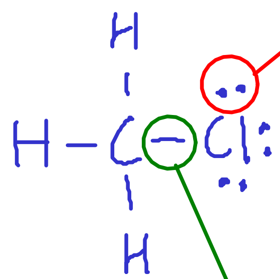
Lewis dot structures for molecules

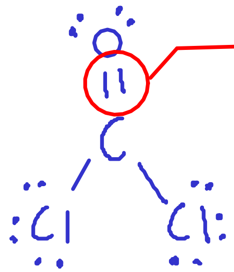
In the dot structure of a molecule,
- SHARED valence electrons are shown with dashes - one per pair.
- UNSHARED valence electrons ("lone pairs") are represented by dots.



lone pair (these electrons are not shared with another atom)

bonding pair (these electrons are shared between C and Cl)

Multiple pairs of shared electrons are represented by multiple dashes:

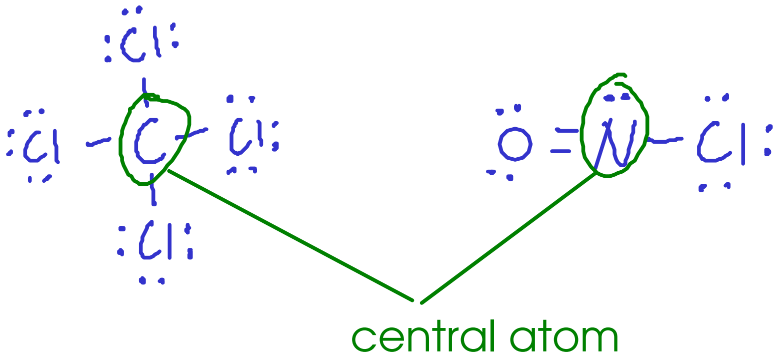


double bond - two shared electron pairs

triple bond - three shared electron pairs

Atoms generally don't share more than three pairs of electrons with a second atom, though they can share more pairs by sharing with several different atoms.

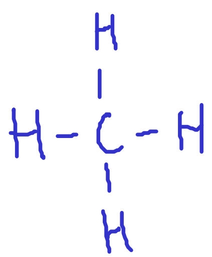
Small molecules generally form around a CENTRAL ATOM.

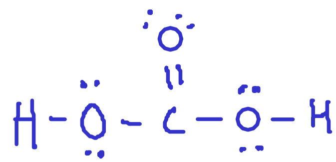


Other atoms in the molecule bond to the central atom.

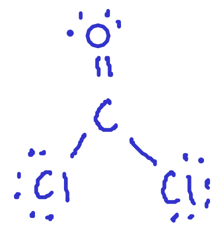
The central atom is usually the atom in the structure which needs to gain the most electrons for its outer shell.

The "octet rule" is a useful guide to figuring out how many electrons an atom will share in a molecule.





Count the electrons for each atom. Remember, each dash represents a pair!



Atoms usually end up with a share in EIGHT VALENCE ELECTRONS in a Lewis structure. This includes bonding pairs and lone pairs.

Hydrogen is different, since its outer shell can hold a maximum of two electrons.

To draw the structure for a simple molecule, first count the total number of valence electrons for all the atoms in the molecule. To quickly determine the number of valence electrons in each atom, use the periodic table!

For "A" groups (the "main group" elements), the number of valence electrons typically equals the "A" group number.

	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIIIB			IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA
1	H																	He
2	Li	Be											B	C	N	O	F	Ne
3	Na	Mg											Al	Si	P	S	Cl	Ar
4	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
5	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
6	Cs	Ba	~~La~~ *	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
7	Fr	Ra	~~Ac~~ *	Rf	Db	Sg	Bh	Hs	Mt									

1 valence electron

2 valence electrons

3 valence electrons

4 valence electrons

5 valence electrons

6 valence electrons

7 valence electrons

8 valence electrons (except helium!)

To choose a central atom, pick the element that needs to gain the most electrons.

HYDROGEN can have a maximum of two electrons, so it's never going to be central.

CH_2O:

C: 4 (Group IVA - needs 4 more electrons to get 8)

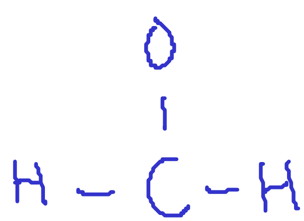
H: 2 x 1 (Group IA, but there are two H atoms. Each needs one more electron.)

O: 6 (Group VIA - needs 2 more electrons to get 8)

12 electrons (total number of valence electrons in the structure)

Carbon is the central atom, since it needs to gain more electrons than either hydrogen or oxygen.

