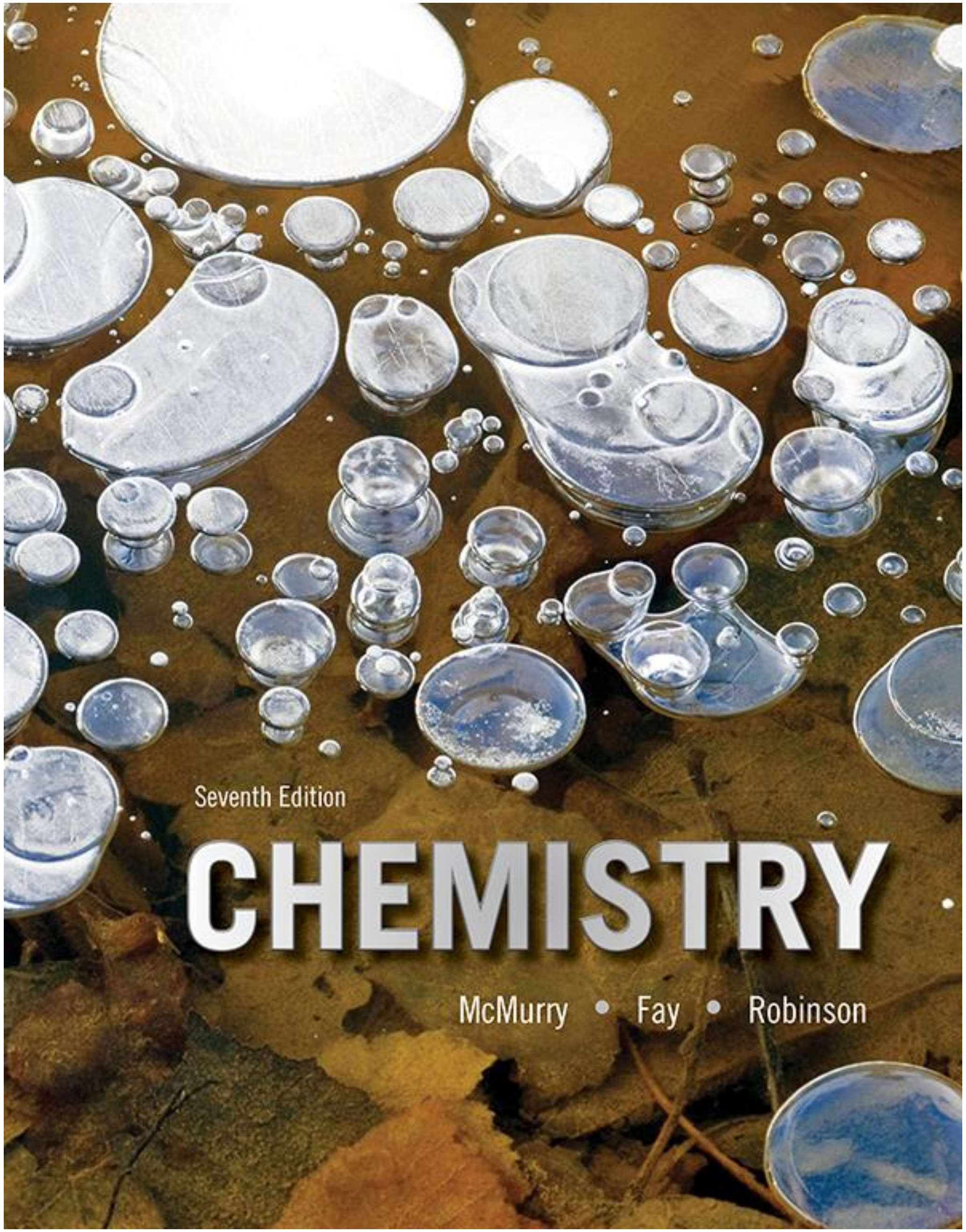

Lecture Presentation

Chapter 7

Covalent Bonding and Electron-Dot Structures

John E. McMurry
Robert C. Fay

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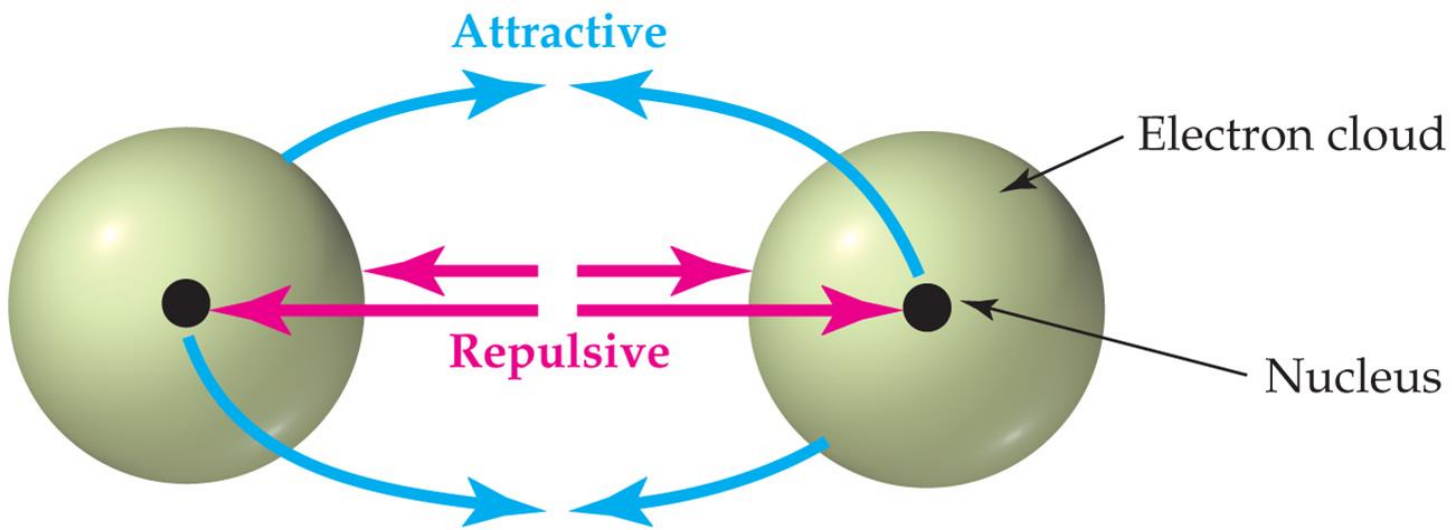
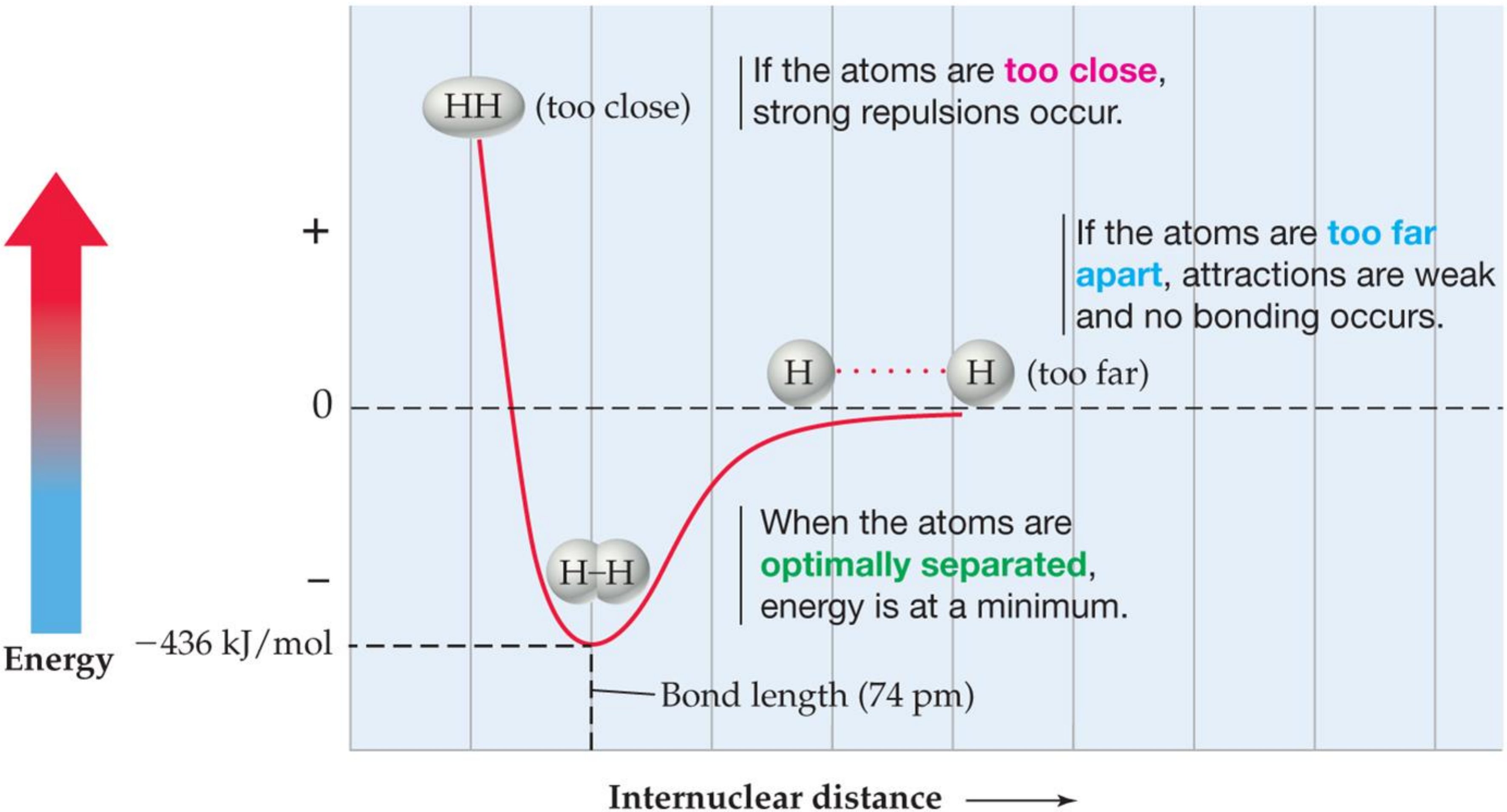
Covalent Bonding in Molecules

Covalent Bond: A bond that results from the sharing of electrons between atoms



The nucleus–electron **attractions** are greater than the nucleus–nucleus and electron–electron **repulsions**, resulting in a net attractive force that binds the atoms together.



If the atoms are too close, strong repulsions occur.
HH
(too close)
+
If the atoms are too far apart, attractions are weak and no bonding occurs.
H
H
(too far)
0
When the atoms are optimally separated, energy is at a minimum.
H–H
–
Energy
−436 kJ/mol
Bond length (74 pm)
Internuclear distance

Strengths of Covalent Bonds

TABLE 7.1 Average Bond Lengths (pm)

