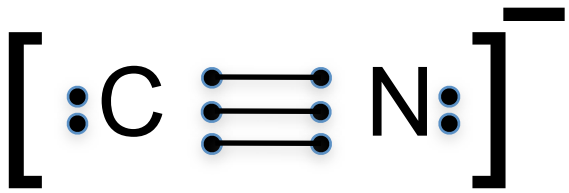
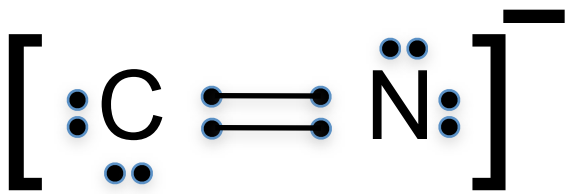
Drawing Lewis Structures

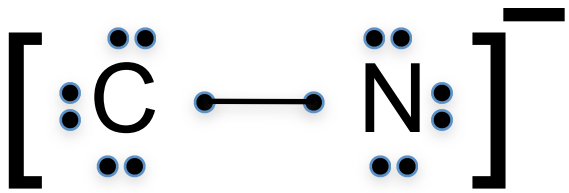
The Octet Rule

1. Determine the arrangement of atoms in the molecule
   - Some elements work only as terminal atoms (H, F, Cl, etc)
   - Some are particularly good as internal atoms (C, N)
2. Determine the total number of valence electrons in all atoms (don't forget charge!)
   - Add numbers of all valence electrons
   - If negatively charged, add the appropriate number of electrons
   - If positively charged, subtract the appropriate number of electrons
3. Place one pair of electrons between each pair of bonded atoms, forming a bond
4. Use remaining electron pairs as lone pairs around each terminal atom, to complete its octet
5. Place electrons around the central atom. If you run out, "share" electron pairs from terminal atoms

Which is the best Lewis structure?

1) 

2) 

3) 

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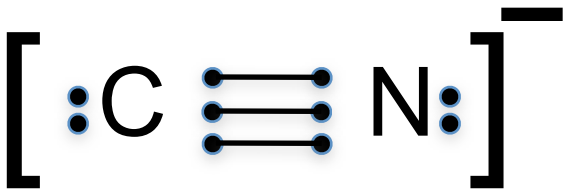
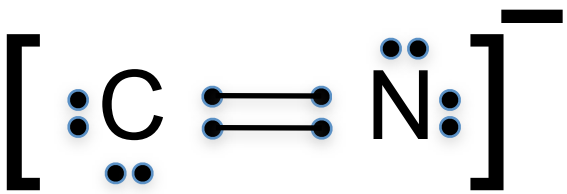
1) 

10 e^-

C – $[He]2s^2 2p^2$

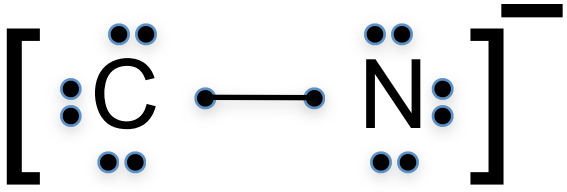
N – $[He]2s^2 2p^3$

2) 

12 e^-

Plus 1 e^-

Too many electrons

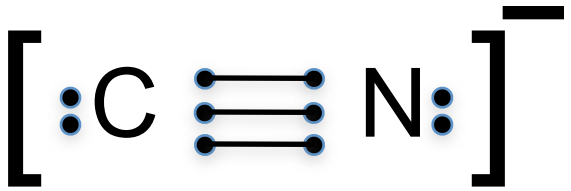
3) 

14 e^-

Too many electrons

Isoelectronic species

(10 electrons)



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Plus 1 e^-

2+3+2+3

2+2+2+4

2+2+2+3**+1**

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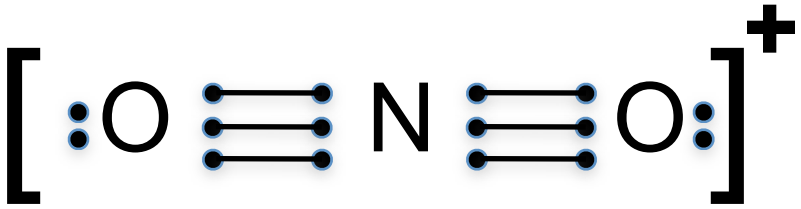
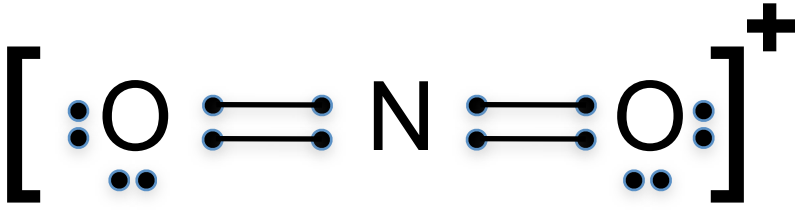
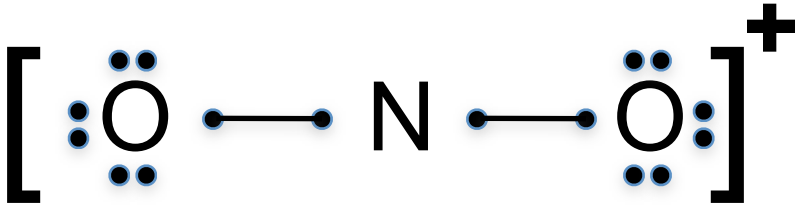
1)  16 e^-

2)  16 e^-

3)  16 e^-

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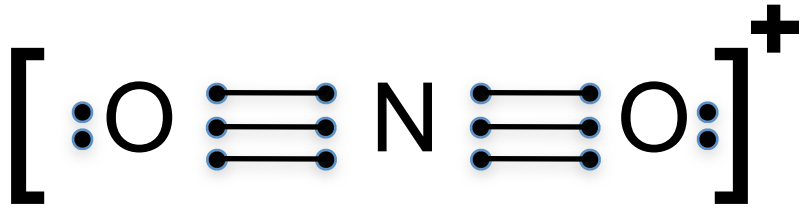
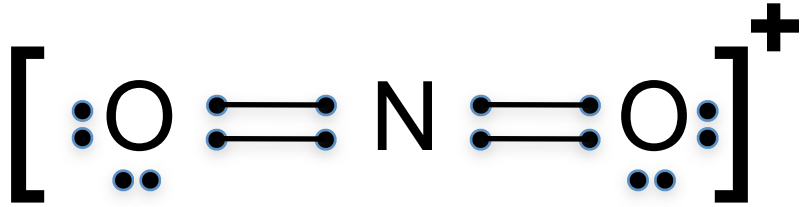
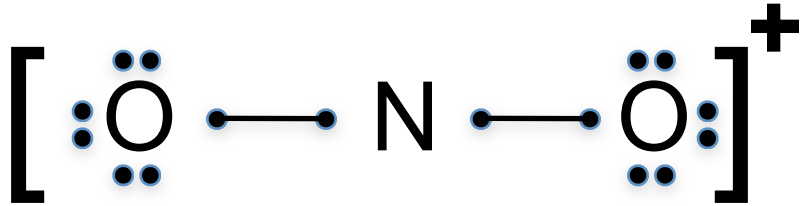
1)  16 e^-

2)  16 e^-

3)  16 e^-

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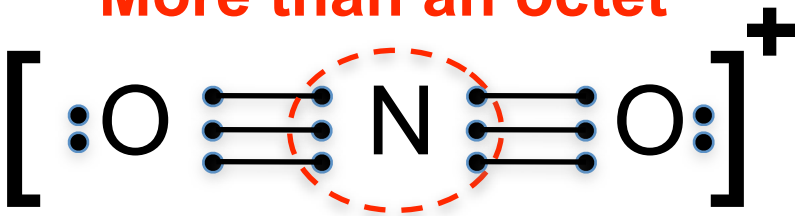
Minus 1 e^-

Which is the best Lewis structure?

More than an octet

1) 

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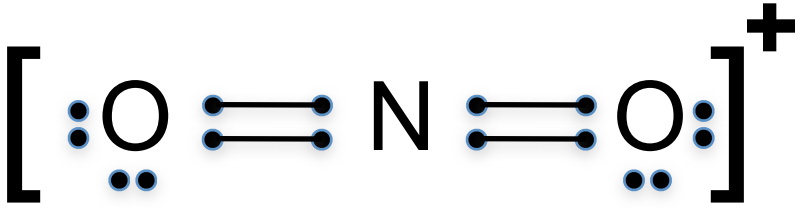
2) 

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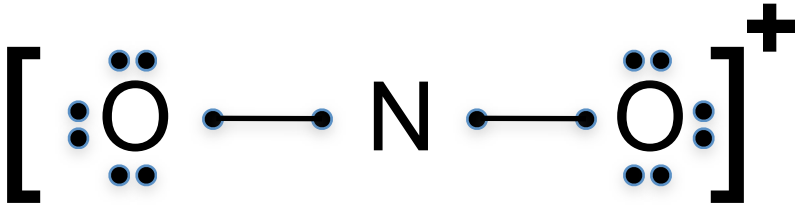
3) 

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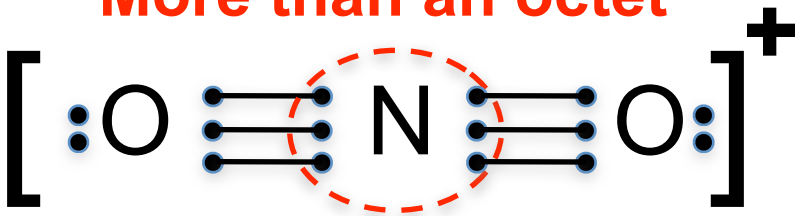
1) 

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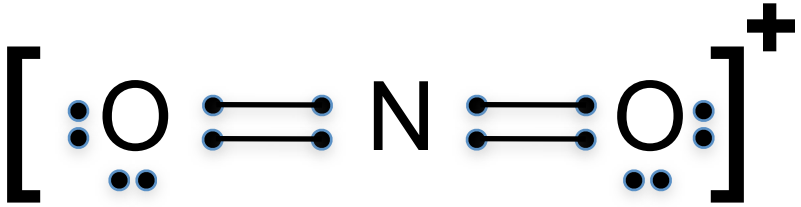
2) 

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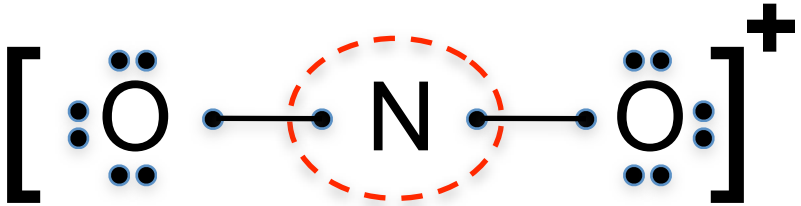
3) 

