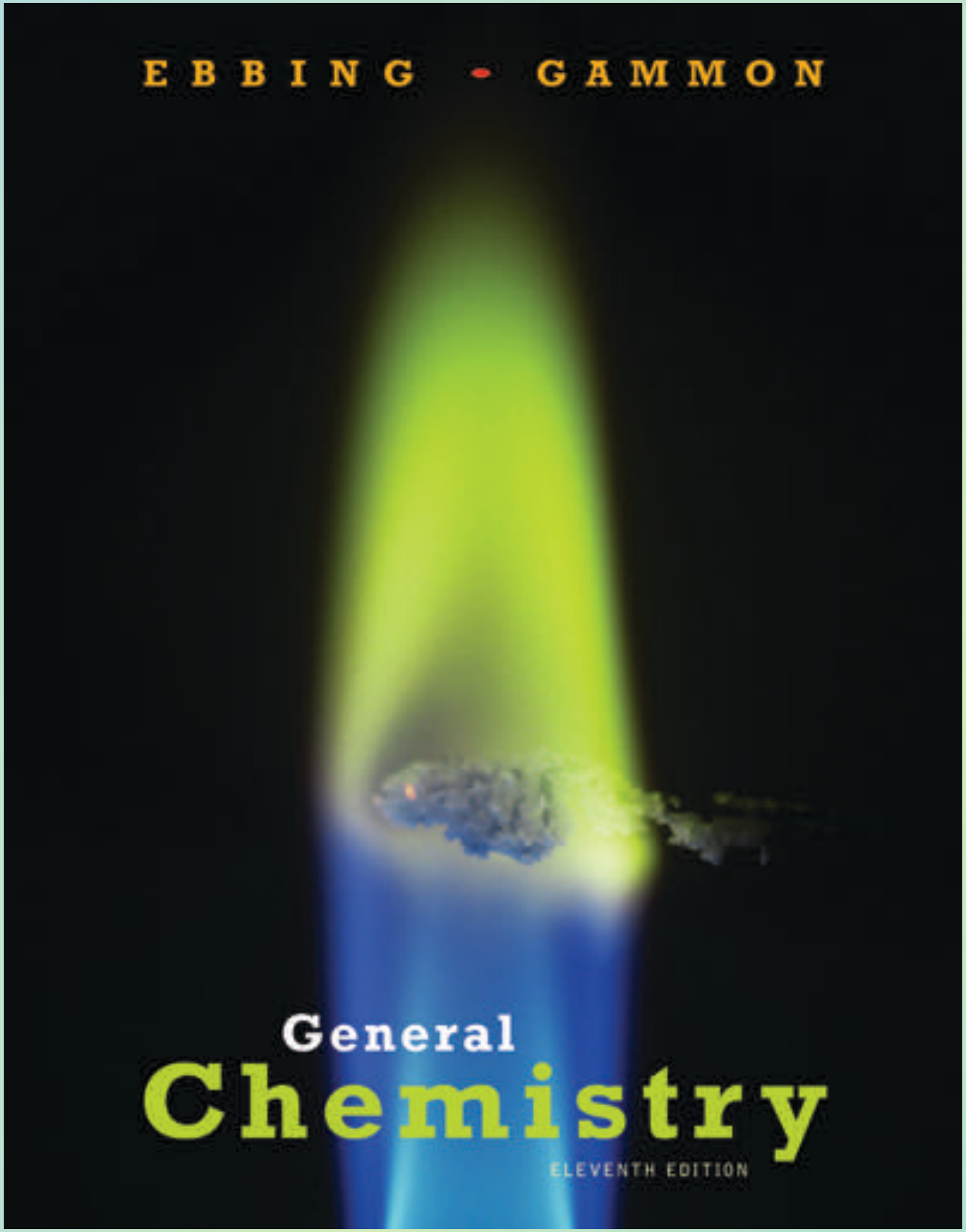


Chapter 9 Ionic and Covalent Bonding

A **Lewis electron-dot symbol** is a symbol in which the electrons in the valence shell of an atom or ion are represented by dots placed around the chemical symbol of the element.

Note – Dots are placed one to each side, until all four sides are occupied.

