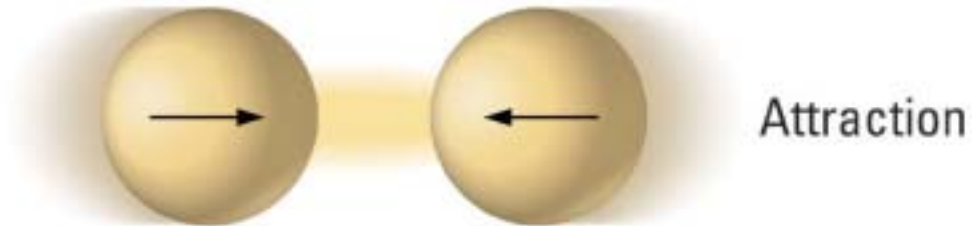


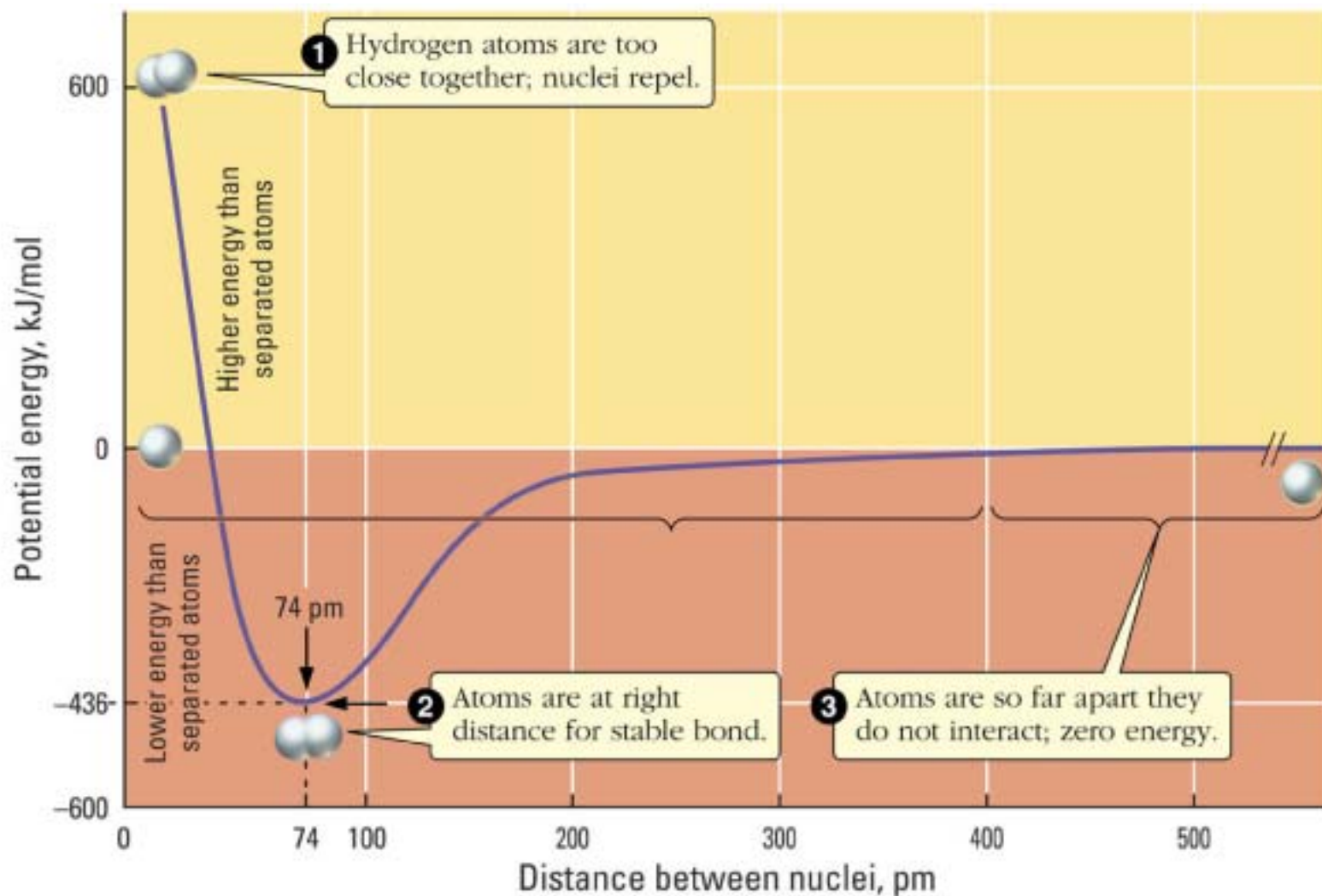
Chapter 8

Covalent Bonding

Atomic Interactions



H-H Bond Formation



Lewis Structures

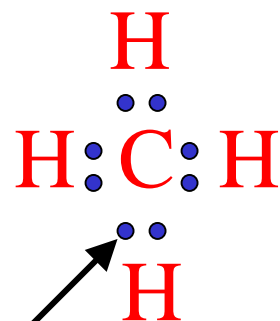
- combinations of Lewis symbols which represent possible bonding