To draw the molecule, first draw a SKELETAL STRUCTURE, attaching all the other atoms to the central atom with single bonds.



Modify the skeletal structure so that it shows all the valence electrons. Distribute electrons around the structure until you have used all the available valence electrons.

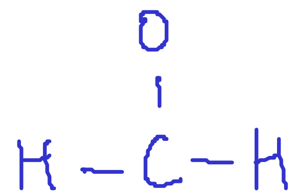
CH_2O:

C: 4

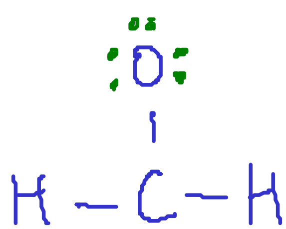
H: 2x1

O: 6

12 electrons



Start with the outer atoms, and if you "fill" them before running out of electrons, move to the central atom.



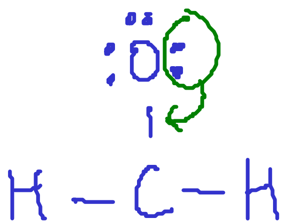
In this example, we could only put electrons on the oxygen atom, since the outer hydrogen atoms were "full" with two electrons.

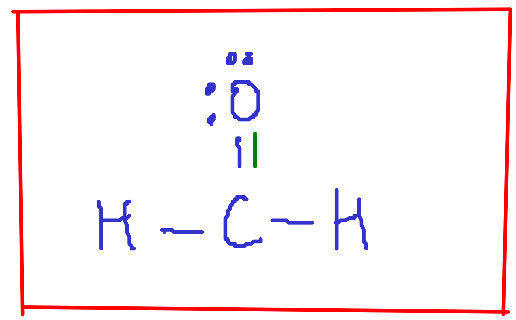
We stop when oxygen is full, because we only have 12 valence electrons to work with.

Count, but remember that each single bond we drew for the skeletal structure represents two electrons.

Each atom in the structure should have EIGHT valence electrons, if it obeys the octet rule. Hydrogen should have TWO valence electrons.

If an atom does not have enough electrons, we can give it a double or triple bond by "relocating" electrons from a lone pair.





Count. Now both oxygen and carbon have eight valence electrons.

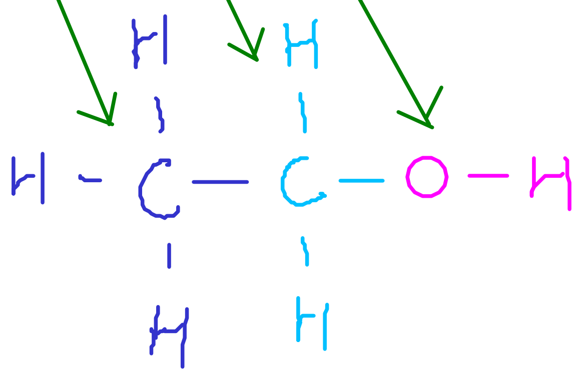
Always check the final structure to make sure it still has the correct total count of valence electrons.

Larger molecules are often made of chains of smaller ones. Sometimes, the chemical formula will hint to this.

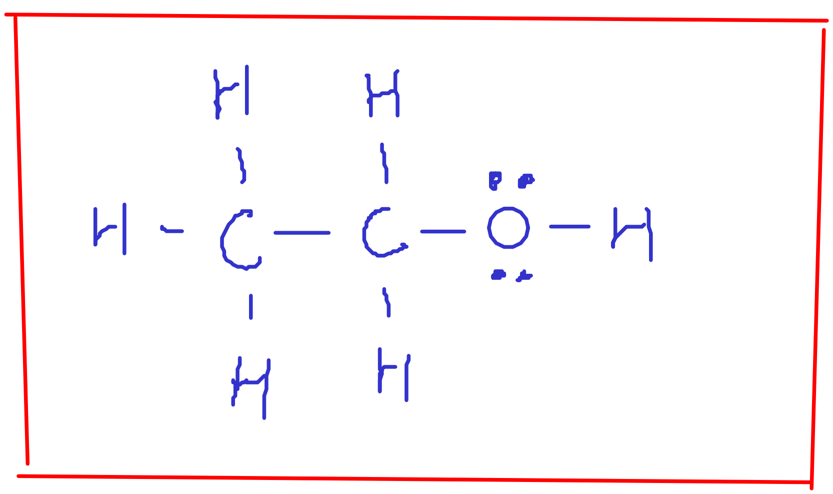
CH_3CH_2OH

$C: 4 \times 2 = 8$
$H: 1 \times 6 = 6$
$O: 6 \times 1 = 6$

20 electrons



this skeletal structure has three central atoms. Each piece of the molecule has own central atom, and is chained to the next one to form the overall molecule.



