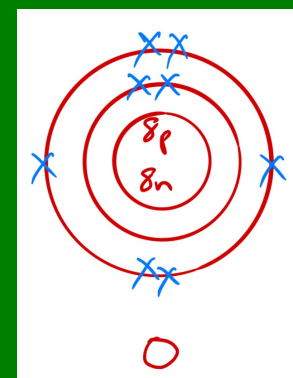
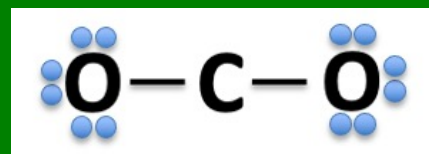
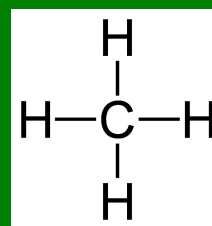


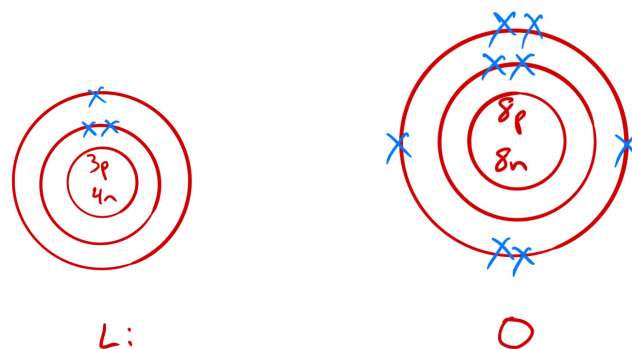
Lewis Structures

- Valence Electrons
- Octet Rule
- Ionic Bonding
- Covalent Bonding
- Polyatomic Ions
- Drawing Lewis Structures
- Determining Charge

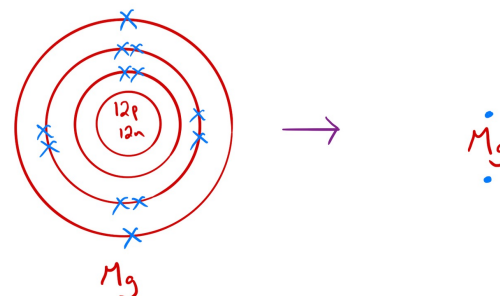


Valence Electrons

- When atoms form compounds, valence electrons are the ones that are involved
- Valence electrons are the electrons found in the outermost orbit



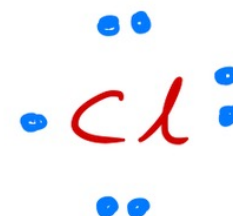
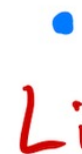
- Lewis structures show atoms depicting their valence electrons only – much easier



Lewis Structures

Rules for drawing lewis structures:

- Symbol of the element is in the center (represents protons, neutrons and inner electrons)
- The symbol is assumed to have 4 sides and the electrons in the outermost energy level are placed as dots on the four sides
- Dots are placed singly (North, South, East, West) and are then paired up



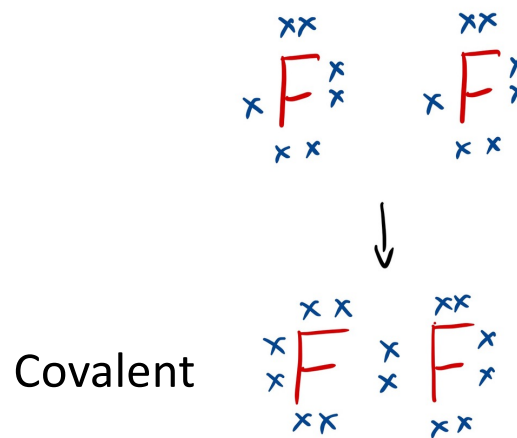
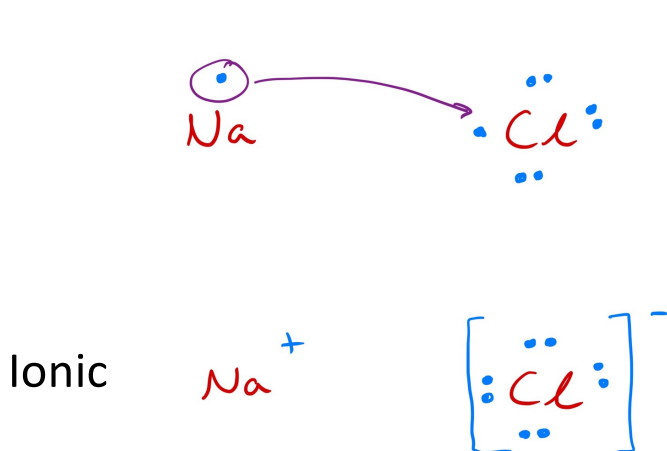
Finding The Number of Valence Electrons

