Bonding

ionic bonding- the transfer of valence electrons between atoms

$$Na_{(g)} \longrightarrow Na_{(g)}^{+1} + 1e$$

$$F_{(g)} + 1e \longrightarrow F_{(g)}^{-1}$$

$$Na_{(g)} + F_{(g)} \longrightarrow Na_{(g)}^{+1} F_{(g)}^{-1}$$

How many electrons gained or lost?

The inert gases all have the electron configuration **ns^2np^6**.

Is This a Favorable Reaction?

$$Na_{(g)} \longrightarrow Na_{(g)}^{+1} + 1e$$

$$F_{(g)} + 1e \longrightarrow F_{(g)}^{-1}$$

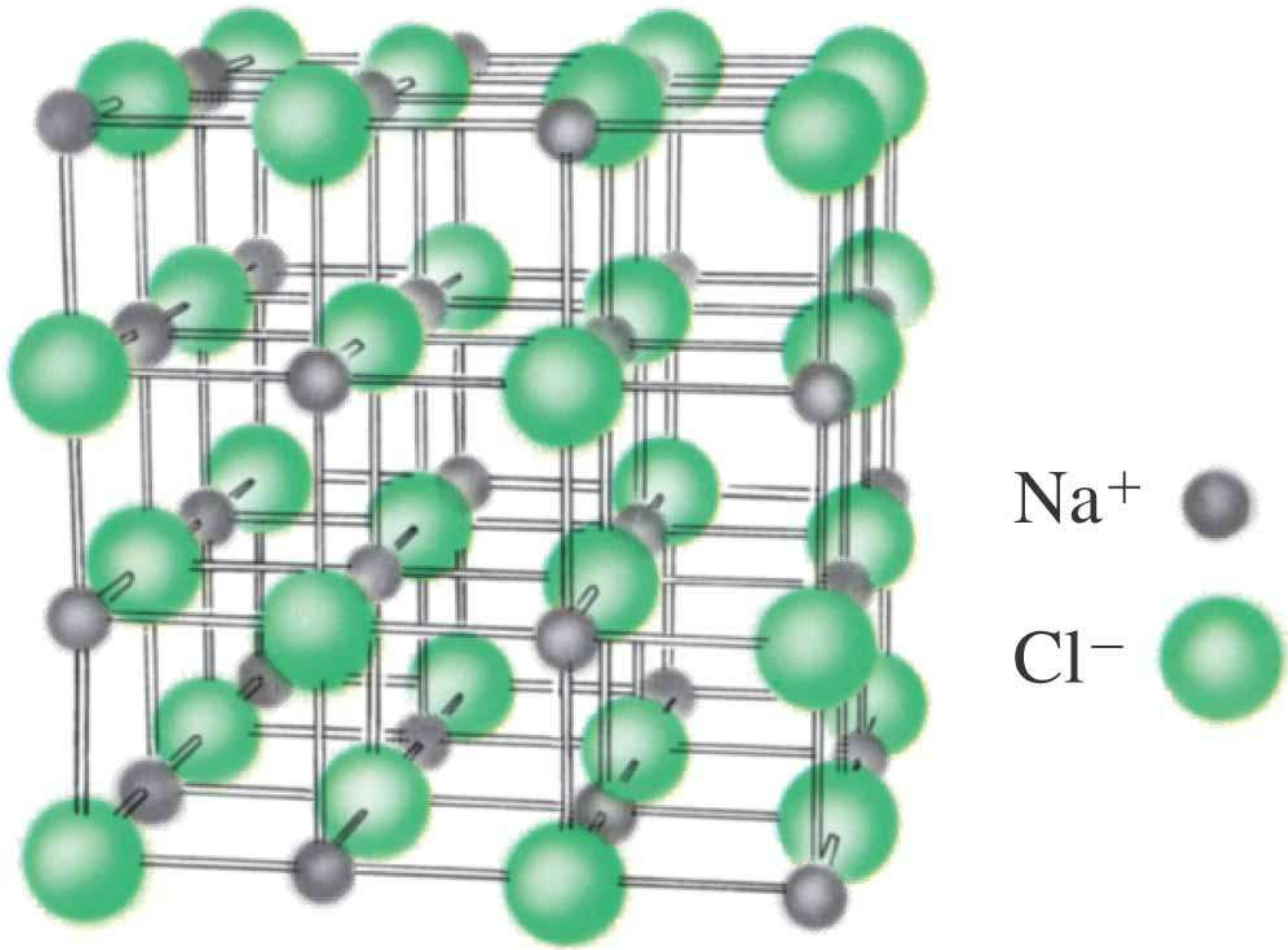
$$Na_{(g)} + F_{(g)} \longrightarrow Na_{(g)}^{+1} F_{(g)}^{-1}$$

