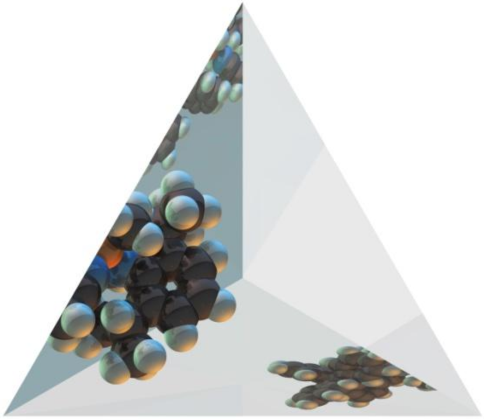


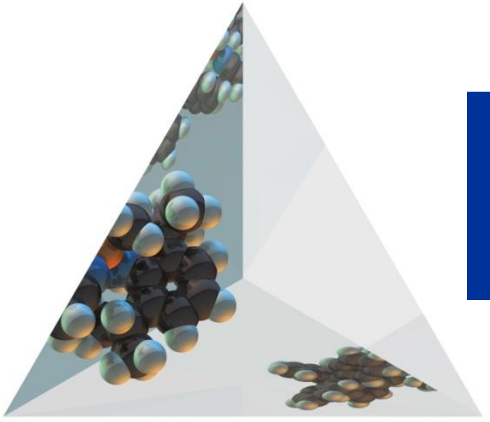
CHEMISTRY

The Central Science

8th Edition

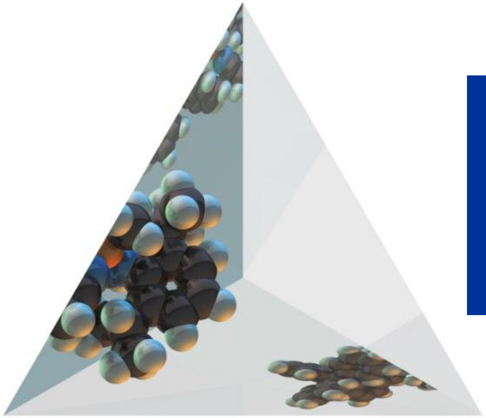
Chapter 8
Basic Concepts of Chemical Bonding

Kozet YAPSAKLI



Chemical Bonds

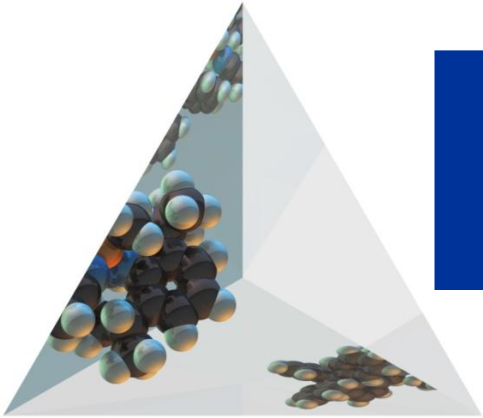
- The forces that hold atoms or ions together in compounds are called ***chemical bonds***.
- Three different types of chemical bonds are introduced: ***ionic***, ***covalent***, and ***metallic***.
- ***Covalent bond*** results from sharing electrons between the atoms (nonmetals).
- ***Ionic bond*** results from the transfer of electrons from a metal to a nonmetal.



Lewis Symbols

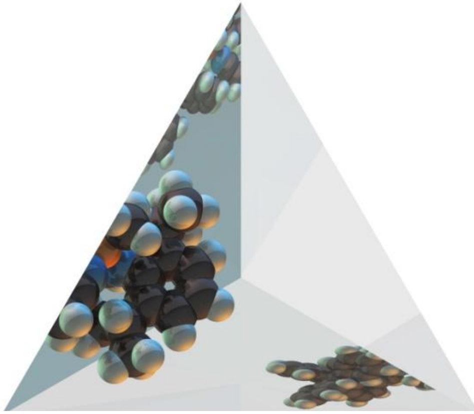
Lewis Symbols

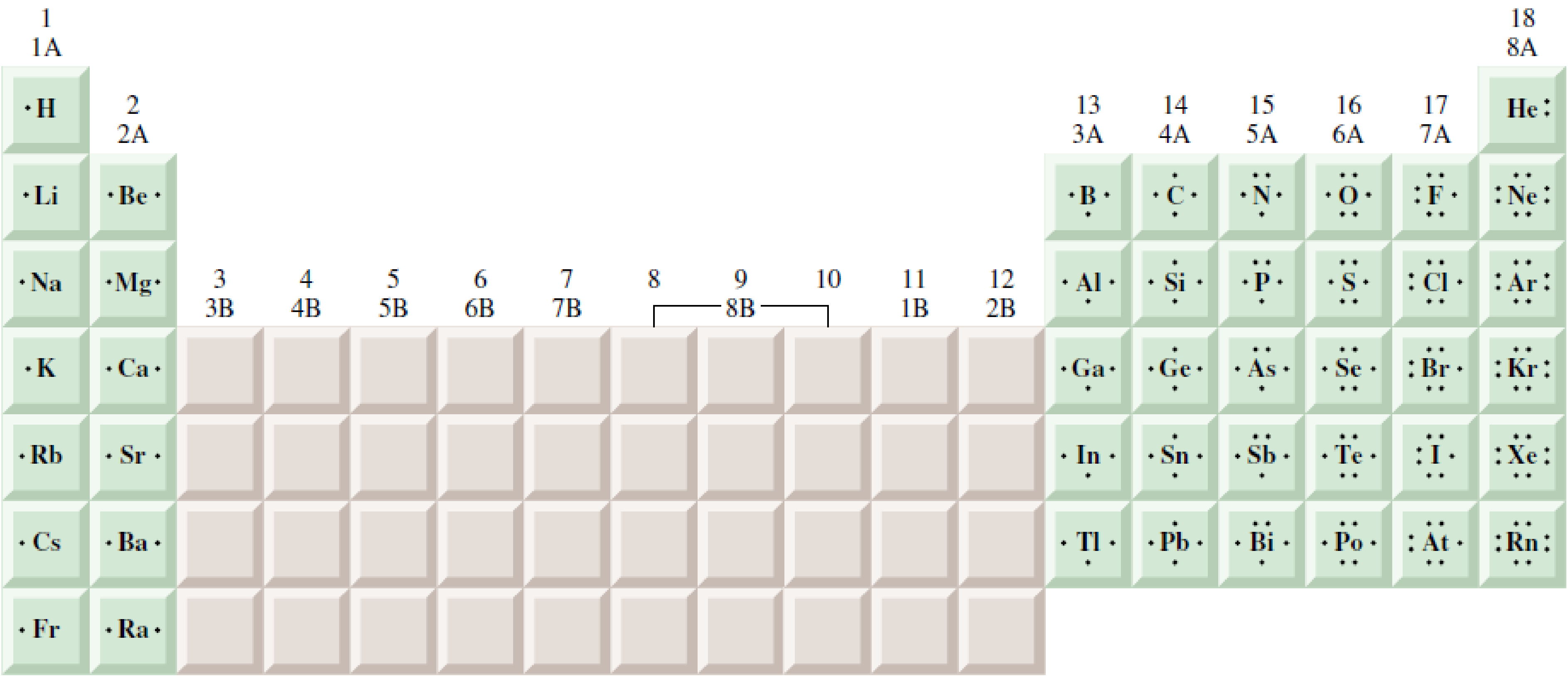
- Lewis Symbol is used to show the valence electrons of an atom or ion.
- Lewis symbol for oxygen, group 6A, shows 6 dots.
- The number of electrons available for bonding are indicated by unpaired dots.
- We generally place the electrons one four sides of a square around the element symbol.

