UNENE Chemistry Primer

Lecture 7:
Concepts of Chemical Bonding

Derek Lister and William Cook
University of New Brunswick

Course Textbook:
Chemistry, The Central Science, 10th edition, Pearson Education Inc., 2006
Theodore L. Brown, H. Eugene LeMay Jr. and Bruce E. Bursten

Chemical Bonds

- Three basic types of bond:
  - Ionic
    - Electrostatic attraction between ions
  - Covalent
    - Sharing of electrons
  - Metallic
    - Metal atoms bonded to several other atoms

Energetics of Ionic Bonding

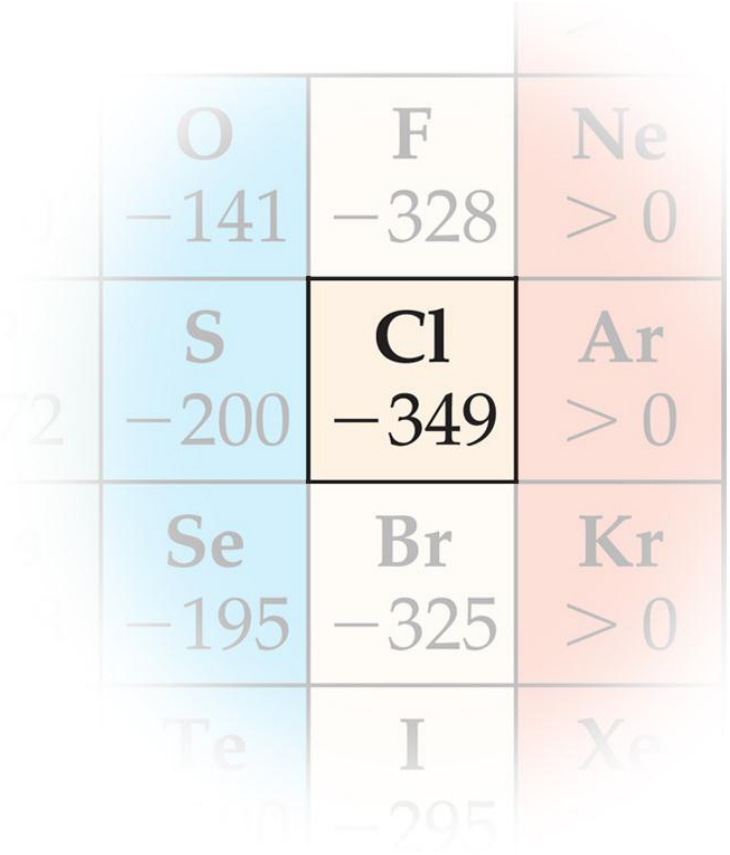
TABLE 7.2 Successive

Element	I_1
Na	495
Mg	738

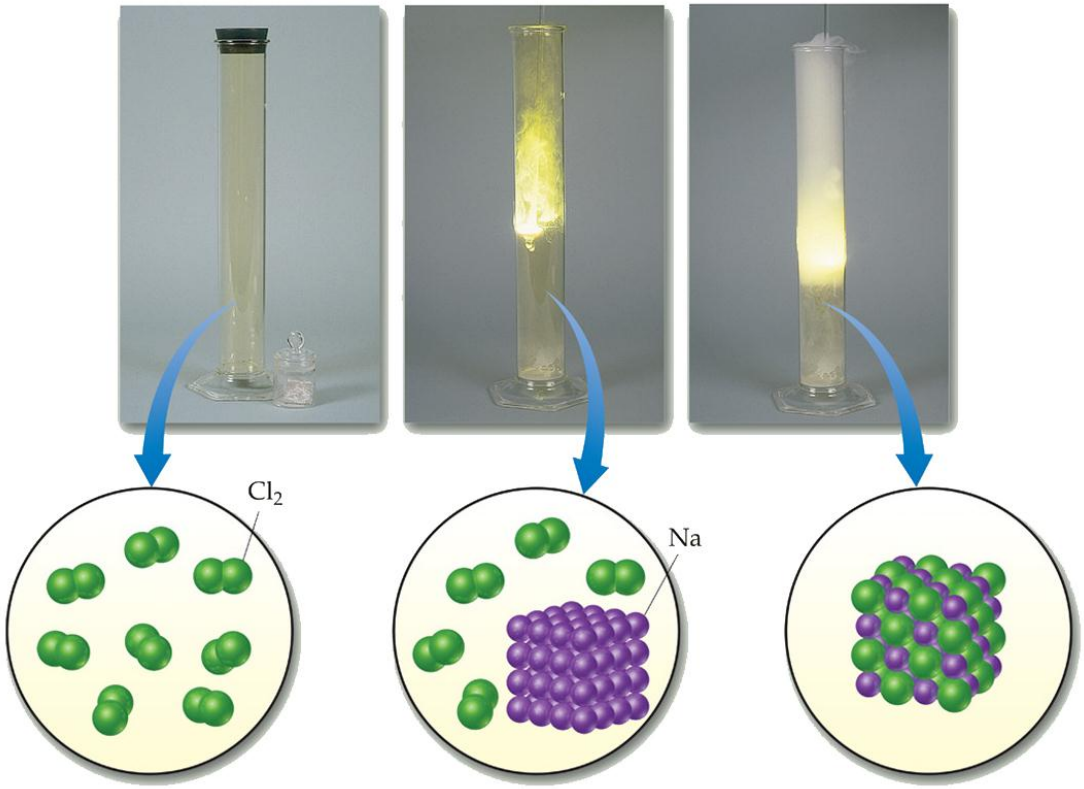
As we saw in the last lecture, it takes 495 kJ/mol to remove electrons from sodium.