IE = +496 kJ/mol

EA = -328 kJ/mol

The Missing Link

ionic compounds form a crystal lattice



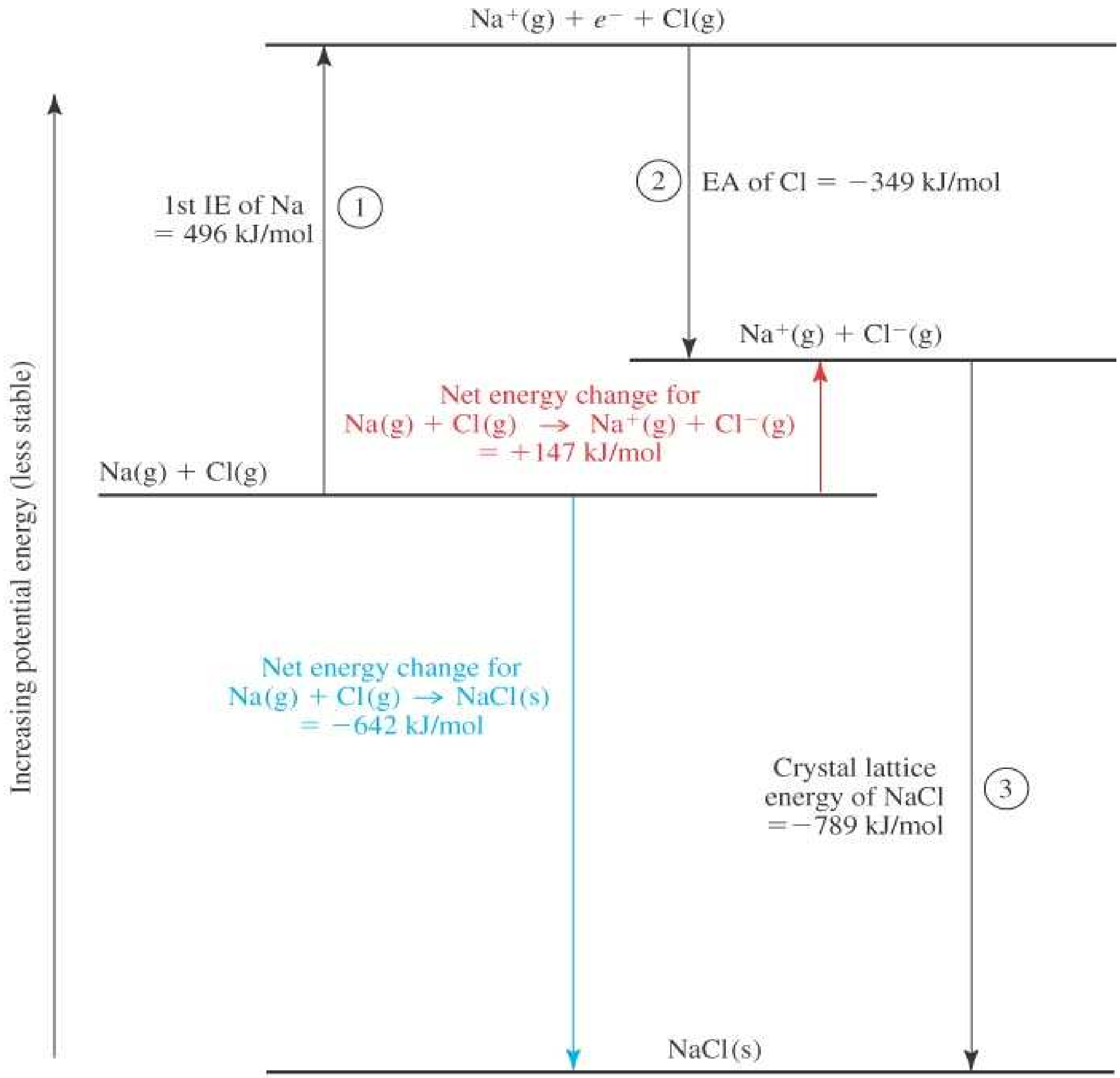
Lattice Energy

$$Na_{(g)} + F_{(g)} \longrightarrow Na_{(g)}^{+1} F_{(g)}^{-1} \quad \Delta H_r = +168 \text{ kJ/mol}$$