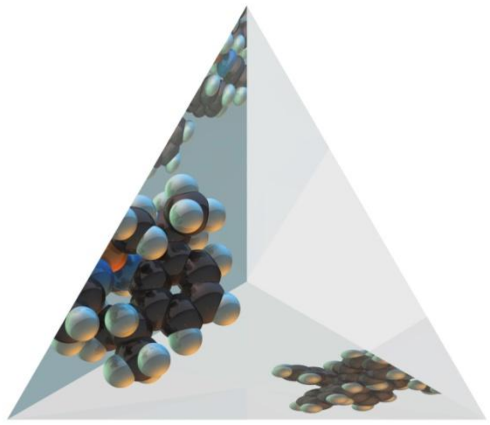


Valence Electrons & Bonding

- Because valence electrons are held most loosely, and
- Because chemical bonding involves the transfer or sharing of electrons between two or more atoms,
- Valence electrons are most important in bonding
- Lewis theory focuses on the behavior of the valence electrons




1
1A
18
8A
H
2
2A
13
3A
14
4A
15
5A
16
6A
17
7A
He
Li
Be
B
C
N
O
F
Ne
Na
Mg
3
3B
4
4B
5
5B
6
6B
7
7B
8
9
8B
10
11
1B
12
2B
Al
Si
P
S
Cl
Ar
K
Ca
Ga
Ge
As
Se
Br
Kr
Rb
Sr
In
Sn
Sb
Te
I
Xe
Cs
Ba
Tl
Pb
Bi
Po
At
Rn
Fr
Ra



Lewis structure

Ionic bonding (transfer of electrons):

$$\mathrm{Na}\times \;+\; \cdot\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathrm{Cl}}}\colon \longrightarrow [\mathrm{Na}]^{+}[\,{}^{\times}_{\cdot}\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathrm{Cl}}}\colon]^{-}$$

Lewis symbols Lewis structure

Covalent bonding (sharing of electrons):

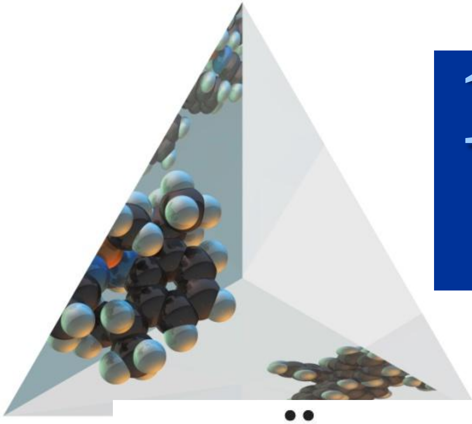
$$\mathrm{H}\times \;+\; \cdot\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathrm{Cl}}}\colon \longrightarrow \mathrm{H}\overset{\cdot}{\times}\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathrm{Cl}}}\colon$$

Lewis symbols Lewis structure

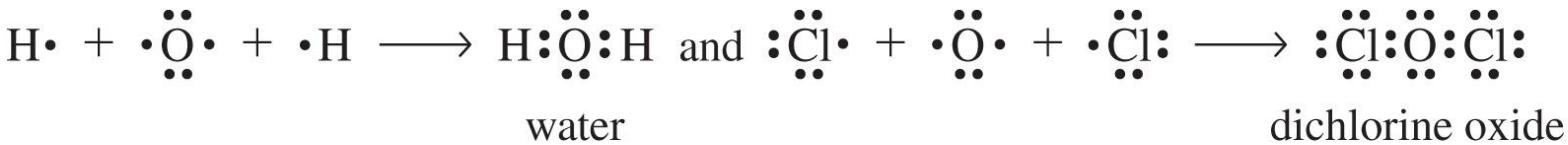
Stable Electron Arrangements and Ion Charge

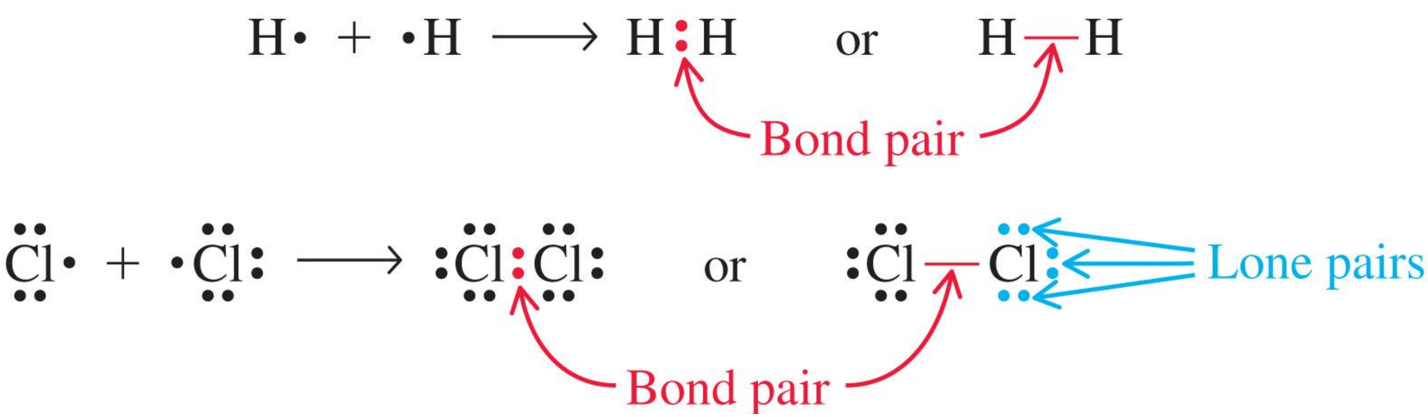
- Metals form cations by losing enough electrons to get the same electron configuration as the previous noble gas
- Nonmetals form anions by gaining enough electrons to get the same electron configuration as the next noble gas
- The noble gas electron configuration must be very stable

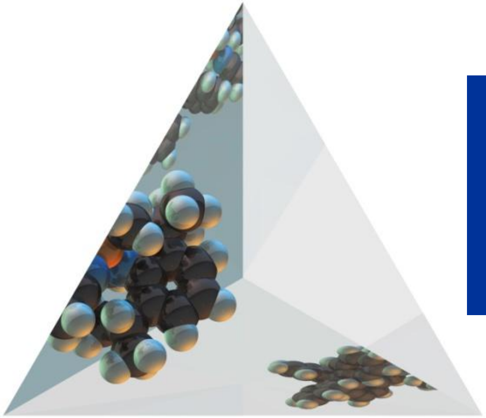
Atom	Atom's Electron Config	Ion	Ion's Electron Config
Na	$[Ne]3s^1$	Na^+	[Ne]
Mg	$[Ne]3s^2$	Mg^{2+}	[Ne]
Al	$[Ne]3s^2 3p^1$	Al^{3+}	[Ne]
O	$[He]2s^2 2p^4$	O^{2-}	[Ne]
F	$[He]2s^2 2p^5$	F^-	[Ne]