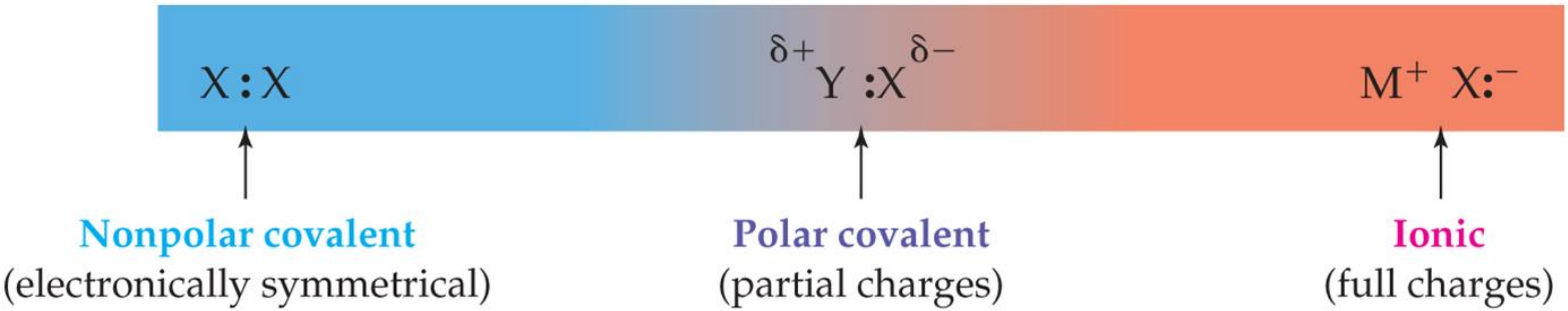
H—H	74[a]	C—H	110	N—H	98	O—F	130	I—I	267[a]
H—C	110	C—C	154	N—C	147	O—Cl	165	S—F	168
H—F	92[a]	C—F	141	N—F	134	O—Br	180	S—Cl	203
H—Cl	127[a]	C—Cl	176	N—Cl	169	O—I	199	S—Br	218
H—Br	142[a]	C—Br	191	N—Br	184	O—N	136	S—S	208
H—I	161[a]	C—I	176	N—N	140	O—O	132		
H—N	98	C—N	147	N—O	136	F—F	141[a]		
H—O	94	C—O	143	O—H	94	Cl—Cl	199[a]		
H—S	132	C—S	181	O—C	143	Br—Br	228[a]		
Multiple covalent bonds[b]									
C=C	134	C≡C	120	C=O	121	O=O	121[a]	N≡N	113[a]

[a]Exact value.

[b]We'll discuss multiple covalent bonds in Section 7.5.

Polar Covalent Bonds: Electronegativity

Electronegativity: The ability of an atom in a molecule to attract the shared electrons in a covalent bond



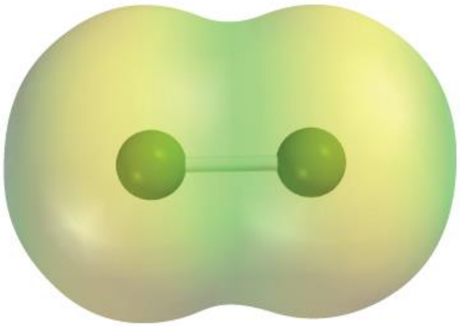
Polar covalent bonds have an unsymmetrical electron distribution in which the bonding electrons, shown as dots, are attracted more strongly by one atom than the other.

Polar Covalent Bonds: Electronegativity

The two bonding electrons, shown here as dots, are symmetrically distributed between the two Cl atoms.

Cl:Cl

A nonpolar covalent bond. Yellow-green represents a neutral atom.

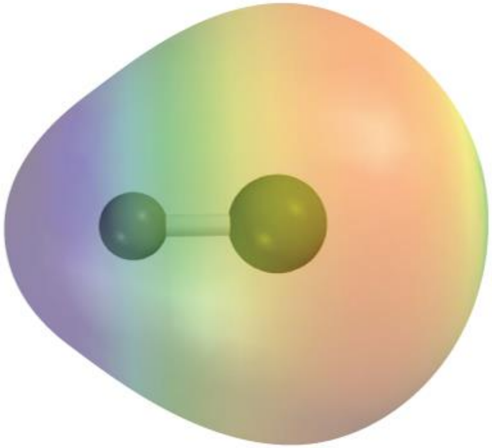


Polar Covalent Bonds: Electronegativity

$^{\delta+}H—Cl^{\delta-}$

$[H\ {:}Cl]$

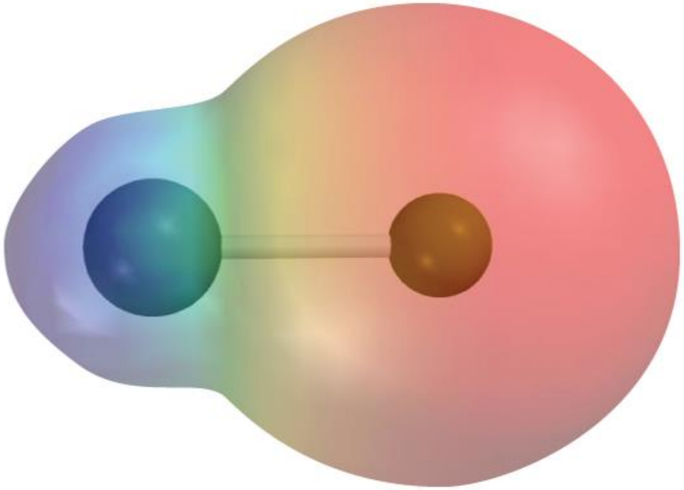
A polar covalent bond. The two bonding electrons (dots) are attracted more strongly by Cl than by H.



Polar Covalent Bonds: Electronegativity

$Na^+ Cl^-$

An ionic bond. Blue indicates a partial positive charge; red indicates a partial negative charge.



Polar Covalent Bonds: Electronegativity

Electronegativity increases from left to right.

Most electronegative

Electronegativity decreases from top to bottom.