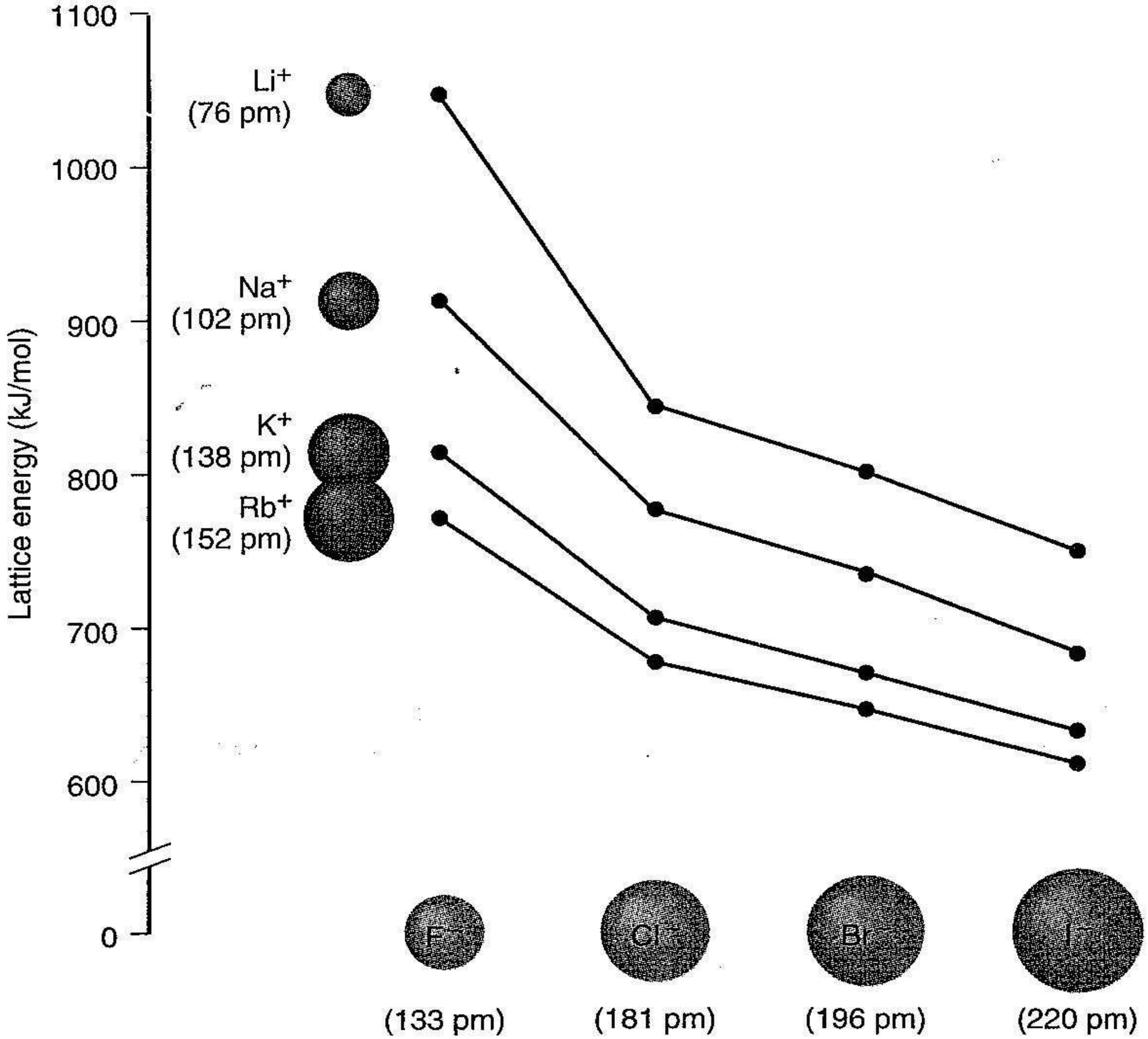
$$Na_{(g)}^{+1} F_{(g)}^{-1} \longrightarrow Na^{+}F^{-}_{(s)} \quad LE = -910 \text{ kJ/mol}$$