Lewis Dot Structures for Atoms

Table 9.1 Lewis Electron-Dot Symbols for Atoms of the Second and Third Periods

Period	1A ns^1	2A ns^2	3A ns^2np^1	4A ns^2np^2	5A ns^2np^3	6A ns^2np^4	7A ns^2np^5	8A ns^2np^6
Second	$\mathrm{Li}\cdot$	$\cdot\mathrm{Be}\cdot$	$\cdot\dot{\mathrm{B}}\cdot$	$\cdot\underset{\cdot}{\dot{\mathrm{C}}}\cdot$	$:\underset{\cdot}{\dot{\mathrm{N}}}\cdot$	$:\underset{\cdot\cdot}{\dot{\mathrm{O}}}\cdot$	$:\underset{\cdot\cdot}{\ddot{\mathrm{F}}}\cdot$	$:\underset{\cdot\cdot}{\ddot{\mathrm{Ne}}}:$
Third	$\mathrm{Na}\cdot$	$\cdot\mathrm{Mg}\cdot$	$\cdot\dot{\mathrm{Al}}\cdot$	$\cdot\underset{\cdot}{\dot{\mathrm{Si}}}\cdot$	$:\underset{\cdot}{\dot{\mathrm{P}}}\cdot$	$:\underset{\cdot\cdot}{\dot{\mathrm{S}}}\cdot$	$:\underset{\cdot\cdot}{\ddot{\mathrm{Cl}}}\cdot$	$:\underset{\cdot\cdot}{\ddot{\mathrm{Ar}}}:$

Lewis Dot Structures for Ions

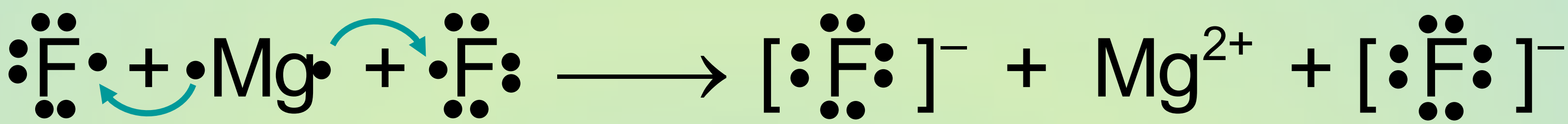
- Write the electron configuration and the Lewis symbol for N^{3-}.

- The electron configuration for N is $[He]2s^22p^{3.}$
- By gaining 3 electrons, the configuration changes to $[He]2s^22p^6$.

Use brackets $\left[:\ddot{\underset{\cdot\cdot}{N}}:\right]^{3-}$ ***Don't forget the charge***

Lewis Dot Structures for Ions

Use Lewis electron-dot symbols to represent the transfer of electrons from magnesium to fluorine atoms to form ions with noble-gas configurations.



Lewis Dot Structures for Molecules

- An electron pair is represented by two dots.
- An electron pair that is *between* two atoms is a bonding pair. It can also be represented by one line for each bonding pair.
- Electron pairs that are not bonding are non-bonding or lone pair electrons.

Follow the Octet Rule

- In forming covalent bonds, atoms tend toward having a full eight electrons in their valence shell. This tendency is called the **octet rule.**
- Hydrogen is an exception to the octet rule; it has two electrons in its valence shell.

Writing Lewis Dot Structures for Molecules

1. Calculate the number of valence electrons.
2. Write the skeleton structure of the molecule or ion.
3. Distribute electrons to the atoms surrounding the central atom or atoms to satisfy the octet rule.
4. Distribute the remaining electrons as pairs to the central atom(s).
5. H is never at the center
6. Element with lowest electronegativity is generally at the center

Covalent Bond

A **single bond** is a covalent bond in which one pair of electrons is shared by two atoms.

A **double bond** is a covalent bond in which two pairs of electrons are shared by two atoms.