- Find the group number the element is in while skipping the transition metals
- **The group number = the number of valence electrons**

[illegible]

Octet Rule - Bonding

- Many atoms react with each other in a way that leaves them with 8 electrons in the outer shell (orbit)
 - This is a stable configuration
- This occurs in ionic bonding when metallic ions bond with non-metallic ions to create neutral molecules and in covalent bonding where atoms share electrons to reach 8



Ionic vs. Covalent

	Ionic	Covalent
Atoms Involved	<u>Metal and a nonmetal</u> <u>(e.g., Na, F)</u>	Two nonmetals (e.g., H, N)
Electron Movement	Electrons are transferred to the anion	<u>Electrons are shared</u>

Lewis Structures: Ionic Bonds

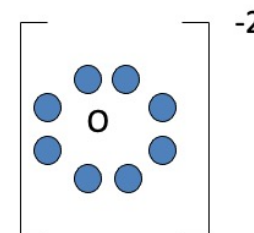
- Cation: positive charge - lost electrons to reach lower stable orbit:

- To draw - don't draw electrons around the symbol. this represents that you have gone down an orbit
- Put the charge outside in the upper right corner



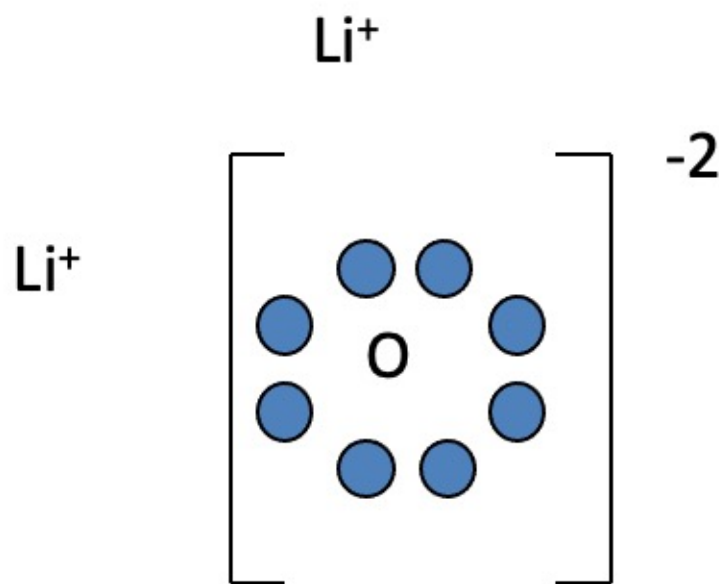
- Anion: negative charge - gained electrons:

- To draw - place the electrons around the symbol
- Put the symbol in a [] and write the charge outside in the upper right corner



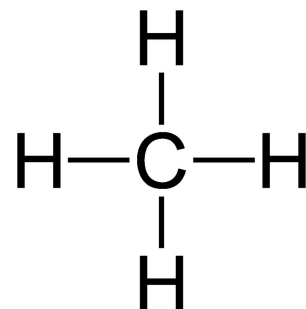
Li_2O – Lewis Structure

- Each lithium atom has given up 1 electron (Li^+)
- Each oxygen atom has gained 2 electrons (O^{2-})



Lewis Structures: Covalent Bonds

- Instead of using two dots to indicate the two electrons that comprise the covalent bond, a line is substituted for the two dots



- Any electron pairs not used in the bonding are drawn separately

Drawing Lewis Structures

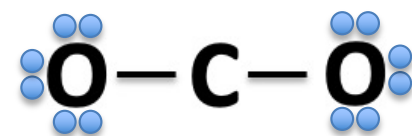
Example: CO₂

1. Determine the number of electrons to be used to connect the atoms then find the sum of the valence electrons for all atoms in the molecule
 - C = 4 valence electrons, O = 6
 - **Therefore, $1(4) + 2(6) = 4 + 12 = 16$ valence electrons**
2. Write the symbols for the atoms involved to show which atoms are connected - they are usually given in their correct order
 - E.g., HCN, carbon is the central atom