16 e^-

Less than an octet

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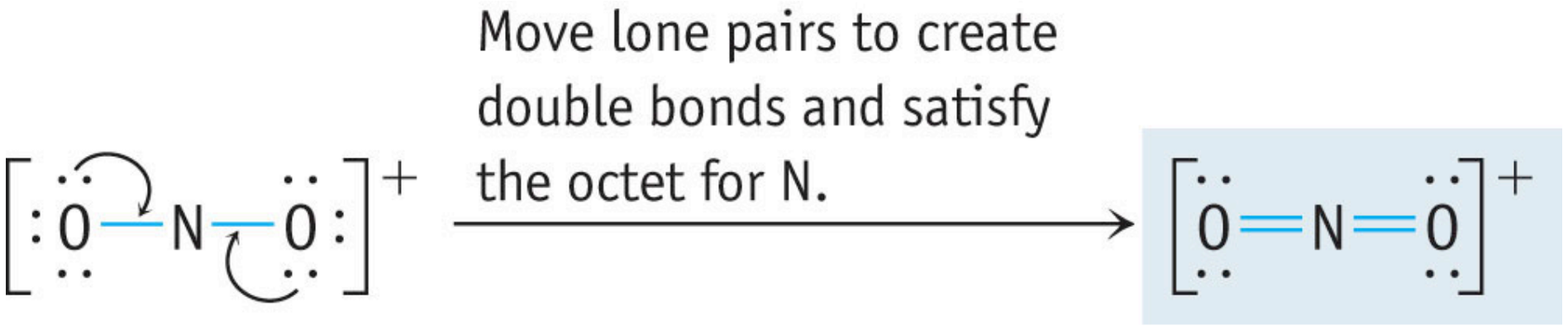
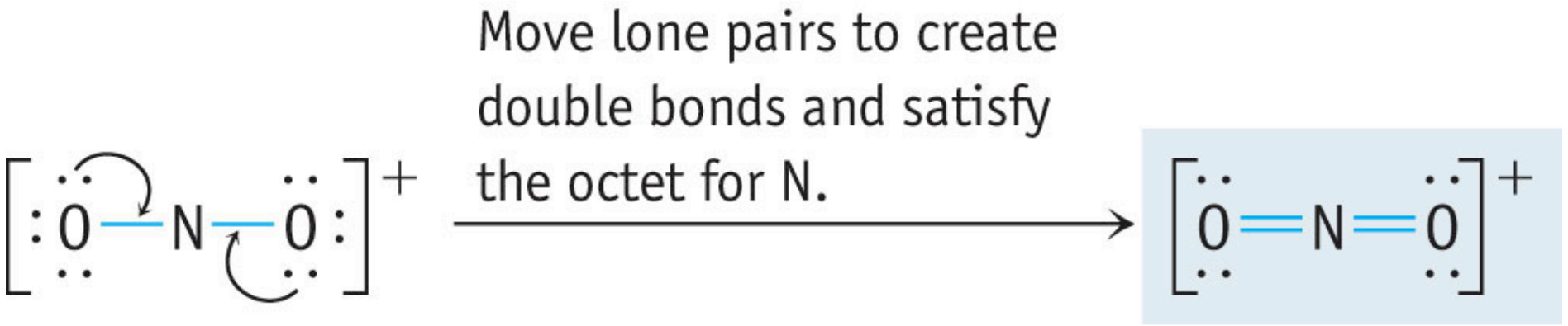
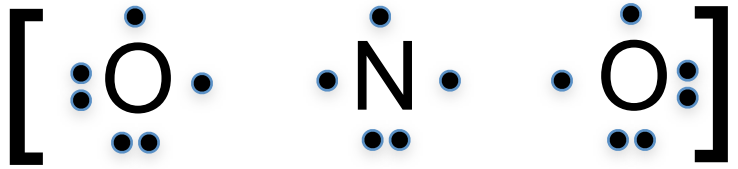
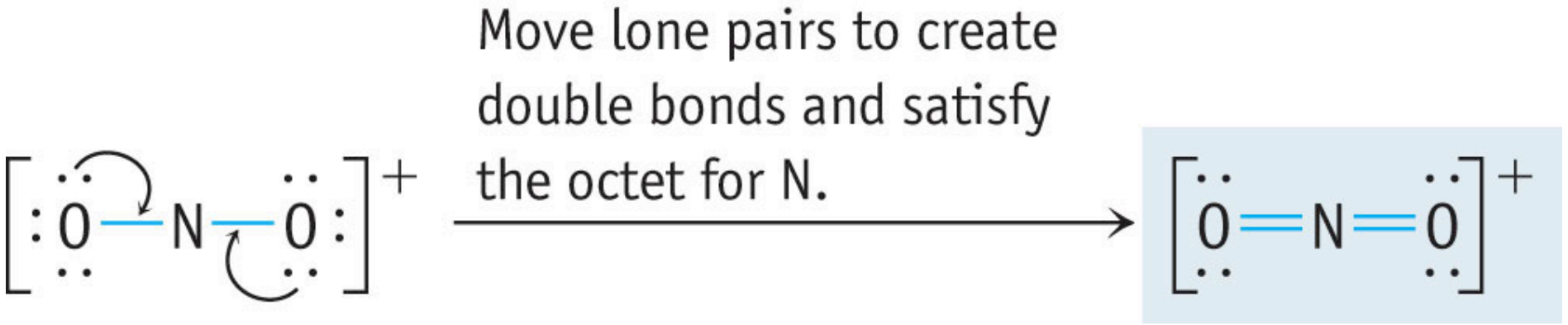
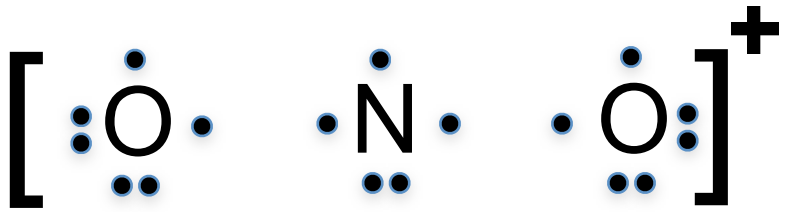
Minus 1 e^-