10-2 Covalent Bonding: An Introduction





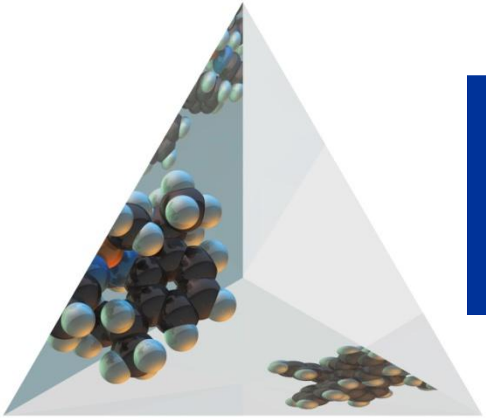


Lewis Symbols

Lewis Symbols

TABLE 8.1 Lewis Symbols

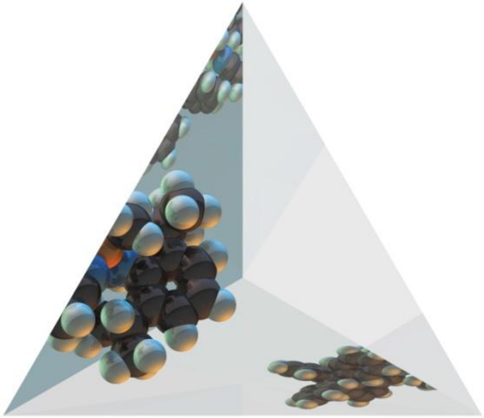
Element	Electron Configuration	Lewis Symbol	Element	Electron Configuration	Lewis Symbol
Li	[He]$2s^1$	$\mathrm{Li}\cdot$	Na	[Ne]$3s^1$	$\mathrm{Na}\cdot$
Be	[He]$2s^2$	$\cdot\mathrm{Be}\cdot$	Mg	[Ne]$3s^2$	$\cdot\mathrm{Mg}\cdot$
B	[He]$2s^2 2p^1$	$\cdot\dot{\mathrm{B}}\cdot$	Al	[Ne]$3s^2 3p^1$	$\cdot\dot{\mathrm{Al}}\cdot$
C	[He]$2s^2 2p^2$	$\cdot\underset{\cdot}{\dot{\mathrm{C}}}\cdot$	Si	[Ne]$3s^2 3p^2$	$\cdot\underset{\cdot}{\dot{\mathrm{Si}}}\cdot$
N	[He]$2s^2 2p^3$	$\cdot\underset{\cdot}{\dot{\mathrm{N}}}:$	P	[Ne]$3s^2 3p^3$	$\cdot\underset{\cdot}{\dot{\mathrm{P}}}:$
O	[He]$2s^2 2p^4$	$:\underset{\cdot}{\dot{\mathrm{O}}}:$	S	[Ne]$3s^2 3p^4$	$:\underset{\cdot}{\dot{\mathrm{S}}}:$
F	[He]$2s^2 2p^5$	$\cdot\underset{\cdot\cdot}{\ddot{\mathrm{F}}}:$	Cl	[Ne]$3s^2 3p^5$	$\cdot\underset{\cdot\cdot}{\ddot{\mathrm{Cl}}}:$
Ne	[He]$2s^2 2p^6$	$:\underset{\cdot\cdot}{\ddot{\mathrm{Ne}}}:$	Ar	[Ne]$3s^2 3p^6$	$:\underset{\cdot\cdot}{\ddot{\mathrm{Ar}}}:$



Octet Rule

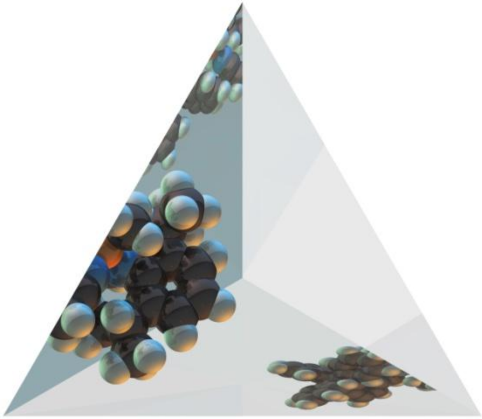
The Octet Rule

- All noble gases except He has an s^2p^6 configuration.
- Atoms tend to lose, gain, or share enough electrons to achieve a noble-gas electron configuration.
- When two chlorine atoms bond to form a chlorine molecule, sharing electrons in the covalent bond.
- These observations are summarized in the ***Octet rule*.**
- There are many exceptions to the octet rule.



Ionic Bonding

- Sodium chloride is an example of an ionic compound.
- The reaction is exothermic.
- Na has lost an electron to become Na^+ and chlorine has gained the electron to become Cl^-.
- There is a regular arrangement of Na^+ and Cl^- in 3D.



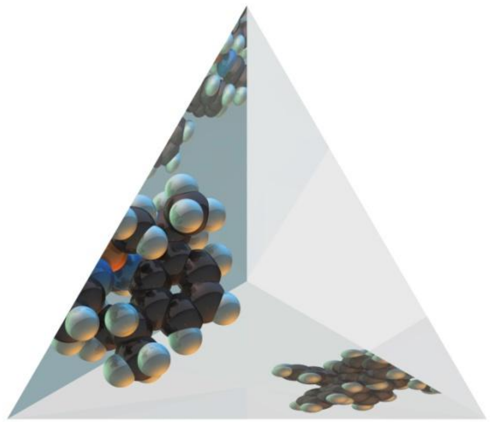
Ionic Bonding

Energetics of Ionic Bond Formation

- **Lattice energy:** the energy required to completely separate an ionic solid into its gaseous ions.
- Lattice energy depends on the charges on the ions and the sizes of the ions:

$$E_l = \kappa \frac{Q_1 Q_2}{d}$$

κ is a constant (8.99×10^{9} J·m/C^2), Q_1 and Q_2 are the charges on the ions, and d is the distance between ions.



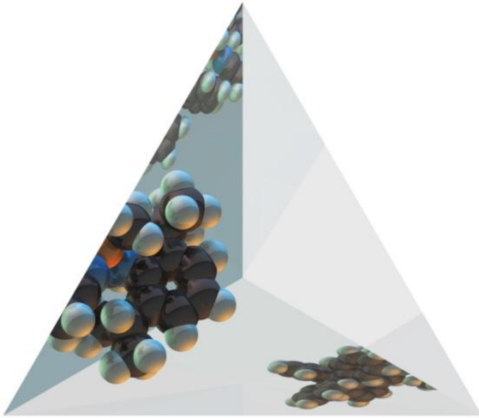
Ionic Bonding

Energetics of Ionic Bond Formation

- Lattice energy increases as
  - The charges on the ions increase
  - The distance between the ions decreases.

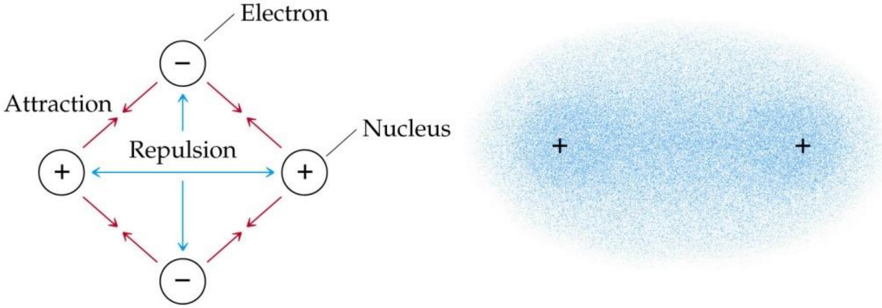
TABLE 8.2 Lattice Energies for Some Ionic Compounds

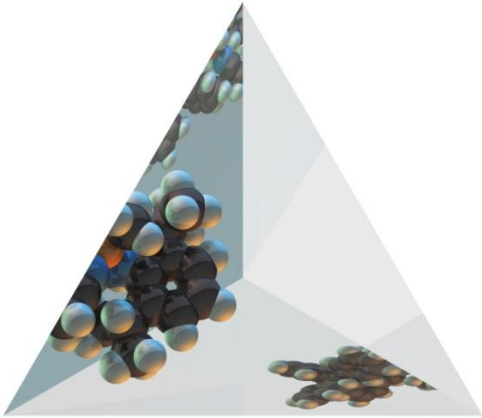
Compound	Lattice Energy (kJ/mol)	Compound	Lattice Energy (kJ/mol)
LiF	1030	$MgCl_2$	2326
LiCl	834	$SrCl_2$	2127
LiI	730		
NaF	910	MgO	3795
NaCl	788	CaO	3414
NaBr	732	SrO	3217
NaI	682		
KF	808	ScN	7547
KCl	701		
KBr	671		
CsCl	657		
CsI	600		



Covalent Bonding

- When two similar atoms bond, none of them wants to lose or gain an electron to form an octet.
- They share pairs of electrons to each obtain an octet.
- Each pair of shared electrons constitutes one chemical bond.
- Example: H + H $\rightarrow$ H_2 has electrons connecting the two H.



