## EXAMPLES FOR DRAWING LEWIS STRUCTURES (see section 4.1 in the textbook for more details and examples)

**Example 1:** Draw the Lewis structure for a **phosphorus tribromide molecule.**

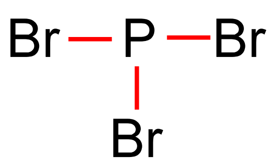
$PBr_3$

Step 1: Count the valence electrons.

The phosphorus atoms has 5 valence electrons, and each bromine atom has 7 valence electrons.

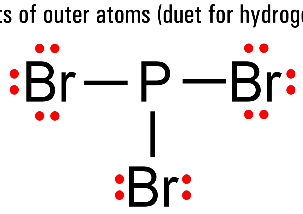
5 + 3(7) = 26 valence electrons

Step 2: Bond outer atoms to central atom (2 e$^-$ per bond).



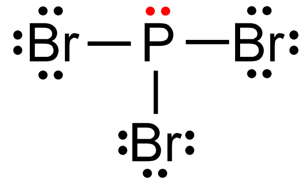
6 electrons used (20 electrons remaining)

Step 3: Fill octets of outer atoms (duet for hydrogen).



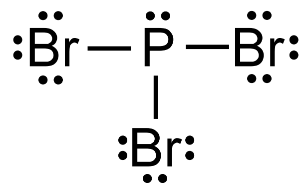
24 electrons used (2 electrons remaining)

Step 4: Pair-up remaining electrons on central atom.



The two remaining electrons are paired up on the phosphorus atom.

Step 5: Reassign outer lone pair to bonding position if needed to obtain octet for central atom (creates multiple bond).



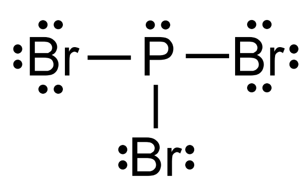
The phosphorus atom is surrounded by 8 electrons (3 bonding pairs and 1 lone pair).

Step 6: Assign formal charges.
The phosphorus atom "owns" 5 electrons (one in each of the three bonds and the lone pair). Since a neutral phosphorus atom has 5 valence electrons, there is no formal charge on the phosphorus atom.

Each bromine atom "owns" 7 electrons (three lone pairs plus one in the bond). Since neutral bromine atoms have 7 valence electrons, there are no formal charges on the bromine atoms.

Step 7: Reduce formal charges if possible.

There are no formal charges to reduce. The final structure is:



**Example 2:** Draw the Lewis structure for a **formaldehyde molecule.**

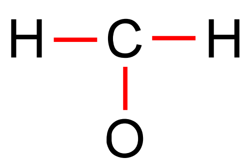
$$CH_2O$$

Step 1: Count the valence electrons.

The carbon atom has 4 valence electrons, each hydrogen atom has 1 valence electron, and the oxygen has 6 valence electrons.

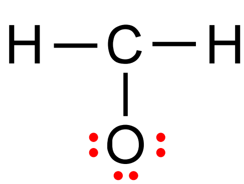
4 + 2(1) + 6 = 12 valence electrons

Step 2: Bond outer atoms to central atom (2 $e^-$ per bond).



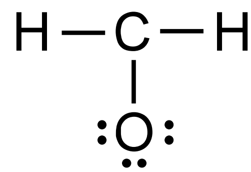
6 electrons used (6 electrons remaining)

Step 3: Fill octets of outer atoms (duet for hydrogen).



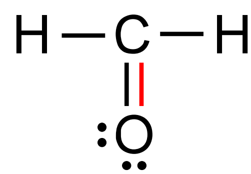
12 electrons used (0 electrons remaining)

Step 4: Pair-up remaining electrons on central atom.



There are no electrons remaining to put on central atom.

Step 5: Reassign outer lone pair to bonding position if needed to obtain octet for central atom (creates multiple bond).



The carbon atom is surrounded by only 6 electrons (3 bonding pairs). One of the lone pairs on the oxygen atom must be reassigned as a second bond between oxygen and carbon.

Step 6: Assign formal charges.

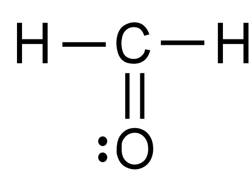
The carbon atom "owns" 4 electrons (one from each bond). Since a neutral carbon atom has 4 valence electrons, there is no formal charge on the carbon atom.

The oxygen atom "owns" 6 electrons (one from each bond plus the two lone pairs). Since a neutral oxygen atom has 6 valence electrons, there is no formal charge on the oxygen atom.

Hydrogen atoms never have a formal charge.

Step 7: Reduce formal charges if possible.

There are no formal charges to reduce. The final structure is:



**Example 3:** Draw the Lewis structure for a **nitrite ion.**

$NO_2^-$

Step 1: Count the valence electrons.

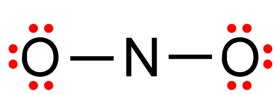
The nitrogen atom has 5 valence electrons, and each oxygen atom has 6 valence electrons. One more electron must be added to account for the 1– charge.

5 + 2(6) + 1 = 18 valence electrons

Step 2: Bond outer atoms to central atom (2 $e^-$ per bond).

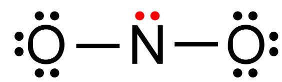
4 electrons used (14 electrons remaining)

Step 3: Fill octets of outer atoms (duet for hydrogen).



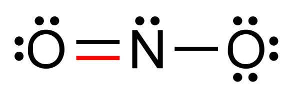
16 electrons used (2 electrons remaining)

Step 4: Pair-up remaining electrons on central atom.



