Hints for Building Lewis Structures

1) Build structures symmetrically when possible, as opposed to linearly. For example: nitrate ions have the oxygen attached to a central nitrogen atom, rather than being attached in a line (O=N=O=O).

An important exception is carbon. Carbon atoms like to form chains when attached to one another. This is because carbon has 4 valence electrons, therefore prefers to have 4 bonds and really doesn’t want to be a terminal atom. For example:

2) Avoid bonding oxygen to oxygen, if you can. Look at nitrate ions (above) as an example.

3) Hydrogen is always terminal. It can only form 1 bond, therefore, will never be in the middle of an atom.

4) Use the formula as a hint. For example: HOCN is listed that was because that is the order in which the atoms are connected. Acetic acid is sometime written as CH$_3$COOH because it’s Lewis structure is:

5) Think about how many bonds each atom wants, to end up with 8 electrons in it’s orbitals. For example: oxygen has 6 valence electrons, so would prefer to form 2 bonds and nitrogen has 5 valence electrons, so would prefer to form 3 bonds, etc.

   - Give as many of the atoms the preferred number of bonds as you can. As an example, HOCN has 2 possible resonance structures, but the first is preferable to the second. Can you see why?
   - If they can’t all have their preferred number of bonds, then the ones that want the least number of bonds are more likely to get what they want. For example, nitric acid has 2 oxygen atoms that have the 2 bonds they want. Another oxygen atom and the nitrogen (which wants 3 bonds), don’t, but for that molecule, that’s the best that can be accomplished.
   - The atom that wants the most bonds is more likely to be the central atom. Look at nitric acid, again, as an example of this.

6) Acids of polyatomic ions almost always have the hydrogen on an oxygen atom. Nitric acid, above, is a good example of this.

7) Anticipate if a molecule has a any multiple bonds by using the 6N + 2 rule. N is the number of atoms in a molecule, not counting hydrogen. 6N + 2 predicts how many valence electrons you would need to have all single bonds. If this number matches the number of valence electrons you have in the molecule, then you have all single bonds. If you actually have 2 less than the predicted number, then you have a double bond. If you have 4 less than the predicted number, then the molecule may have 2 double bonds or a triple bond. Any other difference in electrons indicates a mistake, either in how the valence electrons were counted, or what value you used for N.

<table>
<thead>
<tr>
<th>Example</th>
<th>Valence electrons</th>
<th>6N + 2</th>
<th>Difference</th>
<th>Bonds</th>
</tr>
</thead>
<tbody>
<tr>
<td>CH$_4$</td>
<td>4+4(1)=8</td>
<td>6(1)+2=8</td>
<td>8-8=0</td>
<td>All single</td>
</tr>
<tr>
<td>NO$_3^-$</td>
<td>5+3(6)+1=24</td>
<td>6(4)+2=26</td>
<td>26-24=2</td>
<td>1 Double bond</td>
</tr>
<tr>
<td>N$_2$</td>
<td>2(5)=10</td>
<td>6(2)+2=14</td>
<td>14-10=4</td>
<td>2 Double bond or 1 Triple bond (only 1 TB is possible for N$_2$)</td>
</tr>
</tbody>
</table>