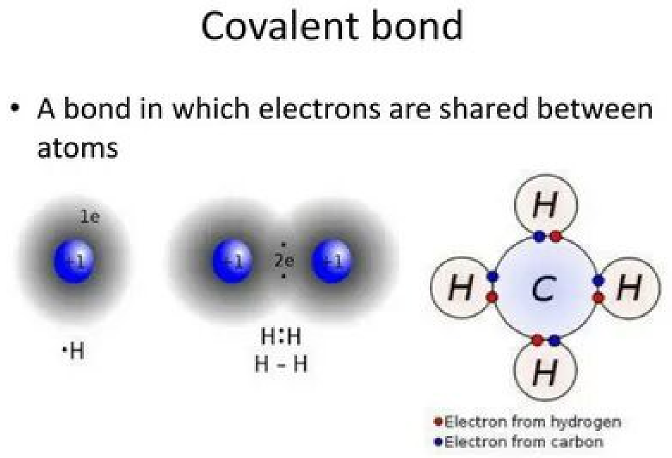
More on Molecular Bonding



Covalent bond

- A bond in which electrons are shared between atoms

by Miles Mathis

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I published a long paper on molecular bonding back in 2011, showing that electron bonding, both ionic and covalent, was a total myth. In the same month I proved the same for the Hydrogen bond. But I am returning to it now to hit it again, since it has admittedly gotten little overt traction in the mainstream. Many highschool and college professors in chemistry, physics, and engineering have written me, telling me they have given their students links to my papers, or even taught them in class—though I assume they have had to do that on the sly. I have even had a couple of science teachers in the army tell me the same thing, which was even more surprising. But for the most part mainstream chemistry is refusing to budge, despite the fact their field just took a huge leap forward out of the abyss in the last fifteen years. My nuclear diagrams and charge field channeling have completely overwritten physical chemistry, and all textbooks should be pulped and replaced.

I showed that again last month with my paper on Ethane, and there I admitted that I was now going through a college textbook on Organic Chemistry chapter by chapter and page by page, rewriting it myself. If no one wants to help me, I will do the entire book alone, as usual.

In that paper on Ethane we got through about p. 6, so I start here again with p. 7. To remind you, this is the 10th edition of Carey/Giuliano from 2017, a college chemistry book I happened to have on my shelf already. In sections on endothermic and exothermic reactions, the authors use NaCl as an example, admitting that

The energy liberated by adding an electron to chlorine is insufficient to override the energy to required to remove an electron from sodium. This analysis, however, fails to consider the force of attraction between the oppositely charged ions, as expressed in terms of the energy released in the formation of solid NaCl from the separated gas-phase ions. This *lattice energy* is 787kJ/mol and is more than sufficient to make the overall process exothermic.

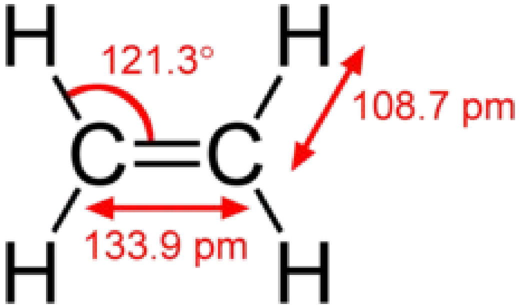
Wow, you may wish to reread that several times, until it dawns on you the enormity of that fudge. Because it begs so many questions, starting with: "If so, then how do the atoms in the gas become ionized to start with? Isn't the mainstream just assuming what it is expected to prove here, and therefore talking in circles?" And even if we are given the ions to start with, without expecting chlorine to ionize sodium directly, the authors have still switched horses midstream, telling you they are going to explain molecular bonding with ionization energies—hence the term ionic bonding—but then on the very next page dumping that and explaining them instead with lattice energies, which aren't really the same thing. Hence the new term. So why not call it lattice bonding instead of ionic bonding? And why would they introduce that important new energy, give it a name, then dump it immediately? We learn absolutely nothing more about lattice energy in the following pages. It is a very strange progression of argument, if nothing else, and should throw up a red flag to even to the most casual reader.