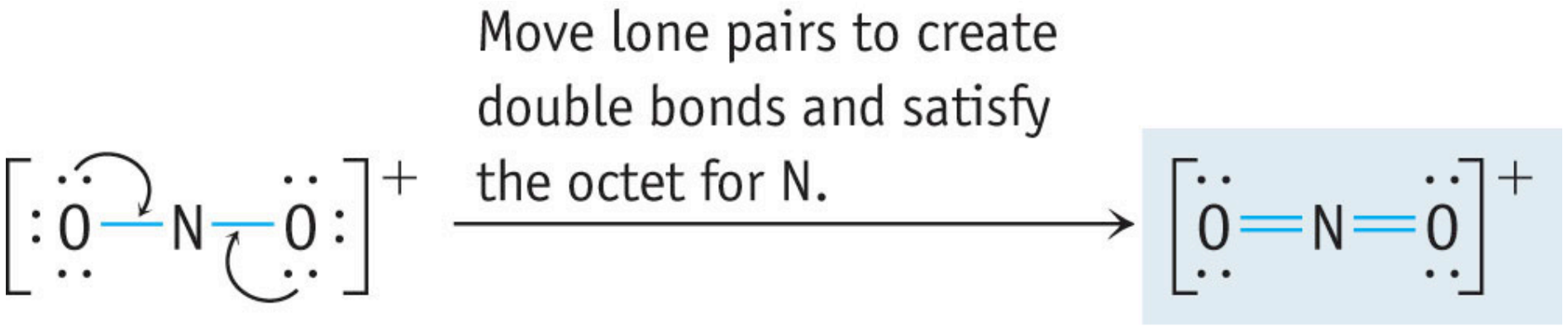


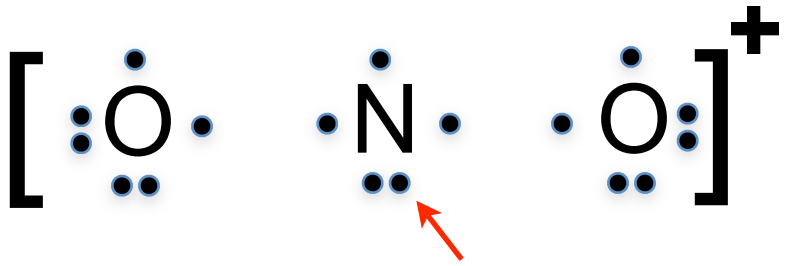


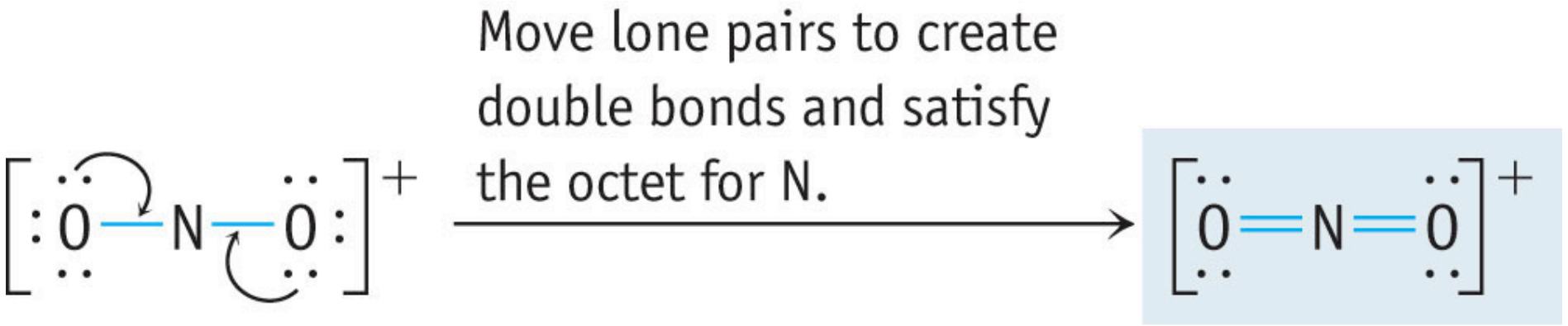


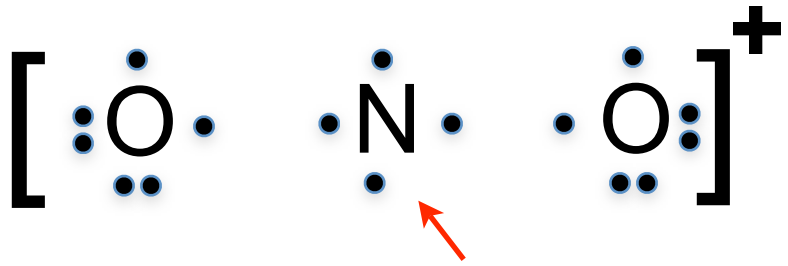


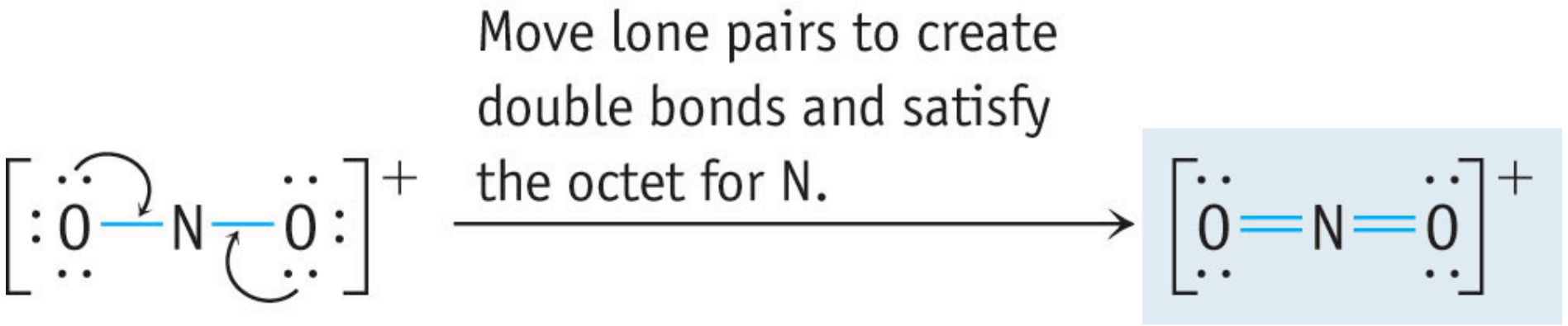


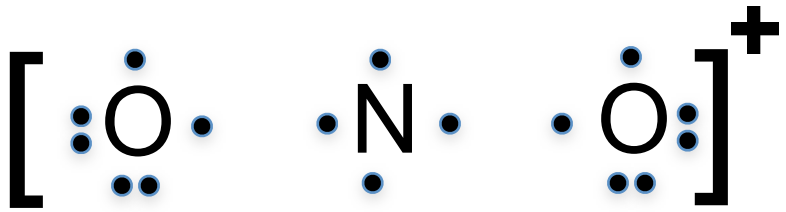


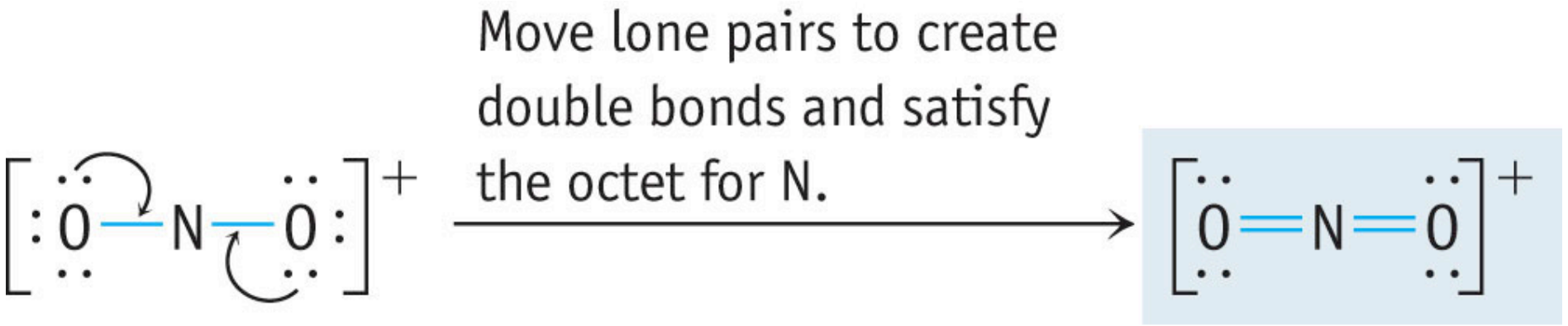


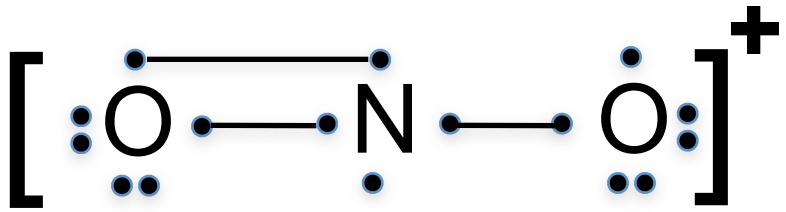


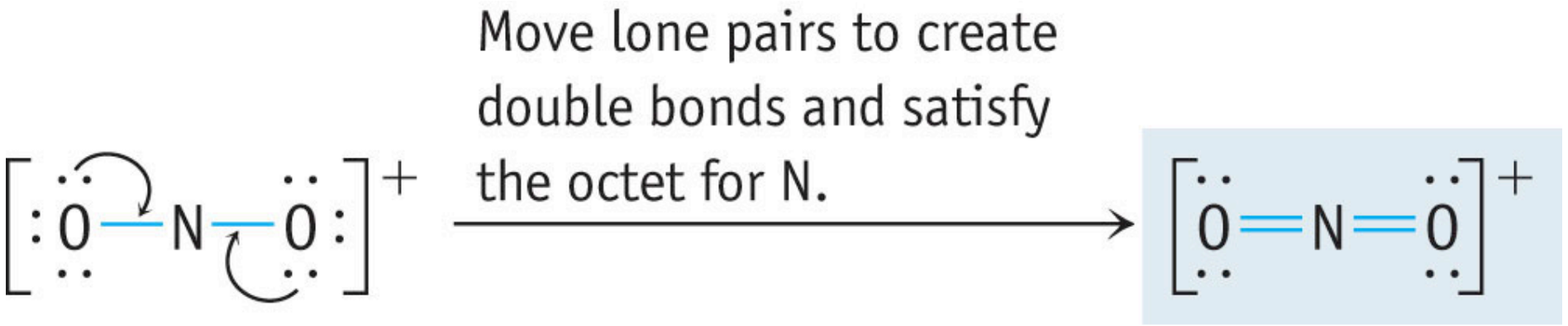


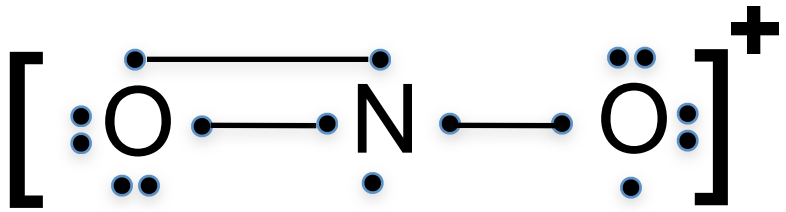


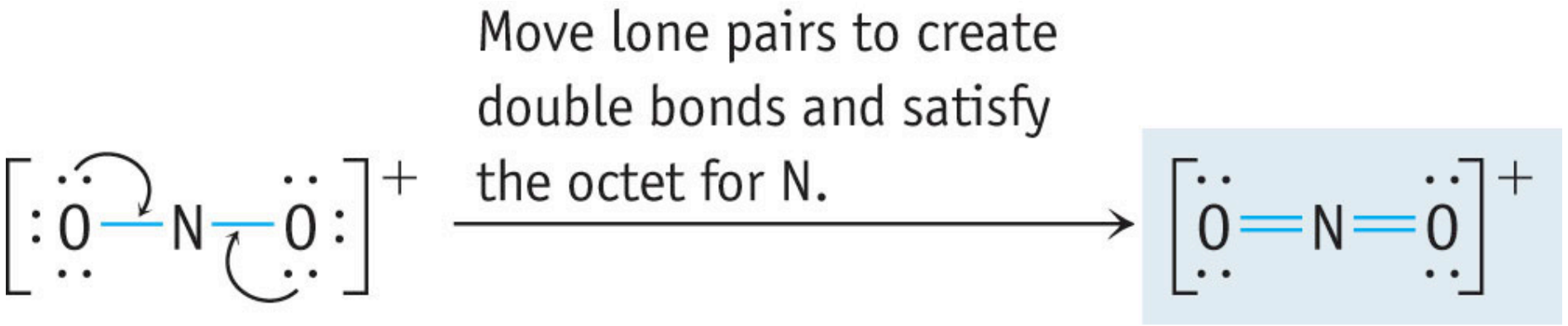


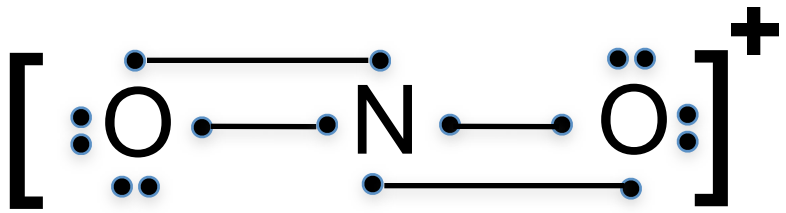


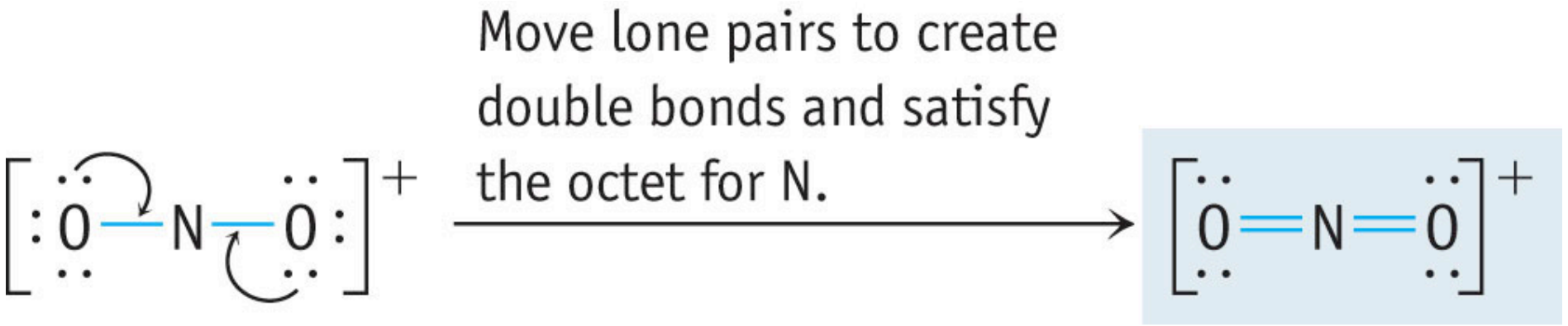


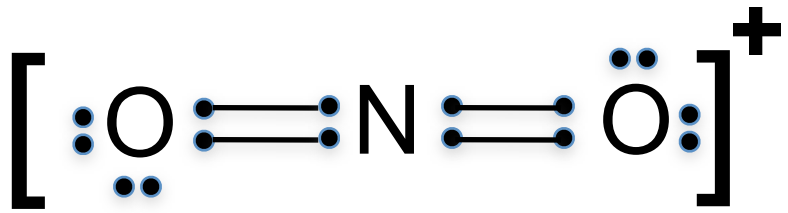




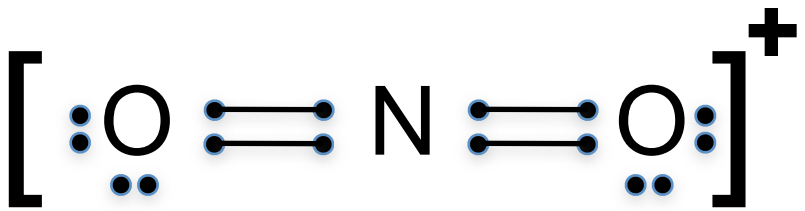








- Octet Rule - double count shared electrons
  - assumes both atoms get both electrons



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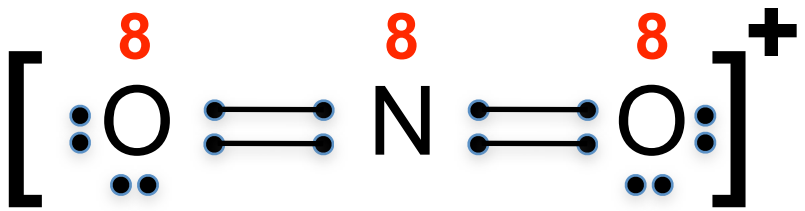

Formal Charge

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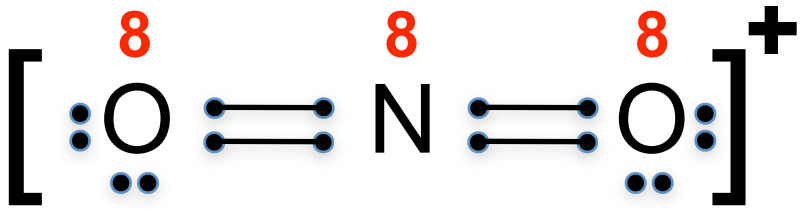
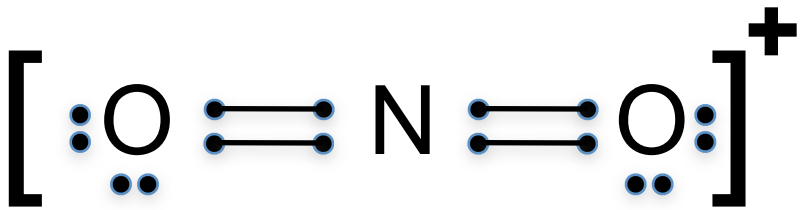


- Formal Charge - Evenly split electrons in a bond
  - one to one atom, one to the other



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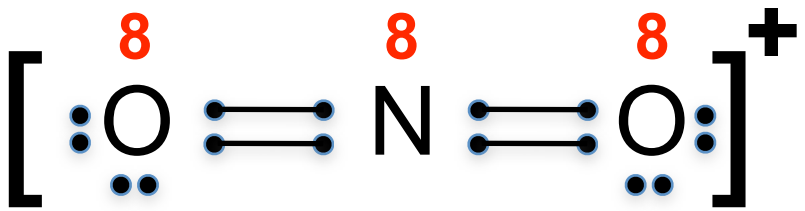

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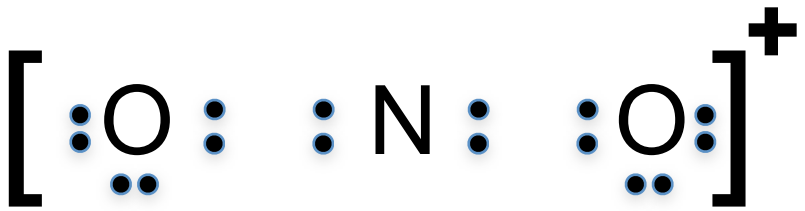

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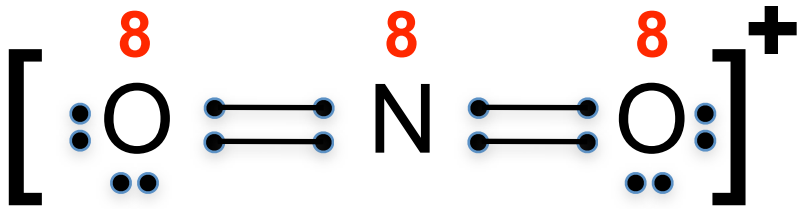

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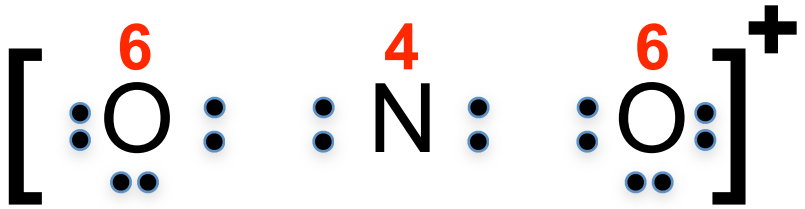

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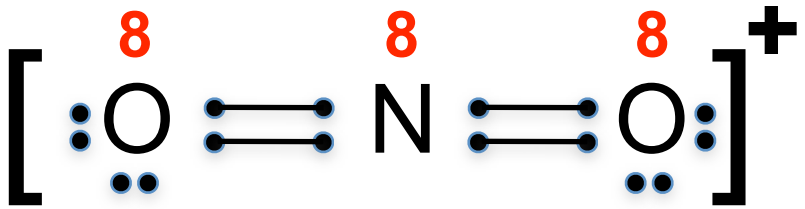

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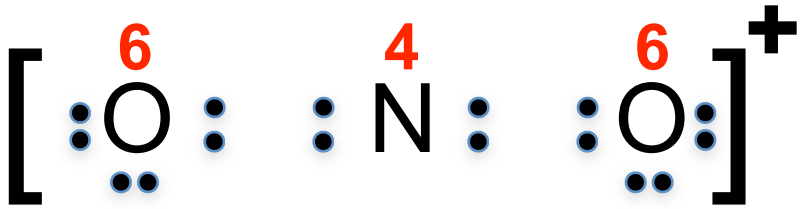

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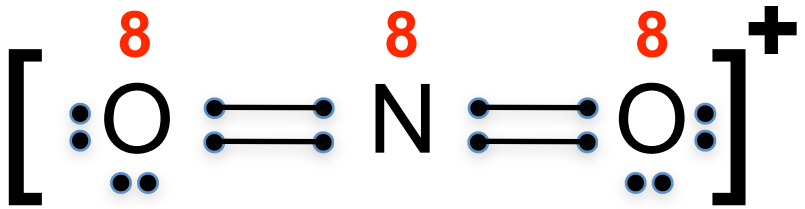

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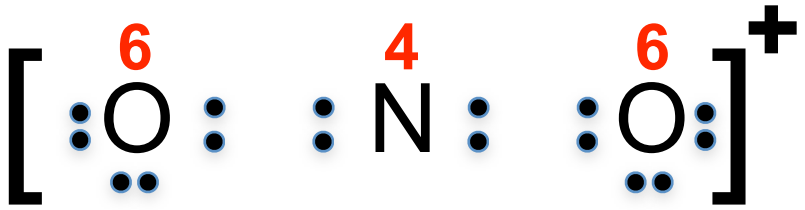

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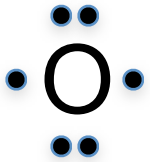

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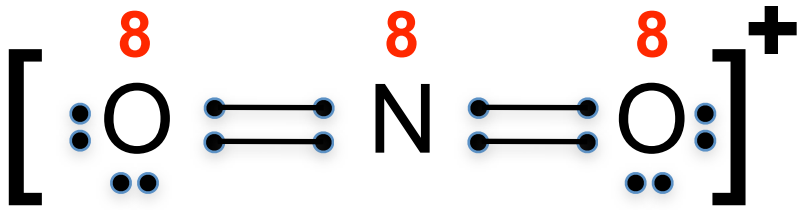

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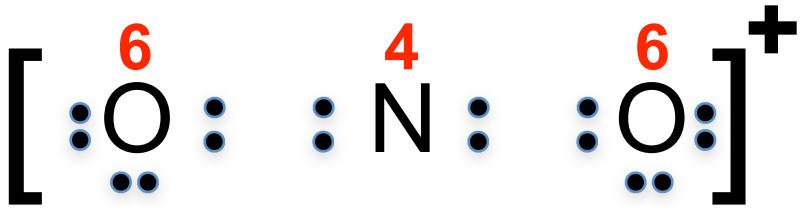
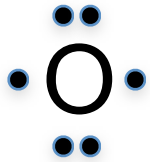
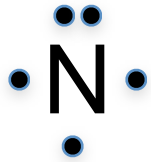