Octet Rule

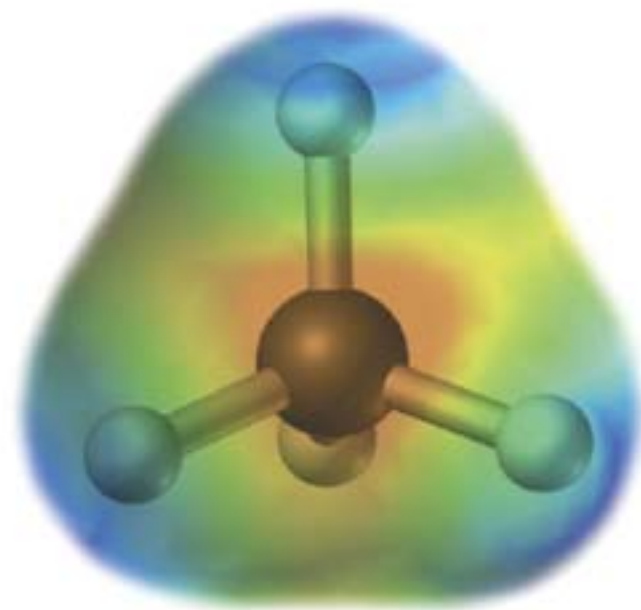
In compound formation an atom gains or loses electrons, or shares pairs of electrons, until its valence shell has eight electrons.

Lewis Structures

CH₄ methane

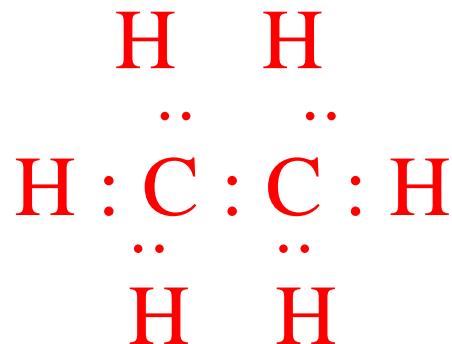


bonding pair
of electrons

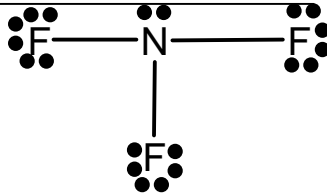
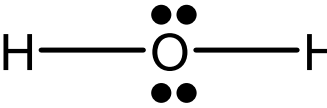
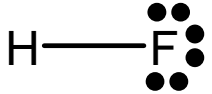


Lewis Structures

C_2H_6 ethane



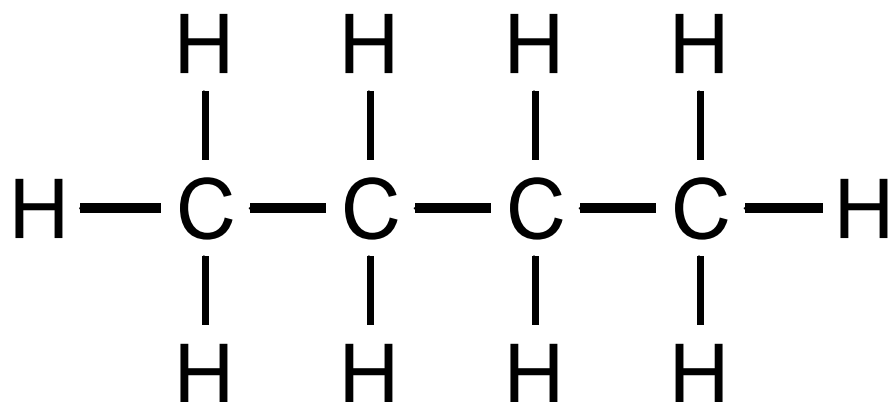
Selected Lewis Structures

Group number	Number of valence electrons	Number of electrons shared to complete an octet (8-group number)	Example
4A	4	4	C in CH ₄
5A	5	3	N in NF ₃ 
6A	6	2	O in H ₂ O 
7A	7	1	F in HF 

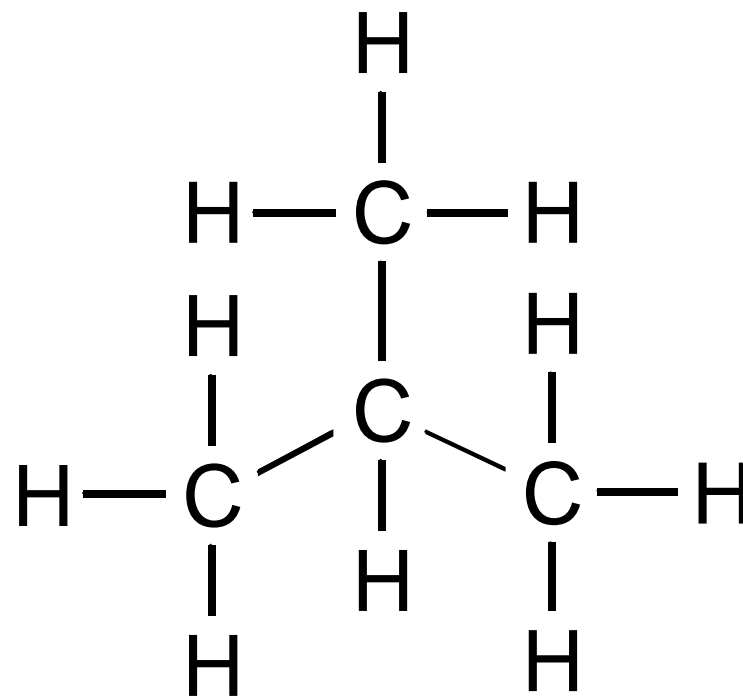
Guidelines for Writing Lewis Structures

1. Add up the number of valence electrons in the molecule or ion, adding electrons for each negative charge on ion and removing electrons for each positive charge.
2. Create a skeleton structure using the chemical symbols for the atoms and a line for bonding pairs of electrons.
3. Satisfy the Octet Rule for each atom, starting with the central atom.
4. All remaining electrons should be placed around central atom
5. Add multiple bonds to central atom if less than eight electrons are around central atom.