And that lattice energy fudge gets bigger the longer and harder you look at it. If we go Wiki or any other textbook for more on lattice energy, we find that it doesn't have much to do with the problem at hand, of normal atoms bonding under the conditions they actually bond in. Lattice energy is mainly a theoretical value, being the energy thought to be released when a gas at 0 K becomes a crystalline compound or solid, for whatever reason. It can't even be measured, for pretty obvious reasons, and has to be calculated from the Born-Haber cycle. But of course our Na and Cl didn't start out at 0 K, and they probably didn't start out as a gas. Salt is formed in the crust from solid or liquid atoms. So bringing up the lattice energy here is a huge cheat. It is a bit shocking a college textbook would bring it up in this fashion.

But as I showed you in that older paper, they have to rush you through this, and can't have any rigor, since none of this makes sense from the first word. At a stretch, we might be able to understand how an electron might be blown off a sodium nucleus, given an energetic field, but why would that same field (the one in the gas) cause chlorine to take an extra electron it didn't previously want? The chlorine atom was already charge balanced, so what exactly would cause it to take on a new electron and *release* energy? That is not only upside down to entropy and all other expectation, it is upside down to the definition of charge. The electron is already bound to the nucleus by the same opposite charges that supposedly create this lattice energy and all molecular bonds, remember, so how does the nucleus suddenly attract another electron to suit this whacky theory? Wouldn't it have to add another proton first, that would then attract that electron? The mainstream has never explained how or why the same field that is knocking electrons off sodium would be gluing them on chlorine, or what that glue consists of. It can't be electromagnetic, because by their own previous statements in this chapter, the electromagnetic force requires opposite charges. Without an increase in positive charge, chlorine cannot attract that extra electron, so we need to be told what is attracting it if not positive charge. You see how this is all just a push. There is no mechanics, it is just a bunch of words—words which, when you read them closely, all collapse into nonsense.

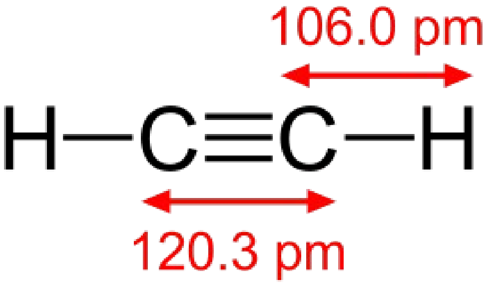
We see that again on p. 10, where we are looking at the Lewis structure of Ethylene. We are told the two Carbons double bond to satisfy the Octet rule. In Acetylene they triple bond for the same reason. But wait, that doesn't explain anything and is again circular. Nothing is mechanical there, and you haven't been told how or why any of these bonds are occurring. In answer, you are told Carbon wants to fill orbitals to be more like Neon, but that is *ad hoc* or *post hoc*. They are just cataloging what appears to happen and then making up a rule afterwards, but the rule doesn't explain anything, it just names it. Supposing Ethylene and Acetylene really do have stronger bonds there than normal, we should be told why they can't just leave the extra electrons hanging. If they do that, they certainly don't do it to match some non-mechanical *post hoc* rule.

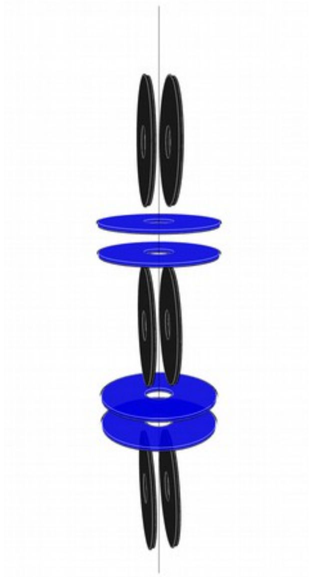
I have already done Acetylene, so let's take a quick look at Ethylene.