Brought to the party

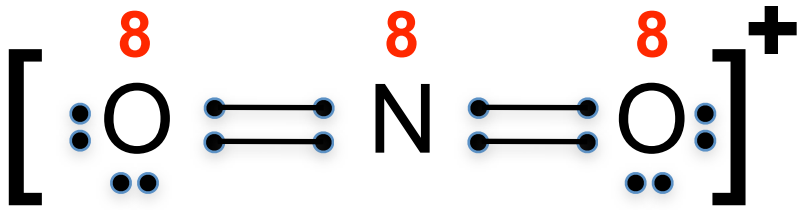




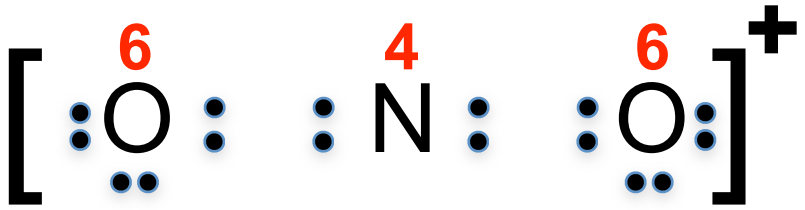
Formal Charge

versus Octet Rule

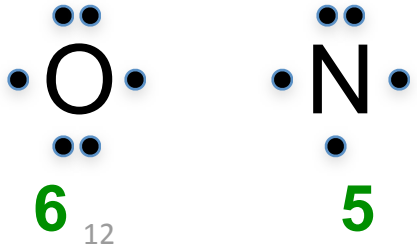
- Octet Rule - double count shared electrons
  - assumes both atoms get both electrons



- Formal Charge - Evenly split electrons in a bond
  - one to one atom, one to the other



Brought to the party



Formal Charge

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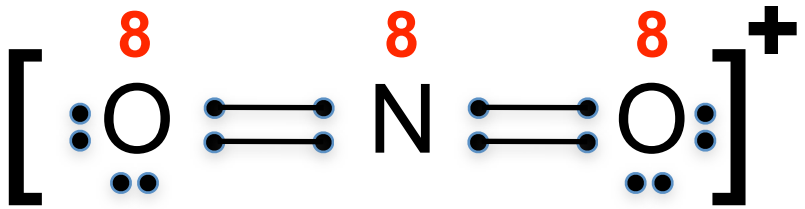

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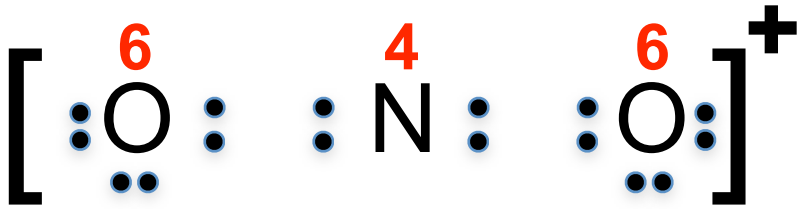
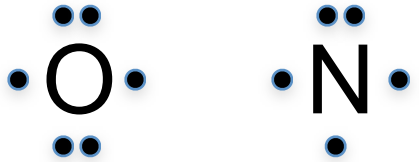


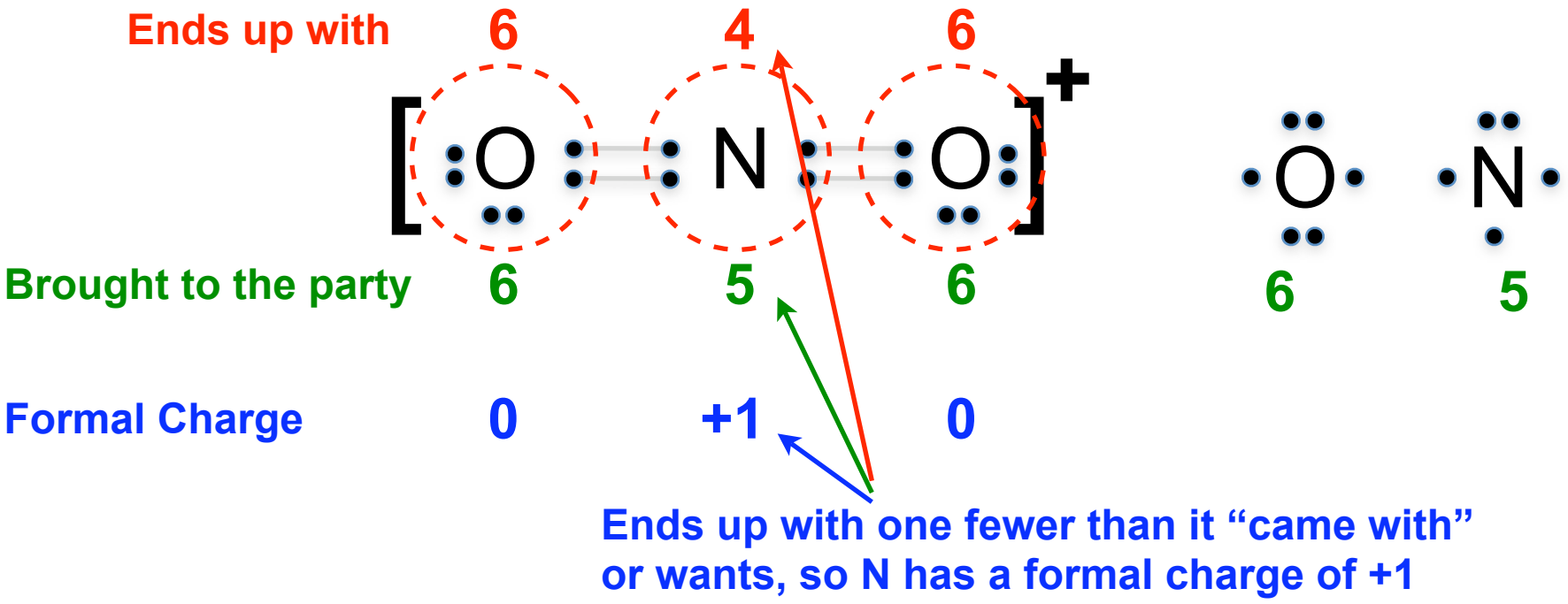
Brought to the party	6	5	6
Formal Charge	0	+1	0



Formal Charge

versus Octet Rule

Number of electrons brought to the party __ Number of electrons ending up with



Formal Charge

versus Octet Rule

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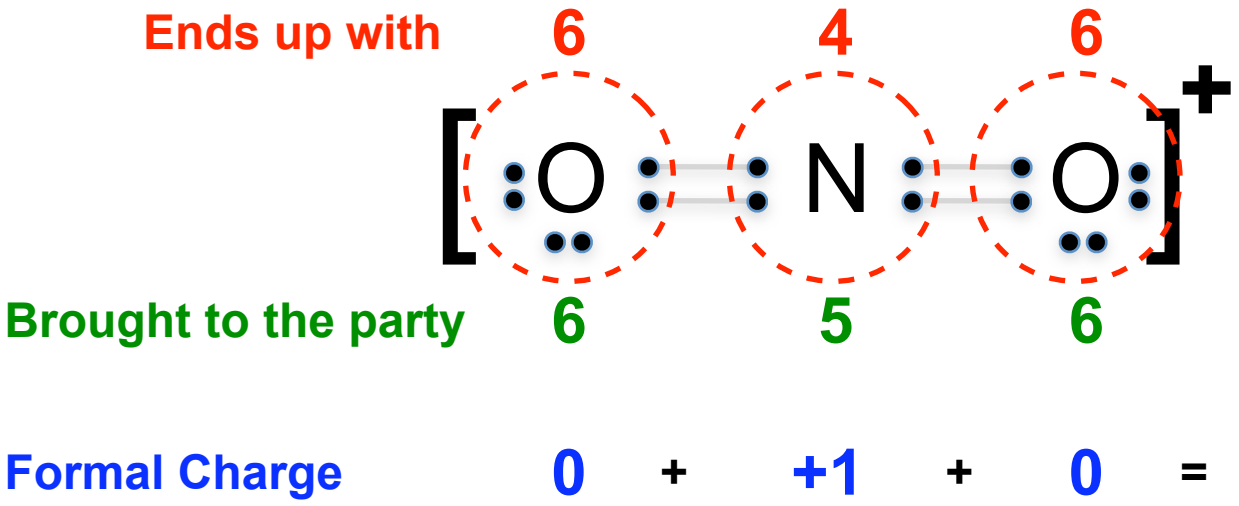


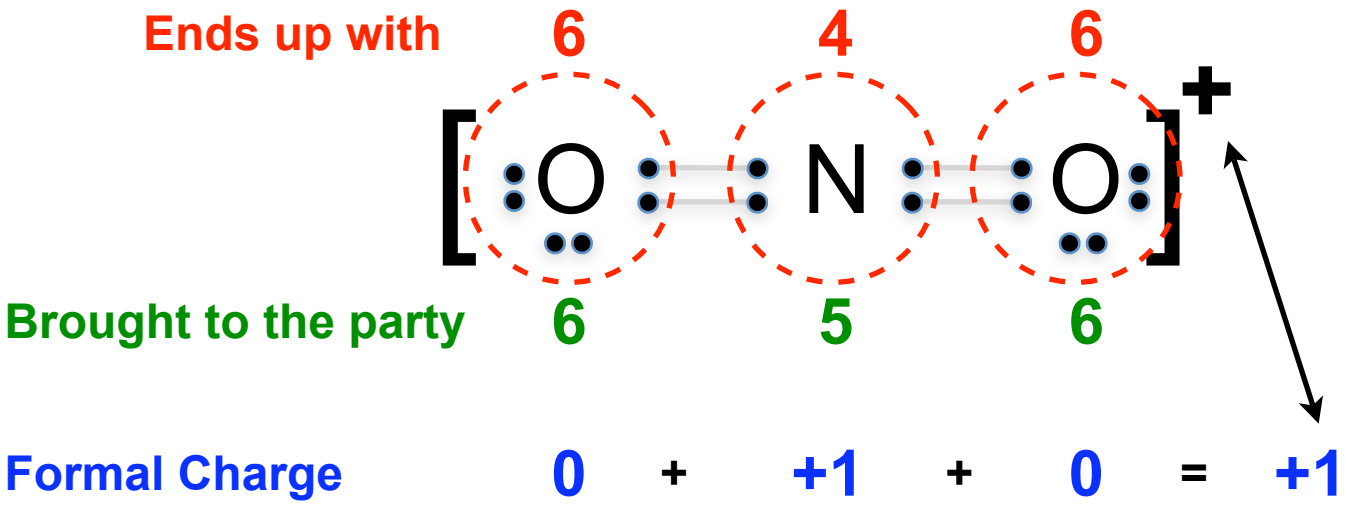
Formal Charge **0 + +1 + 0 = +1**

Note: sum of atom formal charges must equal the overall charge on the molecule

Formal Charge

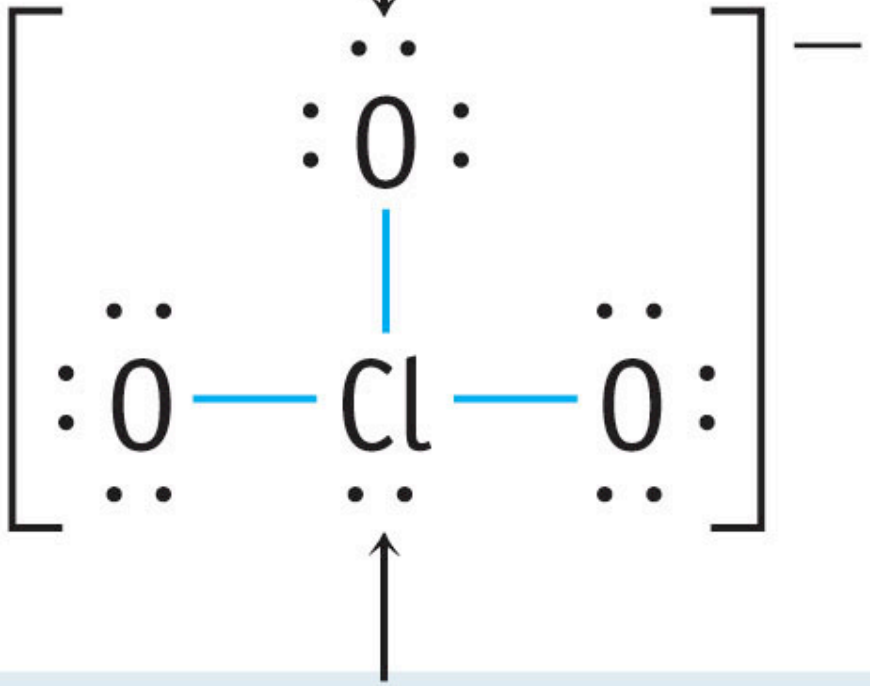
versus Octet Rule

Number of electrons brought to the party __ Number of electrons ending up with



Note: sum of atom formal charges must equal the overall charge on the molecule

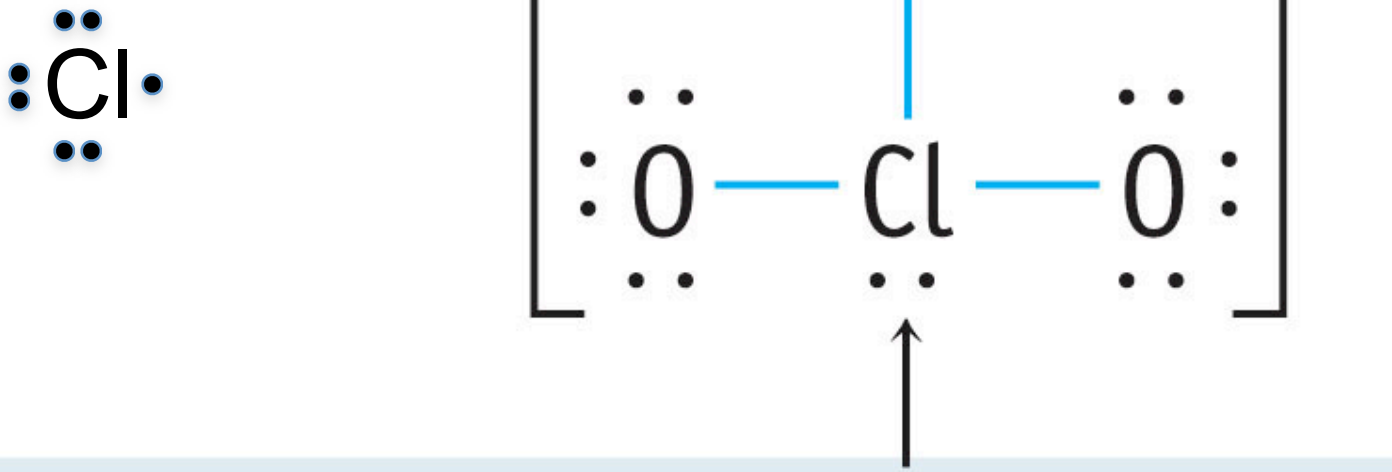
$$\text{Formal charge} = -1 = 6 - \left[6 + \tfrac{1}{2}(2)\right]$$



