

# Toward a Dynamic Lewis Notation

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**Abstract:** Because the presently used Lewis diagrams have their roots in static electron theories of the early twentieth century, this paper proposes substituting a double-arrow symbolism for the customary double dots. This means that students should be taught that a bond line (or “stick” as some call it) means a double arrow, not a double dot. Lone pairs are to be represented by double-curved arrows and later abbreviated to the old familiar double dot. There are obvious advantages to this new notation when atoms have different electronegativities. The double arrows can suggest the shifting of the bond pair equilibrium toward the most electronegative atom. The new notation is more compatible with VSEPR theory, and it seems to suggest the loci of largely imaginary MOs and AOs by the configuration of the arrows. With the help of some imagination, the new symbolism can clarify resonance in aromatic molecules and conjugated chains. The most radical suggestion is that resonance forms be abandoned in favor of conjugated blur bonds. The concept of blur bonds has the merit of being easily extendable to the explanation of metal bonding and electron conduction.

The commonly used pictures of molecules created by combining the element's symbols with dot pairs to create an audit of how many valence electrons surround an atom and which of these pairs serve bonding functions are called Lewis structures [1–3]. These pictures have been used for nearly a century, and many generations of students have been asking the same question that I asked when I took general chemistry. If electrons never stand still, why do these drawings, which show them standing in polite pairs above, below, and to the sides of an element symbol, help us to understand the structure of molecules?

## Part I: Evolution of the Lewis diagram

When I was a student in the 1950s, the popular picture of atoms was the nonplanar solar system picture. Electrons were in highly eccentric elliptical orbits around a stationary nucleus (Figure 1). The nucleus, on the other hand, looked like a plum pudding, with protons and neutrons packed together, obviously static. We were taught that earlier in the century some chemists actually believed that the whole atom was a plum pudding, and thus even the electrons were static (like the electrons in our Lewis pictures).

**Static Models of the Atom.** My instructors generally scoffed at the “plum pudding model” of the atom, attributed to Thomson, while at the same time quite willingly ignored any dynamic structure the nucleus might have. What I was never told was that G. N. Lewis himself went through a period when he thought of the electrons as static, that is to say as unmoving relative to the nucleus. Lewis drew the atoms of common elements as cubes [3].

During Lewis' early days as a chemist the concept of electron transfer to form ionic compounds was well accepted. This was largely because many electrical conductivity studies of solutions of electrolytes had been done in the latter part of the 19th century. So, ionic bonding was not mysterious, and it required no model of the atom except a vague concept of shell

structure, which governed the number of valence electrons. Lewis realized that organic compounds could not be well explained using ionic ideas, and his cubical atoms (Figure 2) helped him conceive of the idea of a shared electron pair bonding together two atoms. This idea of two atoms sharing a pair of electrons was so useful that the fact that it originated in the contemplation of a static model was forgotten, and the electron pair sharing process was transformed by Pauling and many others into a central feature of dynamic (read “quantum mechanical”) atoms and their bonds. Lewis abandoned the static electron idea as quantum mechanics advanced, but others of Lewis's generation did not. Irving Langmuir [3], for example, promulgated a quite elaborate static electron scheme that could explain many coordination compounds as well as simple organic compounds (Figure 3).

By the mid-twentieth century most static-electron models were ignored. As a last gasp of static theories, a little known Belgian engineer named Rudy Weckering [4] proposed that the electrons in atoms were static, sitting on the nodes of standing waves (his nod to quantum mechanics). Weckering, from 1930 through 1960, worked out an elaborate scheme he called the nodic field theory (Figure 4). By placing electrons on the nodes of his standing waves it was possible to explain atomic structure for both the representative elements and the transition elements. Bonding was accomplished by atoms sharing their nodes, very like some early Lewis pictures of cubes sharing an edge.

Not only could Weckering's nodic field theory explain all of chemical valencies and bond angles, if scaled down to a very small size the same nodic field could be populated with neutrons and protons. Thus, he explained radioactivity and all nuclear physics. There was even a tiny little nodic field inside of each proton and neutron by which we could construct these particles from the appropriate mesons.

**Dynamic Theories, but with a Static Mnemonic.** Chemistry and physics paid little heed to nodic field theory and Weckering's work, yet chemists still use static Lewis

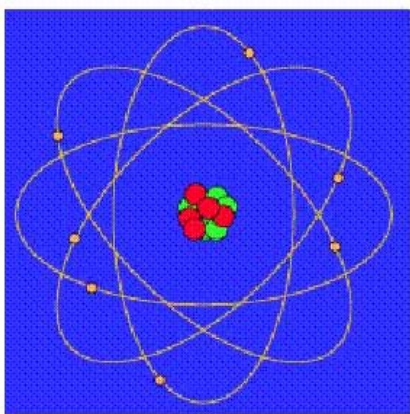


Figure 1. The mid-twentieth century atom picture.

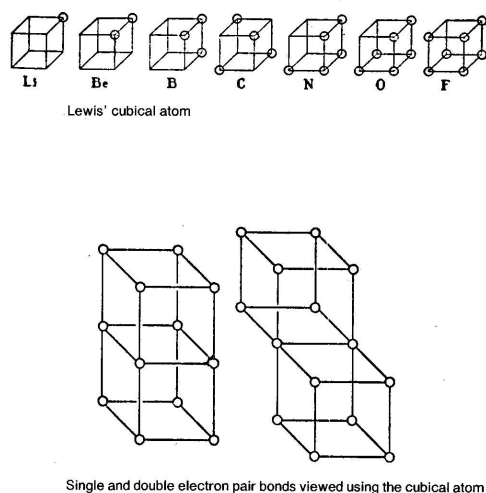


Figure 2. Lewis cubes, reproduced from Jensen, W. B. *J. Chem. Educ.* **1984**, *61*, 194. The lower figures illustrate a double bond and a single bond. (Used with permission from the *Journal of Chemical Education*, Vol. 61, No. 3, 1984, pp. 191–200; copyright © 1984, Division of Chemical Education, Inc.)

diagrams to explain many of the same things that Weckering could explain using his nodic field theory. By continuing to use Lewis pictures today we appear to our students to be suggesting that the electrons can be thought of as stationary on the outer regions of atomic orbitals, much like Weckering's nodes of standing waves.

