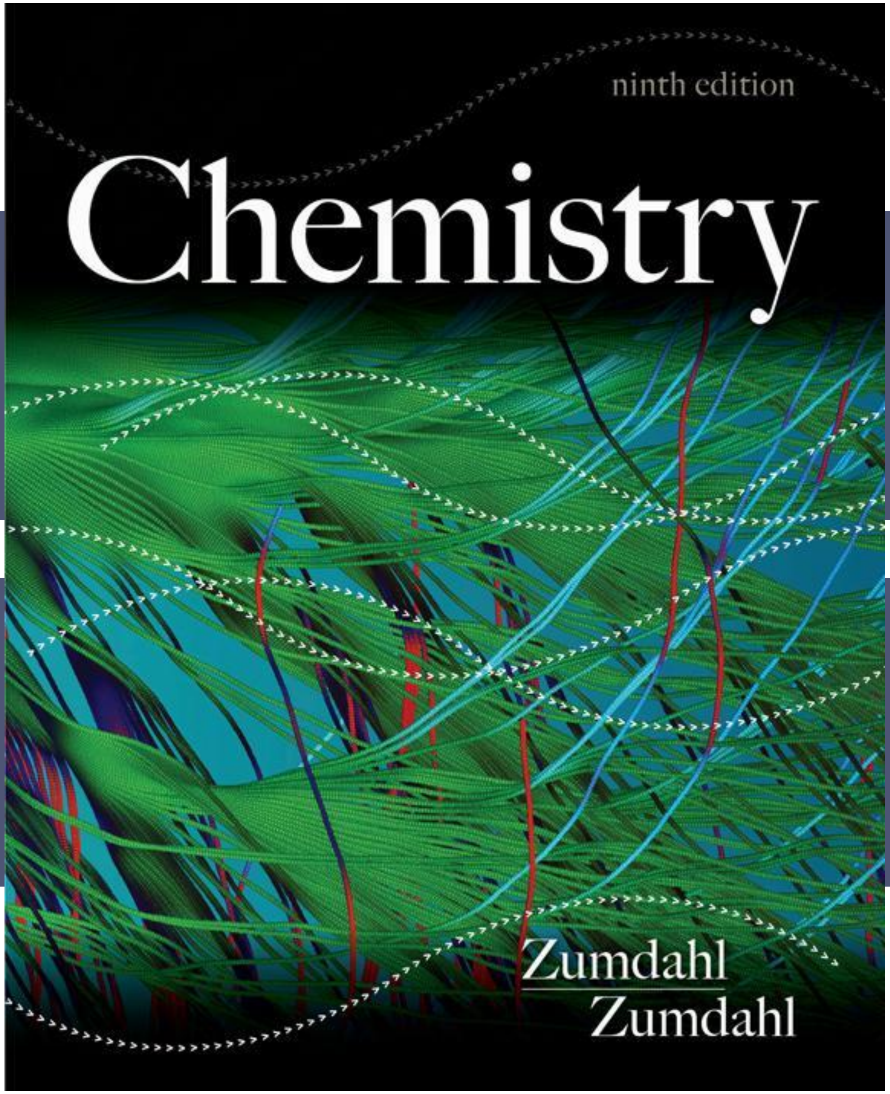


Chapter 8

Bonding:
General Concepts

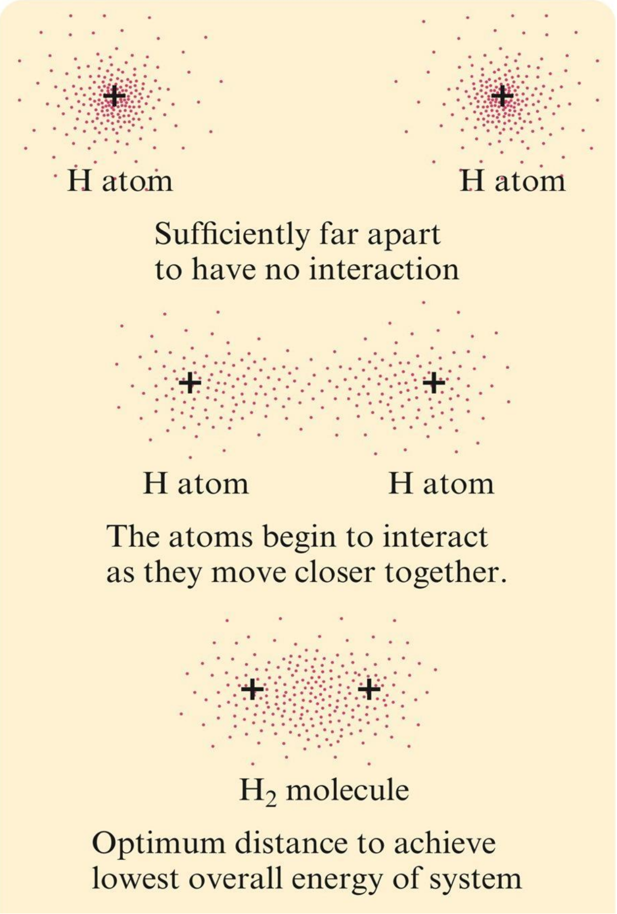
Coulomb's Law

$$E = (2.31 \times 10^{-19}\ \text{J} \cdot \text{nm})\left(\frac{Q_1 Q_2}{r}\right)$$

- E - Units of joules
- r - Distance between ion centers in nanometers
- Q_1 and Q_2 - Numerical ion charges

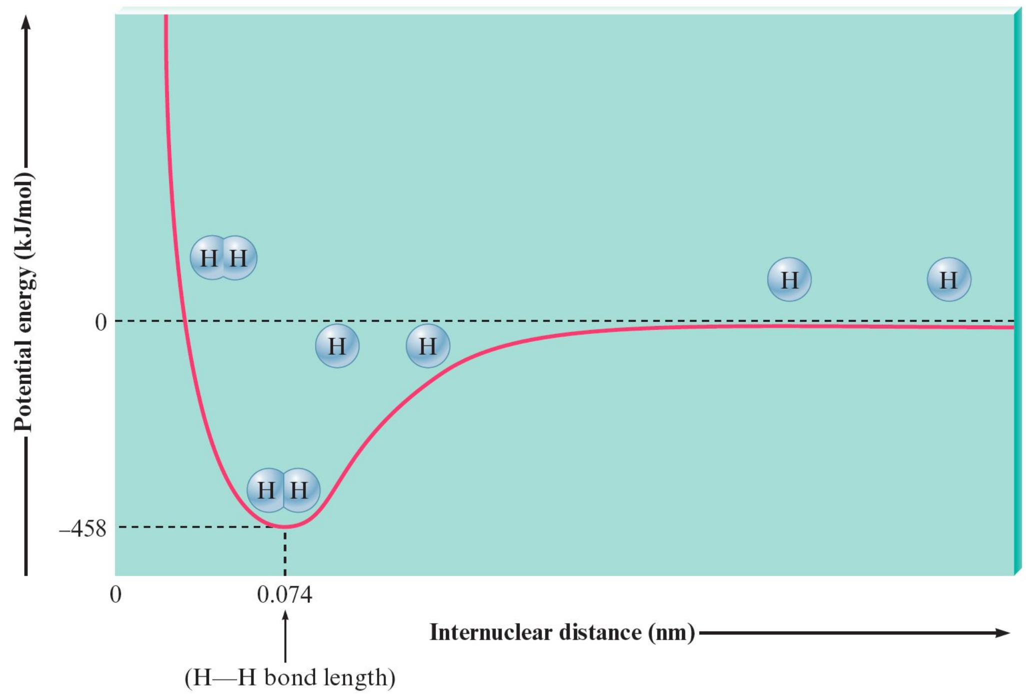
Figure 8.1 (a) - The Interaction of Two Hydrogen Atoms

- Bonds will form if E of is lower than that of atoms



Bond Length

- Distance between two atoms when potential energy is minimal

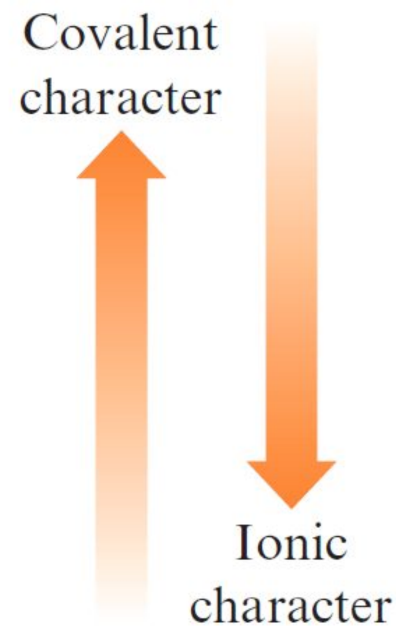


Covalent Bonding

- Equal sharing of electrons between two identical atoms
  - Caused by the mutual attraction of nuclei for shared electrons
- **Polar covalent bond**: Bond in which the electrons are not shared equally because one atom attracts them more strongly than the other
  - Example - Bonding in hydrogen fluoride

