Lewis structure review

1. Count valence electrons

2. Pick central atom and draw skeletal structure
   - central atom is usually the one that needs to gain the most electrons!
   - skeletal structure has all atoms connected to center with single bonds

3. Distribute remaining valence electrons around structure, outer atoms first. Follow octet rule until you run out of electrons.

4. Check octet rule - each atom should have a share in 8 electrons (H gets 2). if not, make double or triple bonds.

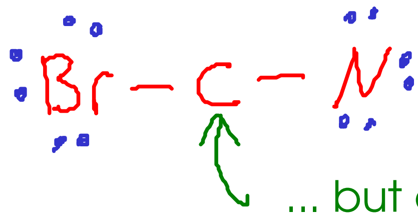
BrCN "cyanogen bromide"

Br : 1x 7
C : 1x 4
N : 1x 5

16 e^-

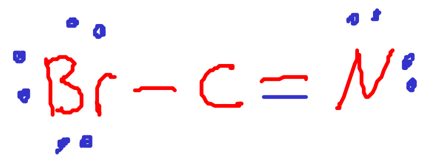
Br – C – N

Choose C as central atom, since it needs more electrons than either Br or N ...

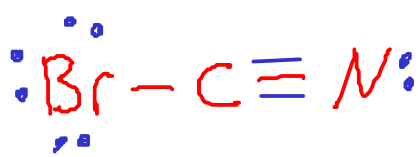


Distribute remaining electrons, stop when we run out...

... but carbon doesn't have enough! (only 4, not 8)



Use a lone pair from N to make a double bond. We picked N because it needed to gain more elecrrons initially. We still need more electrons for carbon, so we make ANOTHER bond!



<-- The final structure!

Polyatomic ions and resonance

NO_3^- "nitrate ion"

N: 1×5

O: $3 \times 6 = 18$

23

$+1$ <-- -1 charge requires one more electron!

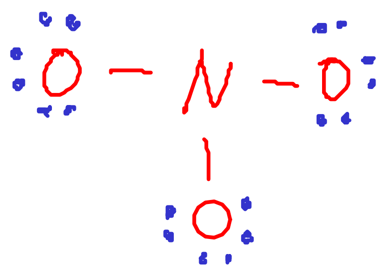
24

For a polyatomic ion, you must adjust the valence electron count to give the molecule the right charge.

Usually, the same kinds of atoms bond the same way in a structure.

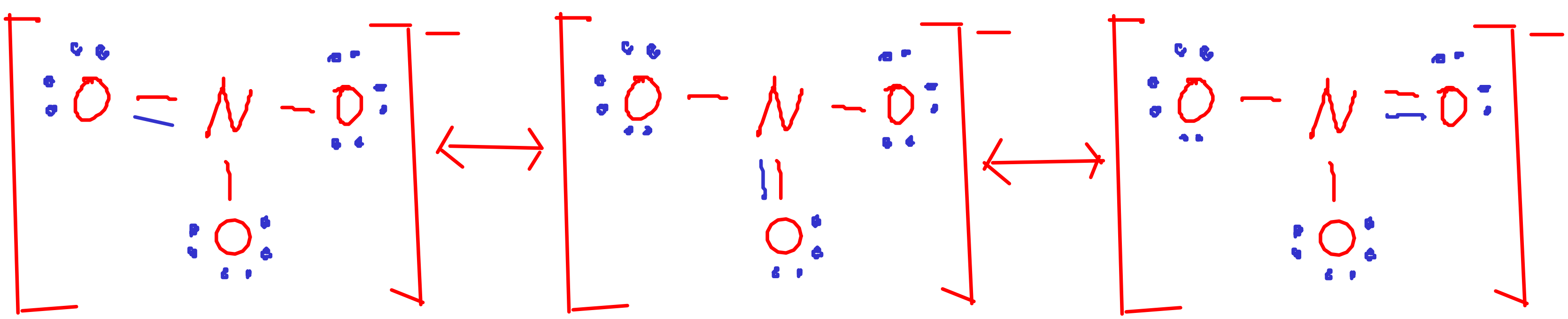
In other words, the three oxygen atoms in this molecule should bond the same way to the central nitrogen atom. But ...

Nitrogen is the central atom.



We used all 24 electrons.

But the nitrogen atom doesn't have enough electrons.



Resonance forms: The "double bond" we drew to give nitrogen eight electrons is really a delocalized bond between all the atoms. All the oxygen atoms are bonded the same way experimentally!

Expanded valence

SF_4 "sulfur tetrafluoride"

S : 1 × 6
F : 4 × 7 = 28

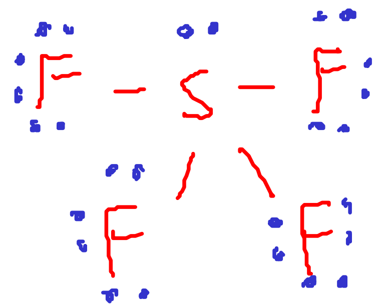
34 e^-

Some atoms may end up with more than eight valence electrons.