Why did Lewis's static symbolism survive while Weckering's went unnoticed? Because Lewis diagrams, if not taken literally, were a somewhat useful and quite simple mnemonic. Counting valence electrons and suggesting the octet rule "explained" why hydrogen is monovalent, oxygen divalent, and nitrogen trivalent. This double-dot scheme does, to a new student, seem to provide order to the chaos of chemical formulas.

But it didn't take students long to realize that there were a significant number of molecules and ions for which the simple Lewis ideas provided little help or predictive power. Chemistry teachers today address this pedagogical problem in several ways:

1. Some teachers prefer to use noncontroversial Lewis structures [5], which work quite well for organic compounds, but then introduce the fundamental ideas of quantum mechanics. This is often followed by the suggestion that quantum calculation techniques, which the students were not prepared to do, would explain the microscopic world satisfactorily to chemists, so students need not worry themselves about it.
2. Other teachers and most modern texts prefer to elaborate the Lewis rules so as to explain more molecules and ions, before finally giving up and tossing the problem to the mysterious quantum mechanical world. This school of chemists, pioneered by Linus Pauling [6], suggests that one can draw multiple Lewis structures to show the location of the bonding electrons and then conclude that none were correct. In effect the students are told that the actual structure was something similar to these drawings, but something that we cannot draw. The drawings conceded to be incorrect were called "resonance structures." Because certain resonance structures seem more believable than others, teachers then utilize the concept of "formal charges," [7–9], a mathematical evaluation procedure for each Lewis structure which produces a formal charge number (not a real electrical charge) for each bonded atom. These numbers, it is suggested, should be as near to zero as possible, and so one could "believe in" those Lewis structures with the lowest formal charge, and pretty much ignore the others.
3. Many teachers augment Lewis structures with three-dimensional special considerations. For example, lone pairs take up more room than bonded pairs. The electrons had a "domain" and were not static in this domain. These steric considerations were responsible for the creation of valence-shell electron-pair repulsion theory (VSEPR) to explain molecular geometry [10–14]. The primary difference between Lewis "dot diagrams" and VSEPR is that VSEPR places the electron pairs in "electron domains," rationalizing their size by whether or not they overlap electron domains of other atoms. In an introductory course even quantum mechanical experts often prefer Lewis structures and VSEPR to attempts at explaining valence bond (VB) and molecular orbital (MO) theories [15].
4. Frustrations with Lewis dots have lead some teachers to either give up on representing all the valence electrons [16], or they give up on Lewis theory and rely on electron density maps generated by computer [17]. It is difficult to see how this increases the student's predictive powers in chemical speciation.

As is evident by the title of this paper, my view is that we should retain Lewis diagrams, but seek to transition the static Lewis pictures into more dynamic ones.

## Part II: Improving Lewis Diagrams

Before explaining the changes that I think would be an improvement to the way we draw Lewis diagrams, I would like to point out that the double-dot Lewis picture of the covalent bond has already been changed. The Lewis picture of methane, to take a simple example, is a C symbol surrounded by four dot pairs, upon which sit the four H symbols. The only place you will find such a picture is in elementary chemistry texts.