3. Draw a single bond between each pair of atoms bonded together



4. Complete the octets of the atoms bonded to the central atom. Remember that hydrogen (and helium) atoms only need 2 electrons



5. Place any remaining electrons on the central atom

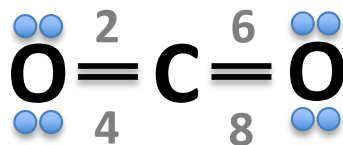
- This example doesn't change. Started with 16 valence electrons and that's what we have now
- If all the atoms have a complete outer shell (8 electrons) then you're done. Ours is not done, why?

6. If there aren't enough electrons to give the central atom an octet, try multiple bonds. Use one of the unshared pairs of electrons on the atoms bonded to the central atom to form double or triple bonds

➤ Our carbon only has 6



7. Continue this process of making multiple bonds between the outer atoms and the central atom until the valence shell of the central atom is complete



Lewis Structures: Polyatomic Ions

Example: PO_4^{3-}

1. Add up the total number of valence electrons

$$\text{P} = 5$$

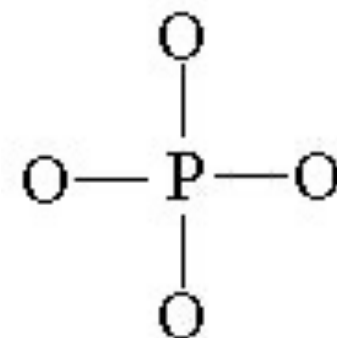
$$\text{O} = 4(6)$$

$$\underline{29 \text{ e}^-}$$

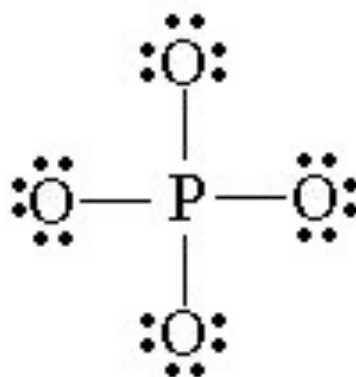
$$+ \underline{3 \text{ e}^-} \text{ (because of the -3 charge)}$$

$$\underline{32 \text{ e}^-}$$

2. Draw the simple structure with single bonds



3. Distribute the valence electrons around each surrounding atom until they have a complete valence shell (8 electrons), including the central atom



4. Ensure all electrons have been used up and



Determining Charge

- Divide covalent bonds (because each atom has 1 of the 2 electrons in a bond)

➤ Each O atom has 7 electrons, P has 4

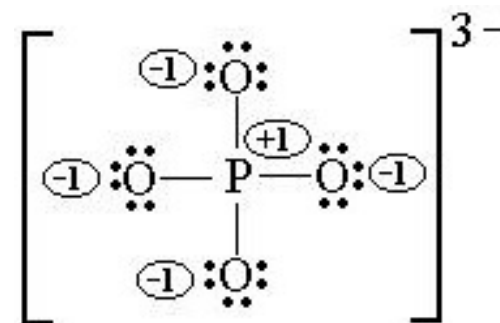
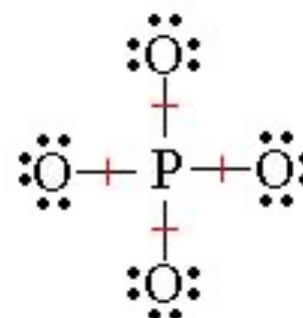
- Compare total electrons around each individual atom from the Lewis structure to the number of valence electrons of neutral atom

➤ E.g., O has 6 valence electrons, $6 - 7 = -1$

P has 5 valence electrons, $5 - 4 = +1$

- Determine the net formal charge of the polyatomic ion and each atom

➤ $(-1) + (-1) + (-1) + (-1) + 1 = 3-$



HAVE/NEED Method

- # of electrons HAVE = (count all valence electrons)
- # of electrons NEED = (how many electrons do you need to fill electron shell)
- Bonding electrons = (Need – Have Columns)
 - Divide this number by 2
- Total # unpaired electrons = (Have – Bonding electrons)
 - Place these around outer atoms to satisfy the octet rule