Energetics of Ionic Bonding

We get 349 kJ/mol back by giving electrons to chlorine.



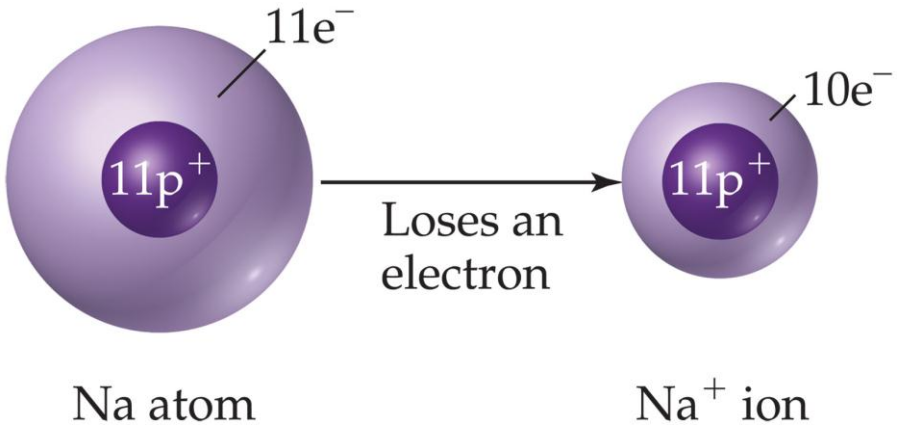
Energetics of Ionic Bonding

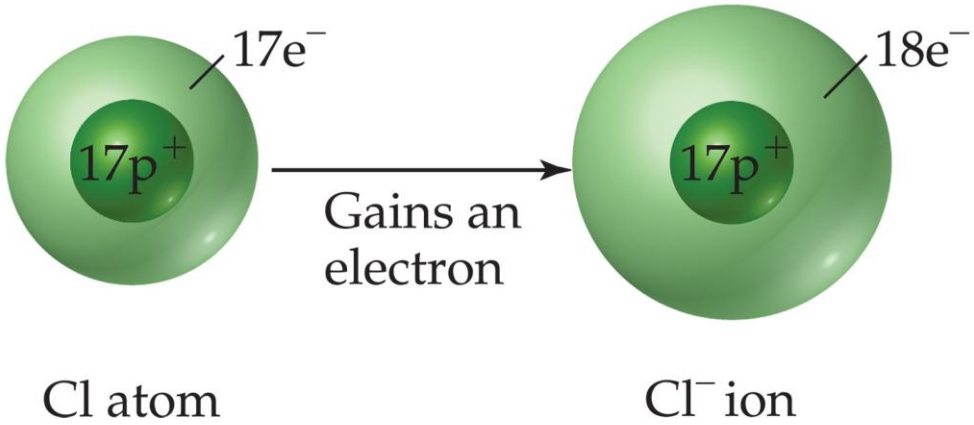


But these numbers don't explain why the reaction of sodium metal and chlorine gas to form sodium chloride is so exothermic.

Energetics of Ionic Bonding

- There must be a third piece to the puzzle.
- What is as yet unaccounted for is the electrostatic attraction between the newly formed sodium cation and chloride anion.





Lattice Energy

- This third piece of the puzzle is the lattice energy:

  The energy required to completely separate a mole of a solid ionic compound into its gaseous ions.

- The energy associated with electrostatic interactions is governed by Coulomb's law:

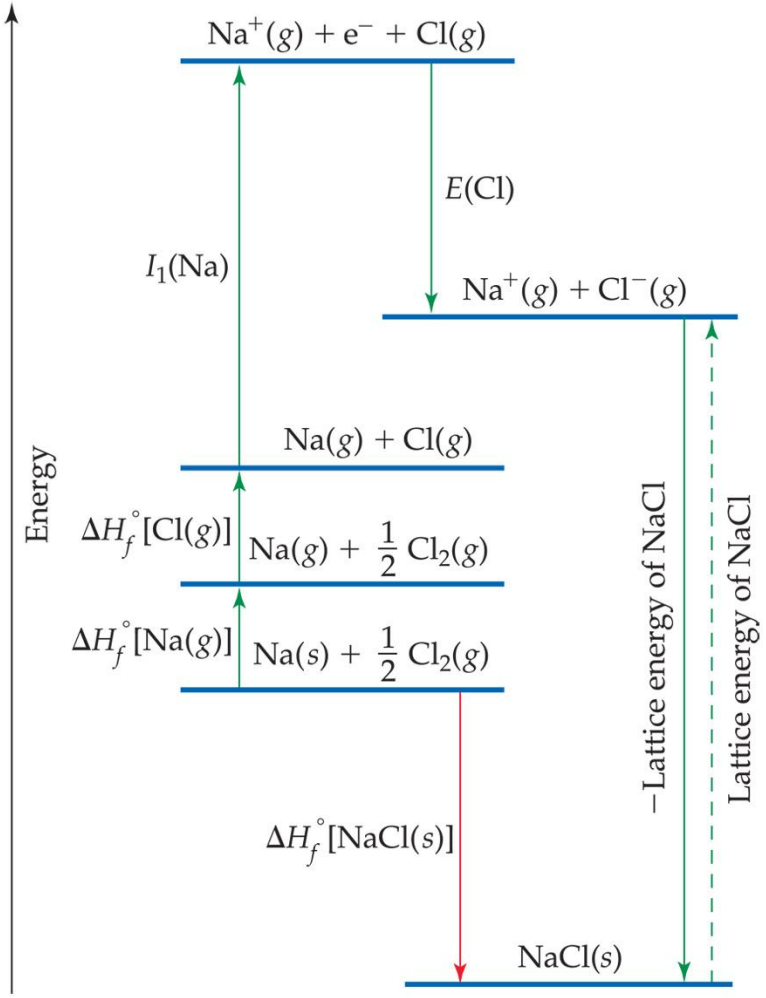
$$E_{el} = k\frac{Q_1 Q_2}{d}$$

Lattice Energy

- Lattice energy, then, increases with the charge on the ions.
- It also increases with decreasing size of ions.