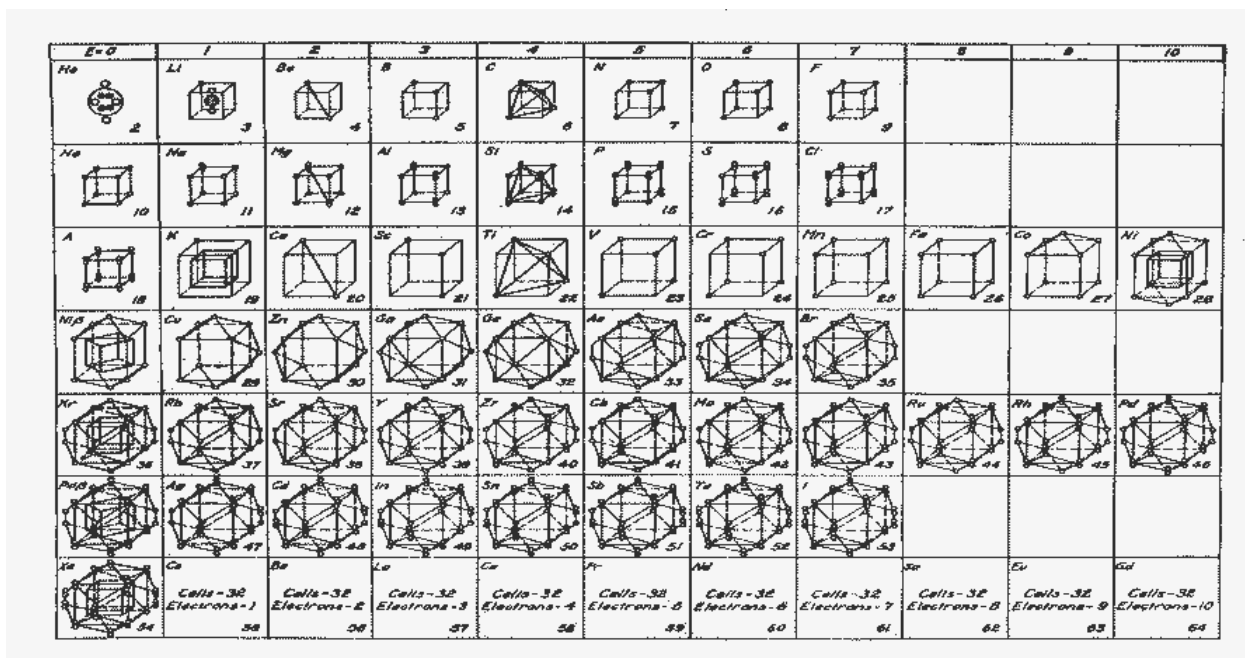


Figure 3. Langmuir cubes, reproduced from Jensen, W. B. *J. Chem. Educ.* 1984, 61, 196. Cubes extended to the transition elements. (Used with permission from the *Journal of Chemical Education*, Vol. 61, No. 3, 1984, pp. 191–200; copyright © 1984, Division of Chemical Education, Inc.)

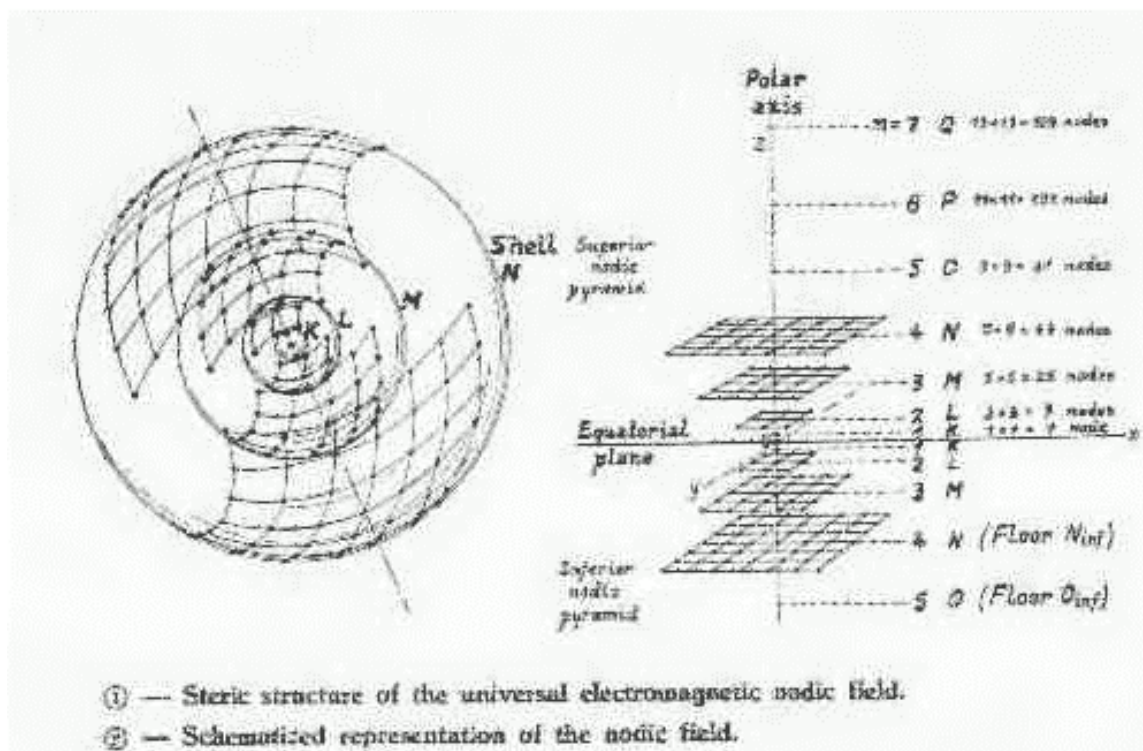
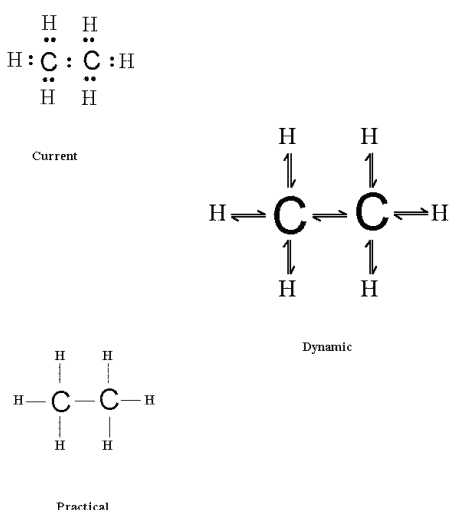


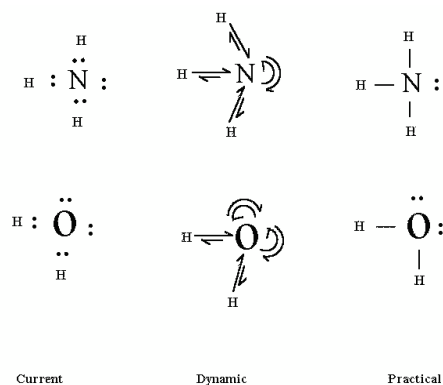
Figure 4. Rudy Weckering's nodic field, reproduced from Weckering, R. *The Nodic Field Atom*, 1st ed.; Camille Hermann: Luxemburg, 1957. The intersections of the lines are the nodes on which the electrons may sit; however, electrons were not permitted to sit on nodes adjacent to another electron. The result was 2,8,18,32 in successive tables of the nodic field.