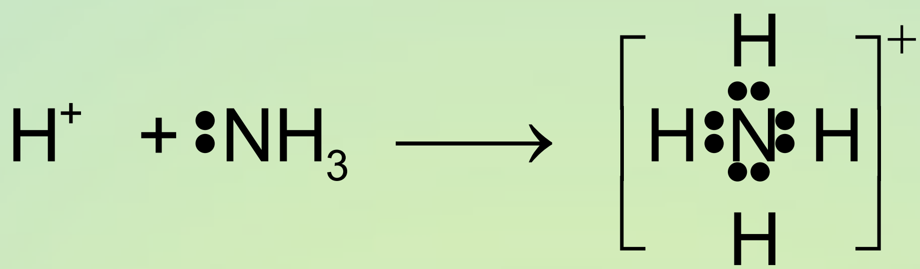
A **triple bond** is a covalent bond in which three pairs of electrons are shared by two atoms.

Double bonds form primarily with C, N, O, and S atoms.

Triple bonds form primarily with C and N atoms.

Coordinate Covalent Bond

This type of covalent bond is formed when both electrons of the bond are donated by one atom.



The two electrons forming the bond with the hydrogen on the left were both donated by the nitrogen. Once shared, they are indistinguishable from the other N—H bonds.

Write the Lewis structure for sulfur dichloride, SCl_2.

1. Determine the total number of valence electrons.

   Valence electrons in S = 6

   Valence electrons in Cl = 7

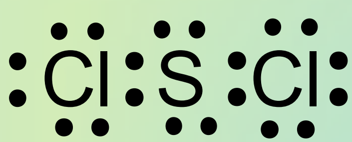
   Total number of valence electrons = 20

2. Write the skeletal structure.

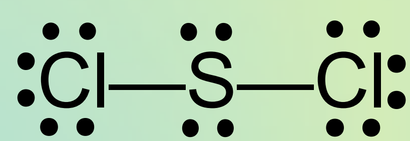
   Cl : S : Cl   or   Cl – S – Cl

3. Subtracting these 8 pairs from the total gives a difference of 2 pairs, which are placed n the central atom (S).

The final Lewis structure is:



or



Determine the Lewis dot structure of $COCl_2$.

Count valence electrons:

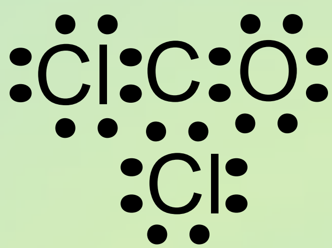
Valence electrons in C = 4

Valence electrons in O = 6

Valence electrons in Cl = 7

Total number of valence electrons = 24

Write the skeletal structure and complete octets on outside atoms.



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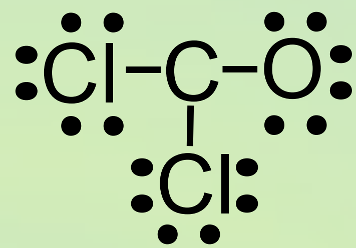
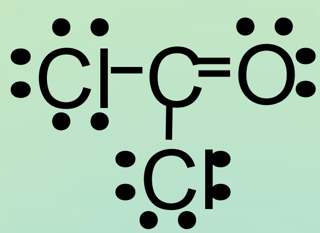


Since there are not enough electrons to give an octet, transfer an electron pair from an outer atom to form a double bond. This gives:

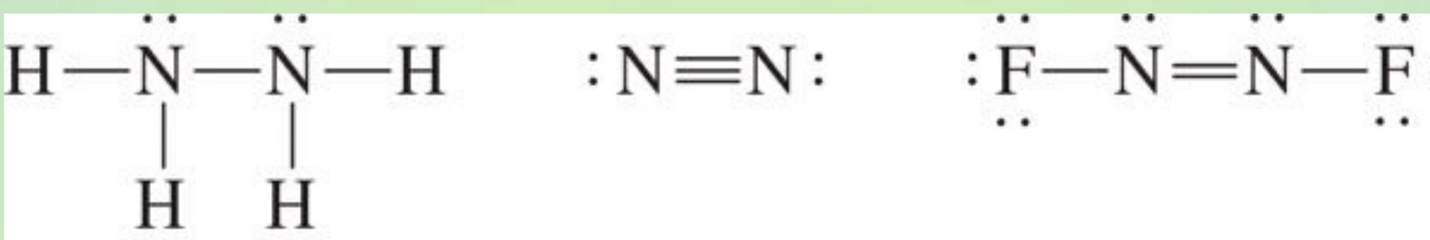


Place the double bond to the most electronegative element.

Sometimes there is more than one central atom.