$$\text{Formal charge} = +2 = 7 - \left[2 + \tfrac{1}{2}(6)\right]$$

Example 8-3, p. 360

$$\text{Formal charge} = -1 = 6 - \left[6 + \tfrac{1}{2}(2)\right]$$



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Example 8-3, p. 360

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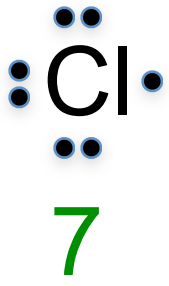
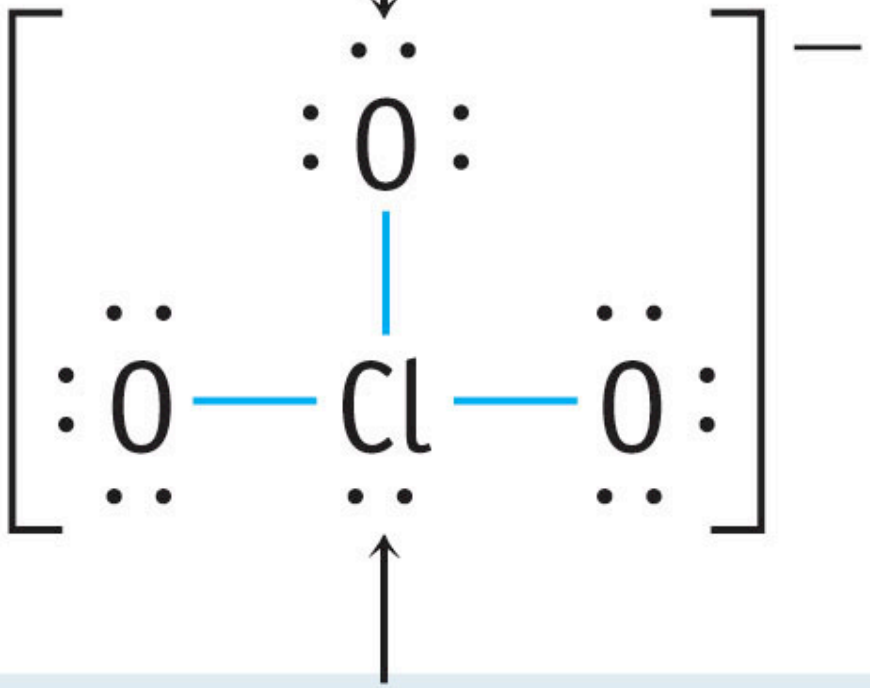


7



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Example 8-3, p. 360

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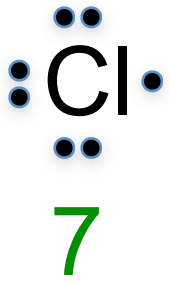

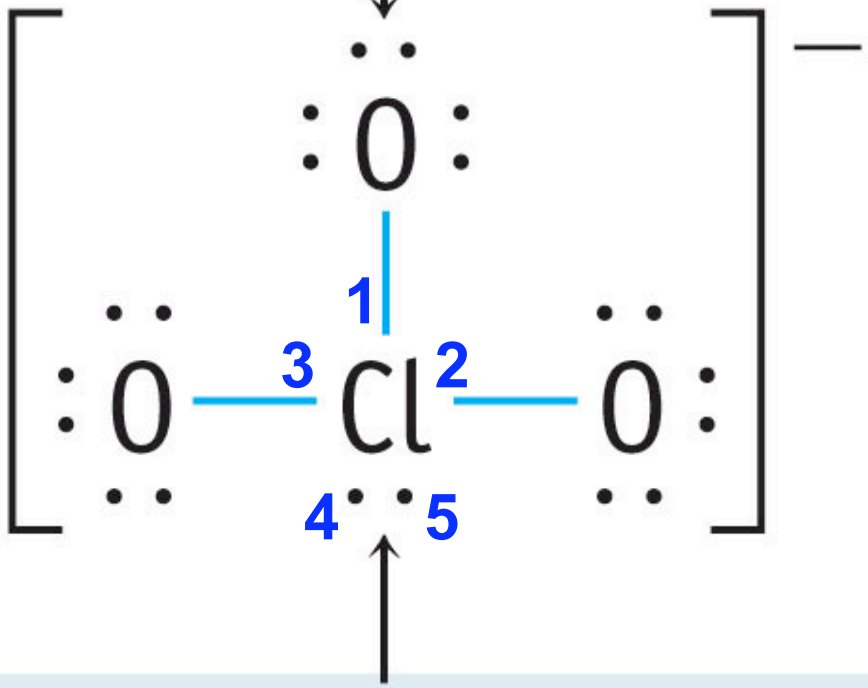

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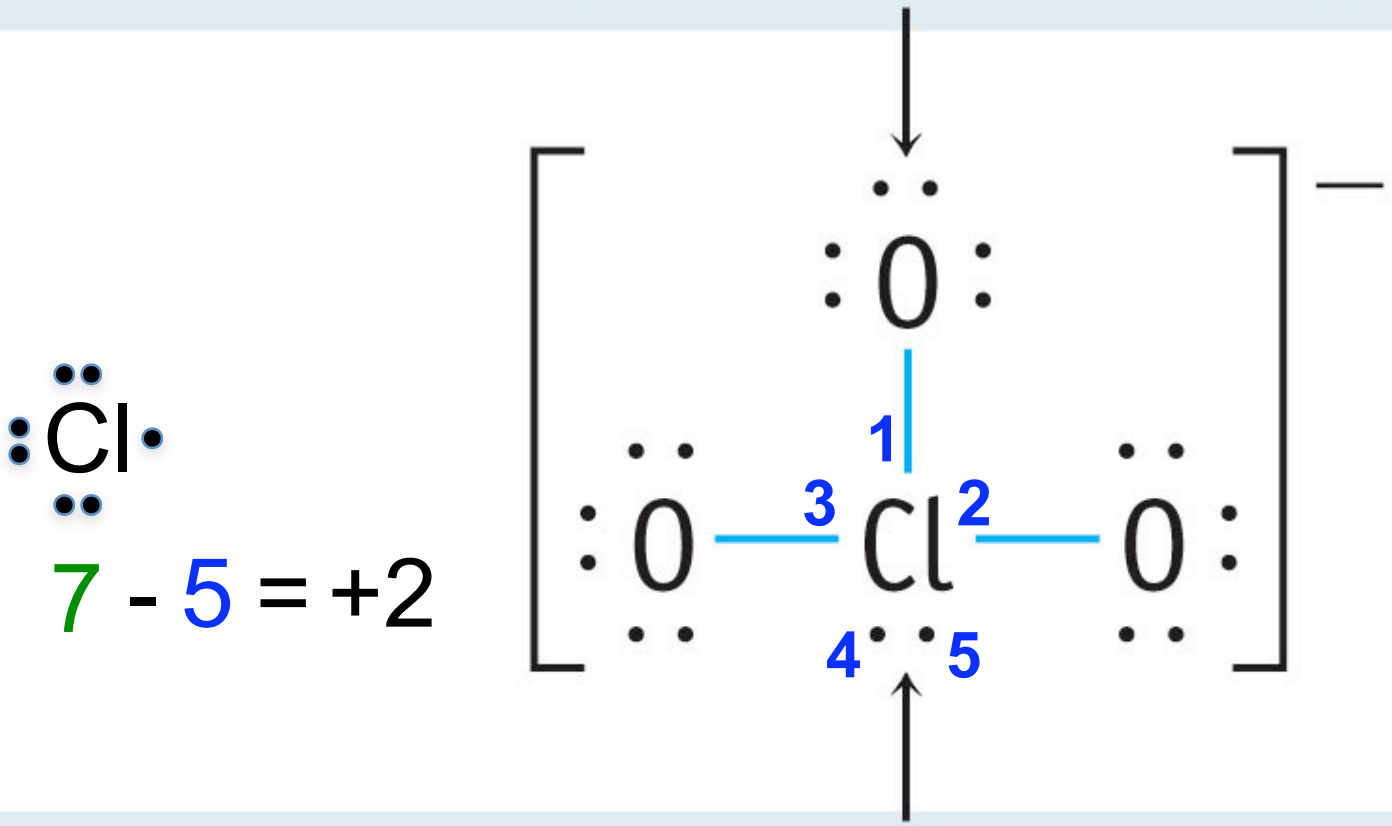

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Example 8-3, p. 360

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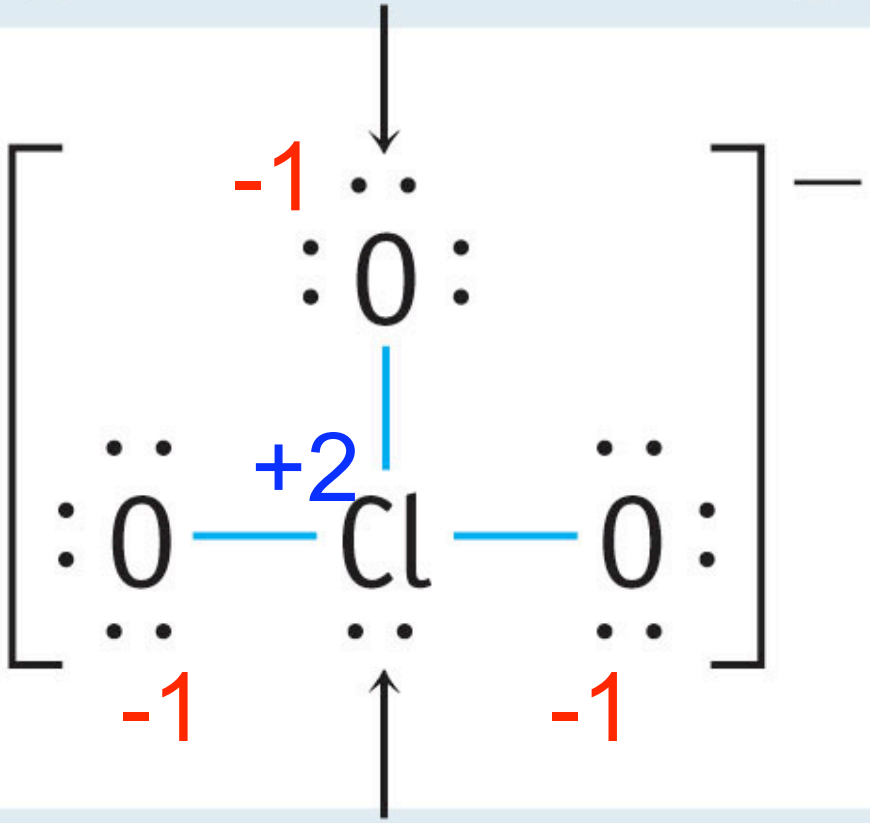