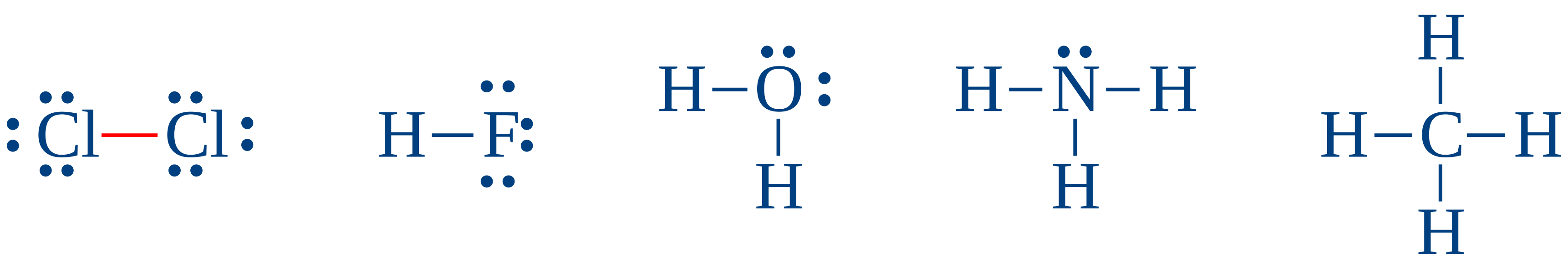


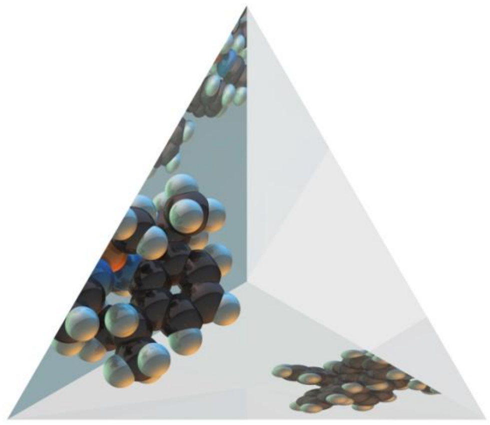
Covalent Bonding

Lewis Structures

- Covalent bonds can be represented by the Lewis symbols of the elements:

- In Lewis structures, each pair of electrons in a bond is represented by a single line:





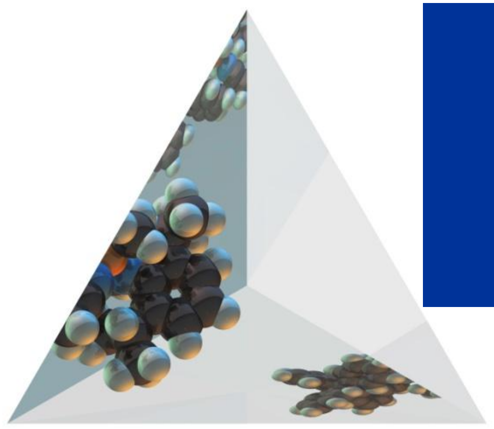
Covalent Bonding

Multiple Bonds

- It is possible for more than one pair of electrons to be shared between two atoms (multiple bonds):
  - One shared pair of electrons = single bond (e.g. H_2);
  - Two shared pairs of electrons = double bond (e.g. O_2);
  - Three shared pairs of electrons = triple bond (e.g. N_2).

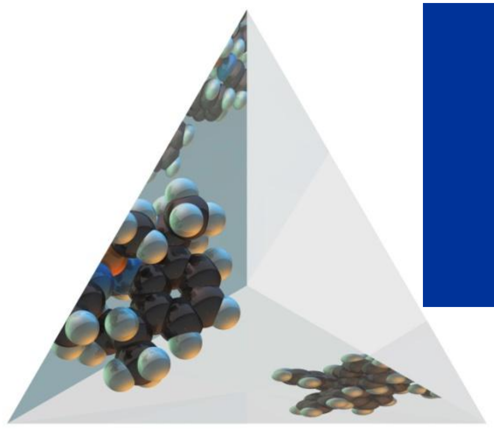
$$\mathrm{H{-}H} \qquad \ddot{\underset{\cdot\cdot}{\mathrm{O}}}{=}\ddot{\underset{\cdot\cdot}{\mathrm{O}}} \qquad :\mathrm{N{\equiv}N}:$$

- Generally, bond distances decrease as we move from single through double to triple bonds.



Bond Polarity and Electronegativity

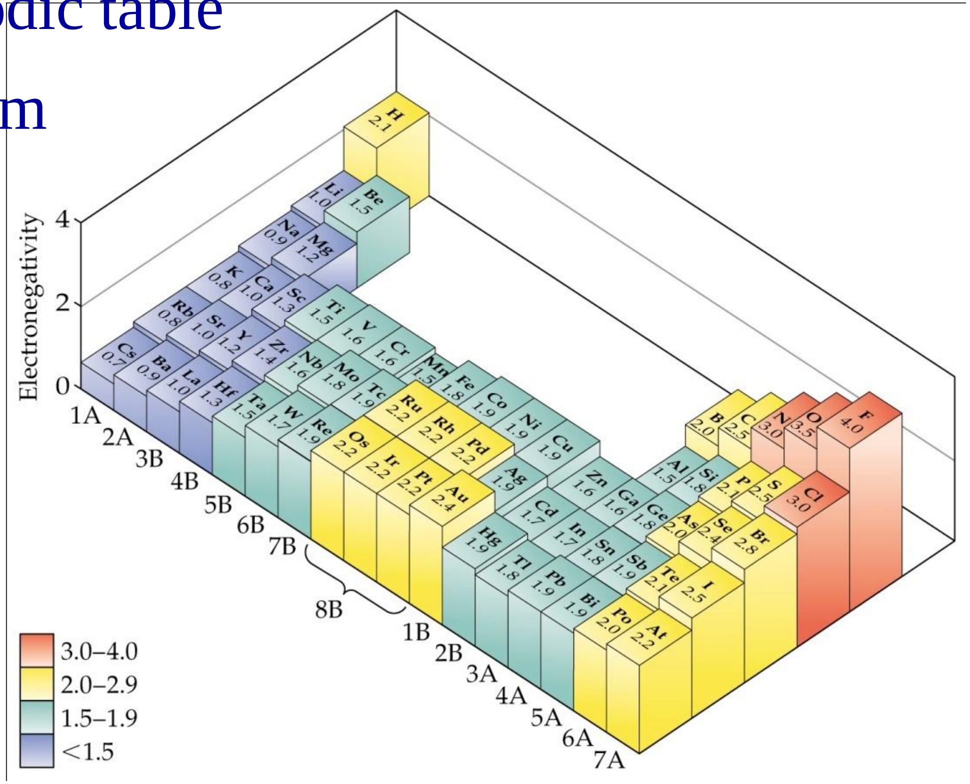
- In a covalent bond, electrons are shared.
- Sharing of electrons to form a covalent bond does not imply equal sharing of electrons (***polar covalent bond***)
- Unequal sharing of electrons results in polar bonds.
- ***Electronegativity*****:** The ability of one atoms *in a molecule* to attract electrons to itself.
- Electronegativity of an atom is related to its ionization energy and electron affinity.