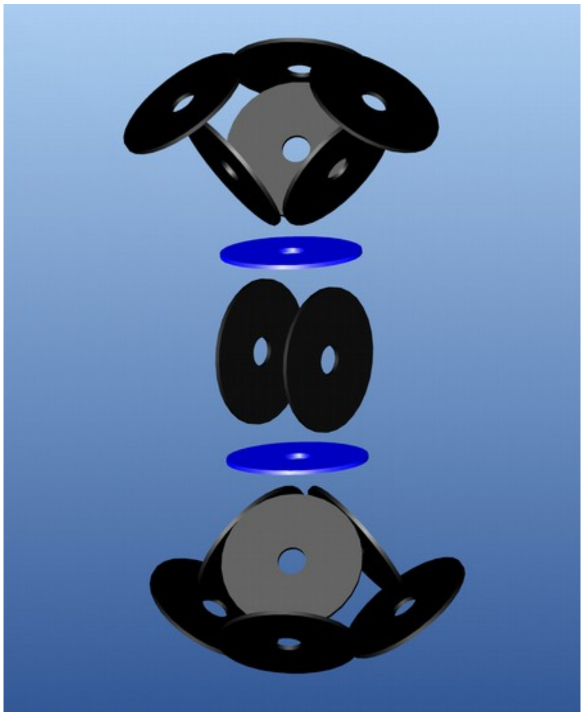
To start with, that is confusing because they draw these Lewis structures in the horizontal, when they should be drawn in the vertical. Then you might be able to make sense of them. The Lewis structures make you think the Carbon-Carbon bond is at nearly right angles to the Hydrogen bonds, when in fact they are all basically on the nuclear pole. By that diagram you would think the Carbons are bonding somehow on the nuclear equator, when they aren't.

To see what I mean, compare the Lewis structure of Acetylene to my nuclear diagram:



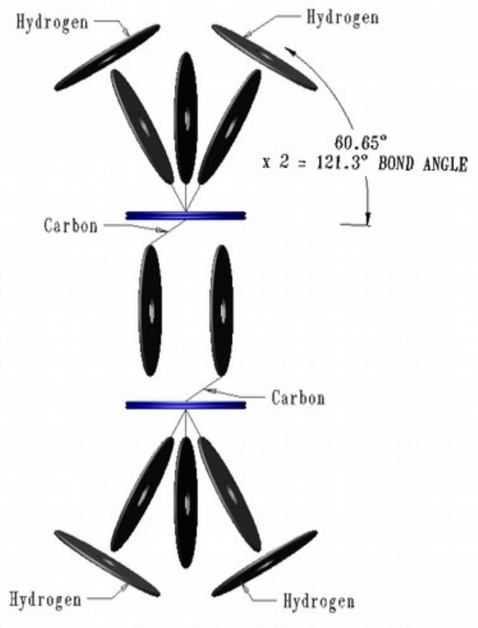


From my diagram you can see that all bonds are on the nuclear poles. So to make sense of Ethylene, we flip it vertical like Ethane:



Then just subtract one of the legs top and bottom to get Ethylene. From that, we again see all the bonds are on the nuclear poles, and the hydrogens then spread out because they are pushing one another apart with their charge emission coming out their equators. All mechanical, all visualizable.

John Wiltbank came through for us again:



Ethylene

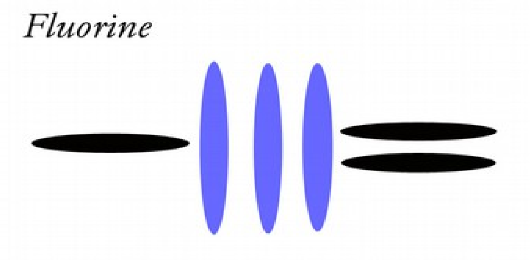
Also click here to see a moving animation of Ethylene.

From that, we can also show why Ethylene has a shorter and stronger bond than Ethane, without any talk of these ridiculous hybridized orbitals. Both have double bonds, in a sense, because both are caused by the alignment of two protons, as you see there. But Ethane, being larger, has to recycle the same ambient field through more atoms, so we would NOT expect the same amount of charge to pass through those central bonds in the same amount of time. You can intuit that difference in charge flux just from the architecture. You might think at first that Ethane could recycle more charge, giving it the stronger bond, and as a matter of maximum charge potential, it can. In other words, if you keep adding charge to Ethane, it can take more than Ethylene. But that isn't what bond strength means in this case. Bond strength means how strong the bond is in *the same* ambient field, so you have to add the same amount of energy to both molecules, you see. In that case, Ethane will have less charge strength per atom, simply because it has more atoms.