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		

Energetics of Ionic Bonding



- By accounting for all three energies (ionization energy, electron affinity, and lattice energy), we can get a good idea of the energetics involved in such a process.

Energetics of Ionic Bonding

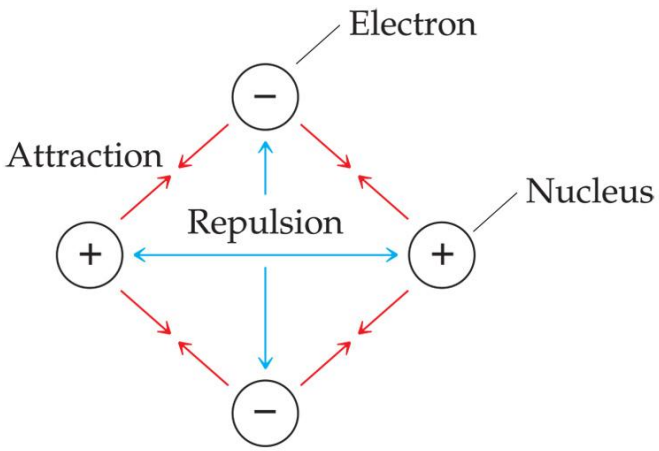
- These phenomena also help explain the "octet rule."

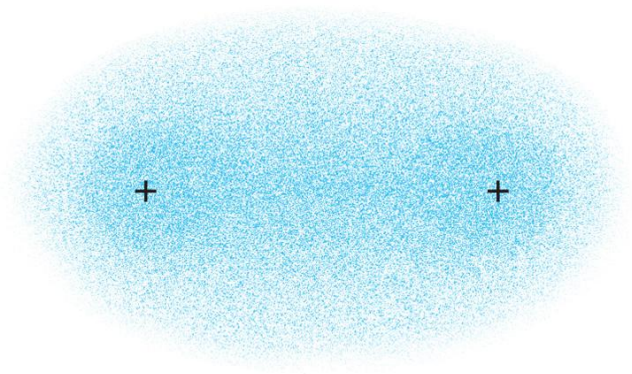
TABLE 7.2 Successive Values of Ionization E

Element	I_1	I_2	I_3
Na	495	4562	
Mg	738	1451	7733
Al	578	1817	2745
Si	786	1577	[illegible]
P	1012	1907	[illegible]

- Metals, for instance, tend to stop losing electrons once they attain a noble gas configuration because energy would be expended that cannot be overcome by lattice energies.

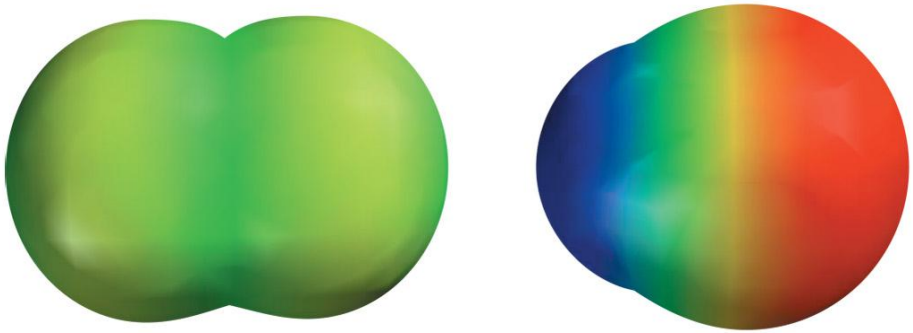
Covalent Bonding





- In these bonds atoms share electrons.
- There are several electrostatic interactions in these bonds:
  - Attractions between electrons and nuclei;
  - Repulsions between electrons;
  - Repulsions between nuclei.