Saturated hydrocarbons

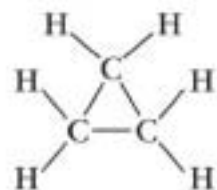


butane

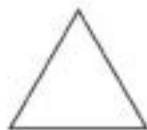


2-methylpropane

Cycloalkanes



or



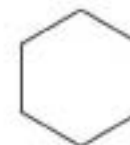
cyclopropane
 C_3H_6



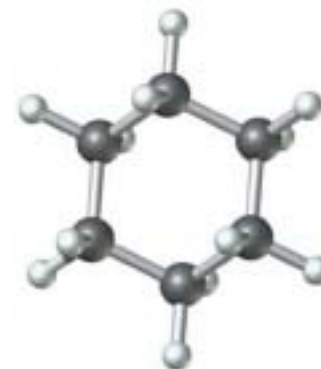
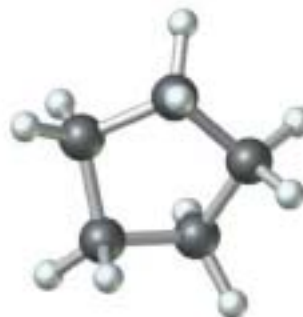
cyclobutane
 C_4H_8



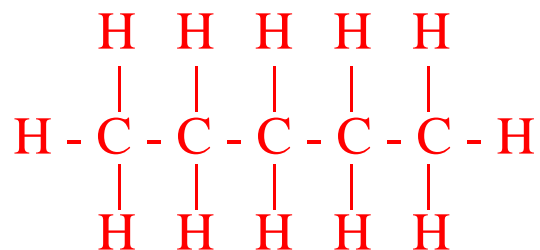
cyclopentane
 C_5H_{10}



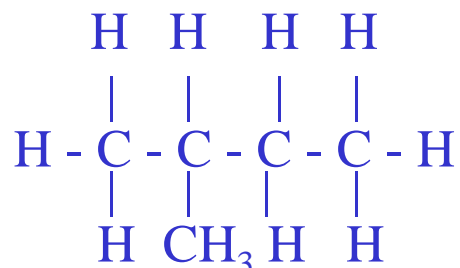
cyclohexane
 C_6H_{12}



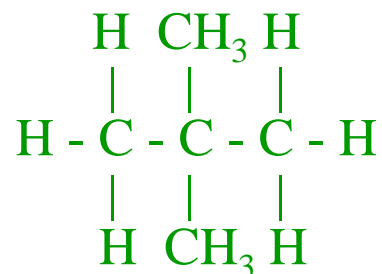
Pentanes



pentane



2-methylbutane



2,2-dimethylpropane

Lewis Structures

CO₂ carbon dioxide

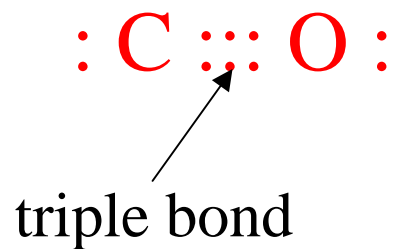


double bonds

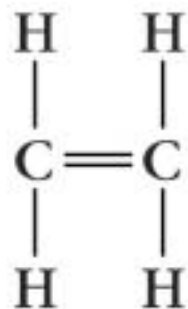


Lewis Structures

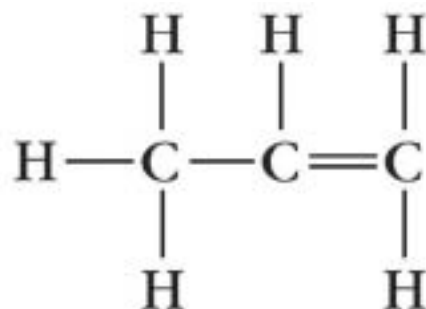
CO carbon monoxide



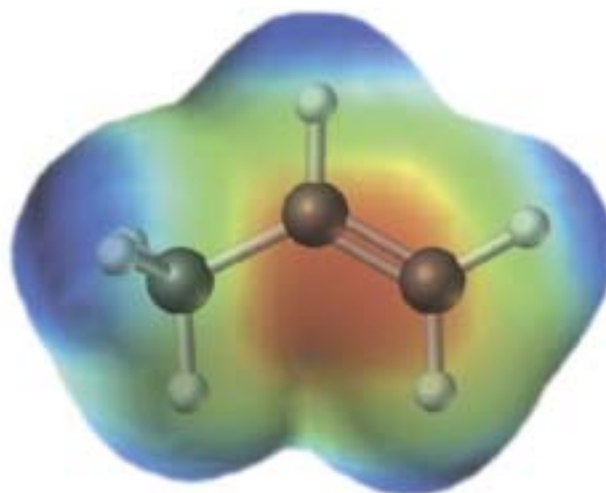
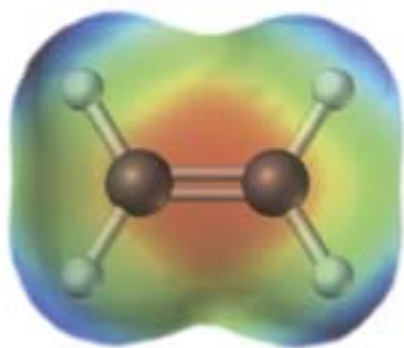
Multiple Covalent Bonds in Hydrocarbons