Ethylene actually has 87% of the bond length of Ethane, and 1.93x the bond strength, so it is not really a double bond. And the "triple bond" of Acetylene is 2.56x, so even less a triple bond. A miss of 17% from triple. But can I calculate those fractions straight from the proton configurations and angles, with no talk of electrons, getting a better result than the mainstream? Of course I can. Anybody could, but nobody does, don't ask me why. Acetylene has 1/3rd the Hydrogens of Ethane. But since it has the same number of Carbons, only about half that differential can be expressed. So, 3/2 = 1.5. But the bond angle of Ethane is 111 while the bond angle of Acetylene is 180. 180/111 = 1.62. 1.5 x 1.62 = 2.43. A miss of 5% from data instead of 17% from expectation. With just two calculations I am able to beat the mainstream, so I hardly need to keep going. And my calculations are based on diagrammable mechanics, while they aren't even able to calculate from data to data like this.

Next we move to the mainstream explanation of electronegativity, which doesn't exist. They can tell you measured values, but not where they come from. For instance, electronegativity generally goes lower as you move left on the Periodic table. But Hydrogen has a value of 2.1, way above Lithium and even above Boron. It is nearly as high as Carbon. This makes no sense in ionic or covalent bonds, since electronegativity is an atom's tendency to attract electrons. Fluorine is the most electronegative, being on the far right side of the Periodic table, while Hydrogen is on the far left side. Fluorine is always tagged as a negative free ion, while Hydrogen is normally tagged as a positive free ion. When it is ionized, it loses its electron, becoming net positive. So how can it also be electronegative in the sense that Fluorine is? Well, basically because atoms can do whatever these nutty theorists want them to, attracting or donating electrons as the need arises, with some fudge to explain it afterwards.

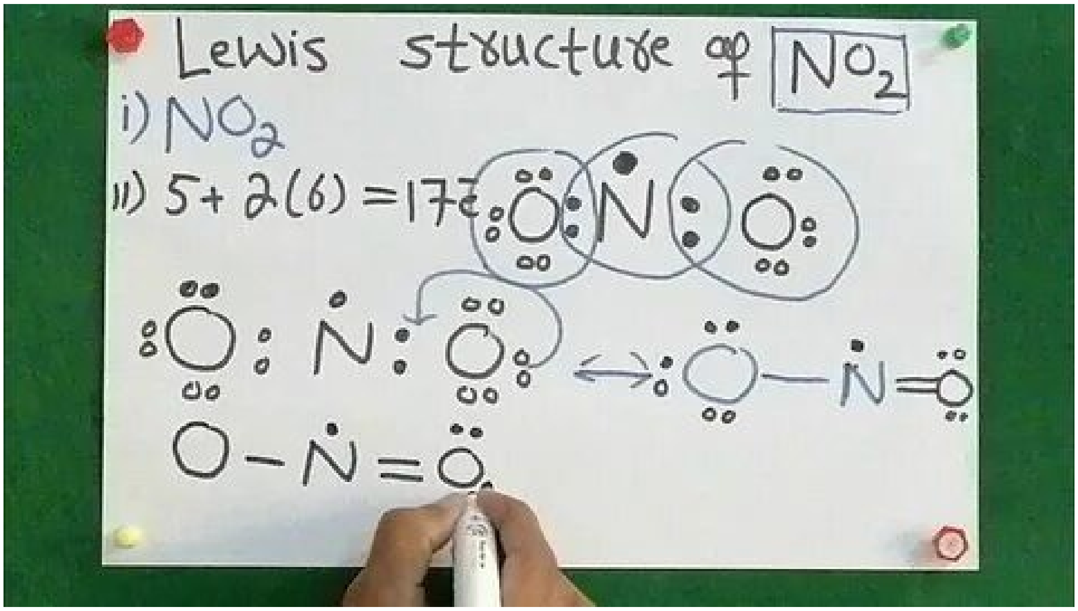
In my charge field, Hydrogen is neither positive nor negative, because none of this is based on electrons or opposite charges. It is based on the charge field, so it is still electromagnetic in a sense, but it isn't based on these manufactured potentials. Hydrogen is universally reactive, especially as an ion, simply because it fits so easily into so many nuclear channels. It would, wouldn't it, since it is basically a free proton—nuclear charge channels being defined by the proton. The only reason a few atoms are more reactive is that they **focus** their charge field better than Hydrogen, as we have seen in many previous papers. The nuclear configuration of Fluorine not only boosts the channel on the pole using the three alphas in the core, it focuses it by having one proton on one end and two on the other, so the charge field knows which way to go, as it were.