Consider the molecules N_2H_4, N_2, and N_2F_2. Determine the molecule with the shortest nitrogen-nitrogen bond and the molecule with the longest nitrogen-nitrogen bond.

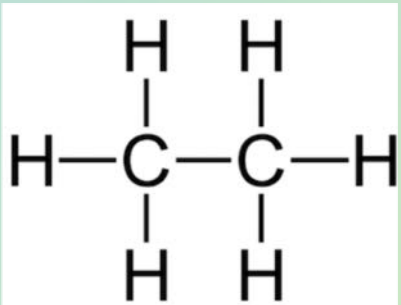
- Write the Lewis formulas.

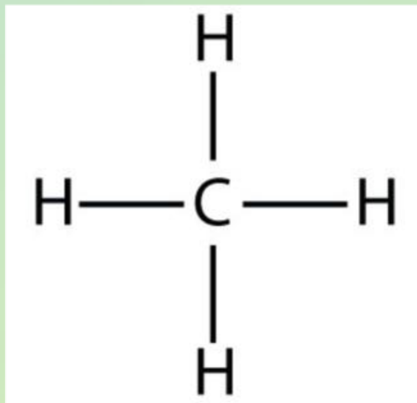


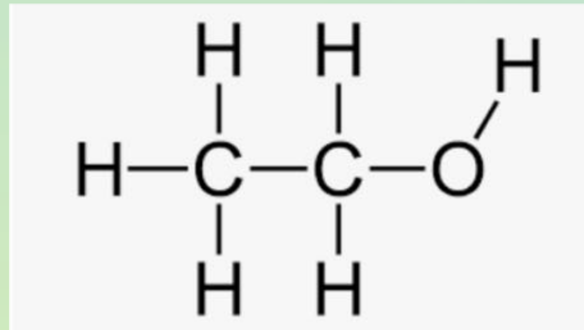
N_2H_4 N_2 N_2F_2

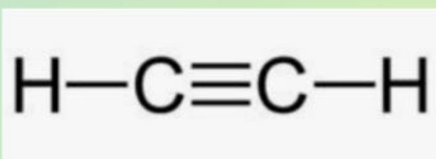
- The nitrogen bond should be shortest; it is a triple bond. The N_2H_4 bond should be longest; it is a single bond.

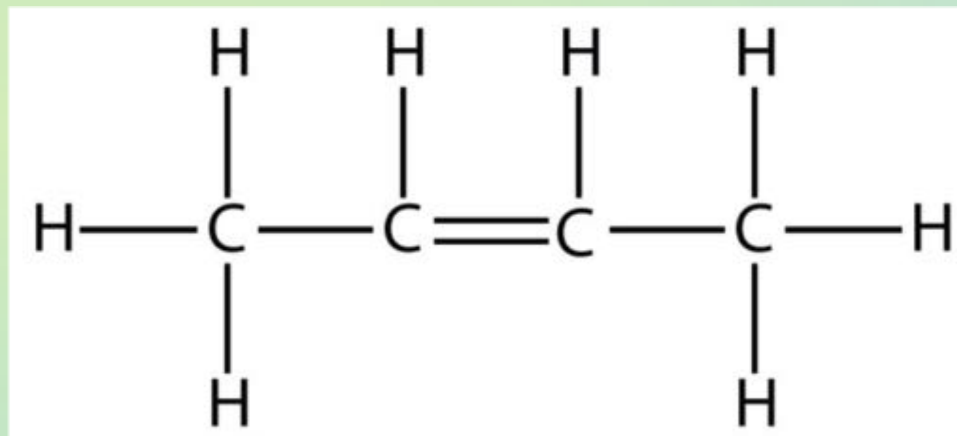
Sometimes there is more than one central atom.