Some molecules have DELOCALIZED BONDING, where the same electrons are shared between more than two atoms. Lewis structures have a problem showing this type of bonding.

NO_3^- "nitrate ion"

$N: 1 \times 5$

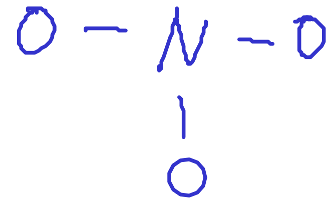
$O: 3 \times 6 = 18$

23

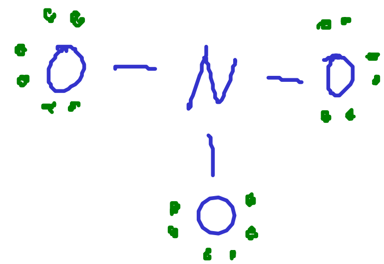
$+\ 1$

24

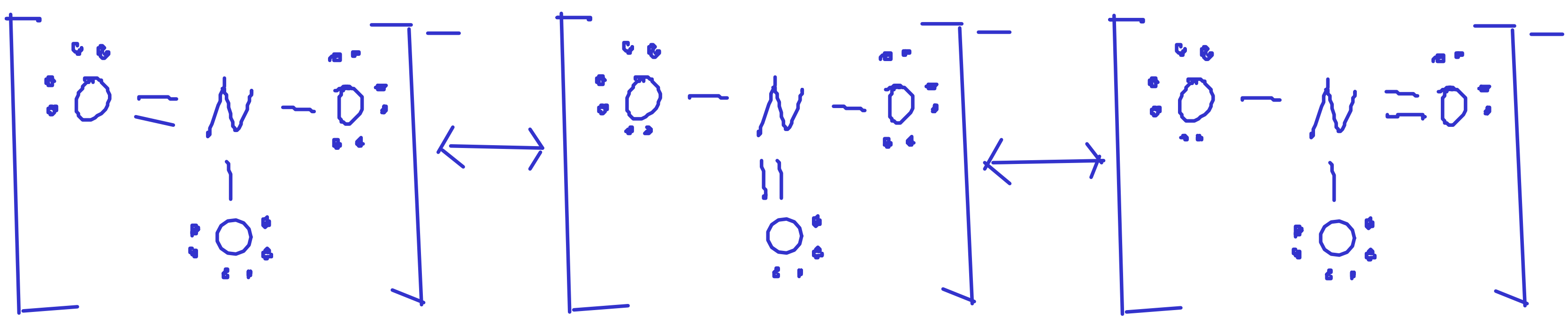
-1 charge requires one more electron!



Skeletal



All 24 electrons used.



Add a double bond to get enough electrons for N.

... but all the oxygen atoms should bond the same way!

So we draw three structures .. called "resonance structures"

There's not really a double bond in the structure that bounces around. The real molecule has some electrons that are shared between all of these atoms - and this is just how we show delocalized bonds with Lewis structures.

Not all atoms obey the octet rule all the time. Some atoms have EXPANDED VALENCE, which means they end up with more than eight valence electrons.

Atoms can fit more than eight electrons in their outer shells only if they have "d" subshells in their outer shell. So, to have expanded valence, an atom must be from period 3 or higher. So, sulfur can do expanded valence, but fluorine (period 2) cannot.

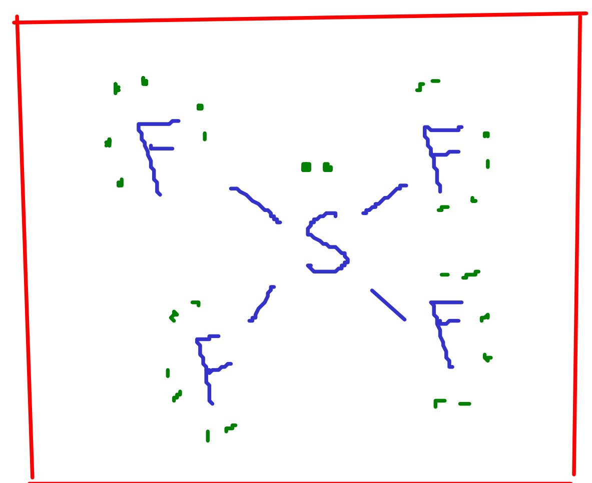
SF_4

$S: 1 \times 6$

$F: 4 \times 7 = 28$

34 electrons

Skeletal structure



To use all 34 electrons, we put the last pair on the central sulfur atom, giving it 10. This is okay for sulfur, as it can accept the extra pair.