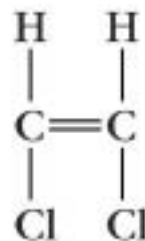
ethylene



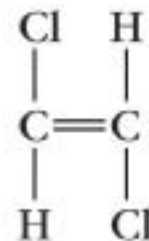
propylene



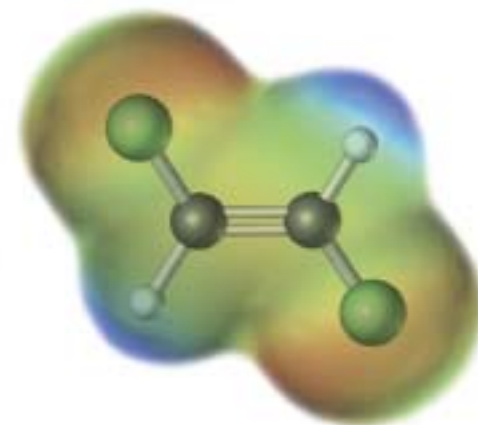
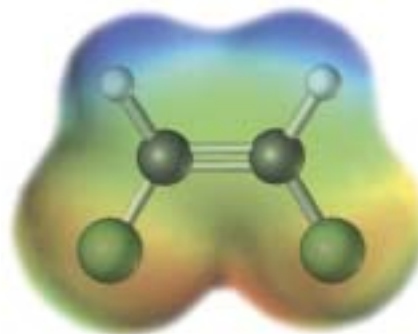
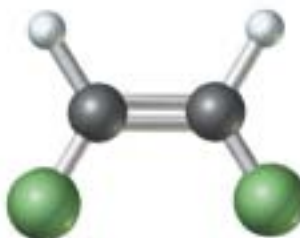
Double Bonds and Isomerism



cis isomer



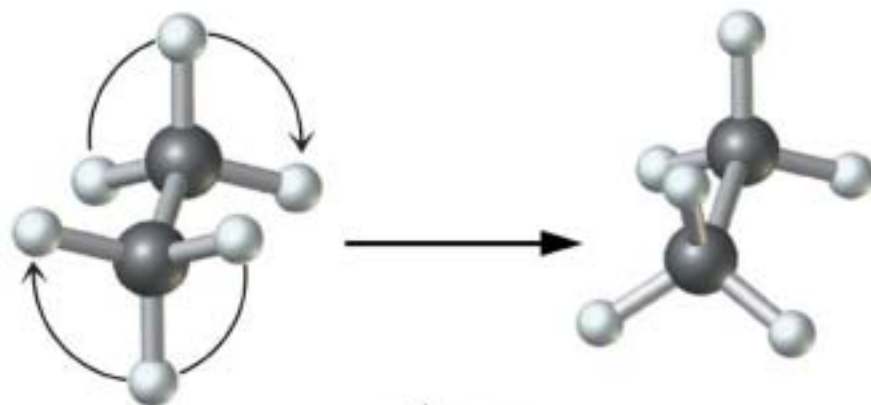
trans isomer



Properties of cis- and trans- 1,2-dichloroethene

Physical Property	<i>cis</i> -1,2-dichloroethene	<i>trans</i> -1,2-dichloroethene
Melting Point	-80.5 °C	-50.0 °C
Boiling Point (at 1 atm)	60.3 °C	47.5 °C
Density (at 20 °C)	1.284 g/ml	1.265 g/ml

Non-rotation around $C = C$



ethane

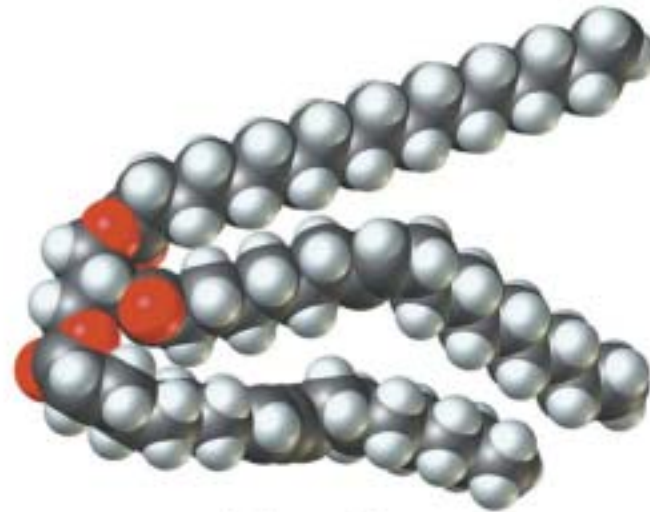
Rotation along the carbon-to-carbon single bond axis occurs freely in ethane...