Charge is moving both ways through all atoms and molecules, but in Fluorine there is less opposing traffic, due to that configuration, so charge can escape that through opposing vortex in a pre-existing channel, while having less resistance.

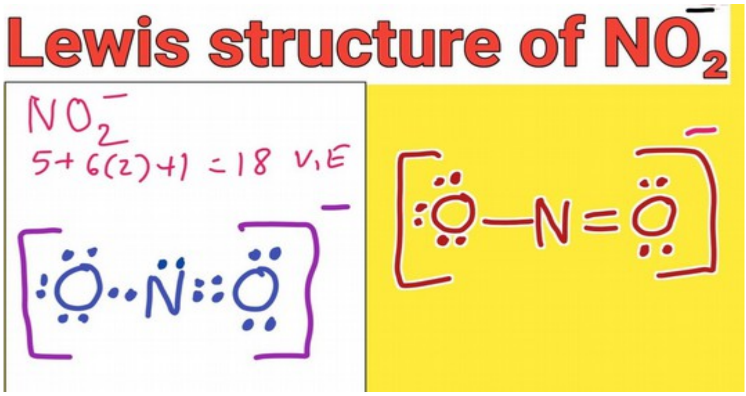
What this tells you is that the mainstream had it upside down from the beginning, as usual. It is not that elements want to accept or donate electrons, it is that they want to accept or donate protons on the nuclear poles that we see in my diagrams. And strictly it isn't even that, since those protons aren't donated or shared. The nuclear diagrams simply allow us to see where the charge channels are, and where the negative and positive plugs are. Those protons are the positive plugs in my diagrams, and the holes they fit into are the negative plugs or sockets. For instance, in the Fluorine above, a second proton can be accepted to your left, fitting next to the existing proton and filling that polar slot. That is where the H goes in HF. Of course it doesn't fit in tight, like the existing proton there, since that would give us variant of Neon (and overload the core with charge, which is why it doesn't exist); but it does exist backed out some amount, giving us a molecular bond rather than a nuclear bond. It follows the existing nuclear charge channel, and the distance it is backed out is the actual bond length.

On p. 13 we move on to formal charges, which again show what a horrible fudge covalent bonding is. They show it with nitromethane, but it is easiest to see with NO2 alone. This is a very simple molecule, but nothing is a bigger mess than the Lewis structure of NO2, even after almost 110 years. To show this, let's look up it up online. One of the first diagrams that comes up is this:



That's from a youtube video. But that's not even right, and doesn't match other mainstream diagrams, as you will see. They are all over the map. According to that diagram, Nitrogen is sharing two of its five, and oxygen is sharing two of its eight. Except that Oxygen only has six valence electrons. The other two are in the 1s and shouldn't be drawn here. That is exactly why he has the number 17 right before that. Nitrogen has five and each Oxygen has six, making 17 available for bonding. Oxygen isn't sharing any of its six in that diagram, so no bond can be created there. To say it another way, none of Oxygen's valence electrons are drawn there on the Nitrogen side, so no bond is drawn or implied in that diagram. It is a hash. He has those overlapping circles to indicate the bond, but that is just a further source of confusion. Purposeful confusion. As you see, he does that so that in the next step he can draw the bond in that position between O and N, where O has no valence electrons. That bond then becomes Oxygen's mysterious seventh valence electron, appearing out of the void with a wave of the wand to give Lewis what he needs to fill octets. A truly criminal cheat, still done right before your eyes at Youtube in many mainstream videos. Prestidigitation.

If you go to Wiki to see what they say about this on the page for Nitrite or Nitrogen Dioxide, you find they conveniently don't publish a Lewis diagram for it. Curious.