$$Na_{(g)} + F_{(g)} \longrightarrow Na^{+}F^{-}_{(s)}.$$

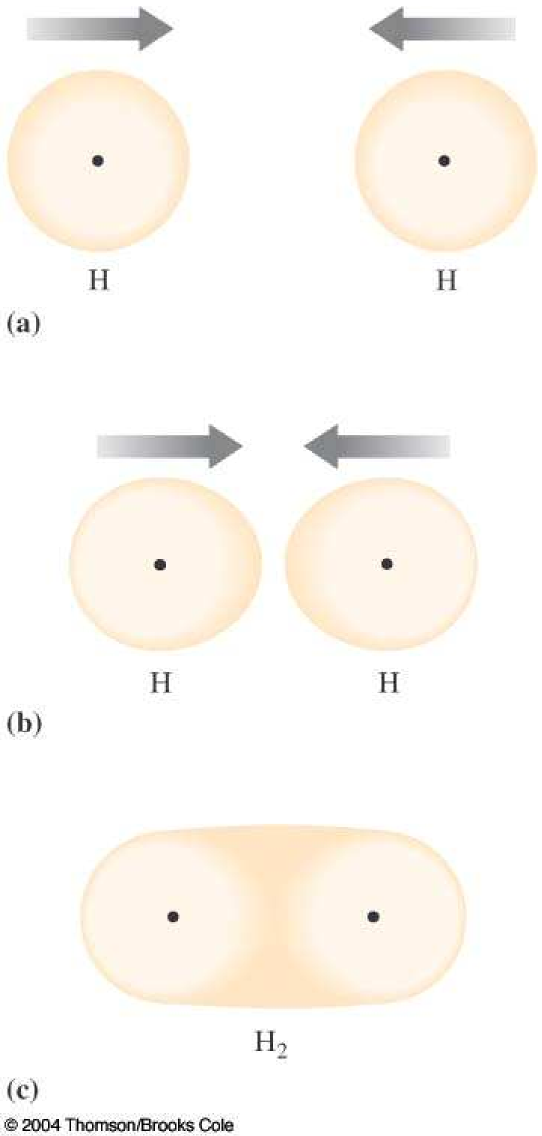
A Graphical Representation

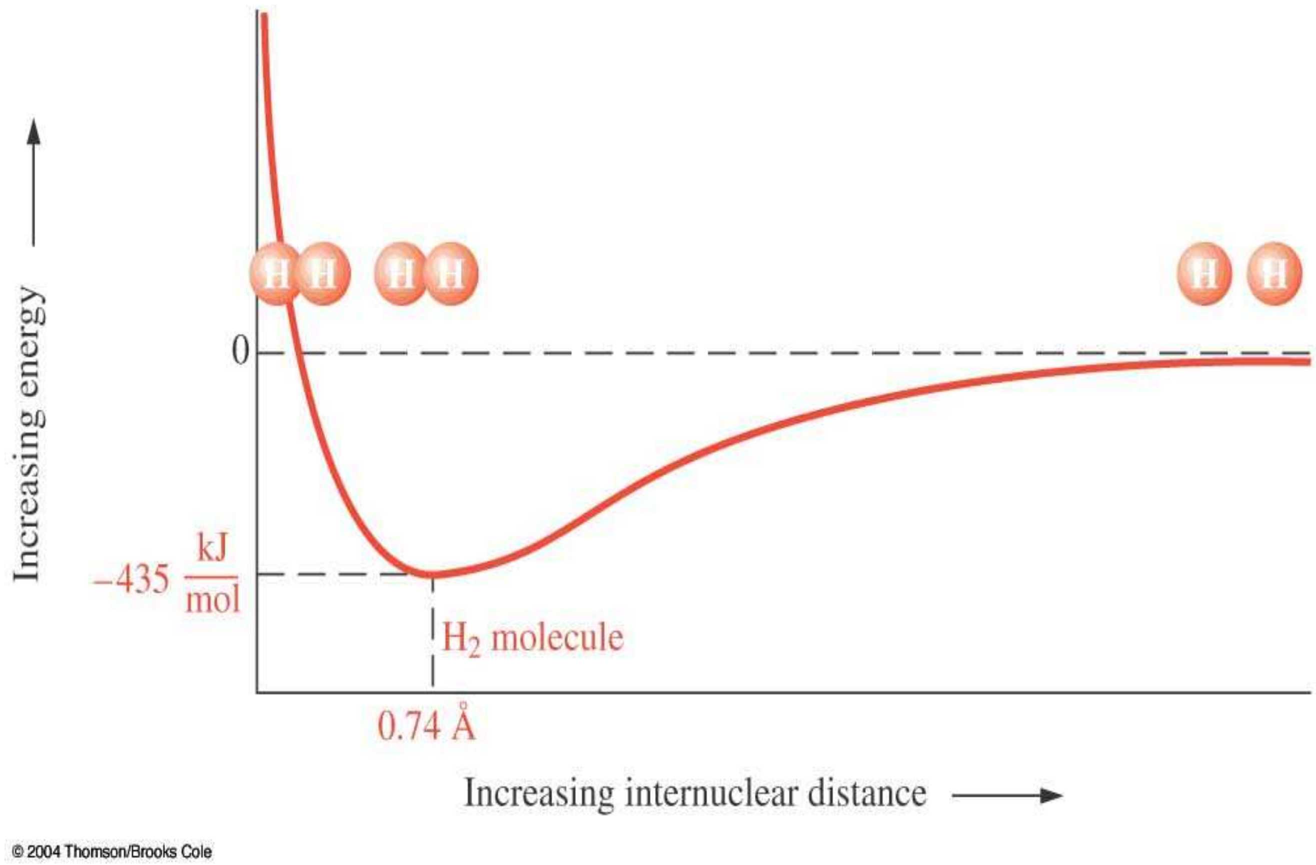


A Table of Lattice Energies



Covalent Bonding





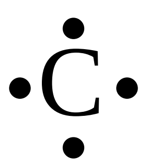
Bonding and Lewis Structures

CF_4

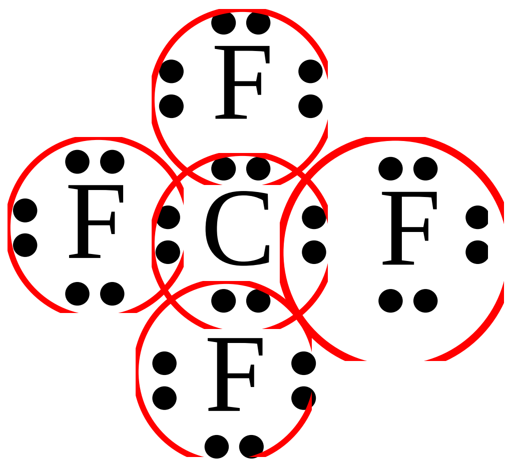
Follow these steps

1. Write the Lewis structures for all atoms

2. Find the "central atom"- the one with the **most unpaired electrons**. Connect other atoms with unpaired electrons



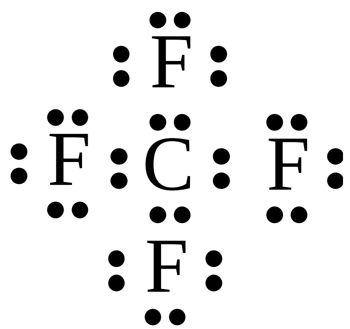
Check for Octets



Note: **All atoms** have **octets** and <u>all electrons</u> are <u>paired</u>!

Finishing Touches

Use lines for covalent bonds



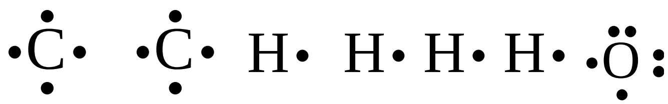
H cannot have an octet!

1st period means 2 electrons fill first energy level

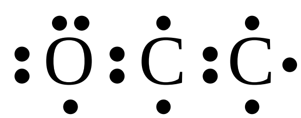
Q1
Draw a Lewis Structure for H_2O

More Complex Structures

C_2H_4O

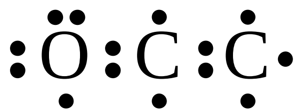
Step 1 

Step 2 Connect the atom with the next most unpaired electrons to the central atom. Continue until only H's are left