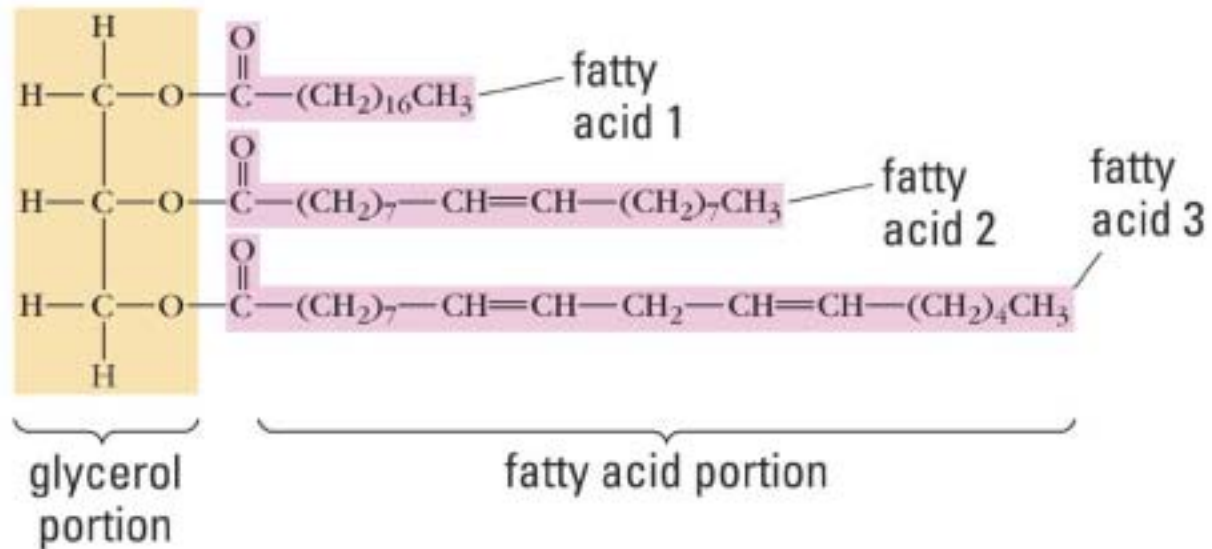
ethylene

...but not in ethylene due to its $C = C$ double bond.

Triglyceride



triglyceride



Fats

saturated fats – fats containing only C–C single bonds

unsaturated fats – fats containing one or more C=C double bonds

monounsaturated fats – fats containing one C=C double bond

polyunsaturated fats – fats containing two or more C=C double bonds

Fatty Acids

animal fat – primarily saturated fatty acids

vegetable fat – primarily unsaturated fatty acids

more unsaturation, lower melting point, more liquid in character

most natural unsaturated fatty acids are cis-configuration, producing lower melting points than corresponding trans- isomers

Bond Energy

TABLE 8.2 Average Bond Energies (in kJ/mol)*

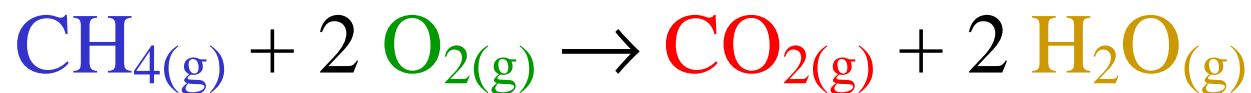
Single bonds											
	I	Br	Cl	S	P	Si	F	O	N	C	H
H	299	366	431	347	322	323	566	467	391	416	436
C	213	385	327	272	264	301	486	336	285	356	
N	—	—	193	—	~ 200	335	272	201	160		
O	201	—	205	—	~ 340	368	190	146			
F	—	—	255	326	490	582	158				
Si	234	310	391	226	—	226					
P	184	264	319	—	209						
S	—	213	255	226							
Cl	209	217	242								
Br	180	193									
I	151										

Multiple bonds

N=N	418	C=C	598
N≡N	946	C≡C	813
C=N	616	C=O (in CO ₂ , O=C=O)	803
C≡N	866	C=O (as in H ₂ C=O)	695
O=O (in O ₂)	498	C≡O	1073

* Data from Cotton, F. A., Wilkinson, G., and Gaus, P. L. *Basic Inorganic Chemistry*, 3rd ed. New York: Wiley, 1995; p. 12.

EXAMPLE: Using bond energy data, calculate the ΔH° for:



$$\begin{aligned}\Delta H^\circ &= ([4(414) + 2(498)]_{\text{broken}} \\ &\quad - [2(715) + 4(464)]_{\text{formed}}) \text{kJ/mol} \\ &= ([1656 + 996] - [1430 + 1856]) \text{kJ/mol} \\ &= ([2652] - [3286]) \text{kJ/mol} \\ &= -634 \text{ kJ/mol}\end{aligned}$$

