Lewis Dot Structure of Hydrogen Fluoride.

![Lewis Dot Structure](image)

**Drawing Lewis Structures**

1. Sum the *valence electrons* from all atoms in the species.
2. Write the *atomic symbols* for the atoms involved so as to show which atoms are connected to which. Draw a *single bond* between each pair of bonded atoms.
3. Complete the *octets* of the atoms bonded to the central atom (i.e. the peripheral atoms).
4. Place *leftover electrons* on the central atom, even if it results in the central atom having more than an octet.
5. If there are not enough electrons to give the central atom an octet, *form multiple bonds* by pulling terminal electrons from a peripheral atom and placing them into the bond with the central atom.

**Question:**

1. **Draw the Lewis structure for ammonia, NH₃.**

**Solution:** Since each H can form only one covalent bond, the arrangement of atoms must be:

![Lewis Structure for NH₃](image)

From the periodic table we see that N has five valence electrons. These, plus one electron from each H, give us a total of eight. Bonding the atoms in the molecule requires the use of six valence electrons, as...
The remaining two valence electrons are then assigned to N to complete its octet:

$$3 \text{H}^\cdot + \text{N}^\cdot \rightarrow \text{H} : \text{N} : \text{H}$$

Ammonia (NH$_3$)

2. Write the Lewis structure for ethane, C$_2$H$_6$.

Solution: A little thought will reveal that the two C atoms must be bonded to each other. (Remember: Hydrogen forms only one bond.) Keeping in mind the octet rule, we predict that besides bonding, to the other C atom, each C forms three bonds to H atoms:

```
\[
\begin{array}{c}
\text{H} \\
\text{H} \\
\text{C} \\
\text{C} \\
\text{H} \\
\text{H}
\end{array}
\]
```

Each C contributes four valence electrons. Bonding all atoms in the molecule with pairs of electrons uses all of these. The Lewis structure is thus

```
\[
\begin{array}{c}
\text{H} \\
\text{C} : \text{C} : \text{H} \\
\text{H} \\
\text{H}
\end{array}
\]
```