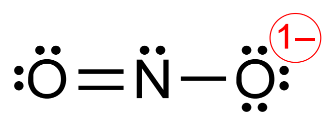
The two remaining electrons are paired up on the nitrogen atom.

Step 5: Reassign outer lone pair to bonding position if needed to obtain octet for central atom (creates multiple bond).



The nitrogen atom is surrounded by only 6 electrons (2 bonding pairs and one lone pair). One of the lone pairs on an oxygen atom must be reassigned as a second bond between an oxygen and the nitrogen.

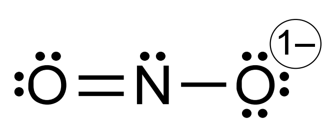
Step 6: Assign formal charges.



The oxygen on the right "owns" 7 electrons (one from the bond and 3 lone pairs). Since a neutral oxygen atom has only 6 valence electrons, there is a formal charge of 1– on the oxygen atom.

Step 7: Reduce formal charges if possible.

The formal charge cannot be reduced. The final structure is:



**Example 4:** Draw the Lewis structure for a **krypton tetrafluoride molecule.**

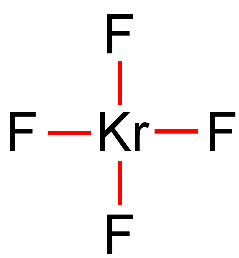
$KrF_4$

Step 1: Count the valence electrons.

The krypton atom has 8 valence electrons, and each fluorine atom has 7 valence electrons.

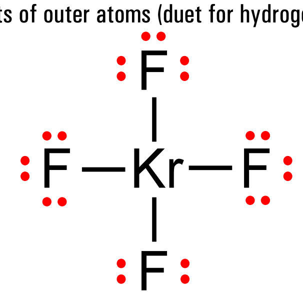
8 + 4(7) = 36 valence electrons

Step 2: Bond outer atoms to central atom (2 $e^-$ per bond).



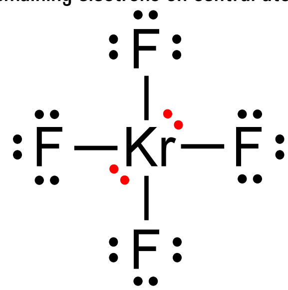
8 electrons used (28 electrons remaining)

Step 3: Fill octets of outer atoms (duet for hydrogen).



32 electrons used (4 electrons remaining)

Step 4: Pair-up remaining electrons on central atom.



The 4 remaining electrons are paired up on the krypton atom.

Step 5: Reassign outer lone pair to bonding position if needed to obtain octet for central atom (creates multiple bond).

The krypton atom is surrounded by 12 electrons (4 bonding pairs and 2 lone pairs). The krypton atom has an "expanded octet."

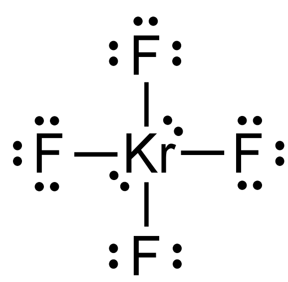
Step 6: Assign formal charges.

The krypton atom "owns" 8 electrons (one from each bond and two lone pairs). Since a neutral krypton atom has 8 valence electrons, there is no formal charge on the carbon atom.

Each fluorine atom "owns" 7 electrons (one from the bond and three lone pairs). Since a neutral fluorine atom has 7 valence electrons, there are no formal charges on the fluorine atoms.

Step 7: Reduce formal charges if possible.

There are no formal charges to reduce. The final structure is:



**Example 5:** Draw the Lewis structure for a **chlorous acid molecule.**

$HClO_2$

Step 1: Count the valence electrons.

The hydrogen atom has 1 valence electron, the chlorine atom has 7 valence electrons, and each oxygen atom has 6 valence electrons.

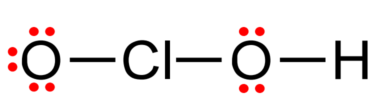
1 + 7 + 2(6) = 20 valence electrons

Step 2: Bond outer atoms to central atom (2 $e^-$ per bond).

Note that in oxoacids ($HNO_3$, $H_2SO_4$, etc) the hydrogen atoms are bonded to oxygen atoms rather than the central atom.

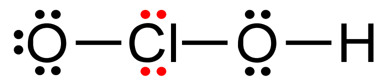
6 electrons used (14 electrons remaining)

Step 3: Fill octets of outer atoms (duet for hydrogen).



16 electrons used (4 electrons remaining)

Step 4: Pair-up remaining electrons on central atom.

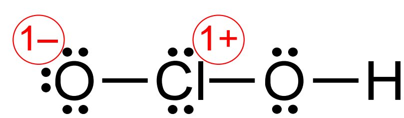


The 4 remaining electrons are paired up on the chlorine atom.

Step 5: Reassign outer lone pair to bonding position if needed to obtain octet for central atom (creates multiple bond).

The chlorine atom is surrounded by 8 electrons (2 bonding pairs and 2 lone pairs).

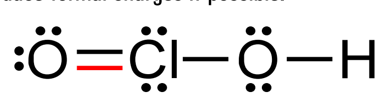
Step 6: Assign formal charges.



The oxygen atom on the left "owns" 7 electrons, and therefore has a formal charge of 1–.

The chlorine atom owns 6 electrons and therefore has a formal charge of 1+.

Step 7: Reduce formal charges if possible.



If one of the lone pairs on the oxygen atom is reassigned as a second bond between the oxygen and the chlorine then the formal charges are eliminated.

The chlorine atom has an "expanded octet."