Chemists, to the best of my knowledge, have replaced the double dots with short lines, or “stick pictures” as some chemists call them. The stick as a representation of the chemical bond is far superior to the Lewis double dot. It suggests connectivity and direction in space in a much simpler way than the double dots can. Sticks are much quicker to draw than dot pairs and are therefore used whenever possible. Of

course, the valence electron count is better depicted by Lewis dots and dictates the number of sticks that can be drawn. The important point I wish to make here is that we currently teach all chemistry students to think of a stick joining atoms as a stationary dot pair in this location. It is this thinking about stick bonds that I hope to change by introducing a new dynamic notation in place of the double dot.



**Figure 5.** Ethane represented as double arrows instead of dots or sticks.



**Figure 6.** Ammonia and water, the representation of lone pairs.

Chemists, subsequent to general chemistry lessons, draw molecule representations almost invariably as sticks and a few lone-pair dots. This is because they are no longer attempting to justify the valence electron behavior but instead have moved on to the electron action at some functional group or reaction site. Chemists no longer bother with the lone-pair double dots unless they are relevant to the reactions of the molecule in question. Even the sticks disappear when the structure of the molecule other than the reaction site is unimportant, as in the use of R to symbolize a molecular fragment.

I advocate that sticks remain the symbol for a covalent bond. I also advocate the double dots remain as a lone-pair symbol. Organic chemists will not need to change their notations representing reaction mechanisms, but what the sticks and the double dots represent should be quite different in the minds of future chemists. I do advocate that what the sticks and the double dots represent should be changed in elementary textbooks. In this paper I am advocating more dynamic symbols to replace the electron-dot symbols as a sigma bond, and more dynamic symbols to represent the lone pair at the student's first encounter with these concepts and, most important of all, I advocate a more dynamic symbolism for multiple bonding.

**Suggestion 1: What the Sticks represent.** The major suggestion is to replace electron dots with arrows and dot pairs

with double arrows (Figure 5). In order to make it very clear just what changes I advocate in this and many of the figures I present three views of the molecule of interest:

- Current: the Lewis current teaching mode figure with double dots, which currently are used for covalent bond representation and which results in chemists thinking of a stick bond as the double dots.
- Dynamic: the dynamic Lewis teaching mode figure with double arrows, which I propose and which in the near future would result in chemists thinking of a stick bond as dynamic electron equilibrium.
- Practical: the practical diagram of the molecule, which will be used by most chemists in depicting or studying functional groups and reaction kinetics. In many cases this form is the form currently in use. Possibly only beginning chemistry teachers will choose to use the dynamic notation.

Why does this improve Lewis diagrams? It is because the double arrows imply a dynamic process and an equilibrium state, rather than two stationary electrons. They imply that the electrons are moving from one atom to another in such a way as to bind the atoms together. This exchange of electron density from one nucleus to another properly represents the covalent bond as an equilibrium state for the pair of electrons. It is, therefore, an appropriate use of the double arrows. The covalent bond is an equilibrium of electrical charge flow until disturbed by an approaching species with which it may react.

The double arrows representing a nonbonding pair of electrons (a lone pair) are restricted to one nucleus but, if quantum mechanics is correct, mainly directed to one conical region with only the tip of the cone encompassing the nucleus. This is the lone-pair domain of VSEPR; therefore, I propose curved double arrows as the lone-pair representation. The double arrows represent the flux of electrons that must be in space and charge equilibrium or else the lone-pair domain would change in size. Reacting the lone pair with a positive center from an approaching species can be pictured as uncurling the double arrows. This, of course, is the familiar Lewis acid–base reaction.

I suggest the curved-double-arrow notation furthers the concept of the lone-pair repulsion, requiring a larger angular diameter for this domain. Figure 6 shows examples of the dynamic notation with arrows for sigma bonds and curved arrows for the lone pairs.

**Why Use Equilibrium Arrows?** Electrons have three properties. They have electric charge, though it is hard to say exactly what that is. They have spin, though it is hard to say exactly what that is. They have mass, and, although it is hard to say exactly what that is, most students think they know. I propose the double arrows represent the electric charge flow within the orbital (whatever that is). A double arrow beside an element symbol means, "There is an orbital encompassing the nucleus of this atom and extending in this direction. This orbital contains two electrons of opposite spin. Because the electrons cannot be stationary, the double arrow represents negative charge flow within this orbital towards and around the nucleus of this atom and away from this atom in a particular direction. The particular direction is suggested by the placement of the double arrows relative to the element symbol." Because Lewis diagrams are often drawn in two

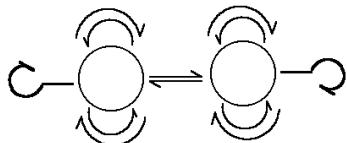
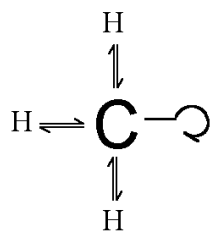


Figure 7. Free radicals and the fishhook:  $\text{CH}_3\cdot$  and  $\text{O}_2$ .

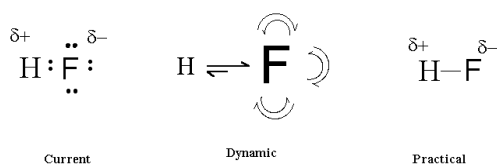


Figure 8. Electronegativity as shown by the double arrows.