Bond Polarity and Electronegativity

Electronegativity

- Electronegativity increases from
  - left to right across the periodic table
  - decreases from top to bottom within a group.



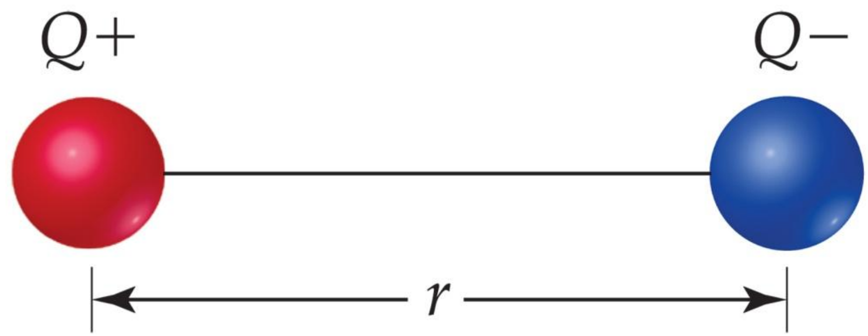
Polar Covalent Bonds

When two atoms share electrons unequally, a bond dipole results.

- The dipole moment, μ, produced by two equal but opposite charges separated by a distance, r, is calculated:

$$\mu = Qr$$

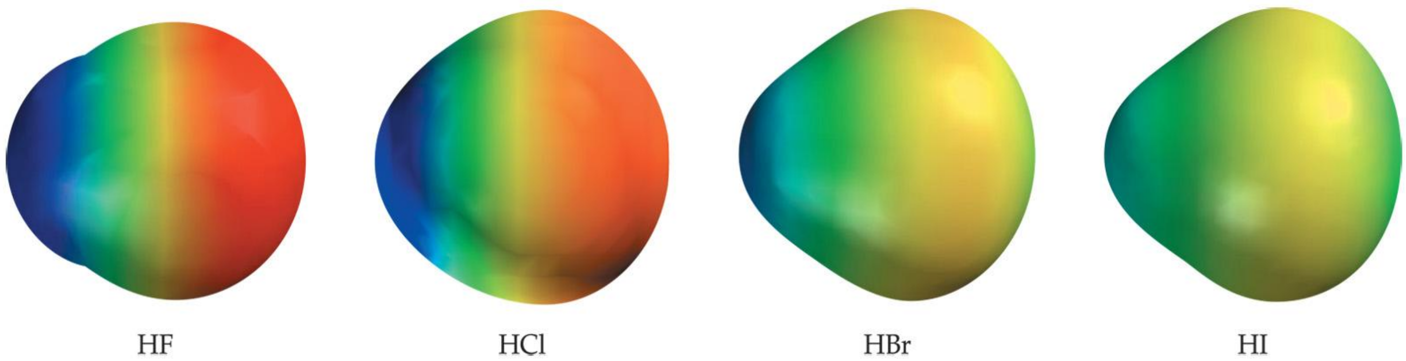
- It is measured in debyes (D).

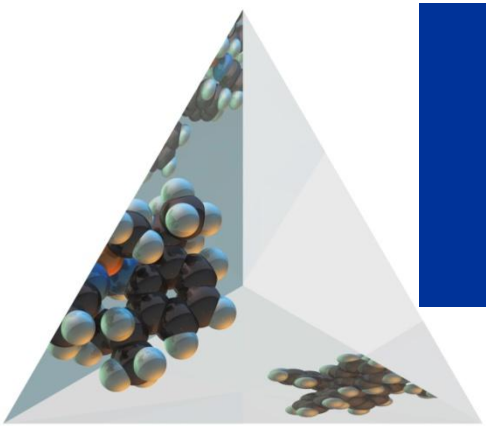


Polar Covalent Bonds

Compound	Bond Length (Å)	Electronegativity Difference	Dipole Moment (D)
HF	0.92	1.9	1.82
HCl	1.27	0.9	1.08
HBr	1.41	0.7	0.82
HI	1.61	0.4	0.44

The greater the difference in electronegativity, the more polar is the bond.



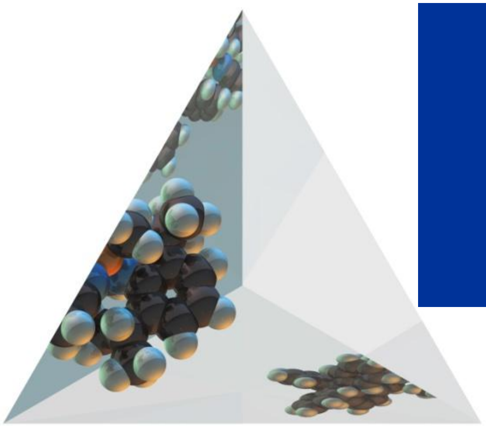


Bond Polarity and Electronegativity

Bond Types and Nomenclature

- In general, the least electronegative element is named first.
- The name of the more electronegative element ends in *-ide*.

Ionic		Molecular	
MgH_2	Magnesium hydride	H_2S	Hydrogen sulfide
FeF_2	Iron(II) fluoride	OF_2	Oxygen difluoride
Mn_2O_3	Manganese(III) oxide	Cl_2O_3	Dichlorine trioxide



Bond Polarity and Electronegativity

Bond Types and Nomenclature

- Ionic compounds are named according to their ions, including the charge on the cation if it is variable.
- Molecular compounds are named with prefixes.

TABLE 2.6 Prefixes Used in Naming Binary Compounds Formed Between Nonmetals

Prefix	Meaning
Mono-	1
Di-	2
Tri-	3
Tetra-	4
Penta-	5
Hexa-	6
Hepta-	7
Octa-	8
Nona-	9
Deca-	10

Lewis Structures

Lewis structures are representations of molecules showing all electrons, bonding and nonbonding.

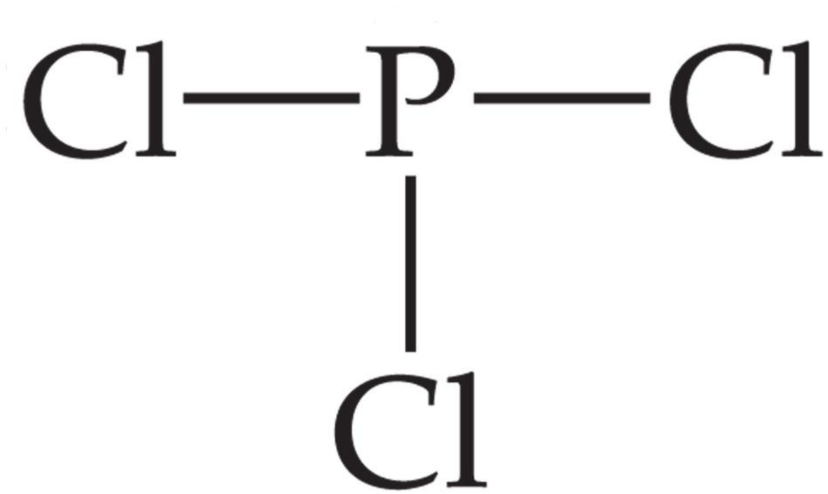
Writing Lewis Structures

PCl_3

$5 + 3(7) = 26$

1. Find the sum of valence electrons of all atoms in the polyatomic ion or molecule.
   - If it is an anion, add one electron for each negative charge.
   - If it is a cation, subtract one electron for each positive charge.