Polar Covalent Bonds



F_2 HF

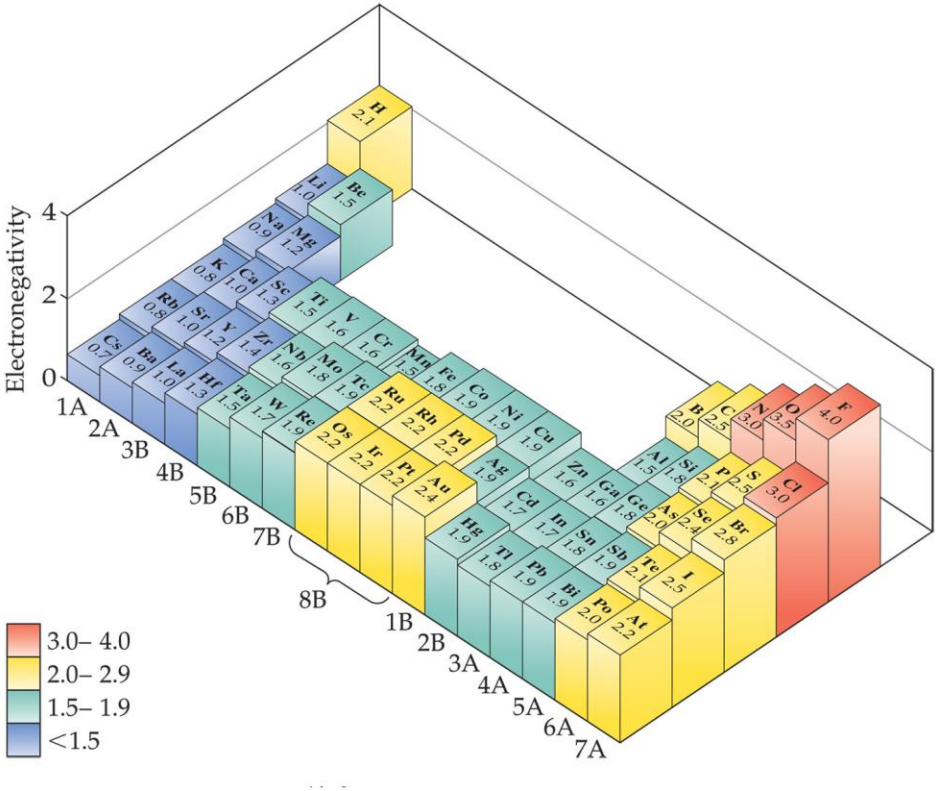
- Although atoms often form compounds by sharing electrons, the electrons are not always shared equally.

- Fluorine pulls harder on the electrons it shares with hydrogen than hydrogen does.

- Therefore, the fluorine end of the molecule has more electron density than the hydrogen end.

Electronegativity:

- The ability of atoms in a molecule to attract electrons to itself.
- On the periodic chart, electronegativity increases as you go…
  - …from left to right across a row.
  - …from the bottom to the top of a column.



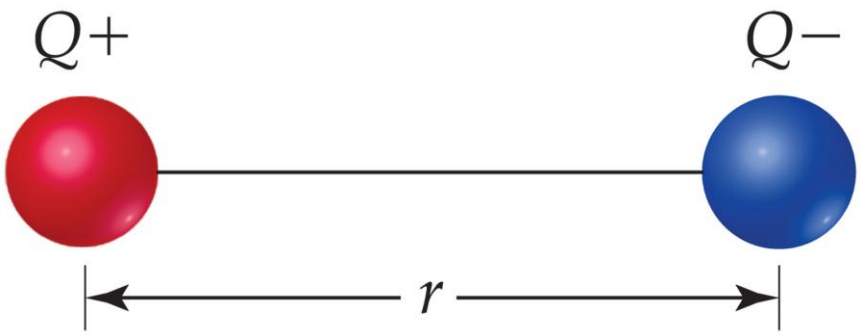
Polar Covalent Bonds

- 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:

$$\mu = Qr$$

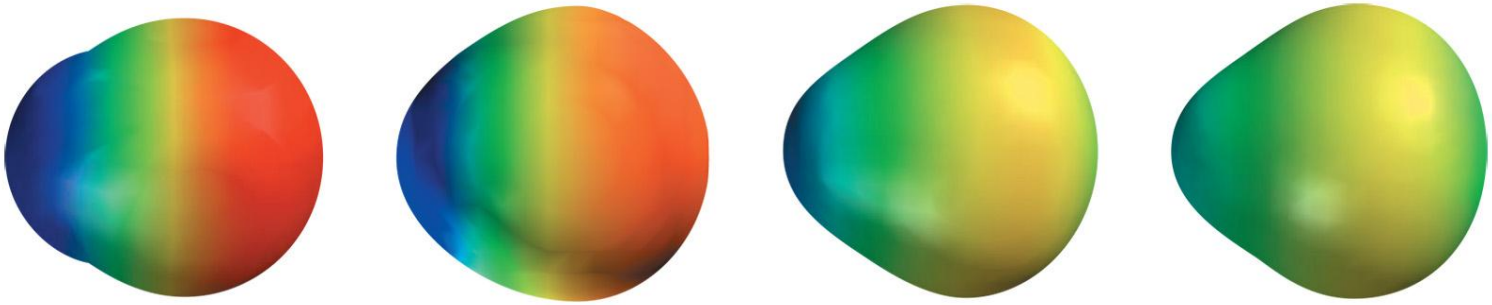
- It is measured in debyes (D).



Polar Covalent Bonds

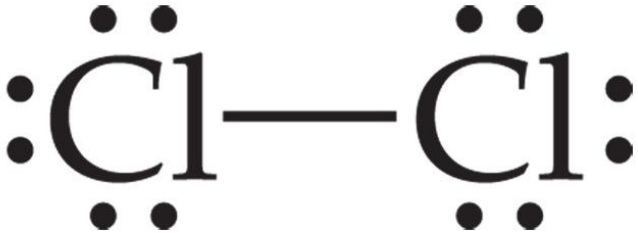
Compound	Bond Length (Å)	Electronegativity Difference	Dipole Moment (D)
HF	0.92	1.9	1.82
HCl	1.27	0.9	1.08
HBr	1.41	0.7	0.82
HI	1.61	0.4	0.44