H 2.1																	He
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe
Cs 0.7	Ba 0.9	Lu 1.1	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.1	Rn

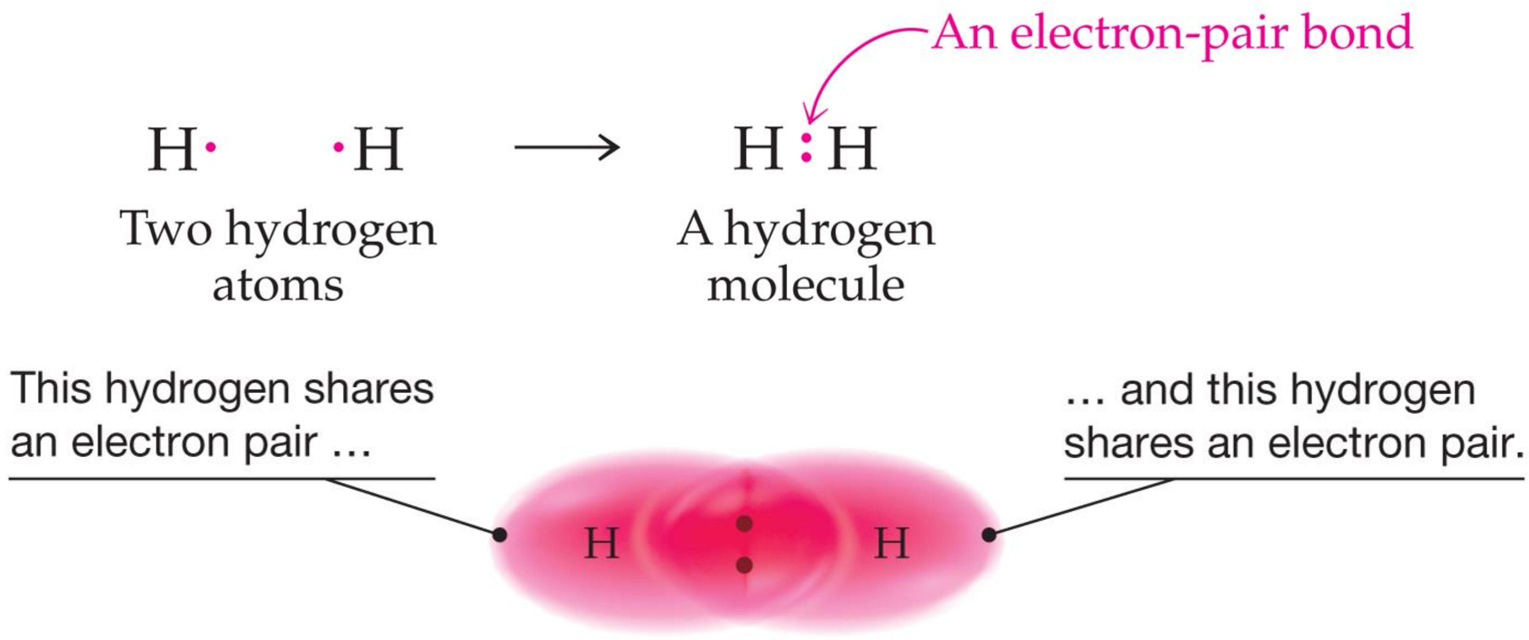
A Comparison of Ionic and Covalent Bonds

TABLE 7.3 Some Physical Properties of NaCl and HCl

Property	NaCl	HCl
Formula mass	58.44 amu	36.46 amu
Physical appearance	White solid	Colorless gas
Type of bond	Ionic	Covalent
Melting point	801 °C	−115 °C
Boiling point	1465 °C	−84.9 °C

Electron-Dot Structures: The Octet Rule

Electron-Dot Structure (Lewis Structure): Represents an atom's valence electrons by dots and indicates by the placement of the dots the way the valence electrons are distributed in a molecule