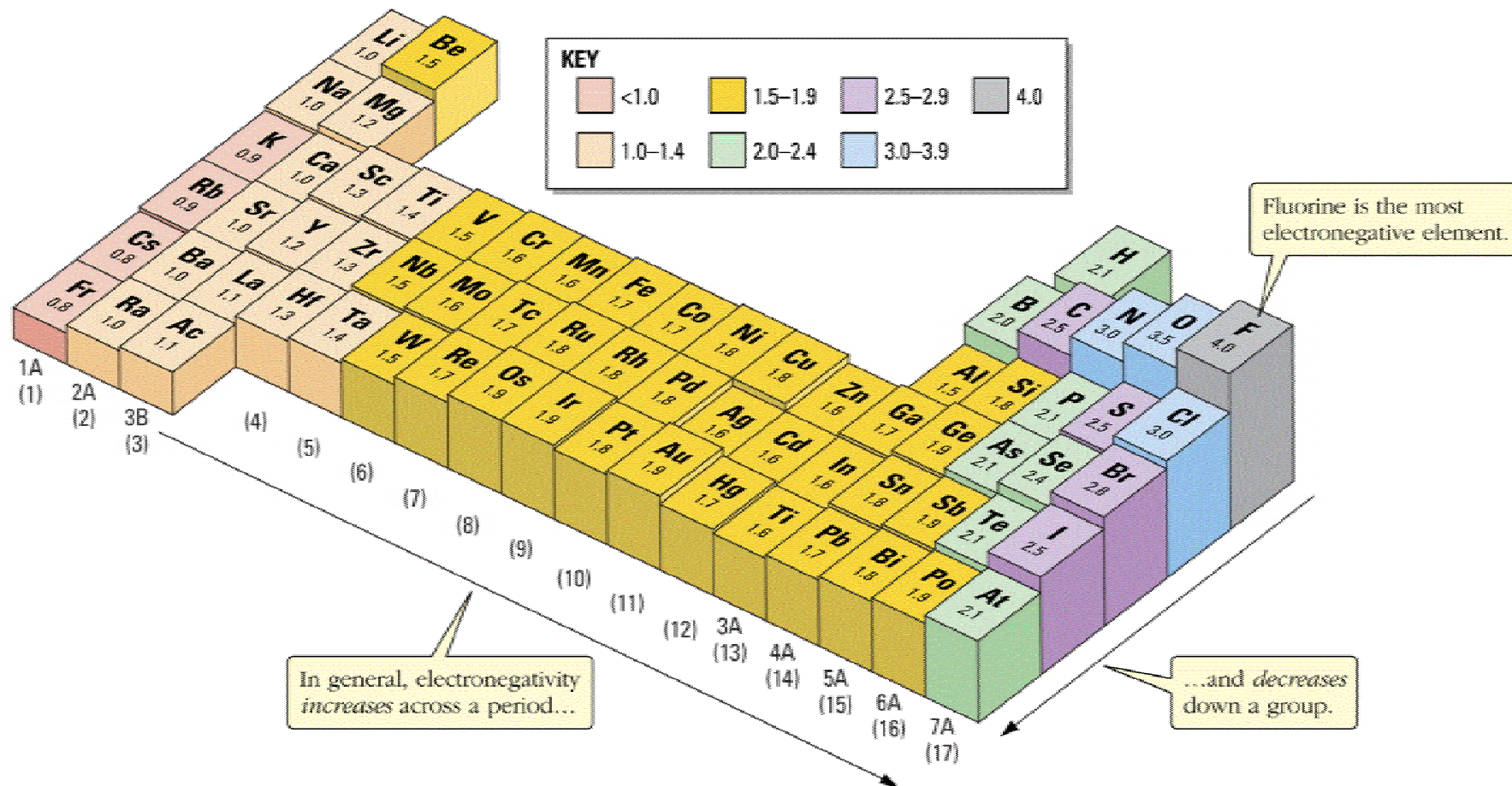
$$\text{from } \Delta H_f^\circ \text{ data: } \Delta H^\circ = -802.3 \text{ kJ/mol}$$