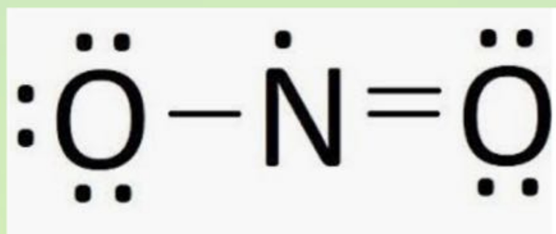




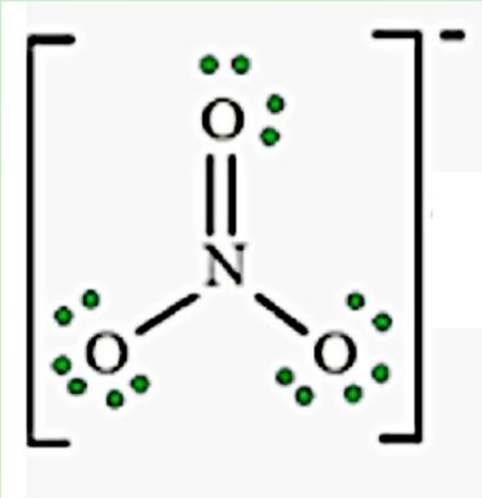
4 bonds to carbon
2 bonds to O
1 bond to hydrogen

Sometimes molecules and ions have an odd number of electrons.

Draw the Lewis structure for nitrogen dioxide.

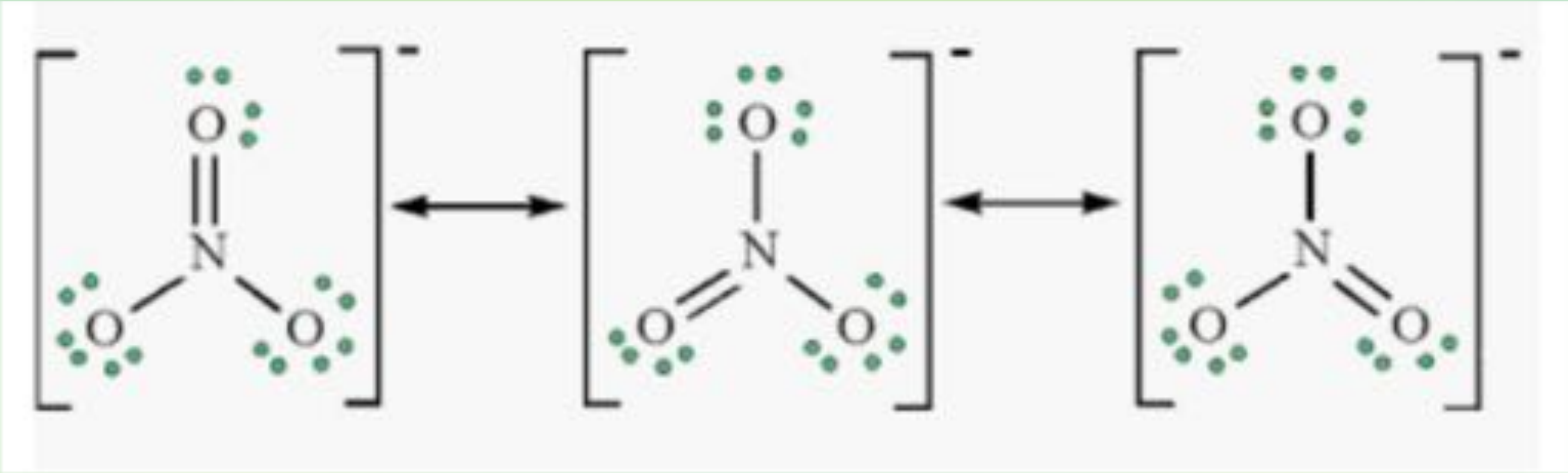


Draw the Lewis structure for nitrate ion.