This one also makes no sense, although at least they don't draw Oxygen with eight valence electrons. But they do draw the left one with seven, which they claim is what gives Nitrite its net negative charge. But the question remains: where did that extra electron come from? Not from Nitrogen, and not from Oxygen itself, since Oxygen can't pull up an electron from its 1s level just to suit a Lewis diagram. And there is no reason for a free electron to arrive from nowhere here, just to make things work out. The truth is, they need it, so that when they draw it the second way there in yellow, it seems to create the shared bond with Nitrogen. Otherwise there would be no sharing and no possible bond. It also fills in Nitrogen's octet, as you see, but only by an obvious fudge. When you first look at that, you should go "wait, none of Oxygen's six valence electrons are being shared, so what is creating that bond?" Well, nothing is, which is why they have to back-manufacture a seventh electron from the void.

They do admit that extra electron should give that Oxygen a net negative charge, or a **formal charge**, so in many diagrams, like the one in my book, they give the Nitrogen a positive formal charge. But Nitrogen would only have a positive formal charge—of the same sort the Oxygen has—if it had one too few electrons. The Oxygen has a negative formal charge because it has one too many, so Nitrogen should have one too few. But it doesn't. In the diagram of nitromethane, it still has five valence electrons, so we are not told why it would have a net positive charge. Yes, it is only bonded through four, with one hanging, but that shouldn't give it a positive charge. Unbonded electrons don't lose their negative charge. In truth they do that only to balance that electron they fudged in from the void in Oxygen, but it makes no sense. If Nitrogen does have a net positive charge there, they would have to explain it in another way.

In the book, they put it this way:

Formal charges correspond to the difference between the number of valence electrons in the neutral free atom and the number of valence electrons in its bonded state. To determine the electron count of an atom in a Lewis formula, we add the total number of electrons in unshared pairs to the one-half the number of electrons in the bonded pairs.

Meaning, we just fudge everything after the fact to fit these Lewis structures, ignoring the number of valence electrons. As we just saw with NO2, they can create any bond they need to match data, then add an electron to the Lewis structure from half that bond, with no explanation necessary. It's a "formal" charge, so you can't ask any questions. It has a name and is in a textbook, so it must be science!

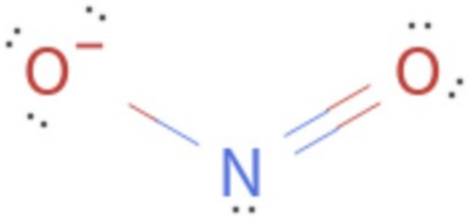
Another problem with NO2 is that although the bonds are drawn as one single and one double, they admit they are about the same length. That makes no sense. As is the fact the negative charge is equally distributed on the two Oxygens. If they have different bonds they should have different lengths and charges. Clearly, they make one double and one single simply to fit the five valence electrons of Nitrogen. Nitrogen needs two to bond out of NO2 later, leaving three, so two have to go to one O and one to other. Plus, in this textbook the charge is given to one Oxygen and not the other, which makes a hash of the Lewis diagram in just one more way—not to mention contradicting Wiki, which says it is split equally.

Can the mainstream explain why Nitrite is angled, and the angle 115? Of course not. The best I could find is the lone pair theory:

The nitrite ion (NO_2^-) has an angular (bent) molecular geometry due to the presence of a lone pair of electrons on both the nitrogen and oxygen atoms. This results in a bond angle of

approximately 115°. Lone pairs affect the bond angle by reducing it. This is because lone pairs occupy more space than bonding pairs and cause greater repulsion, which leads to deviations in molecular geometry. 15 **For example, in water (H2O), the two lone pairs on the oxygen atom repel the bonding pairs, reducing the bond angle to about 104.5°.**

Pathetic as usual



What lone pairs are they talking about? I see a lone pair on N, but five on the two O's. In fact, the left Oxygen is nothing but lone pairs and a fake bond manufactured from air. Do lone pairs occupy more space, or is that the usual fudge? There is actually no data on lone pairs occupying more space, and there couldn't be because this is all garbage. And if they did, we can ask, "exactly HOW do these lone pairs repel bonding pairs or anything else? By what mechanism? Electrons communicate with other electrons how? What is the field there?" Besides, by any logical analysis, lone pairs maximizing repulsion would not cause a bend, but the opposite. The greatest separation of electrons in that molecule would be achieved by a straight line. Doesn't take much perspicuity to see that, since the two Oxygens can obviously achieve more separation than we see there. The further they move away from one another the closer they are to a straight line.