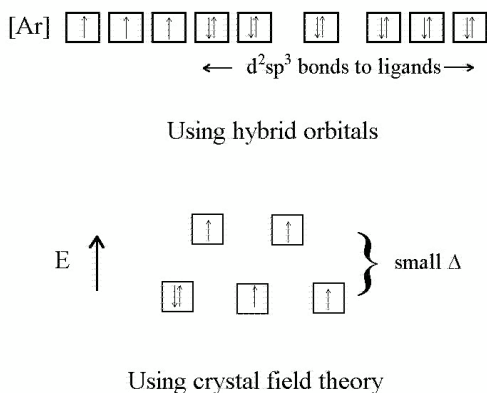


Figure 9. Coordination chemistry and spin states notation unchanged.

dimensions, the directions in question are largely up to the imagination of the beholder.

Neither the curved double arrows nor the straight ones are meant to precisely define the shape of the orbital. Still, double arrows between atom (actually kernel) symbols suggest an orbital of some shape that encompasses these regions. With the use of this double-arrow notation, replacing the double dot, we move closer to a believable representation of the molecular orbital quantum mechanics can sometimes calculate. Certainly the concept of exchange energy, which arises in quantum calculations, seems more meaningful in this dynamic notation for the bond electrons.

Should an outer orbital have only one electron within it, then it could be drawn as a fishhook shaped arrow. This is a “free radical” meant to catch other passing species (Figure 7). I

propose these double arrows not be thought of as the electron’s spin, despite the fact that the two possible values of electron spin are commonly represented by similar symbols. The representation of spin is discussed later in this paper.

**Suggestion 2: Electronegativity and the Double Arrows.** Having the double arrows represent the electric-charge flow within the orbital is especially fortuitous notation when atomic orbitals merge to form a molecular orbital linking two unlike atoms. If the two atoms are identical the double-arrow symbol is unchanged. If the atoms thus joined have different electronegativities the arrow pointing to the more electronegative of these atoms is simply made longer than the contra-arrow. If the relative lengths are approximately scaled to the relative electronegativity values, the negative regions of the molecules become evident without the usual + and – signs (Figure 8).

**Suggestion 3: The Spin State of Electrons.** I do not propose that the double or the single arrows above represent the spin of the electrons. Of course there are two arrows as there are two spin states, and unpaired spins can still be noted by looking at the pairing of arrows in these modified diagrams; however, should the student be told to think of these arrows as spin state arrows, there would arise a problem. The spin states of electrons are limited to two values. If arrows are used to represent spin states, the magnitude of the vectors cannot logically be different one from the other. Using the arrows lengths for electronegativity charge shifting precludes using them to represent spin. Instead, I suggest that the double and single arrows discussed above always be single-pronged arrows, and that any arrows that might be drawn to represent the spin states and their pairing or unpairing be the customary double-pronged arrows.

On the subject of spin states, it should be clear that this new arrow notation used in place of the usual double dots does not enable Lewis pictures to explain any more nor any less about the paramagnetism of molecules. It does not, for example, solve the conundrum of the paramagnetism of oxygen (Figure 7).

**Coordination chemistry.** Lewis diagrams, with either dots or arrows, are two-dimensional figures whether on blackboards or paper. Three dimensionality can be suggested by artistically inclined instructors, but I suggest that it is a mistake to carry such diagrams into octahedral and similarly complex symmetry classes. For coordination chemistry I suggest what is already being done. Draw the atom or ion core. Draw the requisite number of square boxes nearby to represent the hybridized or unhybridized orbitals in question. Populate these boxes with up-and-down double-pronged arrows to represent the spin states of the electrons (Figure 9). My preference for square boxes rather than for circles is based on the fear that circles may mean some kind of circular orbit to certain literal students.

**Suggestion 4: The Special Case of Hydrogen and its Kernel.** The suggestions above all concern the use of arrows rather than dots to represent various aspects of chemical bonding. My next suggestion has to do with the size of the element’s symbol, which presumably represents the entire atom except its valence electrons (its kernel). The element used most often when drawing Lewis structures while teaching general chemistry is hydrogen. Because its kernel is nothing but the nucleus, it seems wrong to write H as large as one would write C, for example. On the other hand, one cannot write H ten thousand times smaller than the other symbols. I

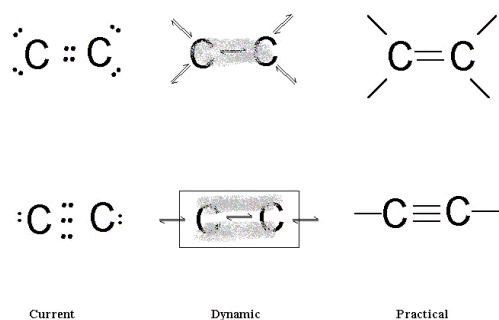


Figure 10. Ethane and ethyne, three symbolism schemes.

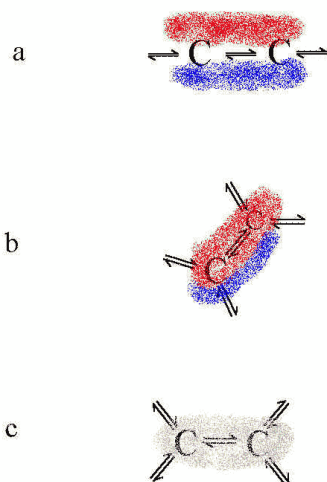


Figure 11. What the student is to imagine: (a) side view of a double bond, (b) perspective view of a double bond, (c) dynamic notation for a double bond.

