Even though the bonding of hydrogen and carbon atoms can generate a remarkable array of molecules, the hydrocarbons are really rather boring (chemically). They take part in a rather limited range of reactions and would not, on their own, be expected to produce anything like life. Of course there are many other elements, and their properties add chemical complexity to molecular behavior. From the perspective of living systems two of the most interesting elements are nitrogen and oxygen. Carbon has six electrons (two core and four valence) and can form four bonds with neighboring atoms. Nitrogen has seven electrons: two core and five valence: 1s², 2s², 2px¹, 2py¹, 2pz¹. So if you are following the rules, you might well assume that nitrogen would be able to form five bonds (after all, it has five valence electrons). But when we look carefully, we never see a nitrogen atom making five bonds, and in all stable compounds it makes only three bonds. We can explain this observation in several ways. One factor is that nitrogen atoms are too small to support five centers of electron density around themselves because the bonds begin to overlap, which is destabilizing, just like we saw with bulky groups around a carbon. Another factor is that there are only four orbitals available in nitrogen in the second quantum shell. If nitrogen were to form five bonds it would have to use orbitals from the next quantum shell (3), but these orbitals are so high in energy that the energy required would not be offset by the energy released upon on bond formation. Together these factors mean that nitrogen, and in fact all elements in the second row of the periodic table, are limited to bonding arrangements with no more than four centers of electron density. As we will see later on, elements in the next row, such as phosphorus (P) and sulfur (S), are larger and have more available orbitals for bonding. These elements can form up to six centers of electron density.

The simplest compound of nitrogen is molecular nitrogen, N₂. The two nitrogen atoms are bonded together by a triple bond, consisting of a σ and two π bonds. Molecular nitrogen, N₂ is a stable (relatively nonreactive) molecular compound, A common nitrogen-containing molecule is ammonia (NH₃), which is analogous to methane (CH₄). In ammonia the nitrogen atom is bonded to three hydrogen atoms. These three bonds involve three of nitrogen’s valence electrons; the remaining two valence electrons occupy a non-bonding orbital and are referred to as a lone pair. Given the molecular hybridization orbital model that we are using this implies that four sp³ orbitals are formed from the nitrogen atom’s 2s and 2p orbitals leading to four electron density centers around the nitrogen. The figure shows three representations of ammonia. The first indicates the N–H bonds but fails to show the lone pair orbital. The second uses the wedge and dash convention and dots to illustrate the geometry of both bonds and the lone pair. The actual shape of the molecule is determined by the arrangements of electron clouds and the bonded atoms. In NH₃ all three bonds are equivalent (N–H) and so must be symmetrical, but the lone pair orbital is different because it takes up more space than bonding pairs, can you imagine why? This has a subtle effect on the shape of the molecule. The angles between the C–H bonds in CH₄ are equal and 109° while the angles between the N–H bonds in NH₃ are slightly smaller, 107.8°. The shape of the molecule itself (as outlined by the atoms) is a triangle-based pyramid rather than a tetrahedron. Finally the Lewis structure (the most abstract representation), indicates the bonds and lone pair electrons but gives an unrealistic depiction of the molecule’s geometry. It is up to the reader to supply the implicit information contained in the structure about bond angles and overall shape.

References

76 However, a nitrogen compound with some structural similarities to diamond has been identified. It was synthesized from N₂ at high pressure and temperatures. In this polymeric nitrogen, each nitrogen is connected to three neighbors via single bonds, in a similar way that diamond has carbons connected to four neighbors. However this polymeric nitrogen is highly unstable and reactive – unlike diamond. [http://www.nature.com/nmat/journal/v3/n8/abs/nmat1146.html](http://www.nature.com/nmat/journal/v3/n8/abs/nmat1146.html)