Writing Lewis Structures

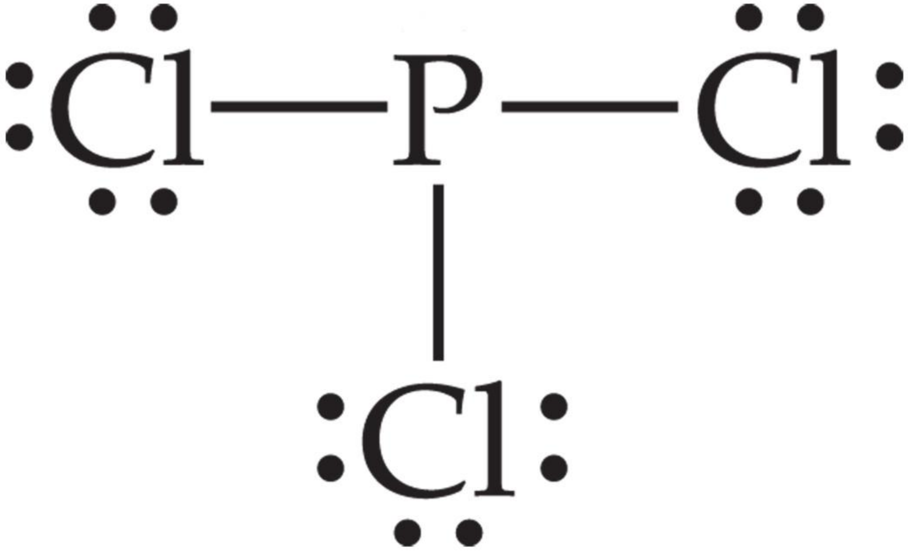


2. The central atom is the *least* electronegative element that isn't hydrogen. Connect the outer atoms to it by single bonds.

Keep track of the electrons:

$26 - 6 = 20$

Writing Lewis Structures

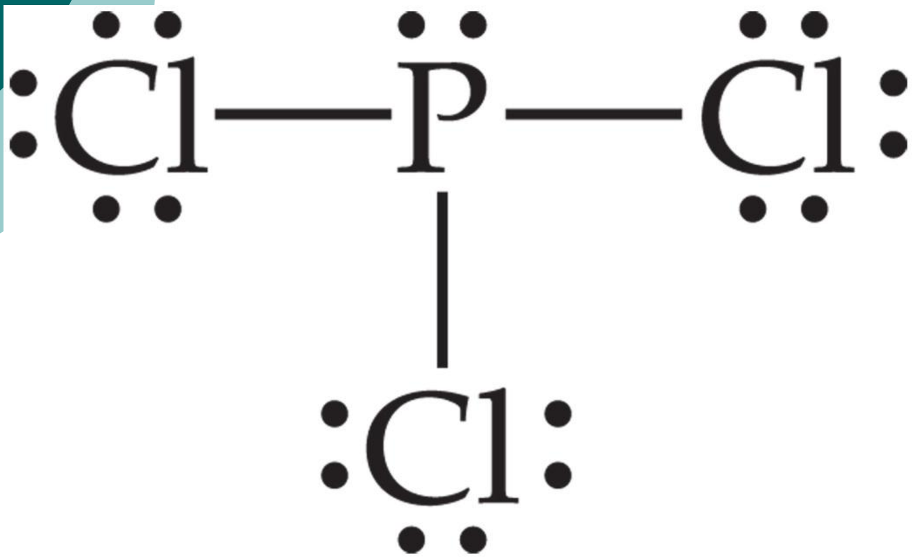


3. Fill the octets of the outer atoms.

Keep track of the electrons:

$26 - 6 = 20 - 18 = 2$

Writing Lewis Structures



4. Fill the octet of the central atom.

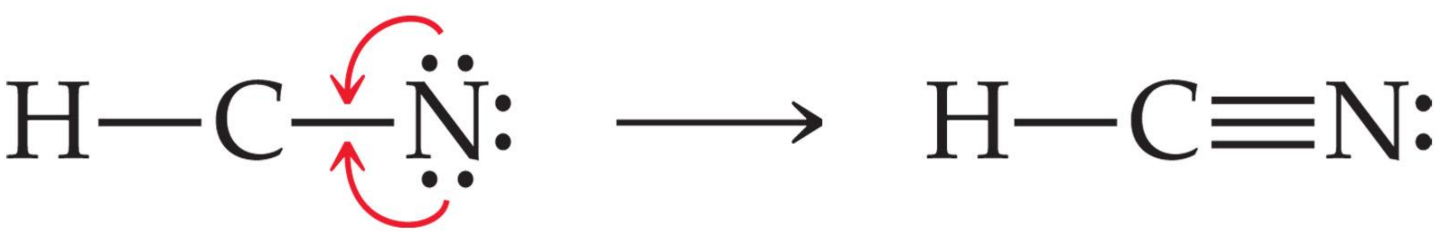
Keep track of the electrons:

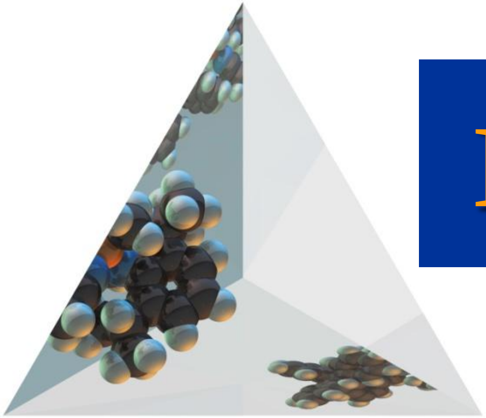
26 – 6 = 20 – 18 = 2 – 2 = 0

Writing Lewis Structures

5. If you run out of electrons before the central atom has an octet…

   …form multiple bonds until it does.

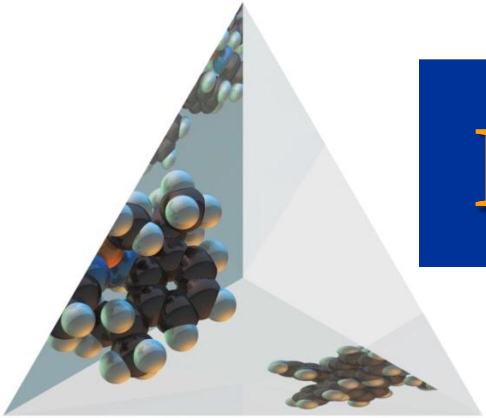




Drawing Lewis Structures

1. Sum the valence electrons from all atoms (periodic table)
2. Write symbols for the atoms and show which atoms are connected to which.
3. Complete the octet for the central atom the complete the octets of the other atoms.
4. Place leftover electrons on the central atom.
5. If there are not enough electrons to give the central atom an octet, try multiple bonds.

DRAW



Drawing Lewis Structures

Formal Charge

- It is possible to draw more than one Lewis structure with the octet rule obeyed for all the atoms.
- To determine which structure is most reasonable, we use formal charge.

N — O — N N — N — O

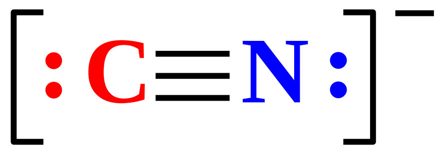
- **Formal charge** is the charge on an atom that it would have if all the atoms had the same electronegativity.

$$\text{Formal charge} = \text{valence electrons} - \left(\text{nonbonding electrons} - \frac{\text{bonding electrons}}{2}\right)$$

Drawing Lewis Structures

Formal Charge

- Consider:

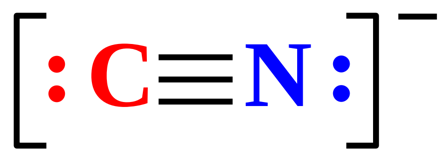


- For C:
  - There are 4 valence electrons (from periodic table).
  - In the Lewis structure there are 2 nonbonding electrons and 3 from the triple bond. There are 5 electrons from the Lewis structure.
  - Formal charge: 4 - 5 = -1.