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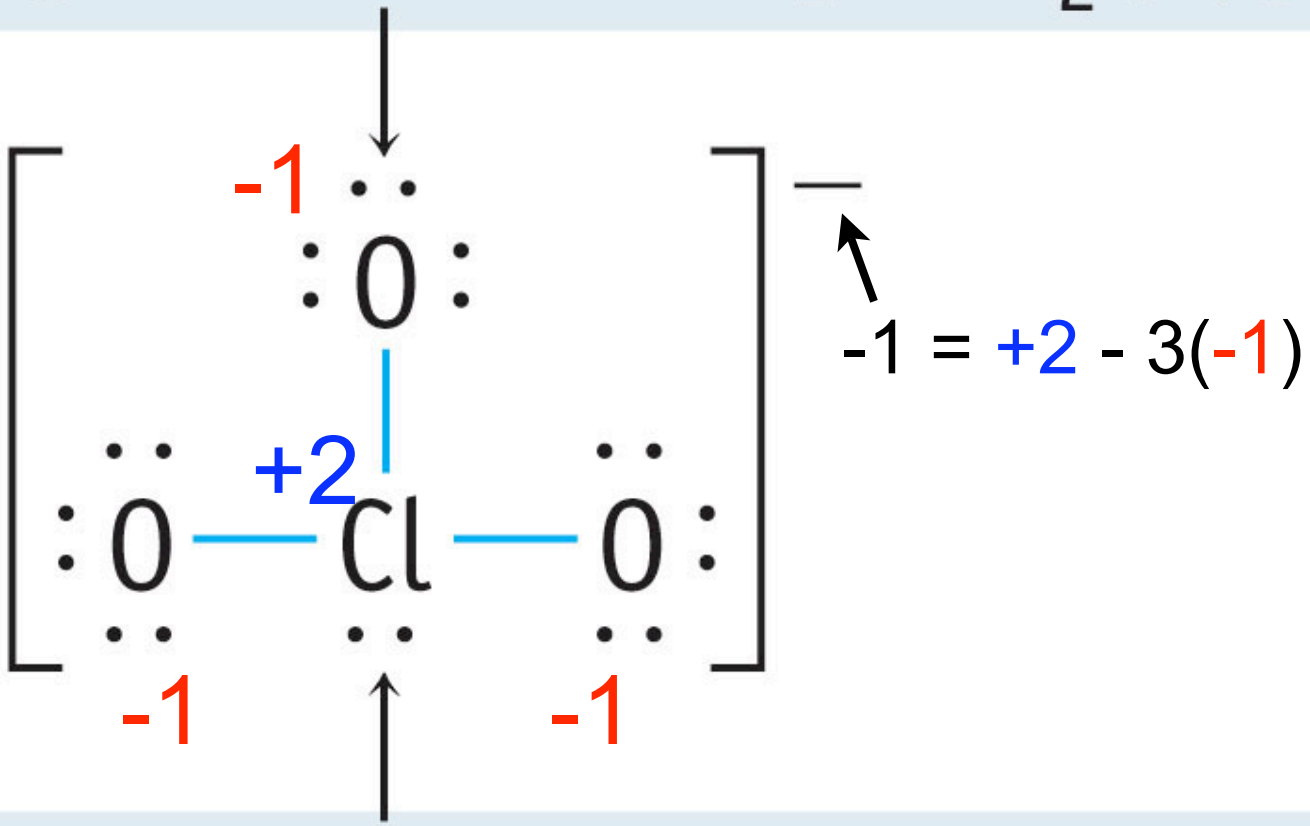
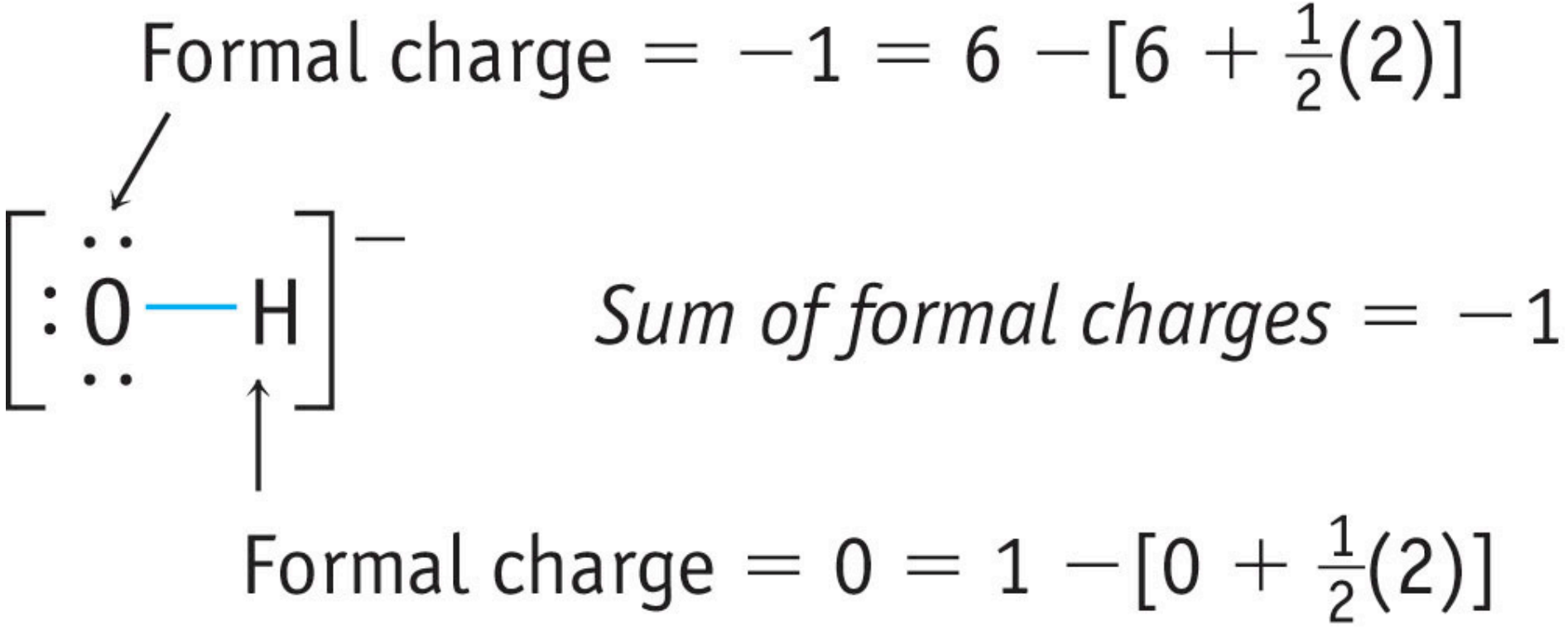


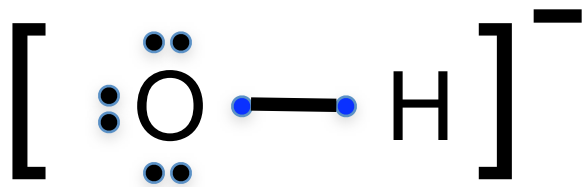
$$\text{Formal charge} = +2 = 7 - [2 + \tfrac{1}{2}(6)]$$