3. Write the Lewis structure for ethylene, C$_2$H$_4$

Solution: Here, as in the previous example, the two C atoms must be bonded to each other.

```
\[
\begin{array}{c}
\text{H} \\
\text{C} \\
\text{C} \\
\text{H} \\
\text{H}
\end{array}
\]
This total number of valence electrons is $2(4) + 4 = 12$. Bonding all atoms in the molecule uses 10 valence electrons:

$$\begin{align*}
\text{H} & : \text{C} & : & \text{C} : \text{H} \\
\text{H} & : & \text{C} & : & \text{C} : \text{H}
\end{align*}$$

Where shall we put the last two electrons? If we put them in the valence shell of one C then the octet rule will be violated for the other.

However, placing them between the two C atoms solves the problem.

$$\begin{align*}
\text{H} & : \text{C} & : & \text{C} : \text{H} \\
\text{H} & : & \text{C} & :: & \text{C} : \text{H}
\end{align*}$$

The two carbons are thus shown to be bonded via a double bond, two shared pairs, or four bonding electrons.

Note that in the Lewis structure for ethylene (in above example) the four electrons of the double bond are counted with each carbon.

$$\begin{align*}
\text{H} & : \text{C} & :: & \text{C} : \text{H} \\
\text{H} & : & \text{C} & :: & \text{C} : \text{H}
\end{align*}$$

Thus the octet rule is obeyed for each carbon. The Lewis structure for ethylene may alternatively be shown as

$$\begin{align*}
\text{H} & : \text{C} & :: & \text{C} : \text{H} \\
\text{H} & : & \text{C} & :: & \text{C} : \text{H}
\end{align*}$$

or

$$\begin{align*}
\text{H} & : \text{C} & :: & \text{C} : \text{H} \\
\text{H} & : & \text{C} & :: & \text{C} : \text{H}
\end{align*}$$

or even

$$\begin{align*}
\text{H} & : \text{C} & :: & \text{C} : \text{H}
\end{align*}$$

Each of these is equivalent to the version shown in above example. Lewis structures are not generally expected to represent the geometrical shape of a molecule, one reason being that they must be drawn in a plane, while the molecule itself may not be planar. A Lewis structure is expected to show (a) which atoms in the molecule are bonded together and
(b) how the valence electrons are assigned in the molecule, that is, which pairs are bonding pairs and which are lone pairs.

4. Write the Lewis structure for the ammonium ion NH$_4^+$

**Solution:** The arrangement of the atoms is

\[ \text{H} \quad \text{H} \quad \text{N} \quad \text{H} \quad \text{H} \]

The total number of valence electrons is \((5 + 4) - 1 = 8\). (Here we have subtracted one electron from the total provided by one N and four H atoms, because the ion has a positive charge and therefore has one less electron than a (neutral) molecule.

**The Lewis structure is**

\[ \text{H} \quad \overset{\text{N}}{\text{H}} \quad \overset{\text{H}}{\text{H}} \]

In the electron counting process one electron of a bonding pair often appears to have come from one of the two bonded atoms, and the other electron from the other atom. Thus we have for the formation of the three covalent bonds in NH$_3$.

\[ \overset{\text{H}}{\text{H}} \quad \overset{\text{N}}{\overset{\text{H}}{\text{H}}} \quad \overset{\text{H}}{\text{H}} \]

or

\[ \overset{\text{H}}{\text{H}} \quad \overset{\text{N}}{\overset{\text{H}}{\text{H}}} \quad \overset{\text{H}}{\text{H}} \]
To keep track of electrons origins small x’s and o’s are sometimes used in Lewis structures. Writing x’s for the electrons from the H atoms and o’s for those from N, we get for NH₃

\[
\begin{array}{c}
\text{H} \\
\text{N} \\
\text{H} \\
\text{H}
\end{array}
\]

A covalent bond in which each electron of the pair appears to have come from each bonded atom is called a normal covalent bond.

In the ammonium ion, again using the x and o symbolism, we can write

\[
\left(\begin{array}{c}
\text{H} \\
\text{H} \\
\text{N} \\
\text{H} \\
\text{H}
\end{array}\right)^+
\]

in this Lewis structure one of the H atoms appears to be bonded to the N by a pair of electrons, both of which originated with the N atom. Such a bond is called a coordinate covalent bond, or sometimes, a dative bond. But all four N—H bonds in \(\text{NH}_4^+\) are identical in all measurable properties. So a coordinate covalent bond is in no way different from a normal covalent bond. There appears to be a difference only when we keep track of the electrons’ origins.

5. Draw the Lewis structure for PCl₃.

Answer:
Step 1: **Sum the valence electrons.** P has 5 and each Cl has 7 for a total of \[5 + (3 \times 7)\] = 26 *valence electrons*

Step 2: Arrange atoms showing connectivity and **draw a single bond between atoms.** NOTE: In a binary (two-element) compound, the first element listed is usually the central one with the others surrounding it.

Step 3: **Complete the octets** on the atoms bonded to the central atom. NOTE: This accounts for 24 of the 26 valence electrons.

Step 4: Place the remaining electrons on the central atom to complete the octet. Since this gives an octet to each atom we are finished.

6. **Draw the Lewis structure for HCN.**

   Answer:
   
   Step 1: Sum the valence electrons. H has 1, C has 4 and N has 5 for a total of \[1 + 4 + 5\] = 10 *valence electrons*
   
   Step 2: Arrange atoms showing connectivity and **draw a single bond between atoms.**

   **NOTE:** Since H can only form one covalent bond it can never be the central atom. The choices are HCN or HNC. Formula is written HCN!!!

   This accounts for 4 *valence electrons*

   Step 3: Complete the octets on the atoms.
bonded to the central atom.  

*BUT:* There are only 6 valence electrons left.  

If we put them on N we do not achieve an octet at C !!  

Step 4: Try using multiple bonding to share the electrons between C and N. A triple bond is required to give an octet to each atom

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**The octet rule and Lewis structures**

Whatever possible, Lewis structures should show the octet rule to be obeyed. There are some molecules, however, in which the octet rule is clearly violated.

Consider the molecule of phosphorus pentachloride, PC15.  

In this molecule phosphorus atom is bonded covalently to five chlorine atoms. The total number of valence electrons is 40 (5 from the P plus 35 from the five Cl atoms). Since the P forms five bonds, the Lewis structure is

![Lewis structure of PC15](image)

Here the valence shell of the phosphorus atom is said to have been expanded in order to accommodate five electron pairs. The expansion of the valence shell of an atom is possible only if the atom has nd or \( n - 1 \) d orbitals which can be added to the ns and three np orbitals normally constituting its valence shell. In the case of PCL5 the 10 bonding electrons are accommodated in the valence shell of phosphorus which has been expanded by the addition of one of phosphorus 3d orbitals. The valence shells of