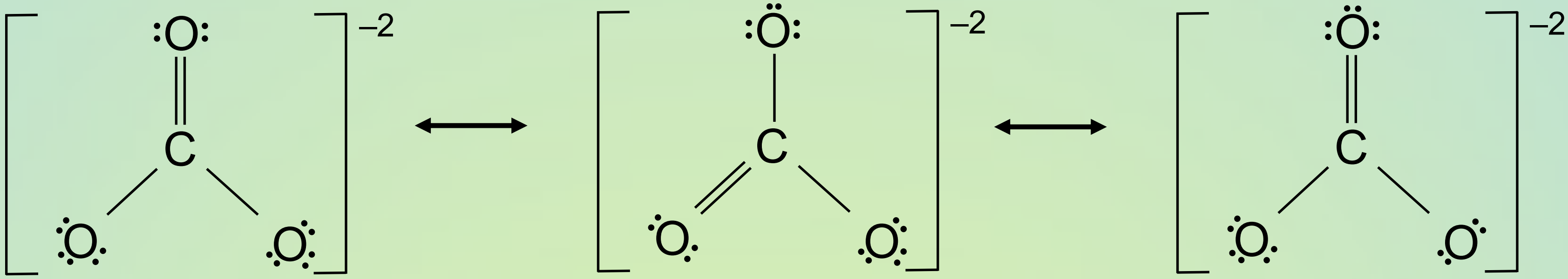
Resonance structures

Draw the Lewis structures for nitrate ion. Nitrate ion has three resonance structures.



Resonance structures represent delocalized bonding.

Draw the Lewis structure of the carbonate ion, CO_3^{2-}.



Use a double-headed arrow to connect resonance structures.

Delocalized bonding is a type of bonding in which a bonding pair of electrons is spread over a number of atoms rather than being localized between two atoms.

Atoms with *d* orbitals can accommodate more than eight electrons.

Determine the electron-dot formula of xenon tetrafluoride, XeF_4.

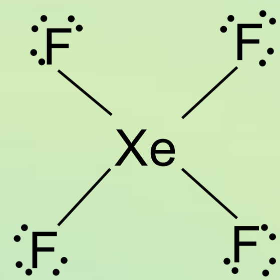
- Determine the total number of valence electrons.

Valence electrons in Xe = 8

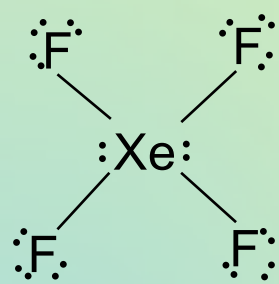
Valence electrons in F = 7

Total number of valence electrons = 36

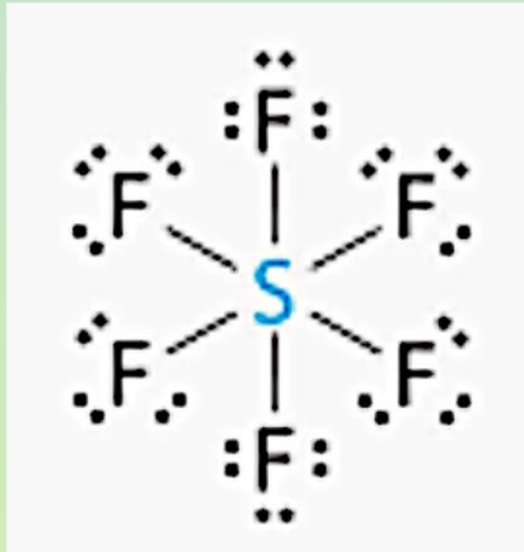
- Draw the electron structure and satisfy the octet rule.



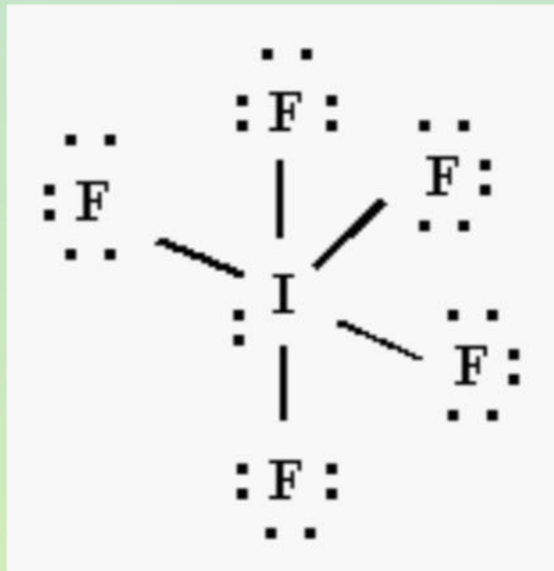
- Add the remaining two pairs of electrons are placed on the Xe atom:



Draw the Lewis structure for sulfur hexafluoride.



Draw the Lewis structure for iodine pentafluoride.



