Lewis Structures

Writing Lewis Structure:

1. Put the least electronegative, or unique element in the center. Connect atoms together, so they are bonded to each other.
2. Count total # of valence electrons. Add 1 for each negative charge, subtract 1 for each positive charge.
3. COMPLETE AN OCTET FOR ALL ATOMS EXCEPT HYDROGEN.
4. *If the # of electrons used to complete the octet was more than the total valence electrons, add a double or tripple bomd.*
5. *If the # of electrons used to complete the octet was less than the total valence electrons, place the left over electrons as unshared electrons, on the central element.*

Formal Charge:

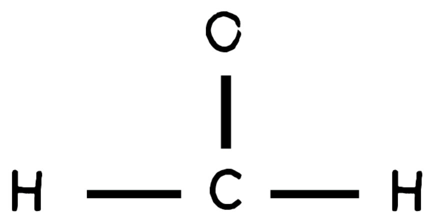
- *An atom's formal charge is the difference between the # of valence electrons in an isolated atom and the # of electrons assigned to that atom is a Lewis Strucure.*
- The sum of the formal charges of the atoms in a molecule or ion must equal the charge on the molecule or ion.

Formal Charge on Lewis Structure:

1. For neutral molecules, no formal charges are preferred.
2. Large formal charges are less plausible.
3. The most plausible structure is the one that has negative formal charges that are placed on the more electronegative atoms.

Example: Write the Lewis structure for CH_2O

Step 1

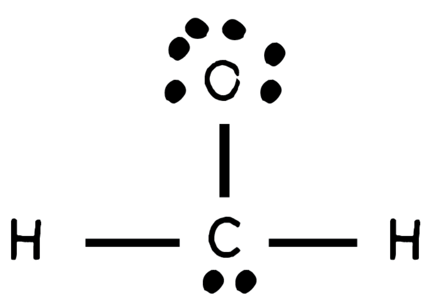


Step 2

$C + 2H + O$

$(1)4 + 2(1) + (1)6 = 12$

STEP 3



Step 4

4 lone pairs + 3 single bonds

$4(2) + 3(2) = 14$

14 is too much so we take off a pair of electrons on oxygen and make the single bond a double bond

Step 5

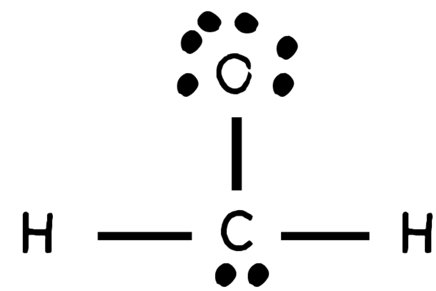
Because there was more electrons in the structure than total valence eloctrons in the molecule, we skip step 5

Check

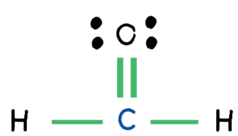
If There are more than one possible structures (such as the one below) then you have to check formal charges of both structures

Formal Charge Structure 1

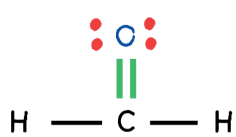
$$\text{Formal Charge of an atom in a Lewis structure} = \text{Total \# of valence electrons in the free atom} - \text{Total \# of nonbonding electrons} - \frac{1}{2}\left(\text{Total \# of bonding electrons}\right)$$



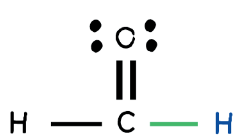
Formal Charge Step 1



Formal charge on C : $4 - 0 - \frac{1}{2}(8) = 0$

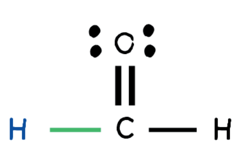


Formal charge on O : $6 - 4 - \frac{1}{2}(4) = 0$



We are going to do this 2 times for the 2 hydrogen atoms present

Formal charge on H : $1 - 0 - \frac{1}{2}(2) = 0$



Formal charge on H : $1 - 0 - \frac{1}{2}(2) = 0$

Formal Charge Step 2

The charge on the molecule is zero, so the sum of the formal charges on each atom should equal zero

↓

(1)Carbon atom + (1)Oxygen atom + (2)Hydrogen atom = Formal Charge

↓

(1)0 + (1)0 + (2)0 = Formal Charge

↓

0 + 0 + 0 = Formal Charge

↓

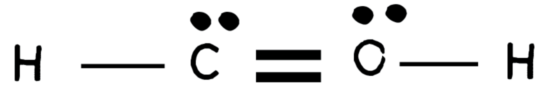
0 = Formal Charge

↓

Now we check the formal charge of the other possible structure

Formal Charge Structure 2

Formal Charge of an atom in a Lewis structure $=$ Total # of valence electrons in the free atom $-$ Total # of nonbonding electrons $- \frac{1}{2}\left(\text{Total \# of bonding electrons}\right)$



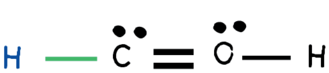
Formal Charge Step 1

Formal charge on C : $4 - 2 - \frac{1}{2}(6) = -1$

Formal charge on O : $6 - 2 - \frac{1}{2}(6) = 1$

We are going to do this 2 times for the 2 hydrogen atoms present, just like we did for the last structure

Formal charge on H : $1 - 0 - \frac{1}{2}(2) = 0$



Formal charge on H : $1 - 0 - \frac{1}{2}(2) = 0$

Formal Charge Step 2

Reminder that the charge on the molecule is zero, so the sum of the formal charges on each atom should equal zero

↓

(1)Carbon atom + (1)Oxygen atom + (2)Hydrogen atom = Formal Charge

↓

(1)(−1) + (1)1 + (2)0 = Formal Charge

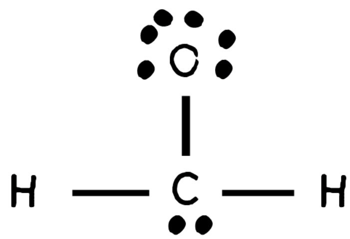
−1 + 1 + 0 = Formal Charge

0 = Formal Charge

↓

The Formal Charge of both structures are 0, so we have to compare individual atoms in the structure

A molecular structure in which all formal charges are zero is preferable to one in which some formal charges are not zero.

	1st structure  Formal Charge of atoms in 1st structure	2nd structure  Formal Charge of atoms in 2nd structure
Carbon	0	-1
oxygen	0	1
Hydrogen	0	0
	Better Structure	