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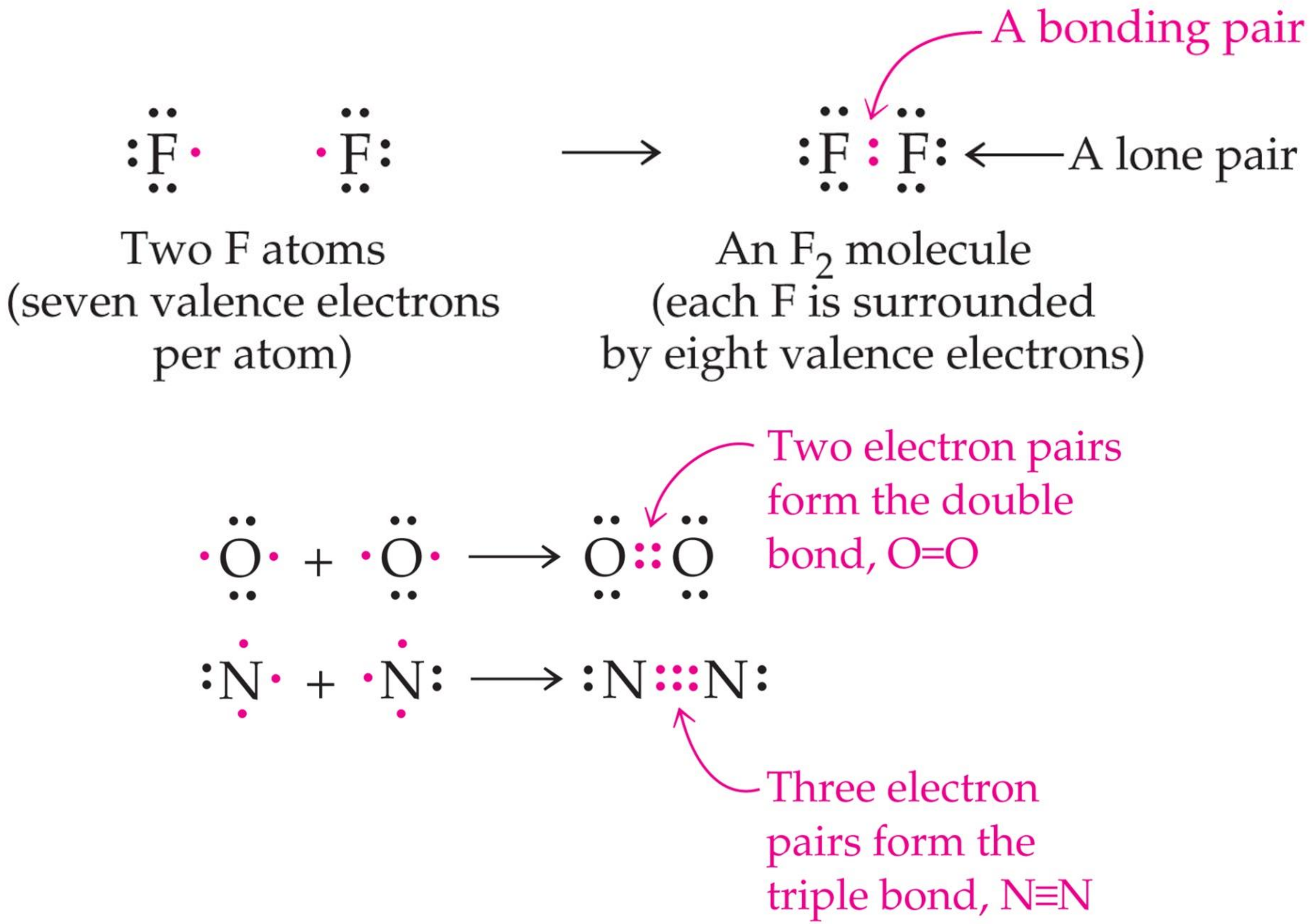
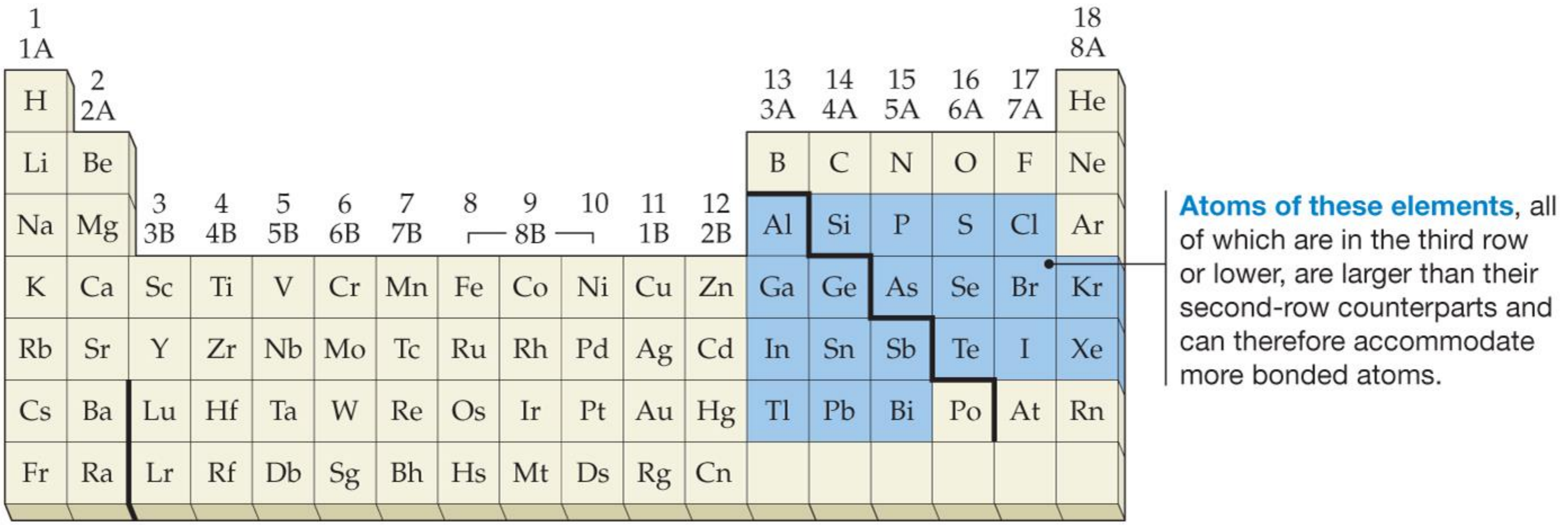
Electron-Dot Structures: The Octet Rule



Procedure for Drawing Electron-Dot Structures

The Octet Rule



Procedure for Drawing Electron-Dot Structures

Step 1: Valence Electrons

- Find the total number of valence electrons for all atoms in the molecule.
- Add one additional electron for each negative charge in an anion, or subtract one for each positive charge in a cation.

Procedure for Drawing Electron-Dot Structures

Step 2: Connect Atoms

- Draw lines to represent bonds between atoms.
- Hydrogen and halogens usually form only one bond.
- Elements in the second row usually form a certain number of bonds based upon the column they occupy.

Procedure for Drawing Electron-Dot Structures

TABLE 7.4 Covalent Bonding for Second-Row Elements

Group	Number of Valence Electrons	Number of Bonds	Example
3A	3	3	BH_3
4A	4	4	CH_4
5A	5	3	NH_3
6A	6	2	H_2O
7A	7	1	HF
8A	8	0	Ne

Procedure for Drawing Electron-Dot Structures

Step 2: Connect Atoms

- Draw lines to represent bonds between atoms.
- Hydrogen and halogens usually form only one bond.
- Elements in the second row usually form the number of bonds given in the next table.
- Elements in third row and lower are often a central atom around which other atoms are grouped and form more bonds than predicted by the octet rule.

Procedure for Drawing Electron-Dot Structures

Step 3: Assign Electrons to the Terminal Atoms

- Subtract the number of electrons used for bonding from the total number calculated in Step 1 to find the number that remain.
- Complete each terminal atom's octet (except for hydrogen's).

Procedure for Drawing Electron-Dot Structures

Step 4: Assign Electrons to the Central Atom

- If unassigned electrons remain after Step 3, place them on the central atom.