The formal charge assigns a charge to each atom based on the number of electrons in bonds and lone pairs.

$$\text{Formal charge} = \text{valence electrons on free atom} - \tfrac{1}{2}\left(\text{number of electrons in bonds}\right) - \left(\text{number of lone-pair electrons}\right)$$

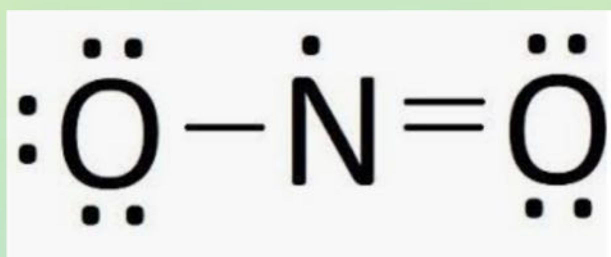
1. Half of the electrons of a bond are assigned to each atom in the bond.
2. Both electrons of a lone pair are assigned to the atom to which the lone pair belongs.
3. The sum of the formal charges on the atoms equals the charge on the formula.
4. Whenever you can write different Lewis structures for a molecule, choose the one having the formal charges closest to zero.
5. When two proposed Lewis formulas have the same magnitudes of formal charges, choose the one having the negative formal charge on the more electronegative atom.
6. When possible, avoid Lewis formulas with like charges on adjacent atoms.

The formal charge on an atom in a Lewis structure assumes bonding electrons are equally shared and lone pairs belong completely to one atom.

Formal charge = valence electrons on free atom – ½ (number of electrons in bonds) – (number of lone-pair electrons)

Calculate the formal charges on atoms in nitrogen dioxide.

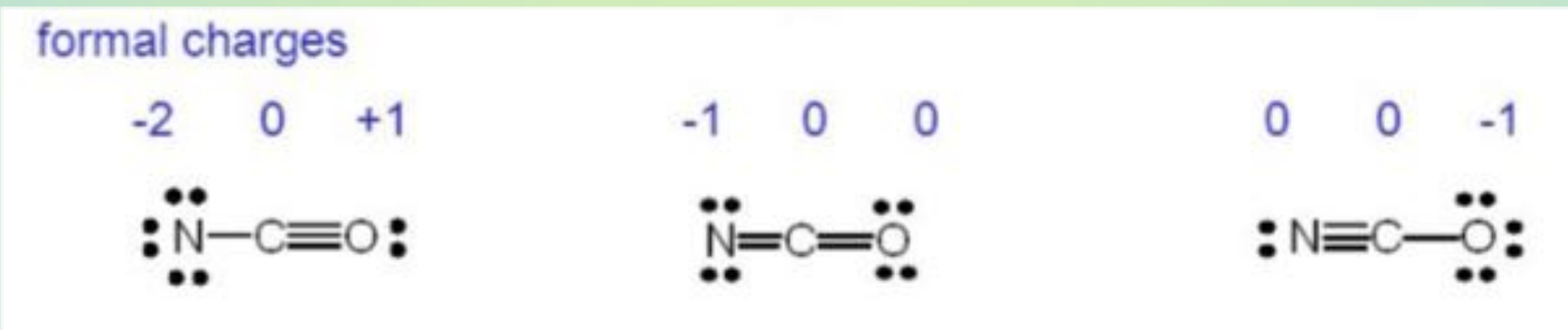
-1 1 0



Use formal charges to determine the best Lewis structure for cyanate ion, NCO^-.

Formal charge = valence electrons on free atom – ½ (number of electrons in bonds)

– (number of lone-pair electrons)



Draw the Lewis structure for sulfate ion.

- Sulfur in sulfate has an expanded octet.
- Sulfate has six resonance structures.
- Sulfate ion is very stable due to its six resonance structures allowing electron delocalization.

Write the Lewis structure for the sulfuric acid molecule, H_2SO_4

- Assume a S is the central atom.
- H is placed on O

- Calculate the formal charge of each atom bonded to an H atom.

$$\text{Formal charge} = \text{valence electrons on free atom} - \tfrac{1}{2}\left(\text{number of electrons in bonds}\right) - \left(\text{number of lone-pair electrons}\right)$$

- Double bonds on O produce lower formal charges on O.

Draw the Lewis structure for phosphate ion. Phosphate ion has an expanded octet and five resonance structures.