Example 8-3, p. 360

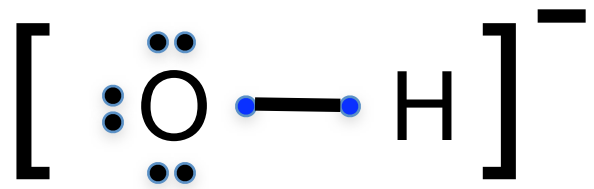


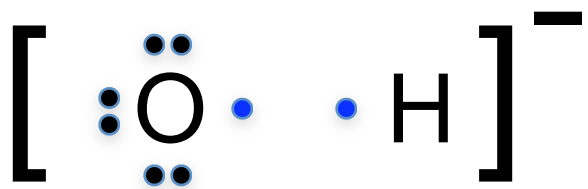
New concept: Oxidation Number

New concept



New concept



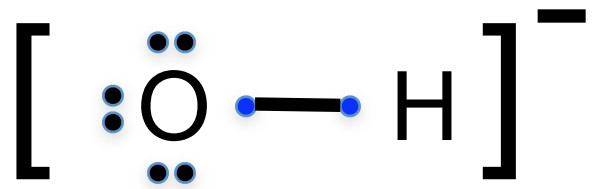


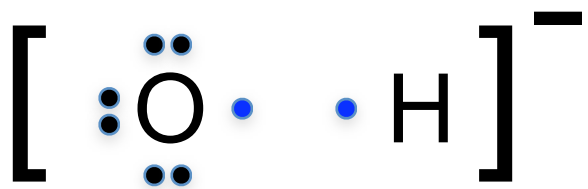
Formal Charge

Covalent assumption

Electrons shared,
one to each

New concept





Formal Charge

Covalent assumption

Electrons shared, one to each

Oxidation Number

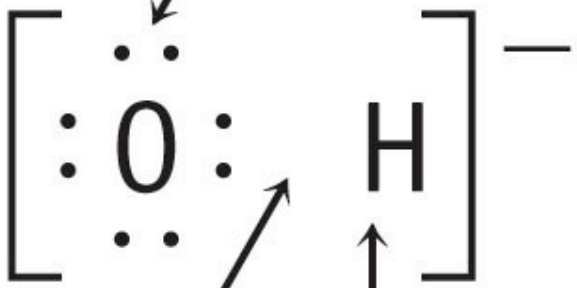
Ionic assumption

Both electrons transferred to the one who wants them more

Brought to the party 6

Ends up with 8

Oxidation number $= -2$



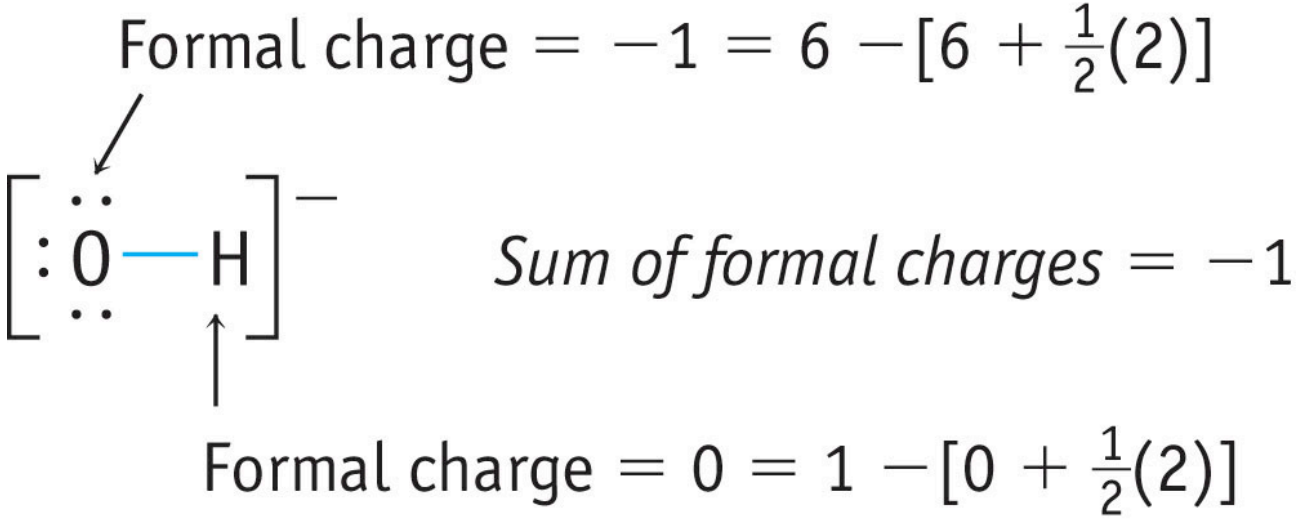
Sum of oxidation numbers $= -1$

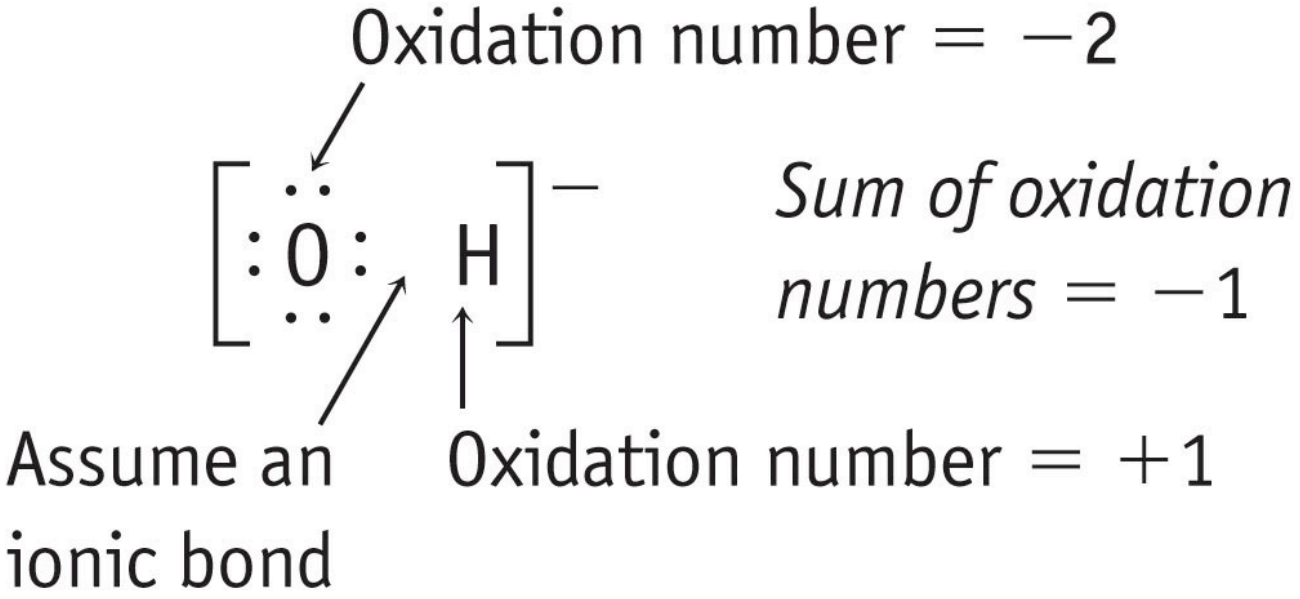
Assume an ionic bond

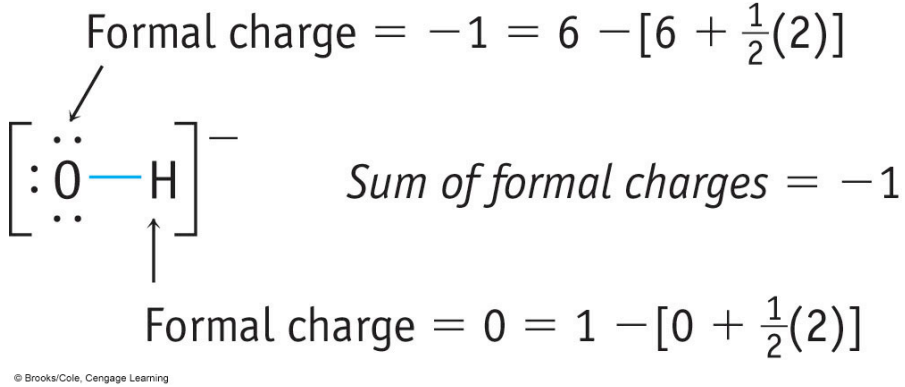
Oxidation number $= +1$

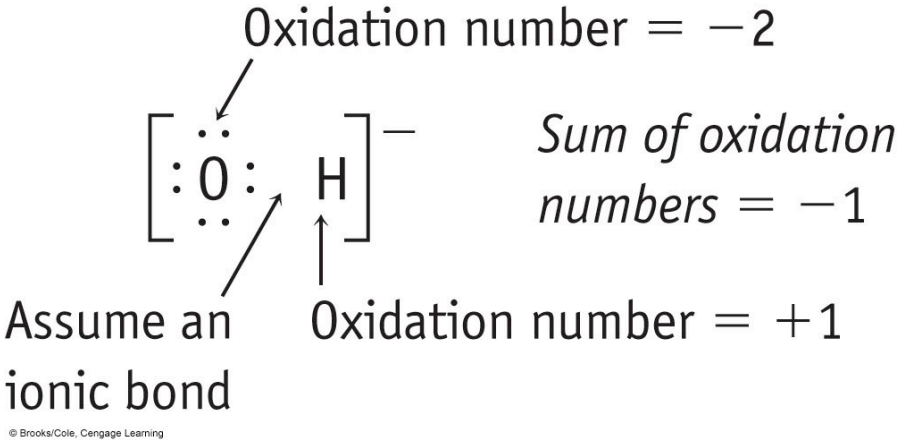
Brought to the party 1

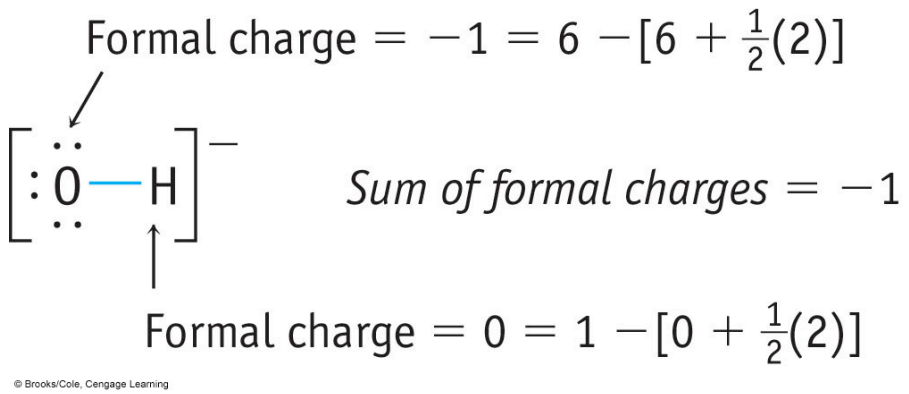
Ends up with 0

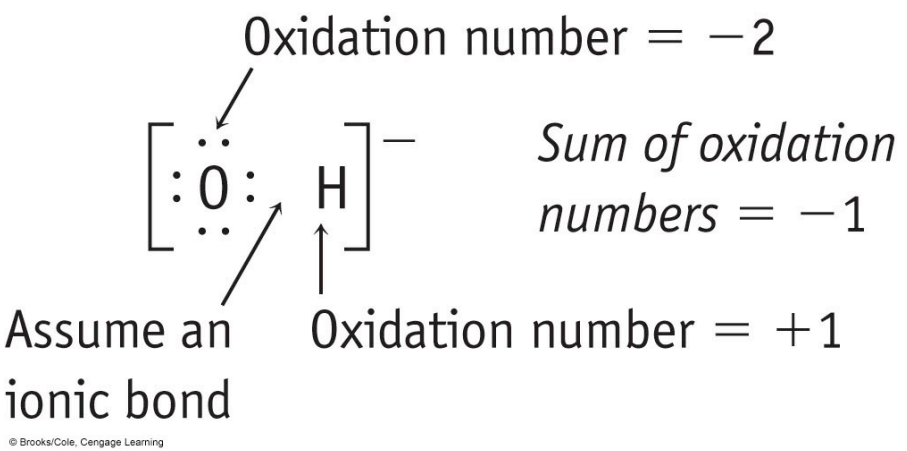




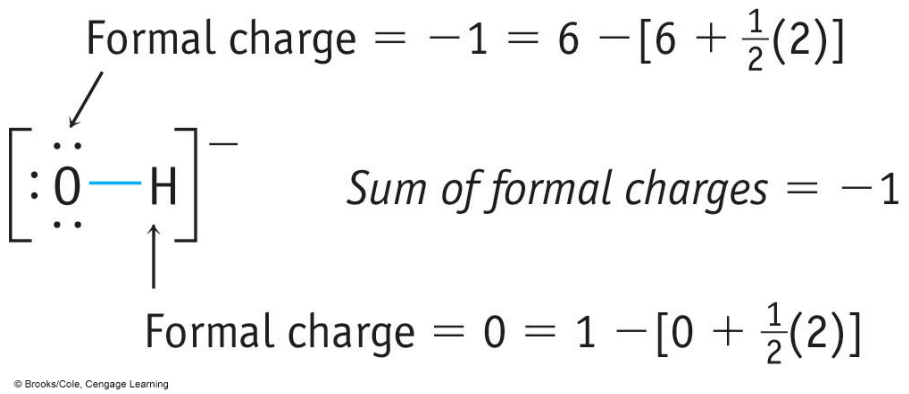


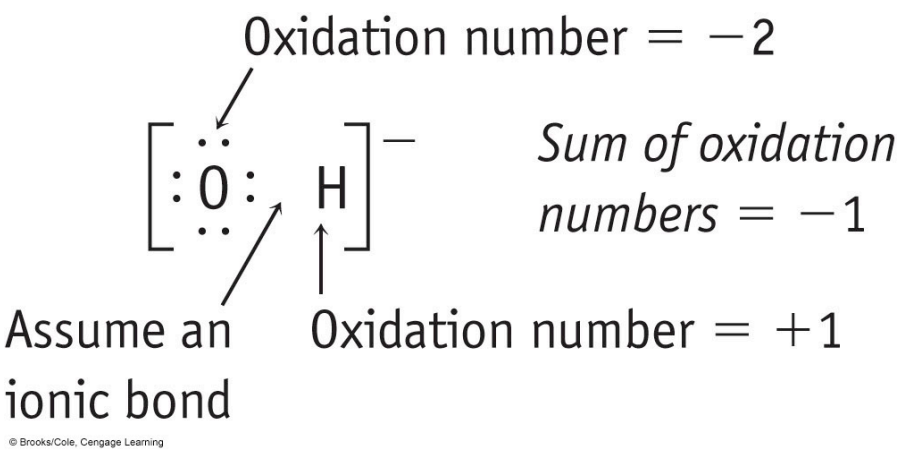






- **Octet Rule** - **double count** shared electrons
  - assumes both atoms get both electrons





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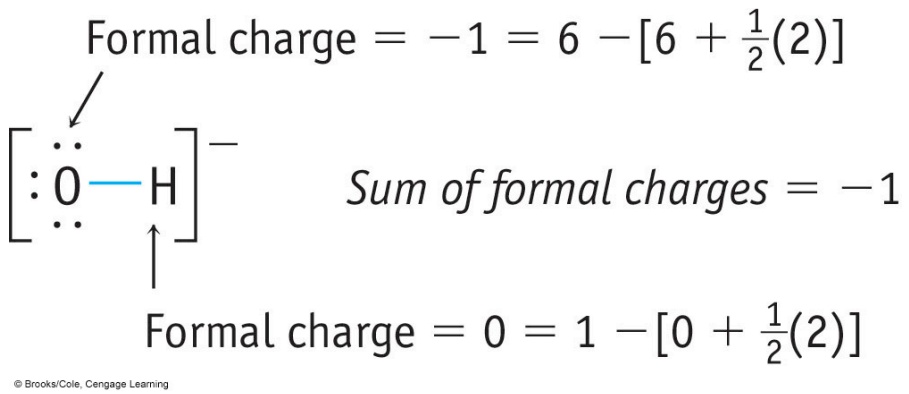
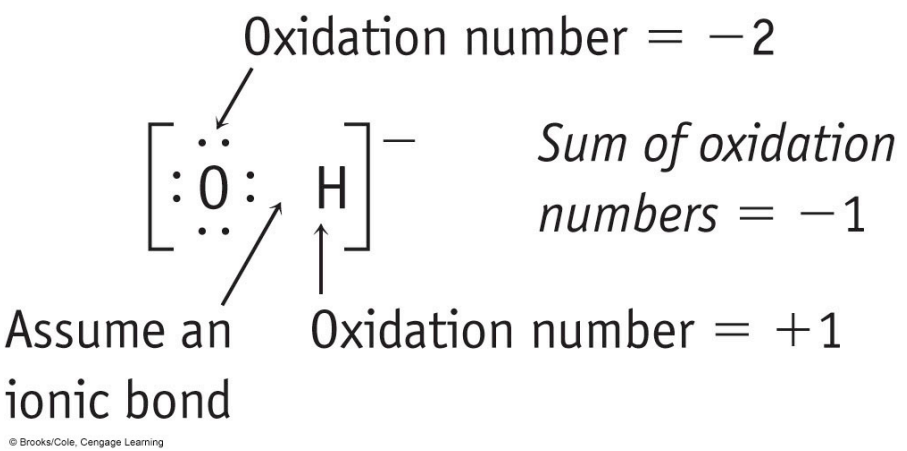




- **Octet Rule** - **double count** shared electrons
  - assumes both atoms get both electrons
- **Formal Charge** - ***Evenly*** split electrons in a bond
  - one to one atom, one to the other
- **Oxidation Number** - ***Unevenly*** split electrons in a bond
  - the atom that wants electrons more, gets both shared electrons (as in ionic bonds). The other one loses out. *An extreme view*.

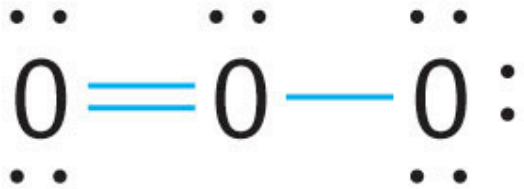
Questions?

Back to Octet Rule

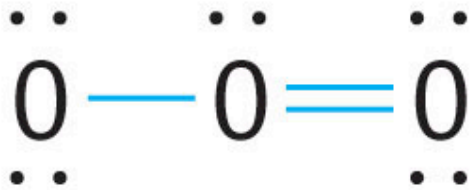
Ozone anion O_3^-

Alternative Ways of Drawing the Ozone Structure

Double bond on the left:



Double bond on the right:



Ozone anion O_3^-

Alternative Ways of Drawing the Ozone Structure

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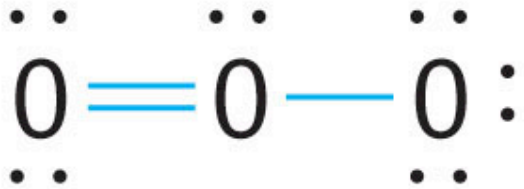
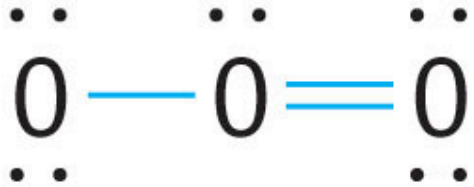
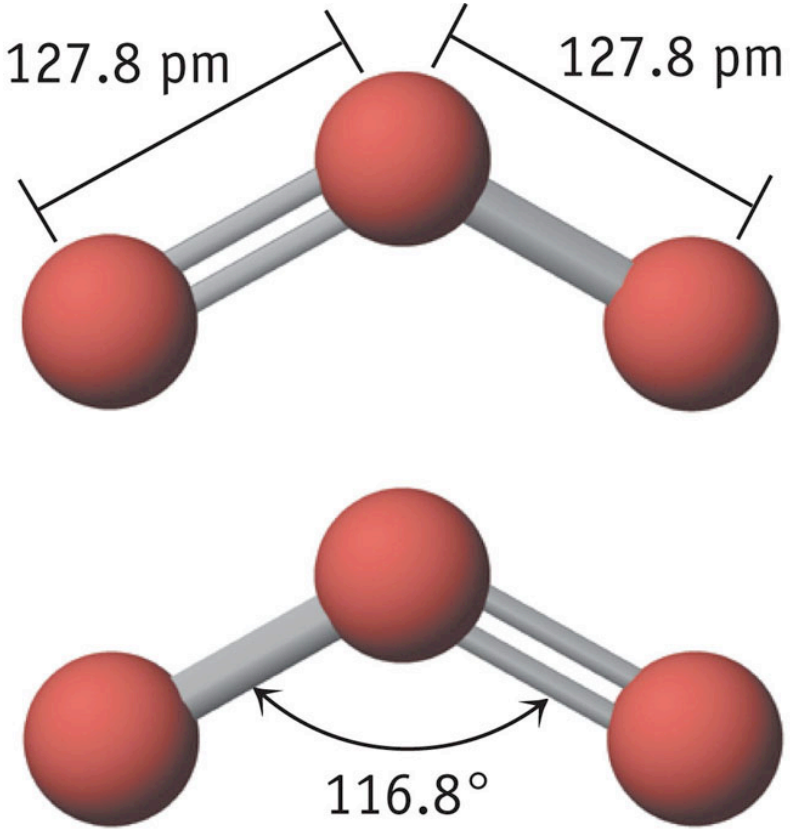


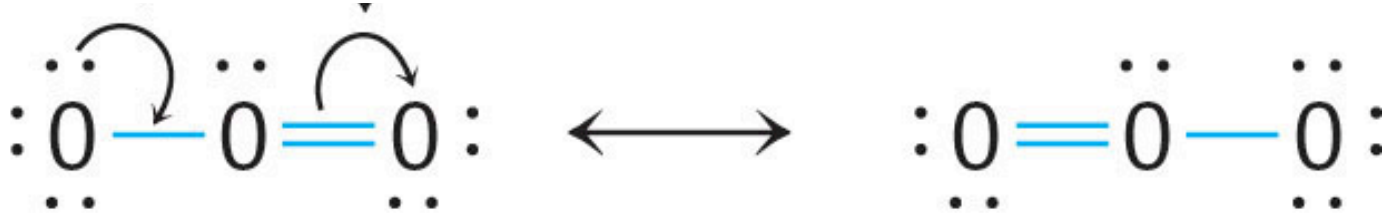
Double bond on the right:



Which side gets the double bond?