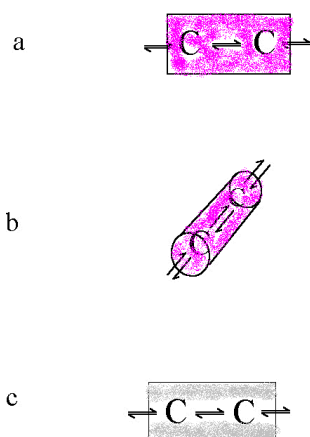


Figure 12. What the student is to imagine: (a) side view of a triple bond, (b) perspective view of a triple bond, (c) dynamic notation for a triple bond.

suggest writing H in Lewis structures as half the size of other symbols. This is noticeable to the students yet not ridiculous. It may serve as a reminder to the student that the  $H^+$  ion is very unlike other ions in its behavior. The exceptional mobility of the aqueous proton in both acid–base reactions and in conductivity experiments on electrolyte solutions is foreshadowed by the use of this somewhat smaller symbolism. You probably noticed this smaller H in the above figures.

Please recognize that the smaller H suggestion above is not a suggestion to show atomic sizes in Lewis notation. Even for experienced chemists it is easy to forget that the element symbols in Lewis notation represent not the atoms, but the atom kernels. The kernels are all of the atoms except their outer-shell electrons. Hydrogen is the only case of a chemically bonded atom whose kernel is a nucleus. This is responsible for its unusual chemical behavior in many reactions; therefore, depicting it as noticeably smaller may remind students of this idiosyncrasy.

### Part III: Multiple Bonds and Resonance

Representing double and triple bonds in a two-dimensional space is difficult in any scheme. Using Lewis dots results in especially meaningless representations. The placement of four or six dots between two atoms suggests an unbelievable localization of electrons. I blame the static nature of Lewis drawings for the numerous attempts to keep double bonds associated with two chosen atoms. As a consequence of these attempts we postulate multiple “resonance forms,” with the “true structure” being something we cannot draw and that is somehow a mixture of these forms. It is my opinion that the fact that we cannot draw it is not a mystery of natural philosophy, but simply the result of our attempts to represent the double bond as two sticks holding two chosen atoms together.

If in this paper I were to advocate drawing the non-sigma bonds by parallel sets of double arrows, which would then appear as double sticks, chemists will simply continue this resonance form confusion. I suggest that instead, we think more of the non-sigma bond as a region of electron flux near the sigma bond which enhances the bond strength, shortens the bond, but is easily delocalized so as to encompass several nuclei in a multi-atom moiety if this results in greater symmetry and low-energy electric charge distribution. The notation we use for this non-sigma bond should clearly indicate that this bond is very unlike the sigma bond in that it is extendable to other regions of the molecule should energy, geometry and symmetry considerations favor this delocalization.

**Suggestion 5.** And so I choose the flat representation shown in Figure 10. I call these “blur bonds,” but they could also be called pi bonds. Most chemists brought up on valence bond theory will probably think of it as overlapping p orbitals remaining after  $sp^2$  hybrid orbitals are formed. However one thinks of it, it is a molecular orbital representation that is intentionally vague as to shape. I propose it be depicted as blurred because it is easily delocalized or merged with other blur bonds. The double bond between two atoms is then double arrows (the sigma bond) accompanied by a blurred region (the blur bond). The students will still need some imagination while using two-dimensional symbols in order to imagine the blur bond as both above and below the framework of atoms. It is crucial that this be thought of as a “sandwich” bond; else cis–trans isomerism would not be explained. Figure 11 attempts to illustrate what is to be imagined. In the case of triple bonds the sandwich of double bonds is replaced by a cylinder of charge containing four electrons. Figure 12 is intended to illustrate this cylinder of electron charge that is the triple bond.

Whether or not this is strictly an accurate representation of the electron domains, it has the advantage of being two-dimensional and of offering a conceptually easy transition to

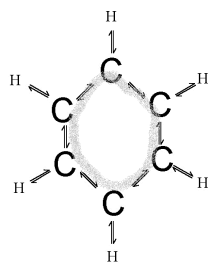


Figure 13. The benzene circle, the parent idea behind the blur bond.

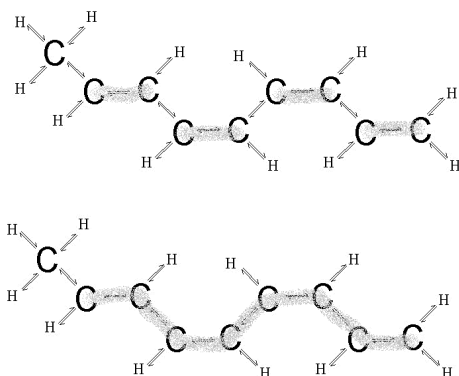


Figure 14. Conjugated chains as blur bonds.

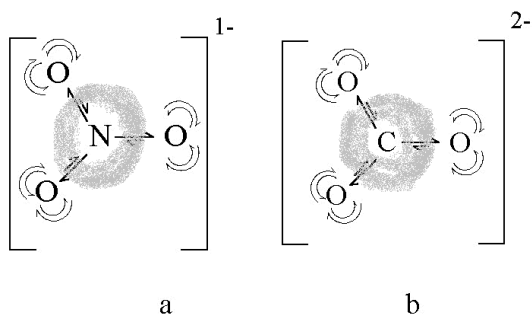


Figure 15. Nitrate and carbonate ions using six-electron blur bonds.

resonance visualization sans resonance forms and formal charges.

**Why blur bonds?** Blur bonds are not new. The name is, but chemists have drawn these bonds for many years. The obvious example is benzene. After drawing Lewis-dot pictures and stick pictures in which the bonds are either single sticks or double sticks, they finally combine the pi bond sticks into a circle. This is exactly what merged blur bonds do in benzene as you can see in Figure 13.