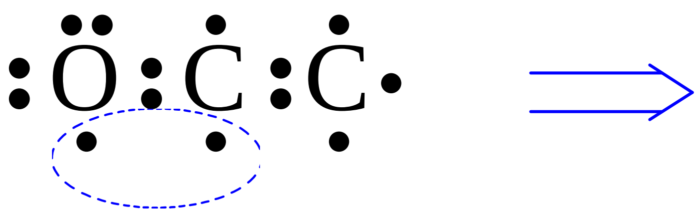
C_2H_4O

Continue



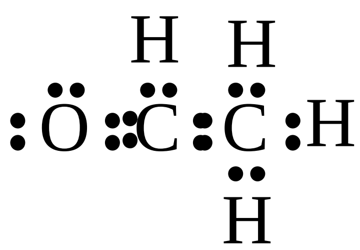
Step 3 Count unpaired electrons and compare to H's left

Step 4 Move unpaired electrons on adjacent atoms between them to form a second covalent bond (or as many as needed to have unpaired electrons equal to H's left)



Last Step

Step 5 Connect H's, draw lines for bonds and include lone pairs



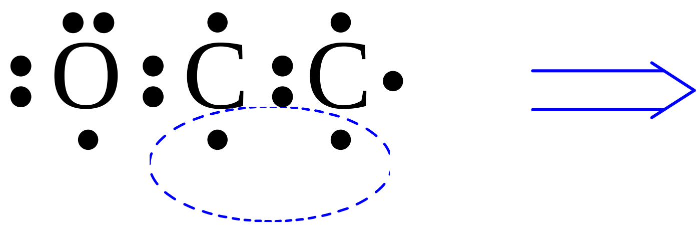
Q2
Provide Lewis structures for the following

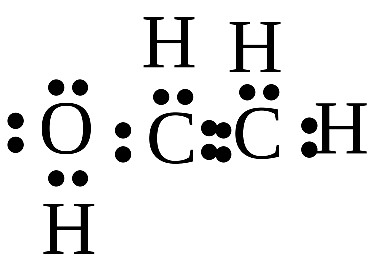
CO_2

C_3H_6

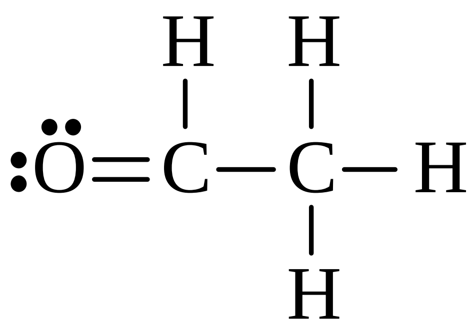
CH_2O_2

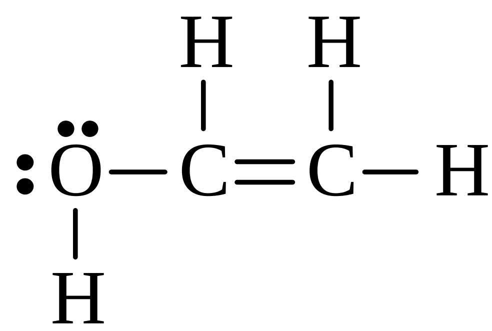
Isomers- a different way of connecting the same atoms



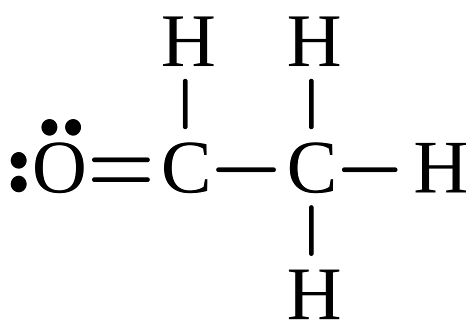


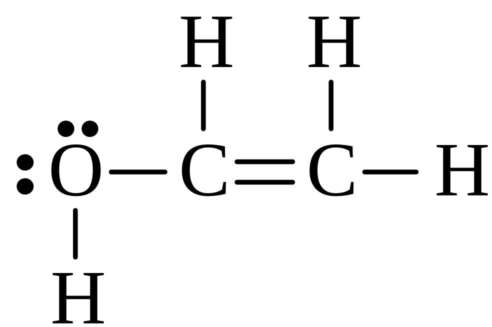
Compare Structures





Differences?





structural isomers- *different* compounds with *different* properties and *different* names having the same formula

Q3
Provide two Lewis structures (isomers) for:

C_4H_8

Q3 cont.
Do the same for:

C_2H_5N

SO_2
a mathematical method

1. give all atoms 8 e(except H)
   3x8=24 (3 atoms, each with octet)

2. count actual valence electrons
   6+6+6=18 (valence electrons for S, O and O)

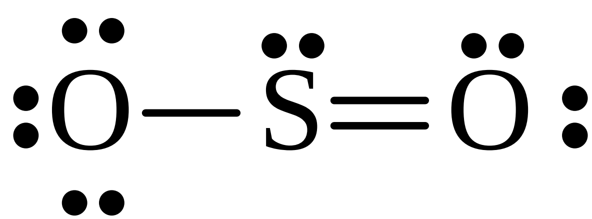
3. subtract for shared electrons
   24-18= 6 electrons shared

Add Electrons to give Octets

Determine Formal Charge

count each bond as "1" and each dot as "1"

compare to number of valence electrons



Resonance Theory

SO_2

3x8=24 (3 atoms, each with octet)

6+6+6=18 (valence electrons for S, O and O)

24-18= 6 electrons shared

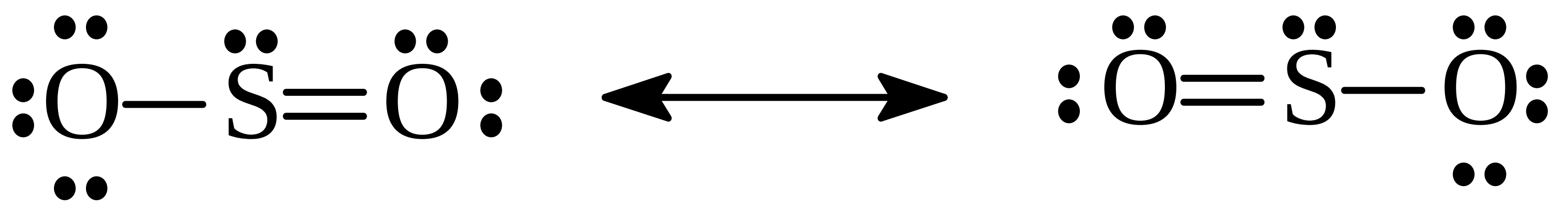
O–S=O

6 electrons shared

O=S–O

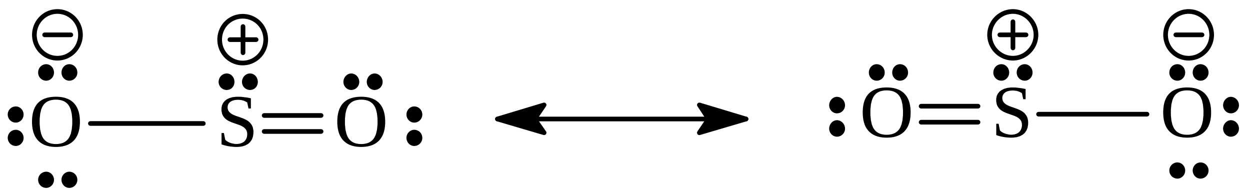
also 6 electrons shared

Finishing the Problem

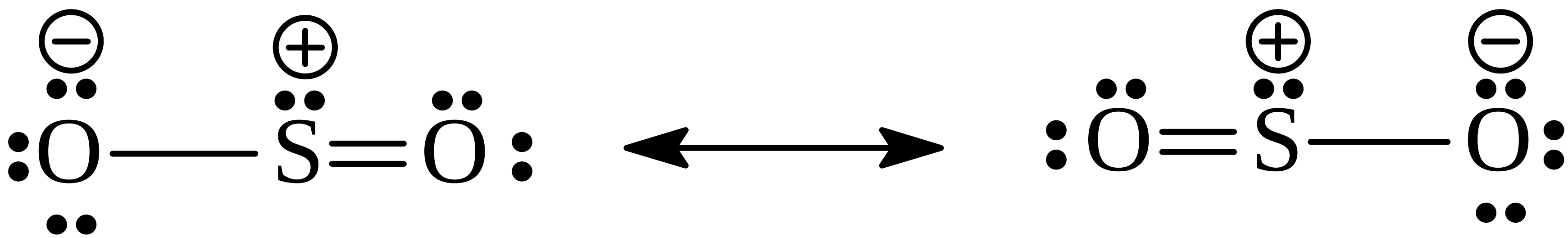


What Does SO_2 Look Like?

