Lewis Structures - *Worksheet*

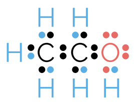
Question 1
Draw the Lewis structure for ethanol, CH_3CH_2OH.

Answer 1
Step 1 - draw all the atoms and their valence electrons

Step 2 - join all the atoms that can form more than one bond
In this example, the condensed formula indicates the connectivity of the atoms. Join the two carbon atoms together and then add the oxygen atom at one end.

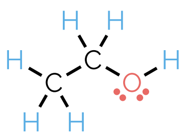
Four electrons are being shared, creating a covalent C–C bond and a covalent C–O bond.

Step 3 - add hydrogen atoms starting with the atoms with the most single electrons
Invariably, this means that you add hydrogen atoms to the carbon atoms first.



Step 4 - check that all that all atoms obey the octet rule
Every atom has a full outer electron shell (8 electrons) so this is an acceptable structure. The structure above is the answer to the question but …

Step 5 - simplify the drawing
Strictly this is not necessary for drawing a Lewis structure but it is good practice! At this stage, we would strongly recommend leaving the lone pair of electrons on the structure.

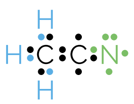


Question 2
Draw the Lewis structure of acetonitrile, CH_3CN.

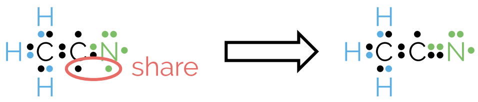
Answer 2
Step 1 - draw all the atoms and their valence electrons

Step 2 - join all the atoms that can form more than one bond
This is the same as the last example. Join the two carbon atoms by a covalent bond and then add the nitrogen atom at one end.

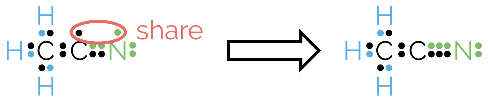
Step 3 - add hydrogen atoms starting with the atoms with the most single electrons
We could argue that you start by adding all the hydrogen atoms to the carbon with the most unpaired electrons, and in this case this would be true, but the real approach we would recommend is trying to match the condensed formula, CH3CN, and this tells us that all the hydrogen atoms are on a single carbon.



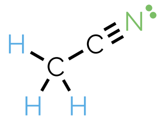
Step 4 - check that all that all atoms obey the octet rule
Not all the atoms do obey the octet rule. One carbon atom and the nitrogen atom both have 6 valence electrons. It is necessary to share single electrons on adjacent atoms to bring each up to a full octet. Sharing one from each atom gives both the nitrogen and the oxygen seven valence electrons.



So, we must share another set of electrons. This finally gives all atoms a full outer electron shell. The Lewis structure of acetonitrile is shown:

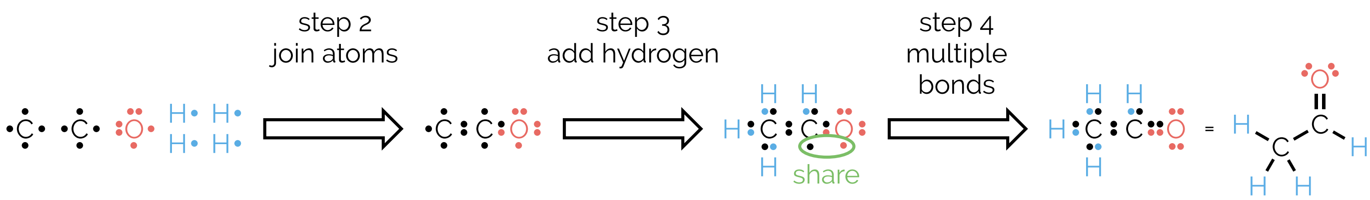


Step 5 - simplify the drawing
The dot diagrams of a Lewis structure are messy and take time to draw. Chemists use a simplified line diagram, replacing each pair of shared electrons (a covalent bond) with a line. For acetonitrile that gives the following structure:



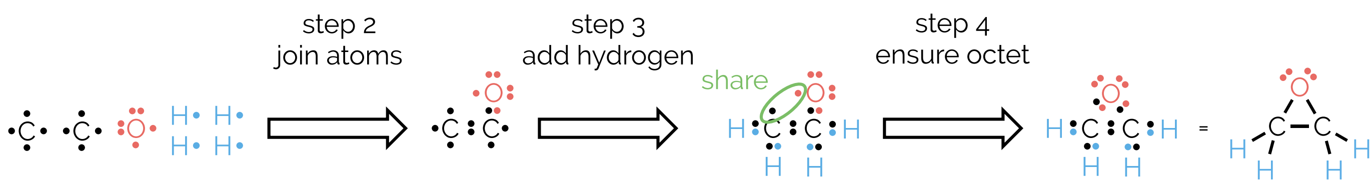
Question 3
There are three allowable Lewis structures for a molecule with the molecular formula C_2H_4O. Draw them.