But I can explain the bend, and have done so before in similar situations. In a unidirectional charge field, that molecule would be straight, but we live in a split field. Meaning, we have both charge and anticharge. Charge is moving both north and south through all molecules, up and down the pole. Again, we should turn that last diagram vertical to make better sense of it, as I showed above. In our local field in the Solar System, we have about twice as much charge as anticharge, or twice as many photons as antiphotons. So the field is very skewed as a matter of spin mechanics. Not only are photons moving the opposite direction of antiphotons through the nucleus, they are spinning opposite. The Nitrogen at the center of this is also spinning about its equator in one direction and not the other, which skews everything again. So basically, opposite charge is releasing into the same ambient field north and south, forcing it left from above and right from below, or looking at both forces from the same vantage, left in diagram from both poles. That acts to fold the molecule, as hopefully you can now see.

It has nothing to do with lone pairs, and actually has nothing to do with electrons at all. And again, you can calculate the bond angle of 115 straight from protons, with no talk of electron bonding. 70% of the molecule's weight is in the Oxygens, but only half that differential can be expressed since we are in a split field. You will say the field isn't split equally, as I just admitted, but that isn't the question here. It only matters how many ways the field is split, and a split gives us the number 2 regardless. Even though more than half the field gets released north, less than half gets released south. So yes, there is more bending from one end. But in a case like this, there is no easy way to tell that. It doesn't change the *amount* of bend, it just changes how much comes from each end. So, we are looking at a 35% bend here. Half of 70. Let's compare that to the known angle. Well, 180-115=65, and 65/180 = .36 or 36%.

Hello, goodbye, shut the door on your way out.

[Notice this also explains the different bonds to the two Oxygens in NO2, without having to draw them as single or double electron bonds. Because the local charge field is split, and is twice as strong in one direction, it will naturally create a stronger bond on one end of this molecule. Remember, all charge isn't moving polar, since both Oxygen and Nitrogen also recycle charge out sideways or equatorially, via their core alphas. This will have the effect of lopsiding the molecule in just the way we see here. More charge will be passing through one end than the other, creating different bonds.]

Why can't the mainstream do these simple calculations? Because they aren't looking at these problems the way I am. It would never occur to them to run the equations I do, in this problem or any other, because they aren't dealing with these as charge field problems. We have seen that over and over, in my solutions to Bode, to planetary tilts, to Core theory, to multi-body problems, and to unified field theory. I am able to solve on post-it notes what they have not been able to solve on supercomputers. The supercomputers cannot solve it because they don't have the proper input. It is that simple. They tell you the universe is chaotic, but it is only their equations and theories that are chaotic. They are chaotic because they are wrong. But instead of admitting that and trying to correct them, as I have, they assume they are right and start pushing them, tagging the amount of push as chaos. It is a science of madmen, frankly.

Nothing shows this more clearly than electron bonding, and particularly these awful Lewis structures. The fact this stuff was set in stone 110 years ago is among the greatest tragedies of science, and surely the greatest tragedy of chemistry. But this is how it has gone, especially in the 20th century, as these people like Lewis bungled their way through a problem like this and then were somehow able to almost immediately cement it in as dogma. What is even more shocking is that five generations of chemists and physicists have bought it without a whimper. Literally millions of students worldwide, and apparently not one before me saw that this was no more than a massive ugly fudge. One generation dutifully taught it and the next dutifully learned it, without ever trying to make sense of it, or trying to improve it. And I guess I can see why, since now that I have, I have gotten little but grief for it. You don't get ahead by questioning anything, or correcting it, you get ahead by continuing to promote your semi-retarded ancestors and their Nobel Prizes of dross.

If you are just getting here, you may ask how I explain more difficult problems without electron orbitals. Well, you can consult many papers on that, including my paper on Balmer/Rydberg series, the Stark Effect, the Hall Effect, Compton Scattering, NMR, and many others.