Examples:

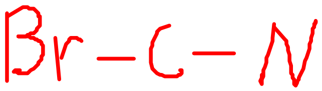
BrCN

Br : 1 x 7

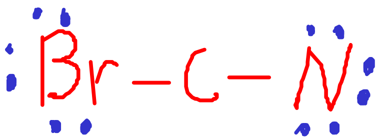
C : 1 x 4

N : 1 x 5

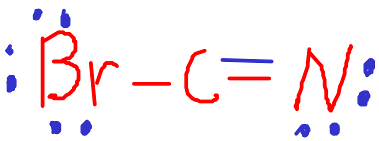
16 electrons



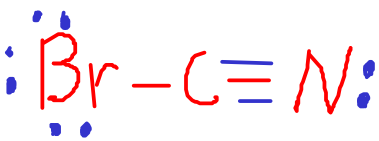
Skeletal



Add electrons, stop when we reach 16



Now carbon has a share in six electrons



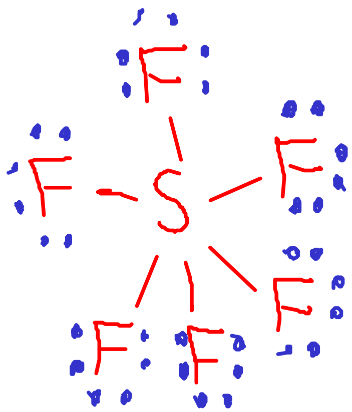
Making a triple bond with nitrogen gives all atoms a share in eight valence electrons...

SF_6 S: $1 \times 6 = 6$
F: $6 \times 7 = 42$

48 electrons

The skeletal structure has twelve electrons in sulfur's outer shell, but since sulfur is period 3, that's OK.

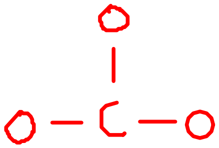
Sulfur hexafluoride is another example of expanded valence. Sulfur ends up with 12 outer electrons.



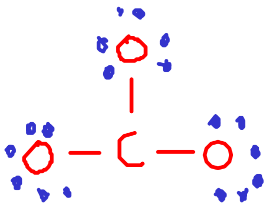
CO_3^{2-}

C – $1 \times 4 = 4$
O – $3 \times 6 = 18$

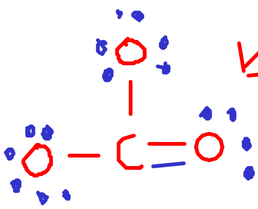
22 electrons
+ 2 charge
24 electrons



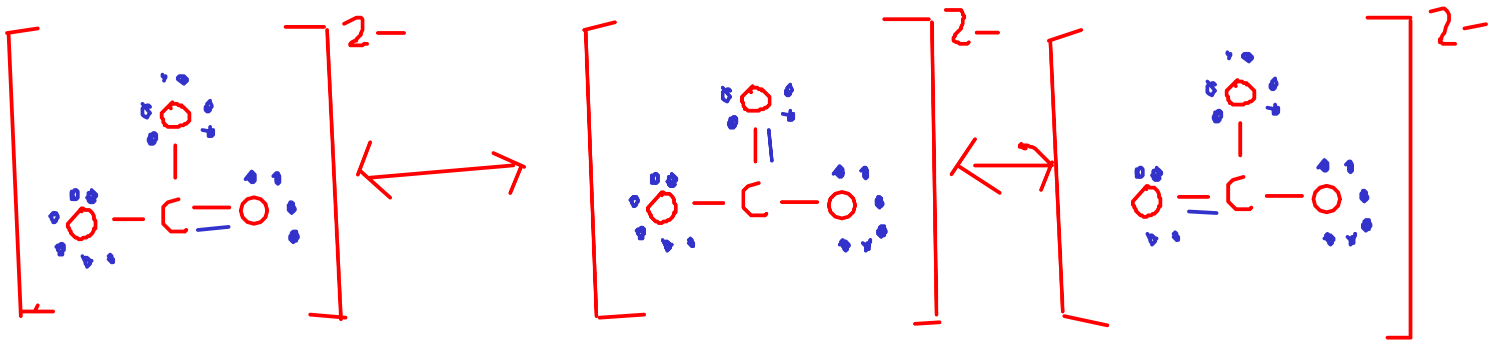
Skeleton



Distribute remaining electrons (total 24)