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



HF HCl HBr HI

Lewis Structures



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

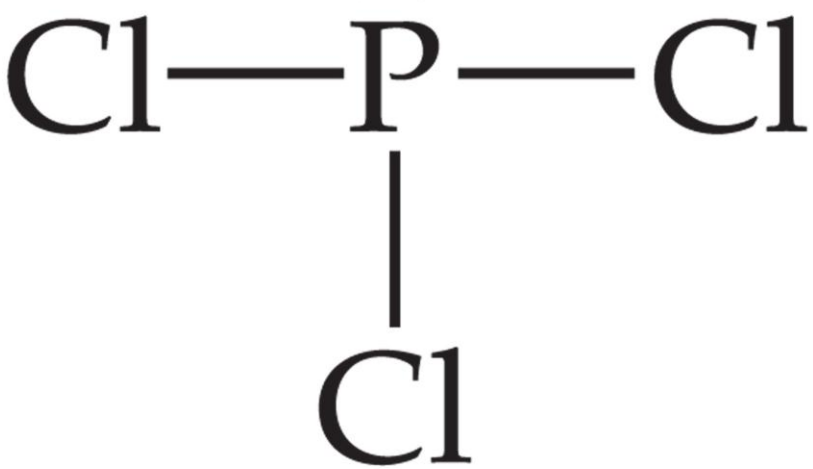
Writing Lewis Structures

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



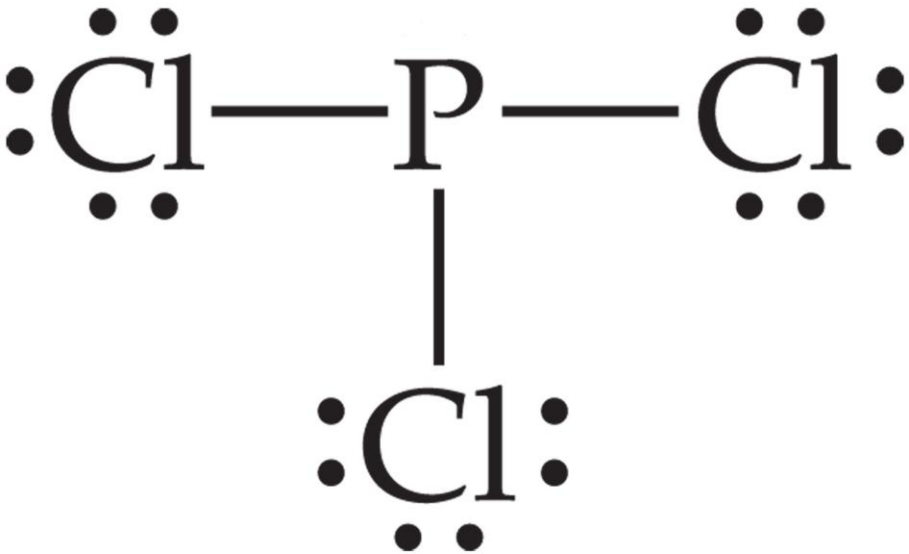
2. The central atom is the *least* electronegative element that isn't hydrogen. Connect the outer atoms to it by single bonds.

Keep track of the electrons:

26 - 6 = 20

Writing Lewis Structures

3. Fill the octets of the outer atoms.

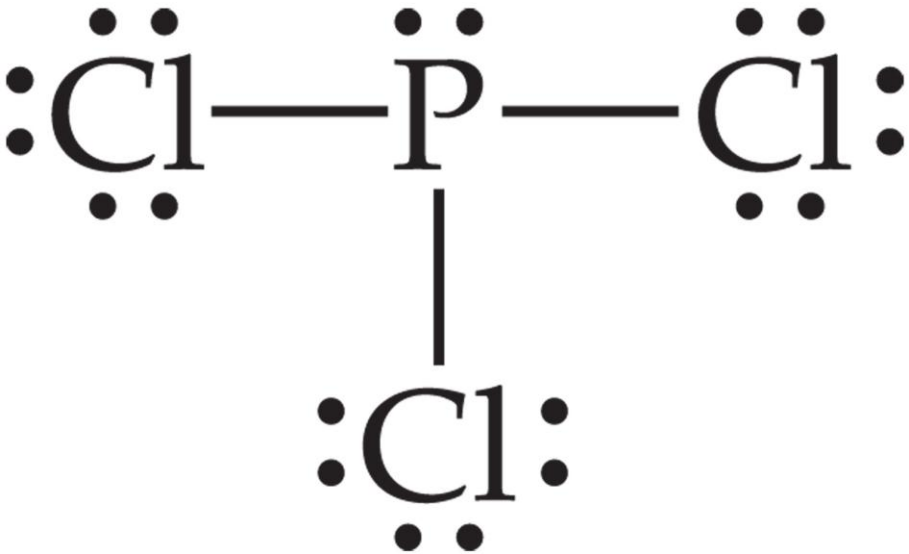


Keep track of the electrons:

26 - 6 = 20; 20 - 18 = 2

Writing Lewis Structures

4. Fill the octet of the central atom.



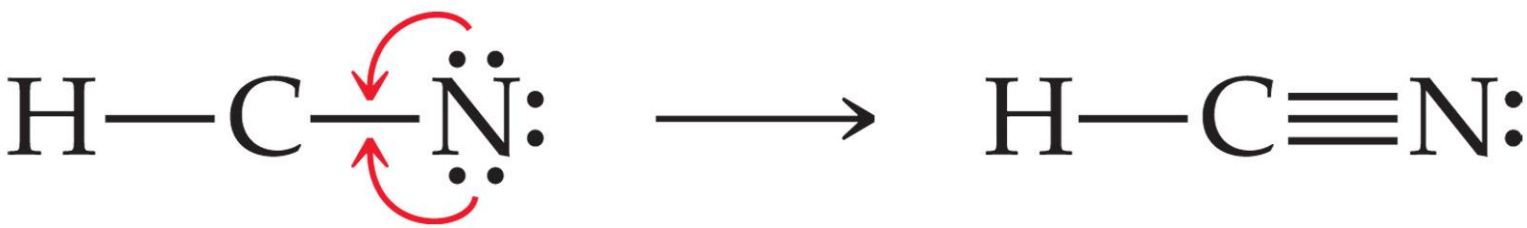
Keep track of the electrons:

26 - 6 = 20; 20 - 18 = 2; 2 - 2 = 0

Writing Lewis Structures

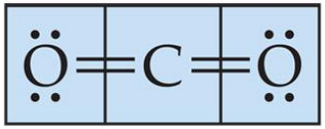
5. If you run out of electrons before the central atom has an octet…

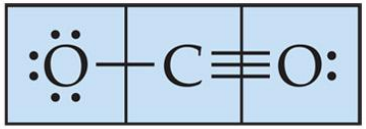
…form multiple bonds until it does.



Writing Lewis Structures

- Then assign formal charges.
  - For each atom, count the electrons in lone pairs and half the electrons it shares with other atoms.
  - Subtract that from the number of valence electrons for that atom: The difference is its formal charge.

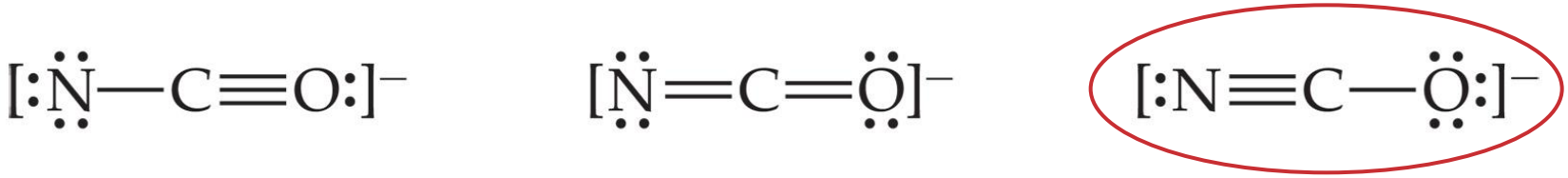




Valence electrons:	6	4	6	6	4	6
−(Electrons assigned to atom):	6	4	6	7	4	5
Formal charge:	0	0	0	−1	0	+1

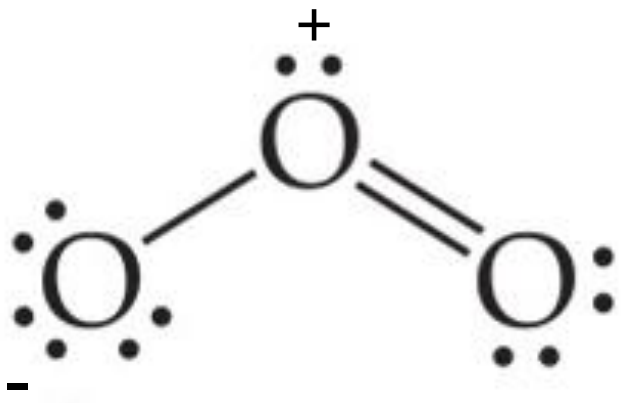
Writing Lewis Structures

- The best Lewis structure:
  - is the one with the fewest charges.
  - puts a negative charge on the most electronegative atom.

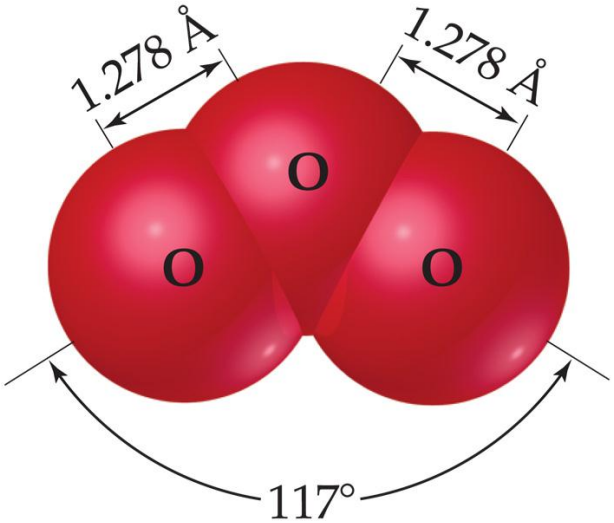


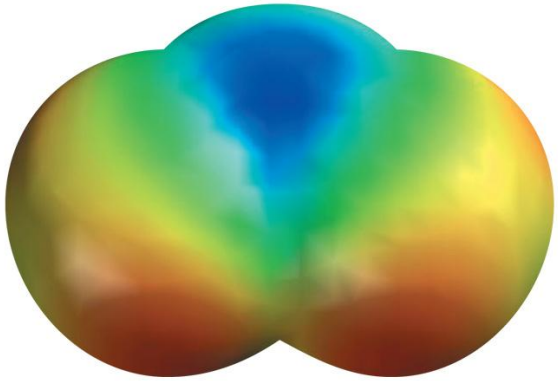
Resonance

This is the Lewis structure we would draw for ozone, O_3.



Resonance





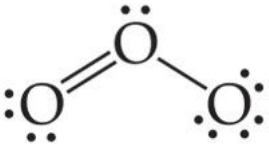
- But this is at odds with the true, observed structure of ozone, in which:
  - both O—O bonds are the same length.
  - both outer oxygens have a charge of -1/2.