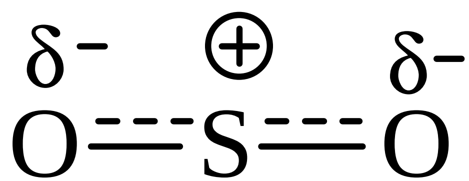
single bonds are longer than double bonds



How do the structures contribute to the hybrid?



The Hybrid Structure of SO_2



Provide all resonance structures for CO_3^{-2}

Q5 & Q6-
Do the same for OPN

Formal charge considerations:

1. structures with fewest charges, best
2. more electronegative elements should have the - charge
3. adjacent atoms should not have same sign charge
4. charges >1 are usually bad

Exceptions to the Octet Rule

boron compounds

BH_3 for example

some nitrogen compounds w/ odd # N's

NO

expanded octets (more than 8e)

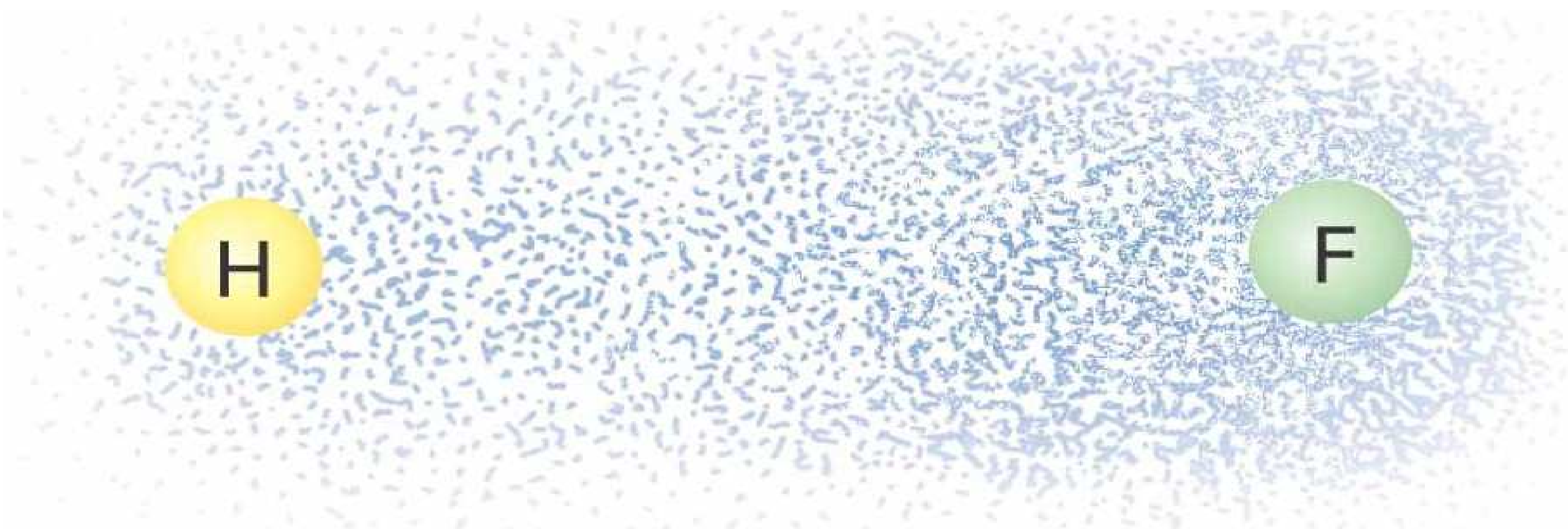
SF_6 PCl_5

will do later

Polar Covalent Bonds

the unequal sharing of electrons within a covalent bond

Polar Covalent Bond



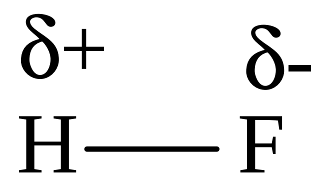
Using the Scale

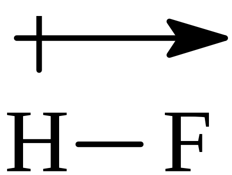
Which bond is more polar?

O—C　　　O—H

O—C	O—H
O = 3.44	O = 3.44
C = 2.55	H = 2.10
0.89	**1.40**

Indicating the “dipole”





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