈ and P₄ allotropes are illustrated but not the more stable infinite-chain and infinite-layer allotropes of these elements. Likewise, though the discrete molecular polycyclic structure of buckyball is shown for carbon, no explicit mention is made within the body of the text of either the infinite-framework structure of diamond or the infinite-layer structure of graphite.

This process of implicitly misleading the student through a failure to mention pertinent facts also extends to the textbook's discussion of phase changes. Since the majority of inorganic solids have nonmolecular structures, they must, of necessity, undergo both a depolymerization and structure change upon either melting or vaporization. Thus not only does the 4/2 infinite-chain structure of solid BeCl₂ change into a discrete triatomic molecular structure on vaporization, the infinite 8/8 metallic framework structure of sodium metal changes into a monoatomic gas, the infinite 6/6 ionic framework structure of sodium chloride changes into a discrete diatomic NaCl molecule, and the infinite 6/3 polar-covalent layer structure of aluminum trichloride changes into a discrete Al₂Cl₆ dimer (which is again the only form mentioned by the textbook). Yet the student is allowed to believe that melting and vaporization consist in nothing more than a change in the freedom of motion of otherwise preexisting discrete molecules common to all three states of matter.

A final problem with the chapter on the solid state is that the student is left with the impression that the unit cell concept is used exclusively for the description of close-packed metallic and ionic solids and that the only important unit cells are cubic. There are no drawings of the other 11 possible unit cell types and no illustrations of the use of unit cells to describe solids containing discrete molecules, infinite chains, or infinite layers. Not only does this massively distorted picture fail provide the student with a proper overview of the true nature and function of the unit cell concept, it conveys the false impression that this concept is somehow a necessary feature of close-packed metallic and ionic bonding and that cubic cells are the norm, when in fact they are the rarest type and constitute only 0.5% of the half million or so crystal structures reported in the Cambridge Structural Database.

6. Coordination Chemistry

Coordination theory is based on three key concepts:

- the central atom or ion concept
- the ligand concept
- the coordination number concept

If so desired, the structure of virtually any inorganic compound or complex ion may be dissected in terms of these three concepts. But once again, our textbook utterly fails to employ them in a consistent and unified manner. Though the concept of coordination numbers is introduced in the chapter on the solid state to describe the structures of close-packed metals and ionic compounds, no mention of it is made in the chapter on VSEPR when describing the structures of discrete molecules, in the chapter on the chemistry of the main-block elements, or in the chapter on complex-ion equilibria in solution, and it does not reappear again until the end of the book in the chapter on the chemistry of the transition metals, where it is used to describe the structures of their complex ions.

Likewise, the terms “central ion” and “ligand” do not appear until the chapter on complex-ion equilibria in solution and are not heard of again until the end of the book and the chapter on transition-metal chemistry. Indeed, it is in this chapter on complex-ion equilibria that the author makes his key blunder when he informs the students that a complex ion is defined as a *metal ion* with a set of surrounding ligands and that therefore both Na^+ and SO_4^{2-} are “simple” ions, whereas $\text{Cr}(\text{NH}_3)_6^{3+}$ is a “complex” ion. This is based on the increasingly common confusion of the term “metal complex” with the term “complex ion.” Though metal complexes do indeed form a very important class of complex ions, they are not the only known kinds of

Table 6. Example Complex Ions of the Nonmetals.

Complex	Name
NH_4^+	ammonium ion
SO_4^{2-}	tetraoxosulfate(VI)
ICl_4^-	tetrachloroiodate(III)
PF_6^{3-}	hexafluorophosphate(V)

Table 7. Example Complex Ions of the Main-Block Metals.

Complex	Name
$\text{Ca}(\text{EDTA})^{2-}$	(ethylenediaminetetraacetato)calcate
$\text{Cd}(\text{NH}_3)_4^{2+}$	tetraamminecadmium
$\text{K}(\text{Cryp})^+$	(2.2.2-cryptando)potassium
$\text{Be}(\text{OH})_4^{2-}$	tetrahydroxoberyllate
$\text{Zn}(\text{CN})_4^{2-}$	tetracyanozincate
AlF_6^{3-}	hexafluoroaluminate
HgI_4^{2-}	tetraiodomercurate(II)

complex ions.

Only in the chapter on the transition metals is the term “coordination compound” or “coordination complex” used and defined as a compound containing at least one complex ion with a metal as its central ion. Though the author does correctly state that the metal in question may come from either the main-block or the transition-block, nowhere in the entire text does he give an example of a main-block coordination complex and so, once again, by omission rather than commission, he leaves the student with the false impression that such complexes are, for all practical purposes, nonexistent.

In reality, a simple ion corresponds to *any charged monatomic species* and a complex ion to *any charged polyatomic species*. As a glance at Werner’s original monograph on coordination chemistry, any recent textbook of either coordination chemistry or inorganic chemistry, or the *IUPAC Redbook* will quickly reveal, there is no requirement that the central ion of a complex must be metallic rather than nonmetallic (Table 6) and main-block coordination complexes, (and especially those corresponding to the oxoanions) are just as common as transition metal complexes (Table 7) (4, 11). What is true is that the central atom is usually the least electronegative atom in the complex.