Real molecule - equal bond lengths!! (not interconverting)

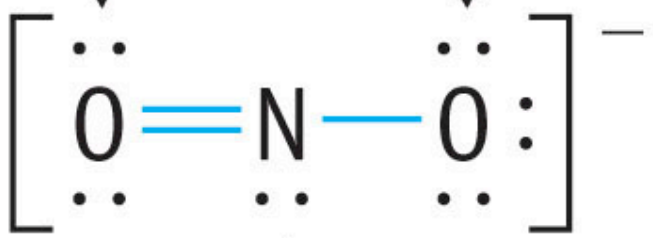




Nitrite anion - NO_2^-

Formal charge = $0 = 6 - [4 + \frac{1}{2}(4)]$

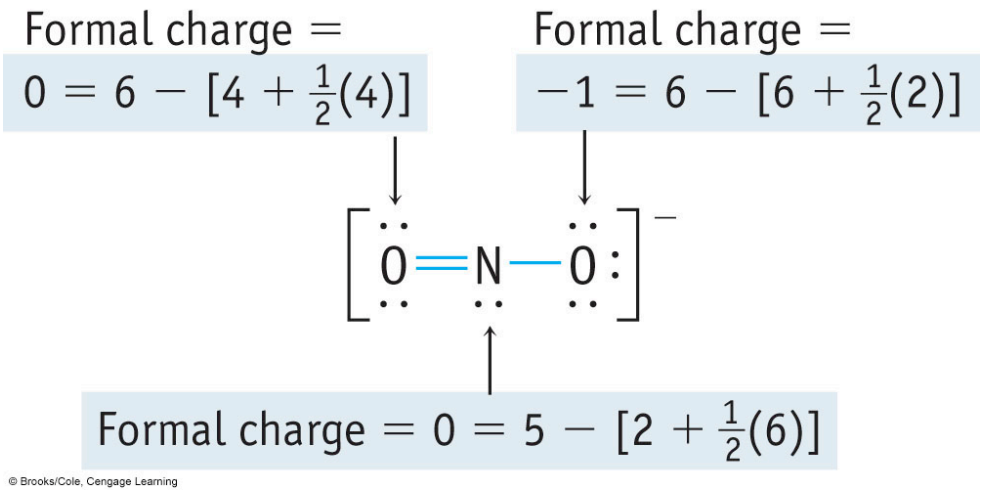
Formal charge = $-1 = 6 - [6 + \frac{1}{2}(2)]$



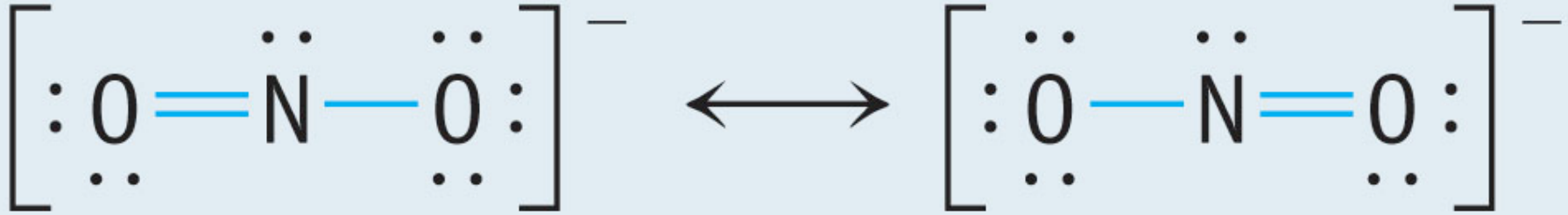
Formal charge = $0 = 5 - [2 + \frac{1}{2}(6)]$

Example 8-4, p. 363

Nitrite anion - NO_2^-



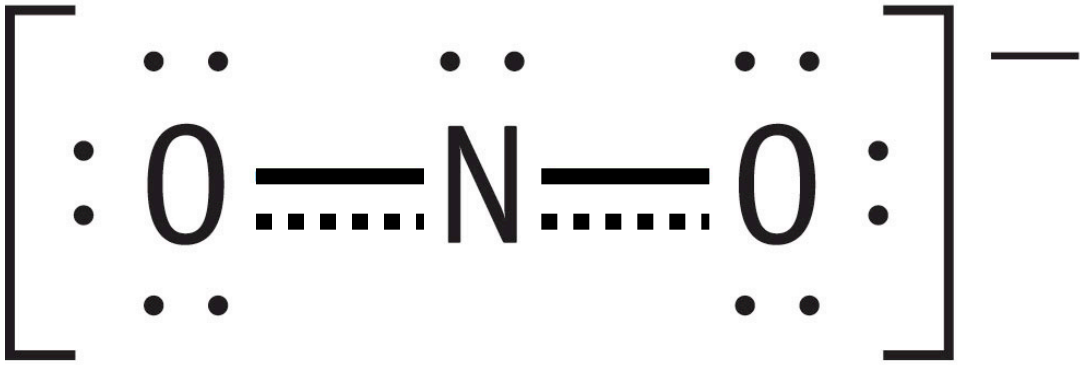
Resonance Structure - two views; neither fully correct



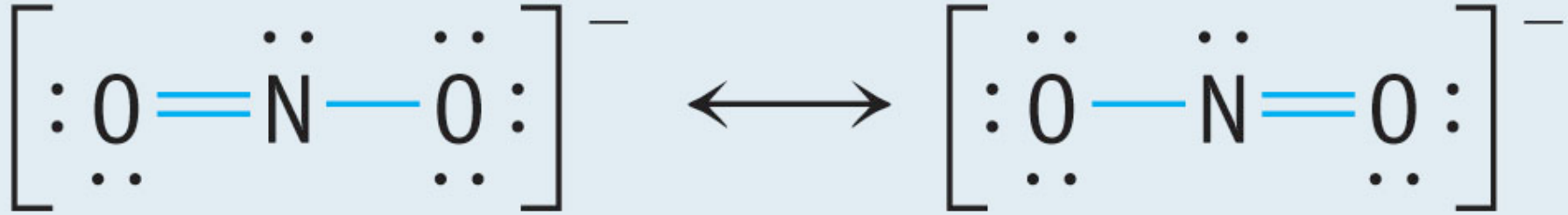
Example 8-4, p. 363

Nitrite anion - NO_2^-

Unified Structure - one view; more correct



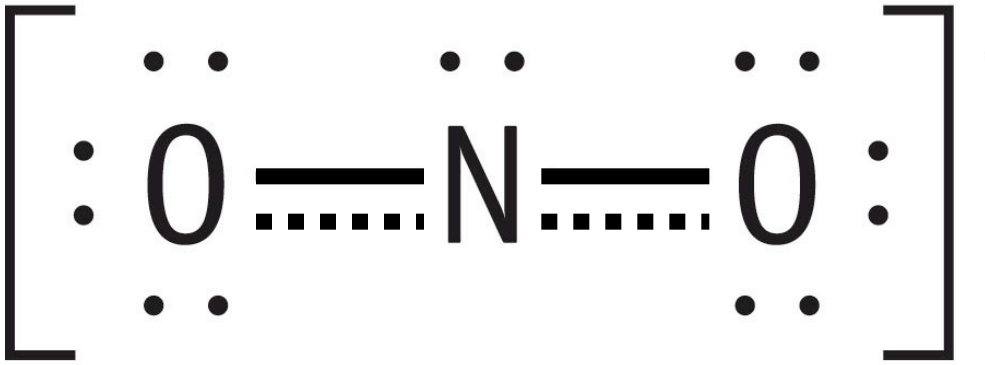
Resonance Structure - two views; neither fully correct



Example 8-4, p. 363

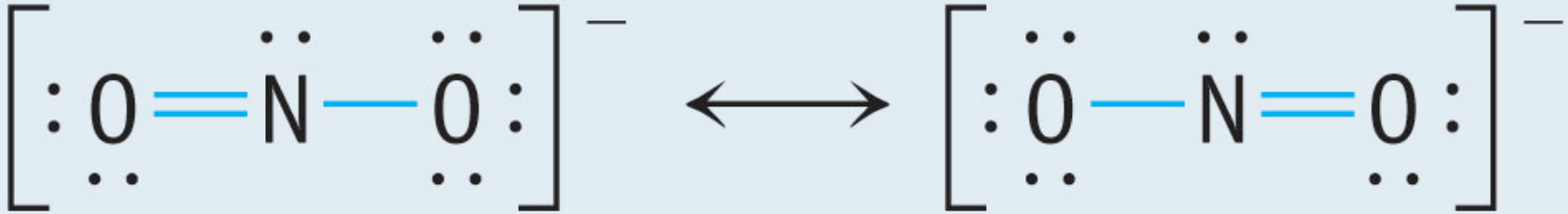
Nitrite anion - NO_2^-

Unified Structure - one view; more correct



But harder to implement the octet rule

Resonance Structure - two views; neither fully correct

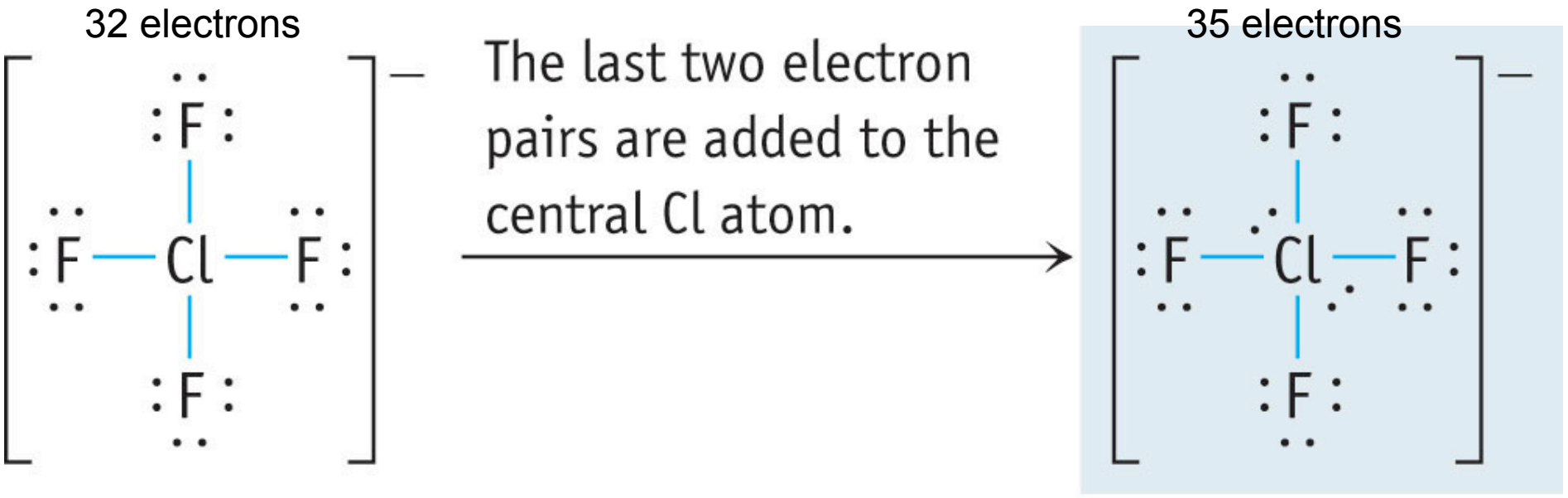


Example 8-4, p. 363

Exceeding the Octet Rule

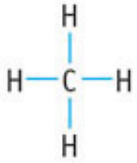
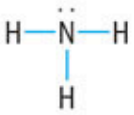
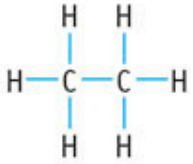
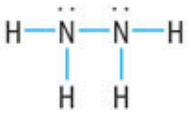
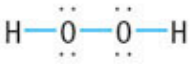
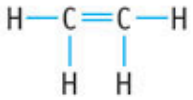
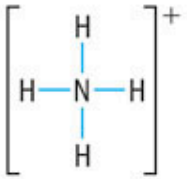
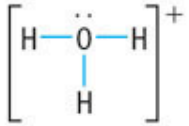
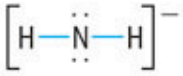
using near-energy d orbitals

5x7 = 35 electrons brought to the party



Example 8-5, p. 366

TABLE 8.3 Lewis Structures of Common Hydrogen-Containing Molecules and Ions of Second-Period Elements

Group 4A		Group 5A		Group 6A		Group 7A	
CH_4 methane		NH_3 ammonia		H_2O water		HF hydrogen fluoride	
C_2H_6 ethane		N_2H_4 hydrazine		H_2O_2 hydrogen peroxide			
C_2H_4 ethylene		NH_4^+ ammonium ion		H_3O^+ hydronium ion			
C_2H_2 acetylene		NH_2^- amide ion		OH^- hydroxide ion