atoms of periods 1 and 2 cannot be expanded because they contain no 1d or 2d orbitals. (The 3d orbital is unavailable for these atoms because it is of such high energy.)

Sometimes the valence shell of an atom in a molecule contains less than an octet. This is the case with boron trifluoride, $\text{BF}_3$. Its Lewis structure is written as

\[
\begin{array}{c}
\text{F} \\
\text{B} \\
\text{F}
\end{array}
\]

Here the valence shell of boron holds only three pairs of electrons, and again the octet rule is violated.

The octet rule is a handy generalization, but exceptions to it are numerous. It must be violated in molecules having an odd number of valence electrons. The Lewis structure for nitric oxide, $\text{NO}$, can be shown as

\[
\begin{array}{c}
\text{N} \\
\text{O}
\end{array}
\]

**Bond polarity**

Identical atoms have identical electronegativities. In the $\text{H}_2$ molecule

\[
\text{H} : \text{H}
\]

the hydrogen atoms attract the electron pair equally. The electronic charge distribution is symmetrical with respect to the two nuclei; that is, it is not pulled closer to one atom than the other. Since one end of the bond is electrostatically just like the other, the bond is said to be nonpolar. (This just means that it does not have different poles, or ends.) For the same reason the bond in
the fluorine molecule is also nonpolar. Atoms with identical electronegativeties form nonpolar covalent bonds.

\[ \cdot \cdot \cdot F : F : \cdot \cdot \cdot \]

Atoms of different elements have different electronegativities. In the hydrogen fluoride molecule.

\[ \cdot \cdot \cdot H : F : \cdot \cdot \cdot \]

because the F atom has a higher electronegativity than the H atom, the electron pair is shared unequally. The electronic charge cloud of the shared pair is pulled closer to the F atom. The resulting bond has negative charge piled up at one end, leaving positive charge at the other. A covalent bond in which the electron pair is shared unequally is said to be a polar covalent bond.

The polarity of a bond, that is, the degree to which an electron pair is unequally shared, depends on the difference between electronegativities of the two bonded atoms. The greater the electronegativity difference, the more polar is the bond.

**Partial ionic character:**

When two atoms of greatly differing electronegativity bond, the result is best classed as an ionic bond. Consider the bond formed between hypothetical atoms A and B. Imagine that B has a high electronegativity. Imagine further that we can alter the electronegativity of A so that it is exactly the same as that of B. Neglecting all electrons but the shared pair, we can schematically show that the bond is
nonpola\(s\) by writing two dots exactly in the middle between the symbols \(A\) and \(B\).

\[
A : B \quad \text{(nonpolar covalent bond)}
\]

Now we turn the dial on the remote controller and gradually decrease the electronegativity of \(A\). We make \(A\) less electronegative, or more electropositive. This decrease is pulling on the pair of electrons, so that their average position moves closer to \(B\) creating partial charges, \(\delta^+\) and \(\delta^-\), on \(A\) and \(B\), respectively:

\[
A^{\delta^+} : B^{\delta^-} \quad \text{(polar covalent bond)}
\]

The bond has now become polar. It becomes increasingly polar as we further decrease the electronegativity of \(A\), until finally the probability of finding the electron pair on \(A\) is very low and on \(B\) very high. The electron pair now largely "belongs" to \(B\). This gives \(B\) a net negative charge and leaves \(A\) with a positive charge, for \(A\) has now transferred an electron to \(B\):

\[
A^+ [:B][-] \quad \text{(ionic bond)}
\]

The ionic bond can be seen to be an extremely polar bond, one in which there is essentially no sharing of electrons.

**Pauling electronegativities**:

The concept of electronegativity was originally proposed in 1932 by the American chemist Linus Pauling. He pointed out that the distribution of the electronic charge cloud of a bonding pair of electrons should be related to the strength of the bond. A bond which is highly polar (has a high
degree of ionic character) should be very strong, as the attraction between the partial negative charge built up on one atom and the partial positive charge left on the other should augment the bond strength. Using measured values of bond energies (bond energy is the energy necessary to break a bond), Pauling devised a set of electronegativity values for most of the elements.

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Numerical values of electronegativities are useful for estimating the polarity or degree of ionic character of a bond. The dividing line between predominately ionic and predominately covalent character works out to be an electronegativity difference of about 1.7.

This fact is sometimes useful for deciding whether to write a covalent Lewis structure or an ionic one. For HCl, for example, we
might write - the ionic structure $\text{H}^+ \left[ \text{:Cl:} \right]^-$ or the covalent structure $\text{H:}:\text{Cl:}$

Since the difference between the electronegativities of Cl and H is $3.0 - 2.1 = 0.9$, the bond is clearly more covalent than ionic, and so the second Lewis structure is appropriate. Numerical values of electronegativities should be used with some caution, however. The values given in the figure are best considered as approximate, because the effective electronegativity of an atom tends to vary somewhat according to the atom's environment.

The variation of electronegativity with atomic number shows the expected periodicity. Note that the effects of nuclear charge, atomic radius, and inner-shell shielding are all apparent. These cause electronegativity to increase across a period (as nuclear charge increased and atomic radius decreases) and decrease down
a group (as radius and the number of inner shells both increase). Thus we find that the most electronegative atom is fluorine (F), at the upper right of the periodic table, and the least electronegative (or most electropositive) is firancium (Fr), at the lower left.

One of the chemical characteristics of a typical metal is a low electronegativity. Thus we find the best metals on the left and the best nonmetals on the right in the periodic table. Note that the transition elements all have fairly low electronegativities; they are metals. The change from metallic to nonmetallic properties occurs to the right of these elements in the periodic table. Lastly, note that it is the metals (at the left) which tend to form positive, simple (monatomic) ions, and the nonmetals (at the right) which tend to form negative ones. (Develop the habit of mentally equating metallic with electropositive and nonmetallic with electronegative.)