I should, however, qualify my earlier statement that the author fails to give examples of main-block coordination complexes. In fact he does give some in the section on Lewis acid-base adducts but fails to identify them as such and also gives some complexes of Zn, Cd and Hg, but incorrectly identifies them as transition-metal complexes. Since these three elements never use d-electrons in their bonding, they are, in fact, main-

block rather than transition-block elements, and hence several of their complexes are included in Table 7 (12, 13).

7. Possible Suspects

What are the reasons for these myths and why do they persist? The role played by copycat textbooks and publisher aversions to novelty have already been mentioned in the introduction. The failure of authors (and teachers) to consult advanced monographs on those subjects which lie outside of their own field of expertise or to even read the pertinent articles dealing with these subjects in the *Journal of Chemical Education* is yet another.

More specific causes may be identified as well. Thus, in the case of the issue of 2c-2e bonds versus the octet rule, I think that the widespread, but false, belief that the former concept is a necessary prerequisite for the application of the VSEPR model for the prediction of molecular geometry plays a major role in resisting the introduction of the concepts of 3c-2e and 3c-4e bonds into the freshman text. But just as the electron-repulsion units in VSEPR may correspond to 1c-2e lone pairs, 2c-2e single bonds, 2c-4e double-bonds, or 2c-6e triple bonds, so they may just as easily correspond to 3c-2e bonds or to 3c-4e bonds.

Likewise, the distorted presentation of the unit cell concept with its overemphasis on close-packed cubic cells is almost certainly a consequence of the current obsession with numerical problem solving in the freshman course. Simple close-packed cubic cells allow the student to do a variety of simple calculations involving densities, radii, and even the calculation of Avogadro's number, which less symmetric cells with looser packings do not. This, in turn, reflects the attitude that, if a topic cannot be made into a calculation, then it is untestable and hence unworthy of inclusion in the textbook.

I see that I have more than used the time allotted and will consequently bring my remarks to a close in the hope that they have stimulated the teachers in the audience to look more closely at what they are teaching and why.

7. References and Notes

1. An invited lecture given as part of the symposium

"Myths of a Chemical Nature" held at the 240th National ACS Meeting, Boston, MA on 26 August 2010.

2. W. B. Jensen, "The Historical Development of the van Arkel Bond-Type Triangle," *Bull. Hist. Chem.*, **1992-1993**, 13-14, 47-59; and W. B. Jensen, "A Quantitative van Arkel Diagram," *J. Chem. Educ.*, **1995**, 72, 395-398.

3. L. Pauling, *Nature of the Chemical Bond*, 3rd ed., Cornell University Press: Ithaca, NY, 1963, p. 99. Earlier versions of this diagram go back at least to the 1940s. It should be noted that this graph is a definition of ionic character and not an empirical correlation as there is no other independent measure of the γ -parameter.

4. G. J. Leigh, Ed., *Nomenclature of Inorganic Chemistry, Recommendation of the International Union of Pure and Applied Chemistry Commission on the Nomenclature of Inorganic Chemistry*, Blackwell Scientific Publications: Oxford, 1990, pp. 47, 64, 66, 67, 132, 149.

5. The application of Stock notation to covalent compounds was first discussed in R. Brasted, "Revised Inorganic (Stock) Nomenclature for the General Chemistry Student," *J. Chem. Educ.*, **1958**, 35, 136-139.

6. W. B. Jensen, "The Origin of the Term "Hypervalent,"" *J. Chem. Educ.*, **2006**, 83, 1751-1752 and the many reviews and papers cited therein.

7. F. Weinhold, C. Landis, *Valency and Bonding*, Cambridge University Press: Cambridge, 2005, Chapter 3.

8. It is true that the formal charges are unreasonable, but the calculation of such charges assumes that the bonds are 100% covalent, whereas those in BF_3 are highly polar. In short, formal charges are too crude to count as valid evidence and most books on bonding theory maintain that back-bonding and octet formation are present.

9. L. Suidan, J. K. Badenhop, E. D. Glendening, F. Weinhold, "Common Textbook and Teaching Misrepresentations of Lewis Structures," *J. Chem. Educ.*, **1995**, 72, 583-586. Once again the textbook's conclusions appear to be driven by an excessively naive use of formal charges.

10. A. F. Wells, *Structural Inorganic Chemistry*, 5th ed., Clarendon Press: Oxford, 1983.

11. A. Werner, *New Ideas on Inorganic Chemistry*, Longmans, Green and Co: London, 1911.

12. E. M. Larsen, *Transitional Elements*, Benjamin: New York, NY, 1965, pp. 1-2.

13. W. B. Jensen, "The Place of Zinc, Cadmium and Mercury in the Periodic Table," *J. Chem. Educ.*, **2003**, 80, 952-961; and W. B. Jensen, "Is Mercury Now a Transition Element?," *J. Chem. Educ.*, **2008**, 85, 1182-1183.