Procedure for Drawing Electron-Dot Structures

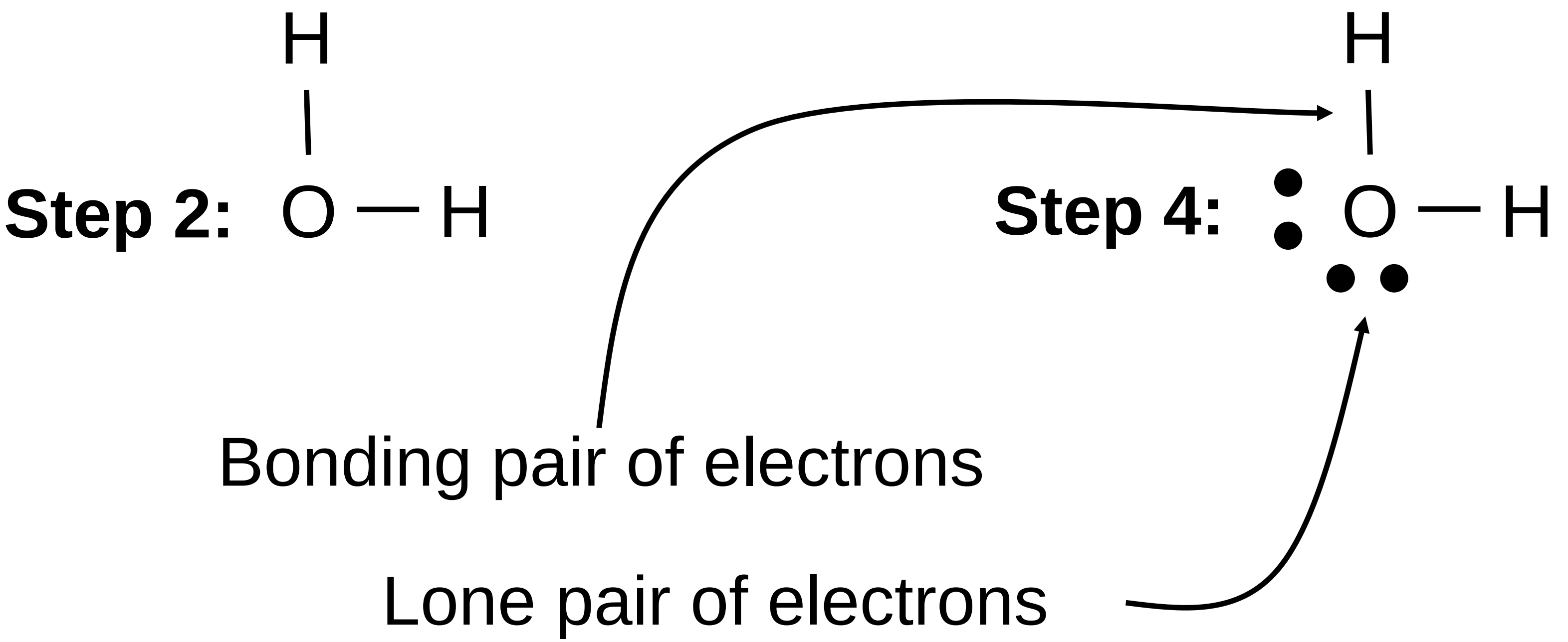
Step 5: Multiple Bonds

- If no unassigned electrons remain after Step 4 but the central atom does not yet have an octet, use one or more lone pairs of electrons from a neighboring atom to form a multiple bond (either a double or a triple).

Procedure for Drawing Electron-Dot Structures

Draw an electron-dot structure for $\mathbf{H_2O}$.

Step 1: 2(1) + 6 = 8 valence electrons



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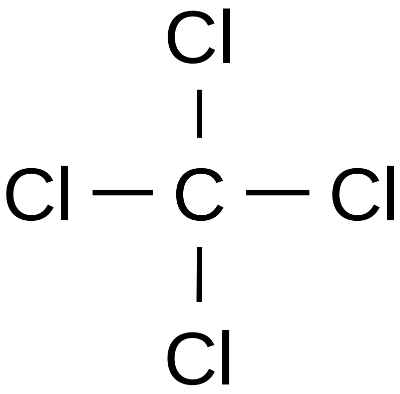
Procedure for Drawing Electron-Dot Structures

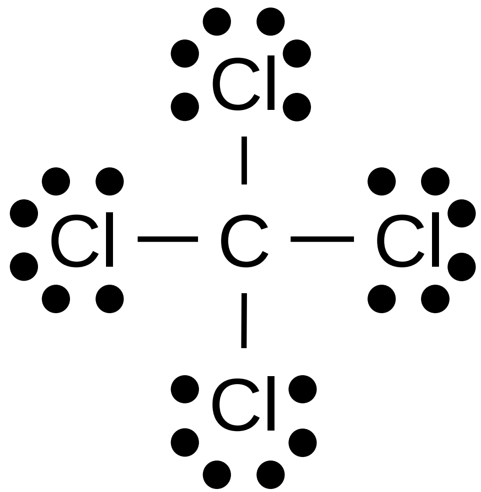
Draw an electron-dot structure for $\mathbf{CCl_4}$.

Step 1: 4 + 4(7) = 32 valence electrons

Step 2:



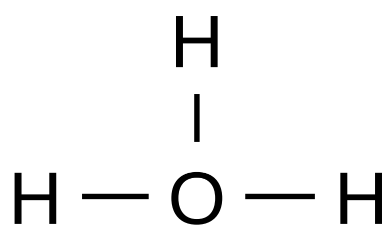
Step 3:

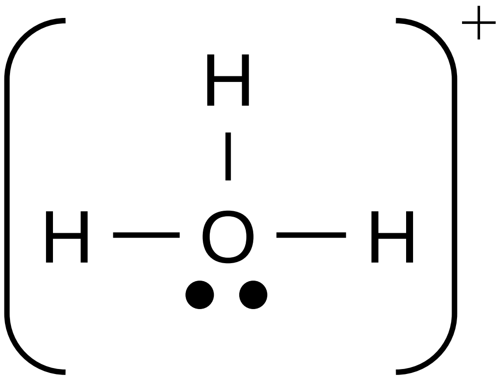


Procedure for Drawing Electron-Dot Structures

Draw an electron-dot structure for $\mathbf{H_3O^{1+}}$.