Electronegativity

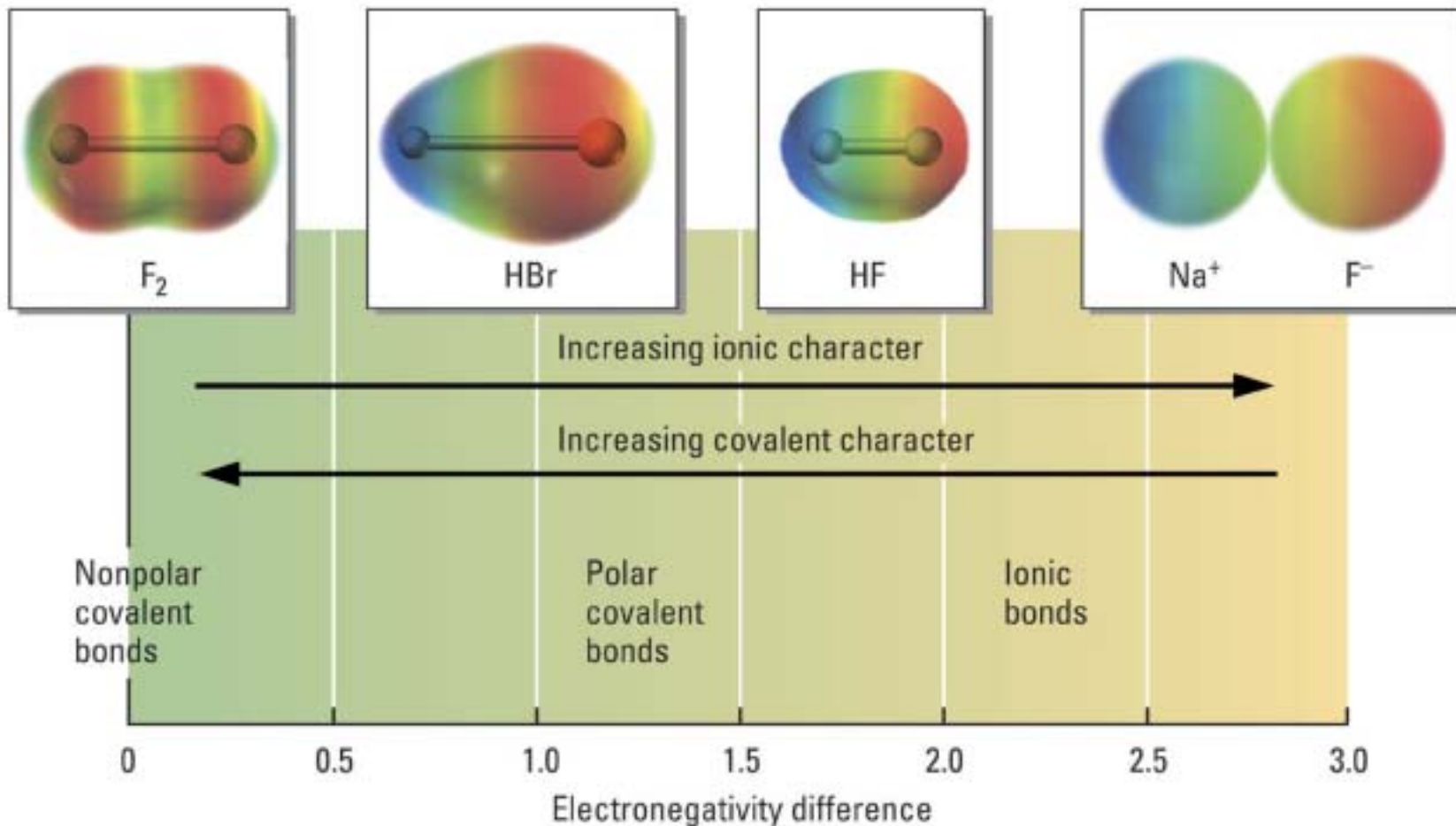
Pauling Scale

- relative attraction of an atom for electrons, its own and those of other atoms
- same trends as ionization energy, increases from lower left corner to the upper right corner
- fluorine: E.N. = 4.0

Trends in Electronegativity



Bond Characteristics



Covalent Bond Properties

electronegativity

nonpolar bonds $\rightarrow \Delta\text{EN} = 0$

polar bonds $\rightarrow \Delta\text{EN} > 0$

ionic bonds $\rightarrow \Delta\text{EN} > 1.5$

Formal Charge

Formal Charge is the charge on an atom if all of the shared electrons were equally shared

- lone pair electrons are assigned to atom about which they are found
- half of the bonding electrons are assigned to each atom of the bonded pair
- the sum of formal charges on a molecule is zero and the magnitude of the charge on ions

Formal Charge Calculations

Formal charge = (# valence electrons on an atom)
- [(number of lone pair electrons)
+ (1/2 number of shared electrons)]

	N	C	O
Valence electrons	5	4	6
Lone pair electrons	2	0	6
1/2 shared electrons	3	4	1
<i>Formal charge</i>	0	0	-1

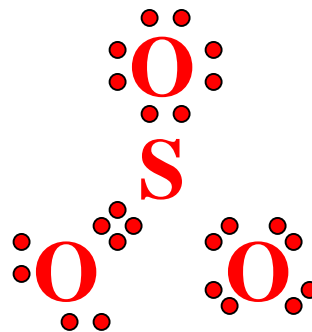
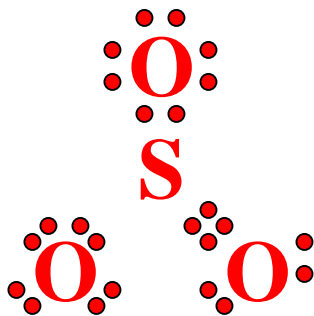
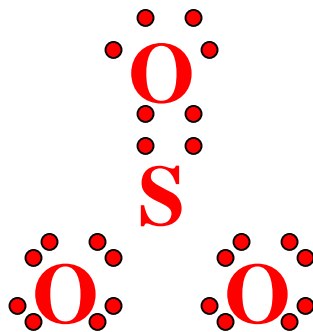
Resonance

Resonance is present when two or more electron dot structures that differ only in the arrangement of the electrons. The *resonant hybrid*, an average structure, best represents the structure