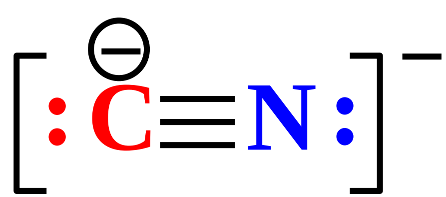
Drawing Lewis Structures

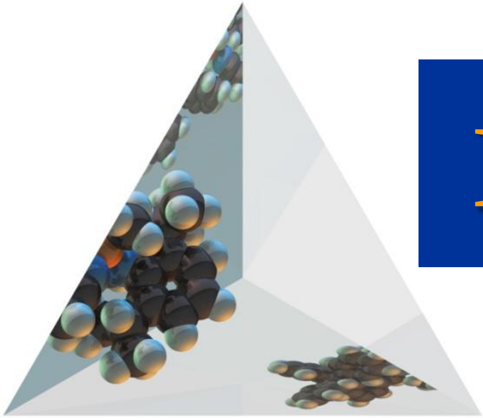
Formal Charge

- Consider:



- For N:
  - There are 5 valence electrons.
  - In the Lewis structure there are 2 nonbonding electrons and 3 from the triple bond. There are 5 electrons from the Lewis.
  - Formal charge = 5 - 5 = 0.
- We write:





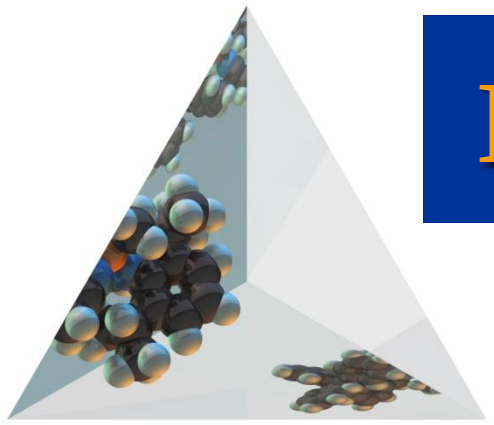
Drawing Lewis Structures

Formal Charge

- The most stable structure has:
  - the lowest formal charge on each atom,
  - the most negative formal charge on the most electronegative atoms.

Resonance Structures

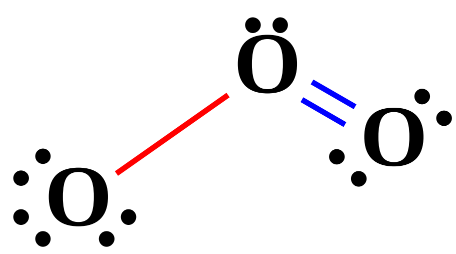
- Some molecules are not well described by Lewis Structures.
- When two or more correct Lewis structures differ only by placement of electrons, they are called ***resonance structures.***

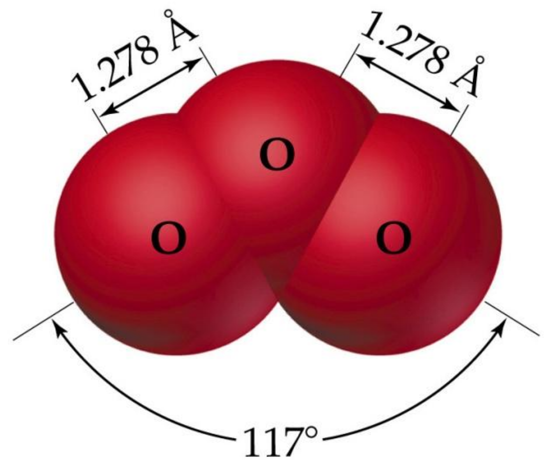


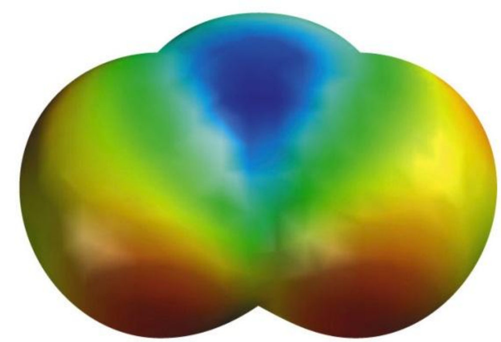
Drawing Lewis Structures

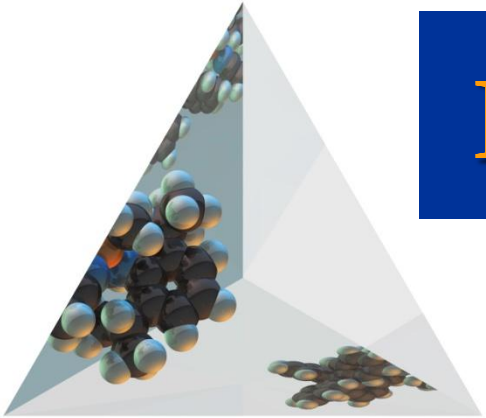
Resonance Structures

- Example: experimentally, ozone has two identical bonds whereas the Lewis Structure requires one single (longer) and one double bond (shorter).





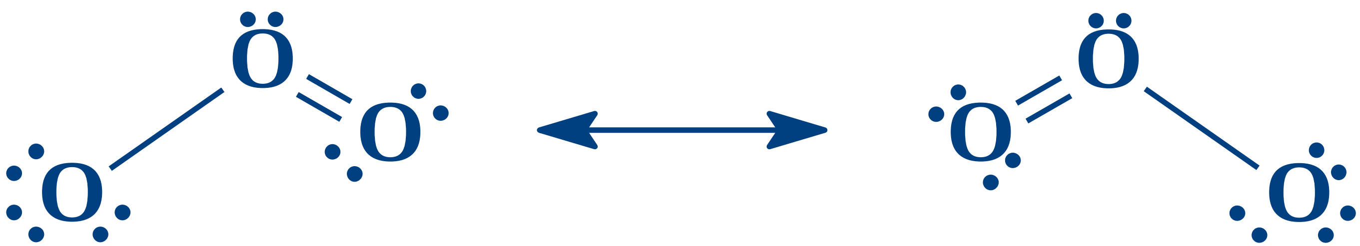




Drawing Lewis Structures

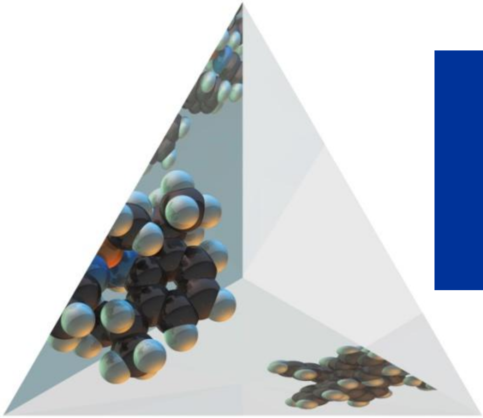
Resonance Structures

- Example: in ozone the extreme possibilities have one double and one single bond. The resonance structure has two identical bonds of intermediate character.



- Common examples: O_3, NO_3^-, SO_4^{2-}, NO_2, and benzene.

DRAW



Lewis Symbols

Draw Lewis structures for the following:

a) SiH_4

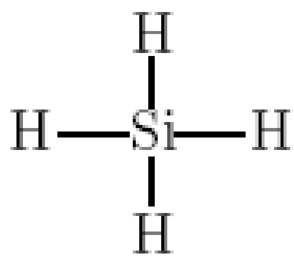
b) CO_2

c) SF_2

d) NH_2OH (N and O are bonded to one another)

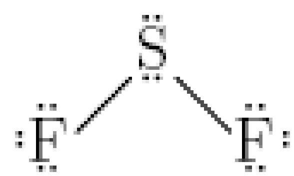
Answers

a) SiH_4

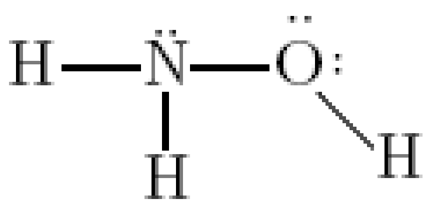


b) CO

c) SF_2



d) NH_2OH



Exceptions (Octet Rule)

“not discussed”

Covalent Bond Strength

- Most simply, the strength of a bond is measured by determining how much energy is required to break the bond.
- This is the bond enthalpy.
- The bond enthalpy for a Cl—Cl bond, *D*(Cl—Cl), is measured to be 242 kJ/mol.