Step 1: 3(1) + 6 – 1 = 8 valence electrons

Step 2: 

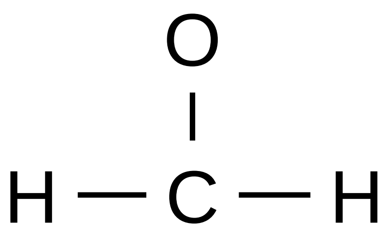
Step 4: 

Procedure for Drawing Electron-Dot Structures

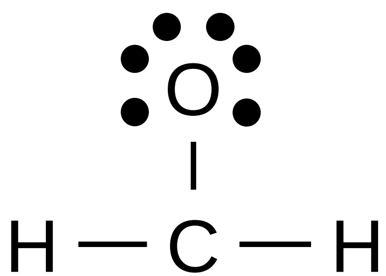
Draw an electron-dot structure for $\mathbf{CH_2O}$.

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

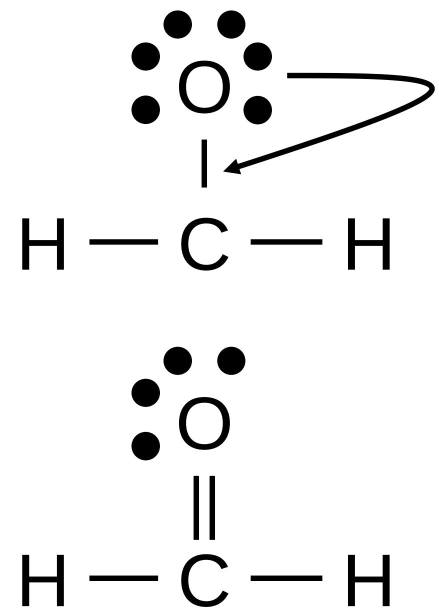
Step 2:



Step 3:



Step 5:



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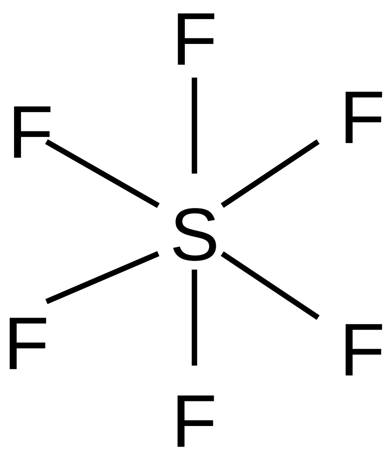
Procedure for Drawing Electron-Dot Structures

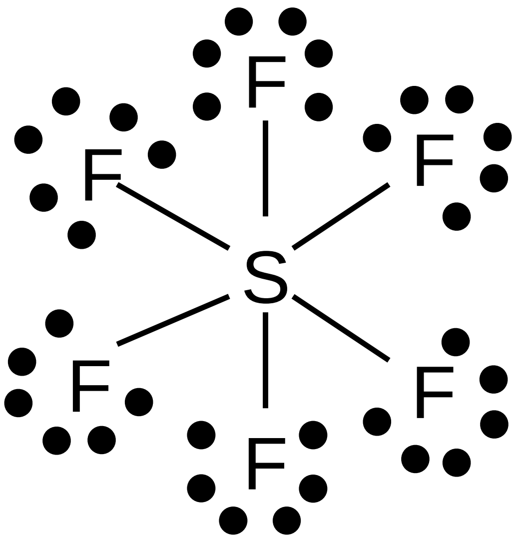
Draw an electron-dot structure for **SF_6**.

Step 1: 6 + 4(7) = 34 valence electrons

Step 2:



Step 3:

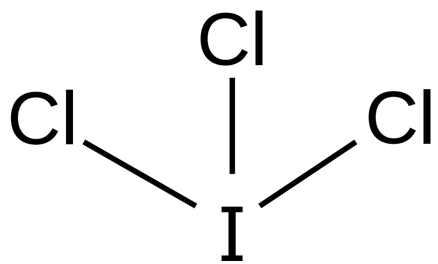


Procedure for Drawing Electron-Dot Structures

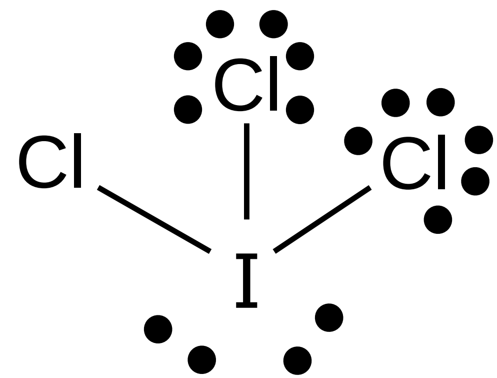
Draw an electron-dot structure for **ICl_3**.

Step 1: 7 + 3(7) = 28 valence electrons

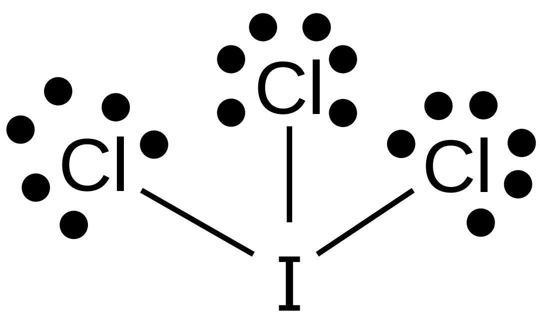
Step 2:



Step 4:



Step 3:



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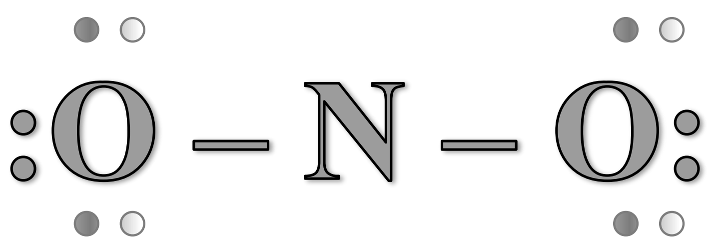
Drawing Electron-Dot Structures for Radicals

- A few substances have an unpaired electron. These are called *radicals*, or *free radicals*.
- Drawing electron-dot structures for radicals follows the steps we have seen.
- There will always be an unfilled octet on one atom.