Resonance

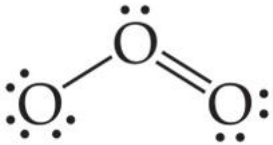
- One Lewis structure cannot accurately depict a molecule such as ozone.

- We use multiple structures, resonance structures, to describe the molecule.

Resonance structure



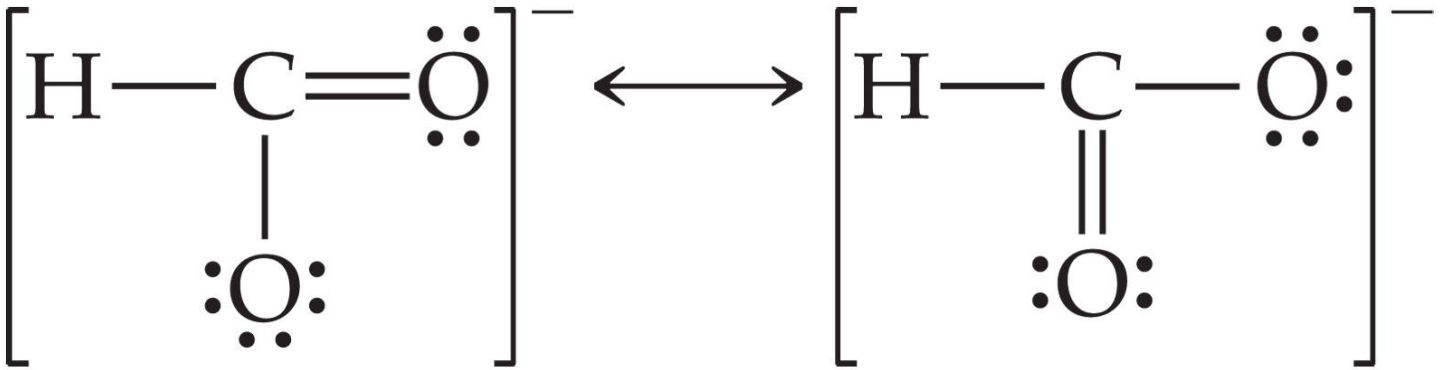
Resonance structure



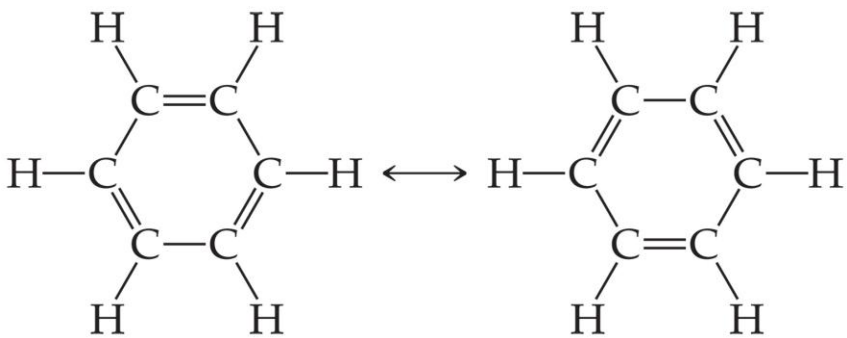
Resonance

- In truth, the electrons that form the second C—O bond in the double bonds below do not always sit between that C and that O, but rather can move among the two oxygen and the carbon.

- They are not localized, but rather are delocalized.



Resonance

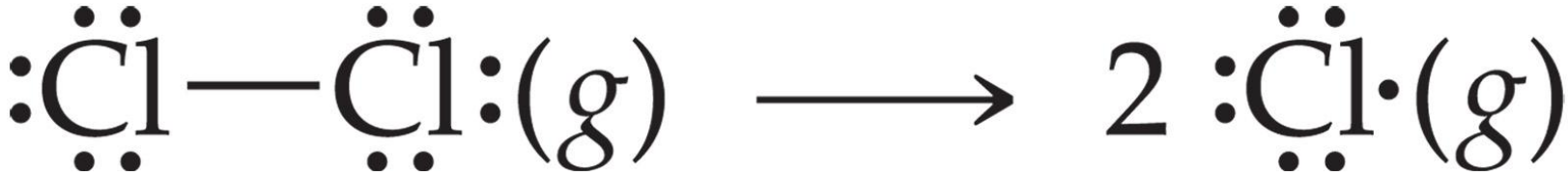


- The organic compound benzene, C_6H_6, has two resonance structures.
- It is commonly depicted as a hexagon with a circle inside to signify the delocalized electrons in the ring.

Exceptions to the Octet Rule

- There are three types of ions or molecules that do not follow the octet rule:
  - Ions or molecules with an odd number of electrons.
  - Ions or molecules with less than an octet.
  - Ions or molecules with more than eight valence electrons (an expanded octet).

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				

Average Bond Enthalpies

NOTE: These are *average* bond enthalpies, not absolute bond enthalpies; the C—H bonds in methane, CH_4, will be a bit different than the C—H bond in chloroform, $CHCl_3$.