The driving force for delocalization (as in aromatics and conjugated chains) might be attributed to “the seeking of symmetry.” To me that seems a reasonable driving force for electron rearrangement, because electrons do have a wave nature. In Figure 13 I have not included the Kekule’ forms dictated by the double-sticks representation, but only the delocalized form. I heartily agree that the double-stick forms are not the real structure. What I disagree with is that the valence-electron structure is something we cannot draw. We have been drawing the benzene circle bond for years. The only

difference between this circle and my blurred circle is the recognition that this is a permitted Lewis structure.

The intent is to suggest that nearby blur bonds in appropriate spatial configurations might rearrange to a delocalized electrons arrangement. The resulting lower-energy delocalized-electron domains are represented as blurred areas, in the figures here. In the classroom they might be drawn on a chalkboard by holding the chalk flat against the board and making a broad sweep [18]. See Figure 14.

**Resonance in polyatomic ions.** I suggest that this double-arrow notation, accompanied by blur bonds, could make resonance structures in many molecules and polyatomic ions unnecessary. Here are some examples of invoking the blur bond, including the rules used to draw these dynamic Lewis configurations.

In drawing these proposed structures the rules are:

1. Blur bonds must be composed of an even number of delocalized electrons. Two-, four-, six-, and eight-electron blur bonds can be postulated depending on the symmetry and nuclear charge distribution in the molecule. With large central atoms it seems likely that ten or more electrons might merge into what is essentially one molecular orbital.
2. When oxygen surrounds the central atom in a group of atoms, oxygen is attached to the central atom by a sigma bond (double arrows). Each oxygen is then considered to have two lone pairs plus one electron to contribute to the delocalized blur bond. If fluorine is the peripheral atom the same thing is true except that now the fluorine can contribute its remaining two electrons to the blur bond. In general, the peripheral atoms keep only two lone pairs, contributing any remaining valence electrons to the blur bond.
3. The exact shapes of the blur bonds and conjugated blur bonds are not important. Quantum mechanics will calculate that. That they distribute the negative charge as symmetrically as possible within the electric charge distribution restraints in a single representation is the important idea, obviating the need for resonance structures.
4. The sigma bonds will be shortened from their normal single bond lengths by an amount related to the fraction of the blur bond in which they participate. When the blur-bond electron number reaches two, the blur bond has become a simple pi double bond.

Examples of Using Blur Bonds To Complete the Electron Audit in Several Molecules and Ions

- Nitrate ion and carbonate ion. Figure 15.
- Phosphoric acid and its ions. Figure 16.
- Sulfuric acid and its ions. Figure 17.
- Boron trifluoride (19). Figure 18.

The structure of ozone is a frequently used example in chemistry texts. It is a good example of the attempt to symbolize a blur bond prior to this paper. Why confuse students with resonance forms of ozone, when a blur bond without a sigma bond bridging the oxygens will suffice as an approximate form (Figure 19). The sigma bond cannot be used here because of bond angle strain and the observed internuclear distances, but because all Lewis drawings are approximations, why not a lone pi rather than “resonance contributions”?

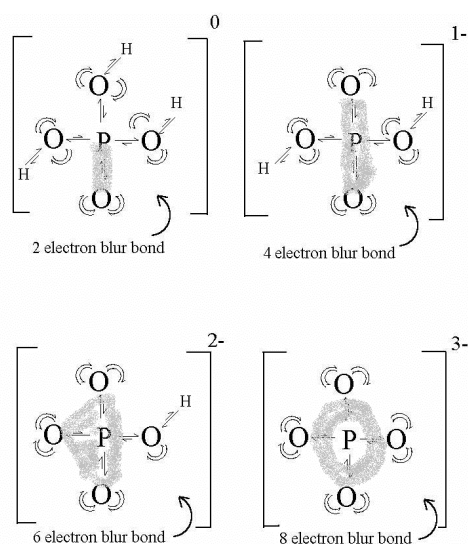


Figure 16. Phosphoric acid and its ions, blur bond representations.

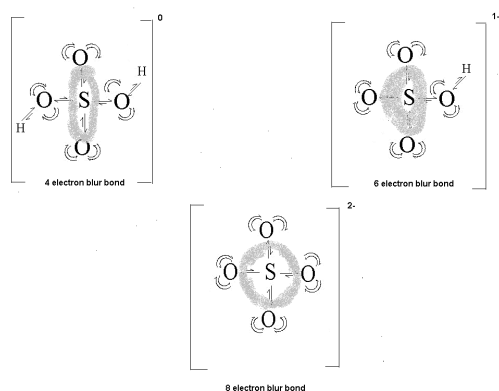


Figure 17. Sulfuric acid and its ions, blur bond representations.

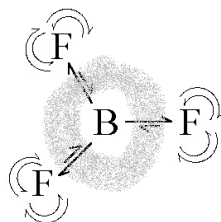


Figure 18. Boron trifluoride and its blur bond. The blur bond contains six electrons.

#### Part IV: Suggestion 6. Metal Conductivity and “the Sea of Electrons.”

Finally, I note the similarity between blur bonds and the usual general chemistry attempt at explaining metallic bonding and the electrical conductivity of most metals. The usual approach is to introduce something called a metallic bond, implying that metals are not bonded by the kinds of chemical bonds previously discussed. Sodium ions surrounded by a “sea of their valence electrons” is a popular example. Then the subject of metallic bonding is usually dropped, or relegated to a materials science special chapter perhaps discussing electron conduction in terms of band theory.