Drawing Electron-Dot Structures for Radicals

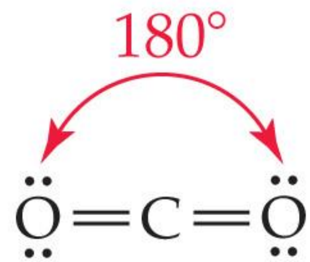
- As an example, NO_2
  - Each oxygen provides 6 electrons.
  - The nitrogen provides 5 electrons.
  - Thus, there are 17 electrons.
  - Note the lone electron on the nitrogen.



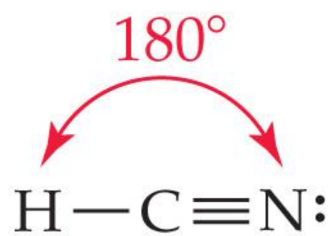
Electron-Dot Structures of Compounds Containing Only Hydrogen and Second-Row Elements

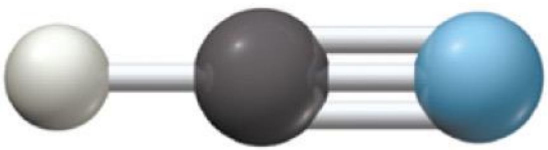
Two Charge Clouds

A CO_2 molecule is linear, with a bond angle of 180°.



An HCN molecule is linear, with a bond angle of 180°.





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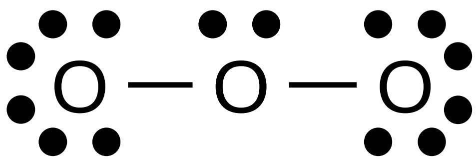
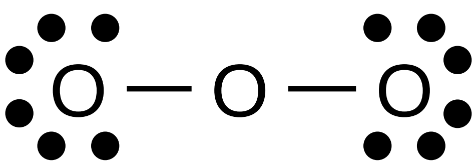
Electron-Dot Structures and Resonance

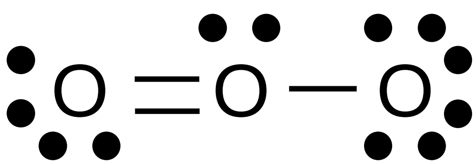
Draw an electron-dot structure for $\mathbf{O_3}$.

Step 1: 3(6) = 18 valence electrons

Step 2: O — O — O

Step 4: 

Step 3: 

Step 5: 

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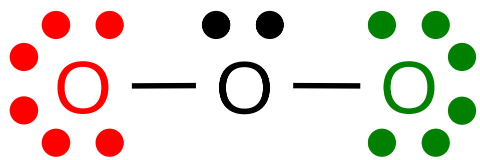
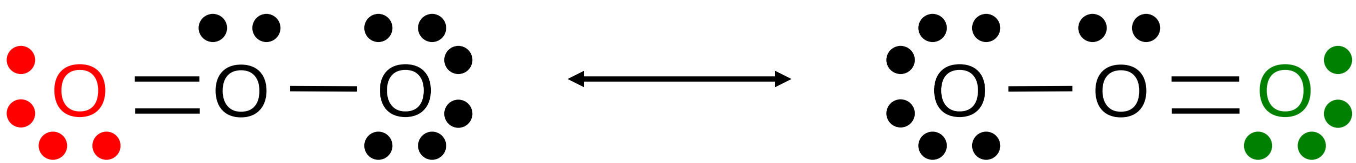
Electron-Dot Structures and Resonance

Move a lone pair from this oxygen?

Step 4: 

Or move a lone pair from this oxygen?

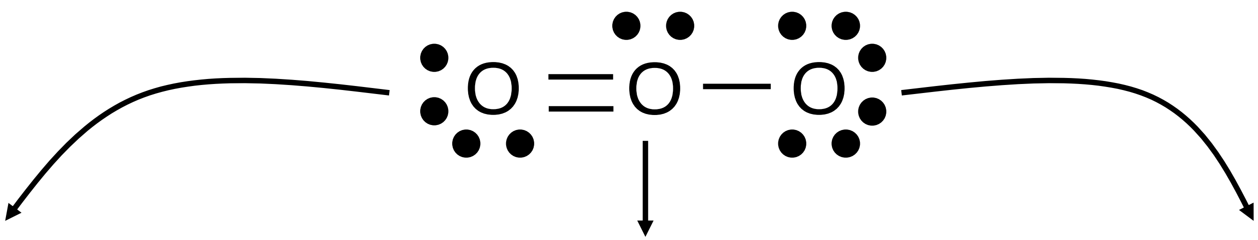


Resonance

Formal Charges

$$\textbf{Formal charge} = \begin{pmatrix}\text{\# of}\\ \text{valence } e^-\\ \text{in free atom}\end{pmatrix} - \frac{1}{2}\begin{pmatrix}\text{\# of}\\ \text{bonding}\\ e^-\end{pmatrix} - \begin{pmatrix}\text{\# of}\\ \text{nonbonding}\\ e^-\end{pmatrix}$$

Calculate the formal charge on each atom in O_3.



$$6 - \frac{1}{2}(4) - 4 = 0 \qquad 6 - \frac{1}{2}(6) - 2 = +1 \qquad 6 - \frac{1}{2}(2) - 6 = -1$$