Single Bonds

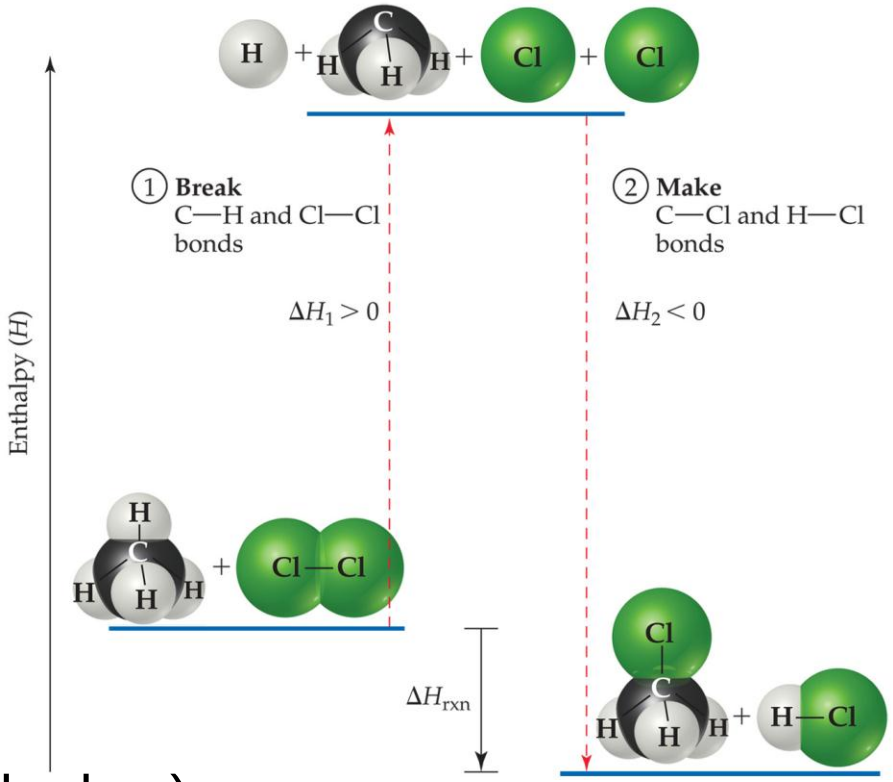
C—H	413	N—H	391	O—H	463	F—F	155
C—C	348	N—N	163	O—O	146		
C—N	293	N—O	201	O—F	190	Cl—F	253
C—O	358	N—F	272	O—Cl	203	Cl—Cl	242
C—F	485	N—Cl	200	O—I	234		
C—Cl	328	N—Br	243			Br—F	237
C—Br	276			S—H	339	Br—Cl	218
C—I	240	H—H	436	S—F	327	Br—Br	193
C—S	259	H—F	567	S—Cl	253		
		H—Cl	431	S—Br	218	I—Cl	208
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

- Yet another way to estimate ⊗*H* for a reaction is to compare the bond enthalpies of bonds broken to the bond enthalpies of the new bonds formed.



$$DH_{rxn} = \sum(\text{bond enthalpies of bond broken}) - \sum(\text{bond enthalpies of bonds formed})$$

Enthalpies of Reaction

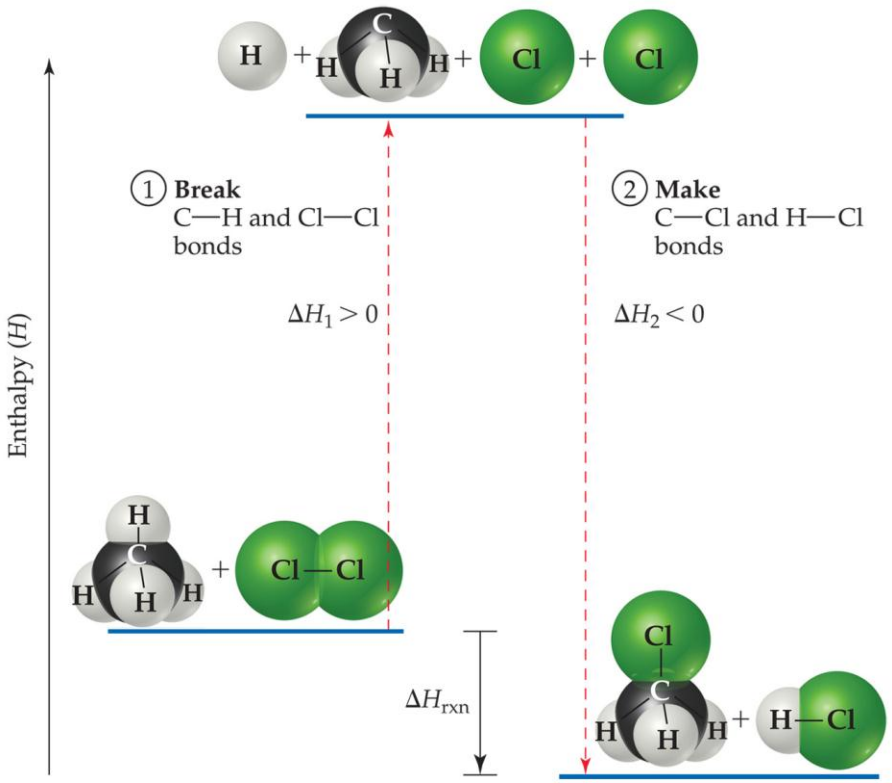
$$CH_{4(g)} + Cl_{2(g)} \rightarrow CH_3Cl_{(g)} + HCl_{(g)}$$

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,

$\otimes H_{rxn}$ = [*D*(C—H) + *D*(Cl—Cl)] – [*D*(C—Cl) + *D*(H—Cl)]

= [(413 kJ) + (242 kJ)] - [(328 kJ) + (431 kJ)]

= (655 kJ) - (759 kJ)

= -104 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.