Atoms that end up with more than eight electrons have what we call "expanded valence". These atoms MUST be in periods 3 or higher on the periodic table.

Atoms from periods 1 and 2 (like H, C, N, O, F) cannot have expanded valence.

Sulfur is our centtal atom.



We ran out of space on the fluorines after 32 electrons, so the remaining pair goes onto sulfur.

.. but sulfur now has a share in 10 electrons, not 8.

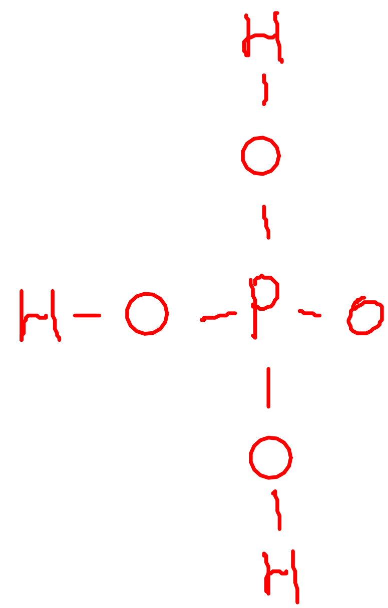
Since sulfur is period 3, it can do expanded valence ... and this structure is okay.

Formal charge

H_3PO_4 "phosphoric acid"

H : $3 \times 1 = 3$
P : $1 \times 5 = 5$
O : $4 \times 6 = 24$
$32 e^-$

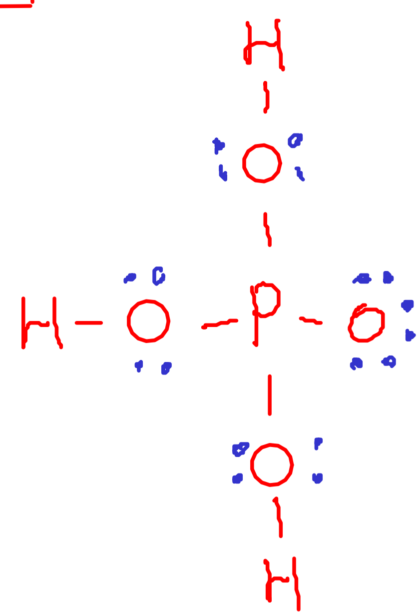
To tell which of two possible structures for a molecule is the better one, calculate the "formal charge" of each atom in the structure.



Phosphoric acid is an OXYACID. An oxyacid must have its acidic hydrogens attached to OXYGEN atoms!

original valence electrons − number of bonds − number of unshared electrons

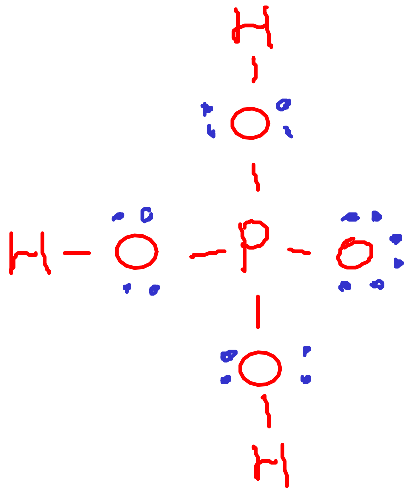
The preferred structure has either
- Lower overall formal charges
- More negative formal charges on electronegative atoms



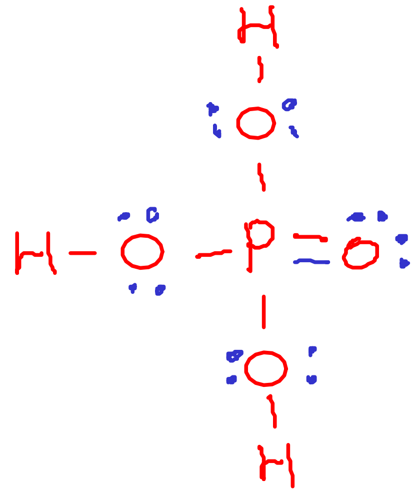
One possible structure for phosphoric acid looks like this...

But ... phosphorus is period 3, so there';s an alternate we can draw!

- Both of these are legitimate Lewis structures. All electrons are accounted for, and all atoms have an acceptable number of electrons.



Structure 1



Structure 2

To decide between these possible structures, we will use FORMAL CHARGE

Structure 1

H: $1 - 1 - 0 = 0$

O: $6 - 2 - 4 = 0$

with H

O: $6 - 1 - 6 = -1$

P: $5 - 4 - 0 = +1$

Structure 2

H: $1 - 1 - 0 = 0$

O: $6 - 2 - 4 = 0$

with H

O: $6 - 2 - 4 = 0$

P: $5 - 5 - 0 = 0$

So, based on FORMAL CHARGE ... Structure 2 is the most likely structure for phosphoric acid! (Lower formal charges)

For more in-depth information about Lewis structures, see the CHM 110 supplemental notes on the web site, or see sections 9.6 - 9.9 in the textbook.