Average Bond Enthalpies

- This table lists the average bond enthalpies for many different types of bonds.
- Average bond enthalpies are positive, because bond breaking is an endothermic process.

Single Bonds

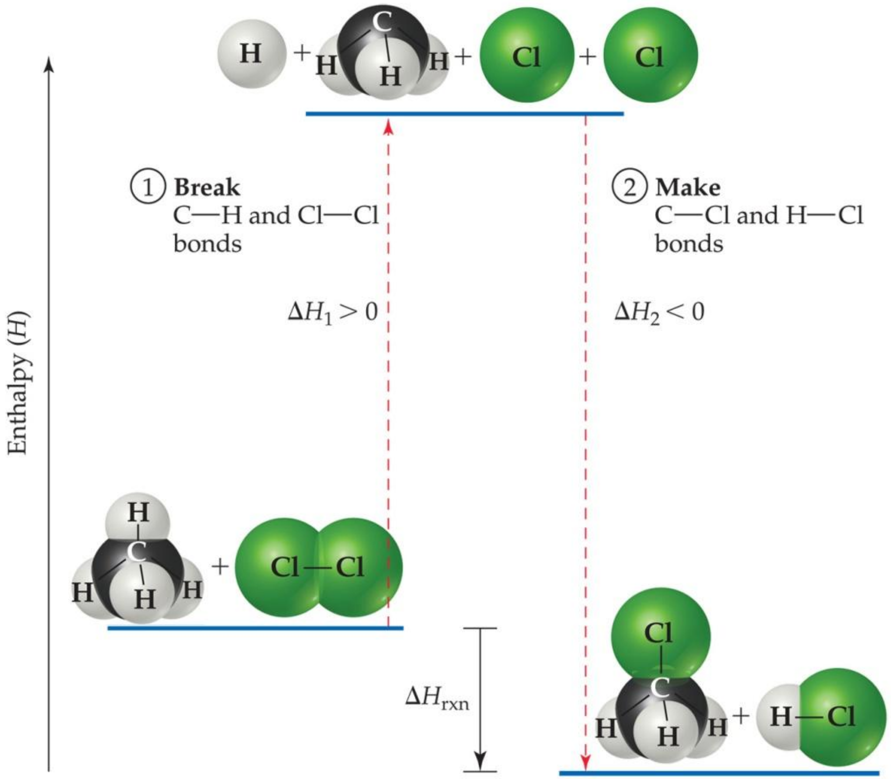
C—H	413	N—H	391	O—H	463	F—F	155
C—C	348	N—N	163	O—O	146		
C—N	293	N—O	201	O—F	190	Cl—F	253
C—O	358	N—F	272	O—Cl	203	Cl—Cl	242
C—F	485	N—Cl	200	O—I	234		
C—Cl	328	N—Br	243			Br—F	237
C—Br	276			S—H	339	Br—Cl	218
C—I	240	H—H	436	S—F	327	Br—Br	193
C—S	259	H—F	567	S—Cl	253		
		H—Cl	431	S—Br	218	I—Cl	208
Si—H	323	H—Br	366	S—S	266	I—Br	175
Si—Si	226	H—I	299			I—I	151
Si—C	301						
Si—O	368						
Si—Cl	464						

Multiple Bonds

C═C	614	N═N	418	O_2	495
C≡C	839	N≡N	941		
C═N	615	N═O	607	S═O	523
C≡N	891			S═S	418
C═O	799				
C≡O	1072				

Enthalpies of Reaction

another way to estimate ΔH for a reaction is to compare the bond enthalpies of bonds broken to the bond enthalpies of the new bonds formed.



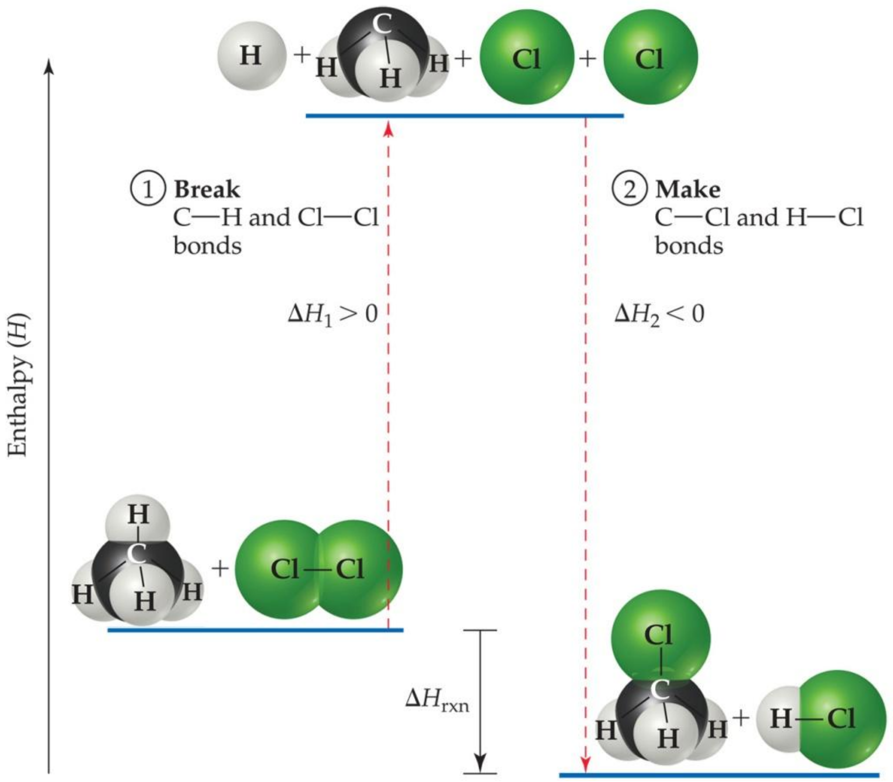
- In other words,

ΔH_{rxn} = Σ(bond enthalpies of bonds broken) – Σ(bond enthalpies of bonds formed)

Enthalpies of Reaction

$$CH_4(g) + Cl_2(g) \longrightarrow CH_3Cl(g) + HCl(g)$$

In this example, one C—H bond and one Cl—Cl bond are broken; one C—Cl and one H—Cl bond are formed.



Enthalpies of Reaction

So,

$$\Delta H_{rxn} = [D(\text{C—H}) + D(\text{Cl—Cl}) - [D(\text{C—Cl}) + D(\text{H—Cl})$$

$$= [(413 \text{ kJ}) + (242 \text{ kJ})] - [(328 \text{ kJ}) + (431 \text{ kJ})]$$

$$= (655 \text{ kJ}) - (759 \text{ kJ})$$

$$= -104 \text{ kJ}$$

Bond Enthalpy and Bond Length

Bond	Bond Length (Å)	Bond	Bond Length (Å)
C—C	1.54	N—N	1.47
C=C	1.34	N=N	1.24
C≡C	1.20	N≡N	1.10
C—N	1.43	N—O	1.36
C=N	1.38	N=O	1.22
C≡N	1.16		
		O—O	1.48
C—O	1.43	O=O	1.21
C=O	1.23		
C≡O	1.13		

- We can also measure an average bond length for different bond types.
- As the number of bonds between two atoms increases, the bond length decreases.