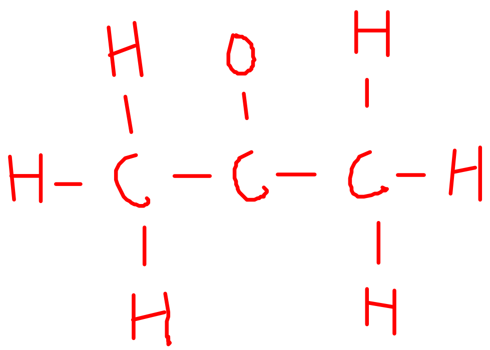
Notice that one oxygen atom is bonded differently from the others, even though all three are attached to the same carbon and to nothing else. This is a hint that the molecule has RESONANCE structures (delocalized bonds)



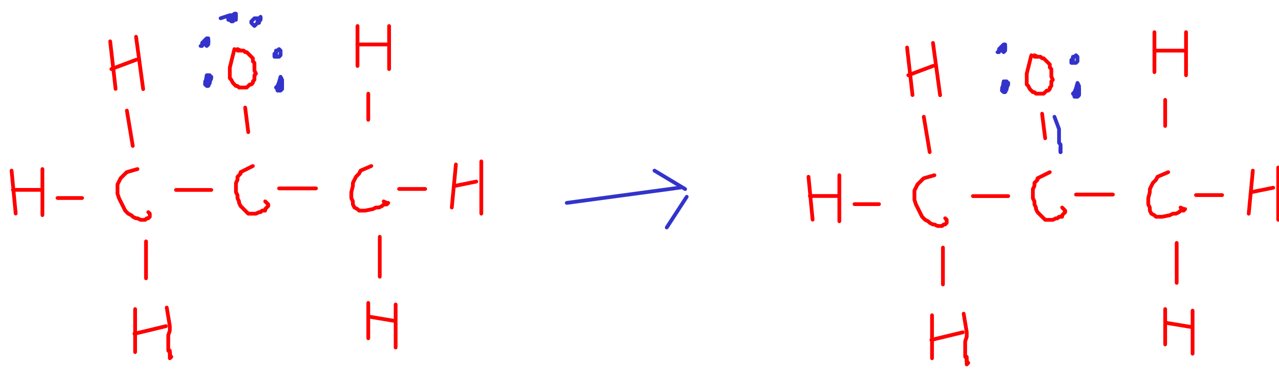
CH_3COCH_3

$C - 3 \times 4 = 12$
$H - 6 \times 1 = 6$
$O - 1 \times 6 = 6$
24 electrons

This is a large molecule. The formula hints that this molecule has three small-molecule pieces...



The skeletal structure has three "central" atoms linked together...



Making a double bond gives carbon enough (eight) outer electrons...

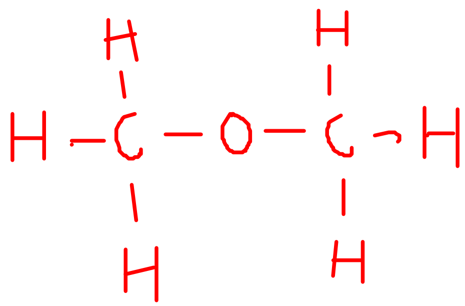
Carbon needs more electrons!

$(CH_3)_2O$

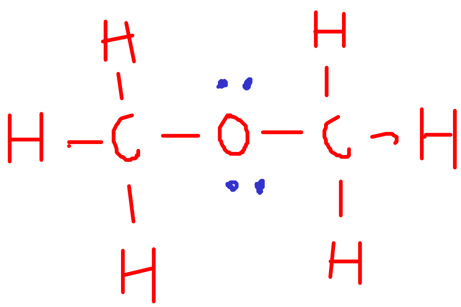
$C - 2 \times 4 = 8$
$H - 6 \times 1 = 6$
$O - 1 \times 6 = \underline{6}$
20 electrons

This molecule is called "dimethyl ether", and is an isomer of ethanol, a molecule we used earlier as an example.

ISOMERS are molecules that have the same molecular formula as each other, but have different arrangements of atoms. Depending on how different the arrangement of atoms is, some isomers may have very different properties - like dimethy ether and ethanol.



The formula hints at this skeleton.



The last four electrons must go onto oxygen, as all the other atoms are "full".