If Lewis structures are taught using the dynamic model proposed in this paper, then no such special “metallic bond” need be postulated. The metallic bond now becomes a crystal-encompassing blur bond consisting of the s-orbital electrons just outside the closed inert gas configuration.

It is a mistake to compare the specific conductances of metals in order to understand metallic conduction. To understand the conductances of electrolyte solutions we do not compare their specific conductances. We compute their molar conductances. For an element that has any electrical conduction, the molar conductances are easily computed by multiplying the specific conductances by the molar volumes of the metal [20]. This quantity allows comparison of the conductances of equal numbers of atoms, rather than equal volumes of them. Some of the trends that can be seen by general chemistry students through a study of this graph (Figure 20) are listed here.

1. The alkali metals are all good conductors per mole. Apparently their one-electron-per-atom blur bond is a good mechanism for electron conduction. At the same time it is also a bond, but because these are soft elements, the one-electron-per-atom blur bond is not a very strong and rigid bond.

2. For the alkaline earths with two s electrons per atom, the conductivity per mole is similar or less. The pairing of the valence electrons increases the bonding (these elements are not as soft as the alkali metals), but does little to increase the electrical conductivity.

3. The third electron in the valence shell causes the electrical conductivity to disappear or greatly decrease in all periods except the third period. Here aluminum has an anomalously high conductivity. One can speculate that, except for aluminum, the s electrons are removed from the blur bonding and incorporated into sigma bonding through the formation of hybrid bonds.

4. In those periods where there are d-electrons or f-electrons involved, it appears that although these electrons may contribute some blur bond conduction, they also create sigma bonding, add rigidity to the lattice, and interfere with the blur bond electrons’ mobility.

5. Because of its delocalization of valence electrons over large numbers of atoms, the case of a metallic bond arises only in condensed phases. Gases can obviously not be metallic, whereas liquids and solids may be. The special case of carbon is a good example of two allotropic forms, one with and one without blur bonds. In graphite conditions favor three electrons in sigma bonds (per atom) and one electron donated to the blur bond in each graphite crystal. In diamond, under high T and P, all the electrons are involved in sigma bonding, leaving no electrons for conduction.

6. When the number of valence electrons reaches ten, a pseudo-inert gas configuration is reached. This results in little (but not zero) electronic conduction (Ni, Pd, Pt), because of the extensive sigma bonding. The next valence electron beyond ten increases the electronic conduction dramatically (Cu, Ag, Au). Presumably this electron is an s electron, the others having occupied d orbitals.

#### Part V: Conclusion

Because the presently used Lewis diagrams have their roots in static electron theories of the early twentieth century, I have proposed substituting a double-arrow symbolism for the customary double dots. This means that students should be



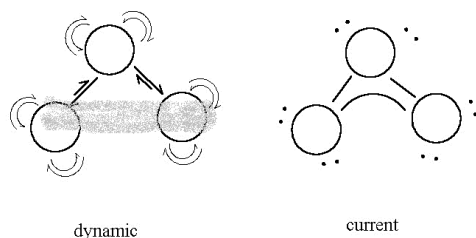


Figure 19. Ozone and the blur bond.

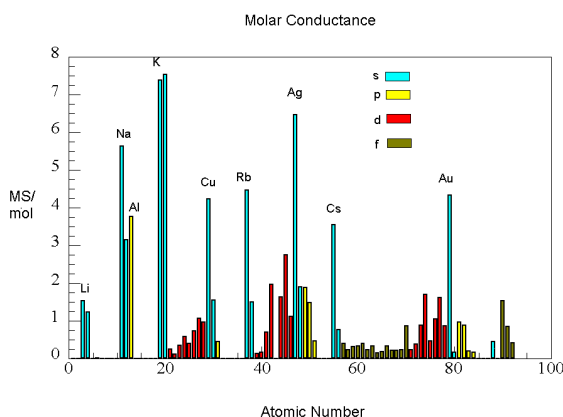


Figure 20. Molar conductance of the elements. Data taken from Emsley, J. *The Elements*; Clarendon Press: Oxford, 1989.

taught that a bond line (or “stick” as some call it) means a double arrow, not a double dot. Lone pairs are to be represented by double curved arrows, and later abbreviated to the old familiar double dot. There are obvious advantages to this new notation when atoms have different electronegativities. The double arrows can suggest the shifting of the bond-pair equilibrium toward the most electronegative atom. The new notation is more compatible with VSEPR theory, and seems to suggest the loci of largely imaginary MOs and AOs by the configuration of the arrows. With the help of some imagination the new symbolism can clarify resonance in aromatic molecules and conjugated chains. The most radical suggestion is that resonance forms be abandoned in favor of conjugated blur bonds. The blur bonds are not proposed as representations of molecular orbitals that will someday be calculated. They are, as Lewis structures have always been, an accounting procedure for the valence electrons. It is my hope that students will find them easier to envision than current Lewis structures. In my opinion the idea that resonance forms are to be “combined in your head into a form which we cannot draw” is an idea chemistry teachers can do without.

The concept of blur bonds has the merit of being easily extendable to the explanation of metal bonding and electron conduction. This blur bond explanation of metals is not meant to supplant the quantum mechanical picture of energy bands in solids. It is simply meant to be a more easily envisioned alternative to band theory, just as Lewis pictures are an alternative to molecular orbital theory.

A transistor is a small object through which one can pass a small electric current and then control this current either with another current or by applying a small gate voltage. It is now possible to arrange experiments in which single molecules can be made to act as transistors [21–23]. Faced with the passage

of an electric current through a molecule, should we not offer our students a more dynamic model than that implied by the stationary electron pair? I rather agree with Galileo, who is reported to have said, “And yet it does move.”

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