Answer 3
In each case you are aiming to make the two carbon atoms and the oxygen obey the octet rule and have 8 valence electrons. As always, start by joining these three atoms together and then adding hydrogen atoms. The first structure you might draw will probably be the aldehyde below:

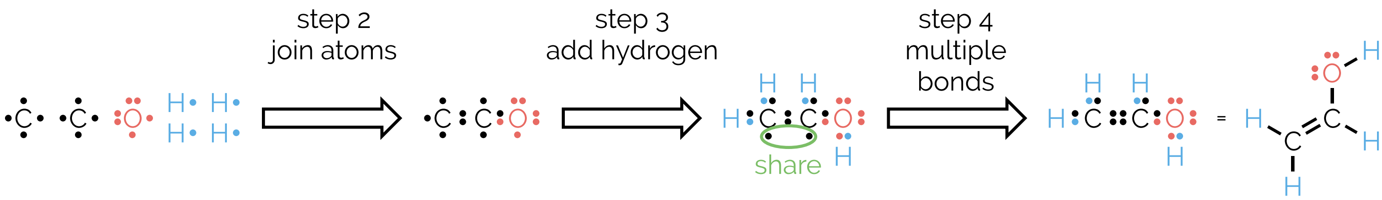


The sequence of steps to achieve this is to first draw all the atoms and their valence electrons. Then link the atoms that can form more than one bond. Next, add the hydrogen atoms starting with the atom with the most unpaired electrons. Finally, ensure all atoms have an octet. This is achieved by sharing the single electrons on carbon and oxygen,

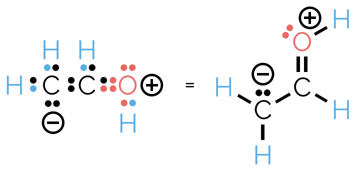
The second possible structure is formed by sharing the hydrogen atoms evenly between the two carbon atoms. This leaves the unshared electrons on the oxygen atoms as before and the other carbon. Sharing these two electrons does not create a multiple bond but a ring. This is oxirane or ethylene oxide.



The third structure sees one of the hydrogen atoms being shared with the oxygen. This leaves the carbon atoms without a full octet. The solution to this is to share four electrons between the two carbon atoms:



In all three of these molecules each atom obeys the octet rule. All three molecules exist (the first and third are related by an equilibrium process known as tautomerisation). There is a fourth way of arranging the electrons that also obeys the octet rule but leads to formal charges (more about these in a future summary). This is a resonance structure of the final molecule (again, more on resonance in a later summary). It is acceptable as it obeys all the guidelines we have provided.

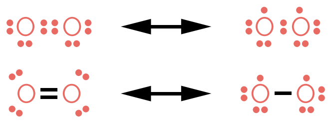


As we have stated earlier, Lewis structures are like training wheels. They are useful to teach you where the electrons are. But it should soon become apparent that there are patterns that can help you draw structures faster. First, neutral atoms always form the same number of bonds: hydrogen always forms one bond; oxygen always forms two; nitrogen always forms three and carbon always forms four bonds. Secondly, in organic chemistry there are always four zones of electrons around an atom. These zones comprise of either two electrons from a bond or two electrons in a lone pair of electrons. Formal charges can change this (but normally only on carbon) and we will have to deal with them at a later stage.*

Atom	Bonds	Lone pairs	Zones
carbon	4	0	4
nitrogen	3	1	4
oxygen	2	2	4
fluorine	1	3	4

Question 4
Draw the Lewis structure of molecular oxygen (dioxygen), O_2.

Answer 4
Yes, this is a trick question and not as easy as it looks. The most common answer, the one that obeys the guidelines we have given you is shown below:



But, the physical properties of dioxygen suggest that it has unpair electrons. In liquid form dioxygen is magnetic. This can be explained by an alternative resonance structure as shown. Why have we included this here? Lewis structures are incredibly useful. They form the basis of > 90% of the material we will cover, but they are just a representation of each molecule, they are not the real structure. As chemists, or scientists, you will need to get happy with working with useful, but not strictly correct, models.

*Alright, this is not strictly true. Group 13 elements (boron and aluminium) can have only three zones of electrons but they are very reactive. Atoms with formal charges can have fewer zones as well.