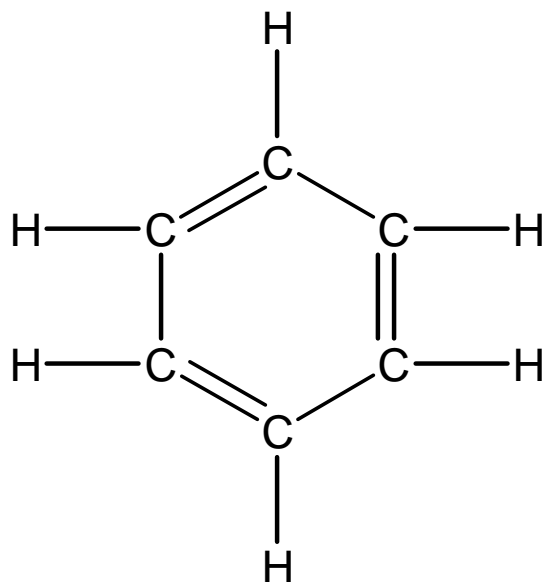
Ozone



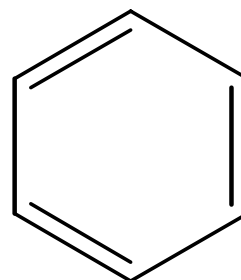
Resonance in SO_3



Aromatics

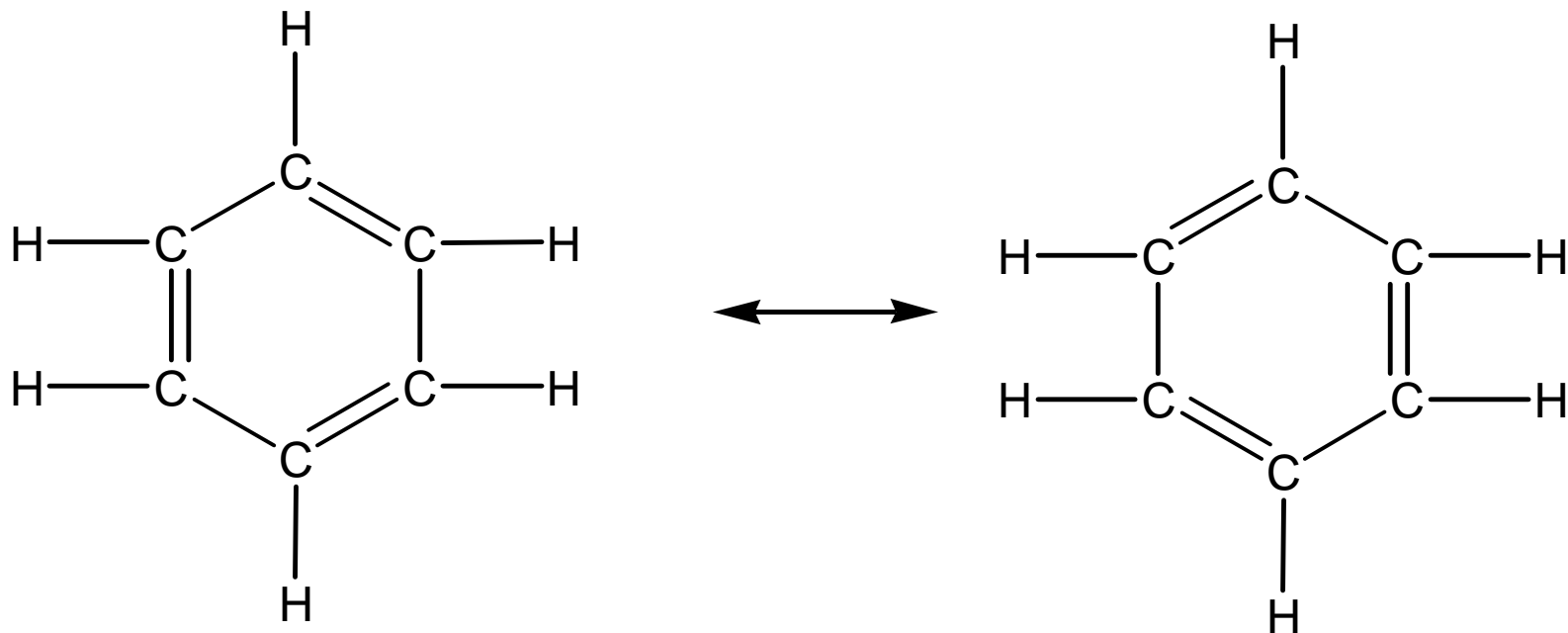


or

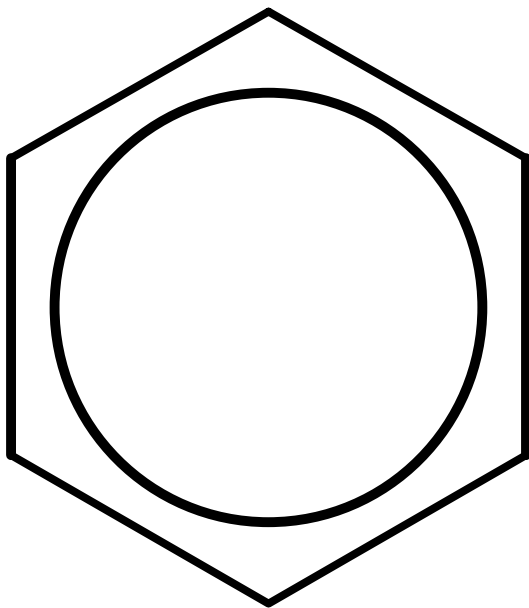


benzene

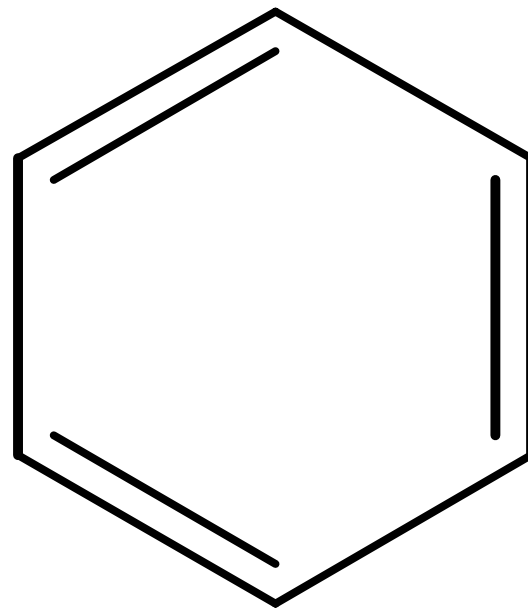
Aromatics



Aromatics



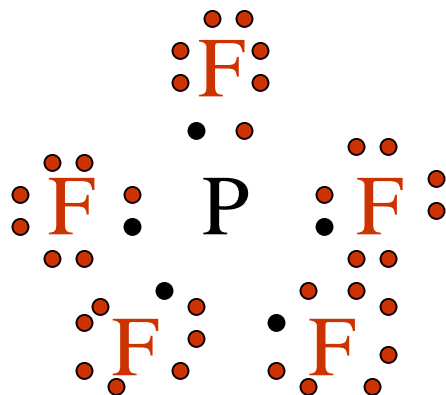
rather than



Exceptions to Octet Rule

expanded octet

PF₅



Exceptions to Octet Rule

NO nitric oxide

less than octet