Table 8.1 - Relationship between Electronegativity and Bond Type

Electronegativity Difference in the Bonding Atoms	Bond Type
Zero	Covalent
Intermediate	Polar covalent
Large	Ionic



Answer the following in your notes, compare with a partner

- Order the following bonds according to polarity:
  - H—H
  - O—H
  - Cl—H
  - S—H
  - F—H

Figure 8.5 - H_2O Molecule

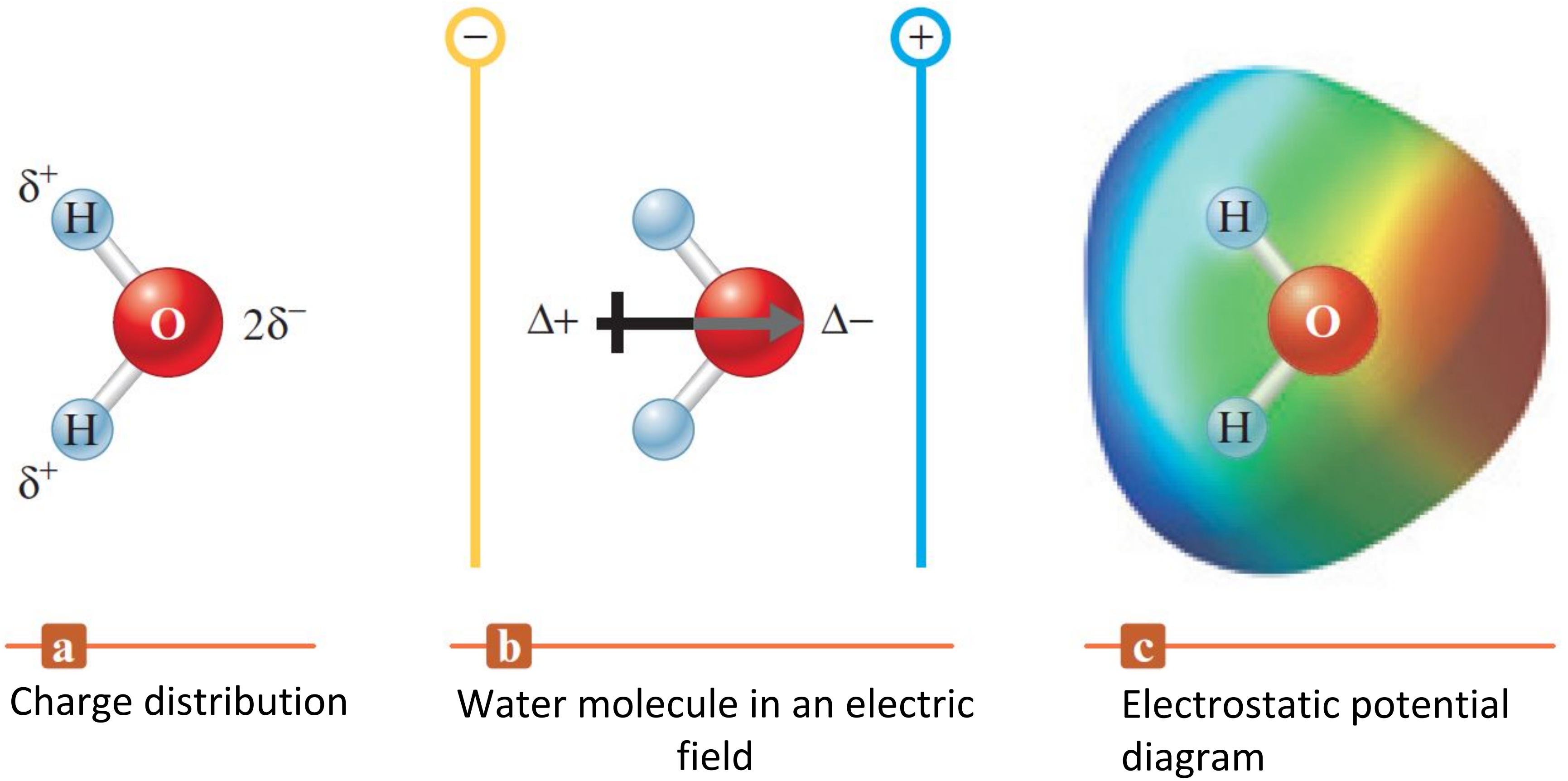
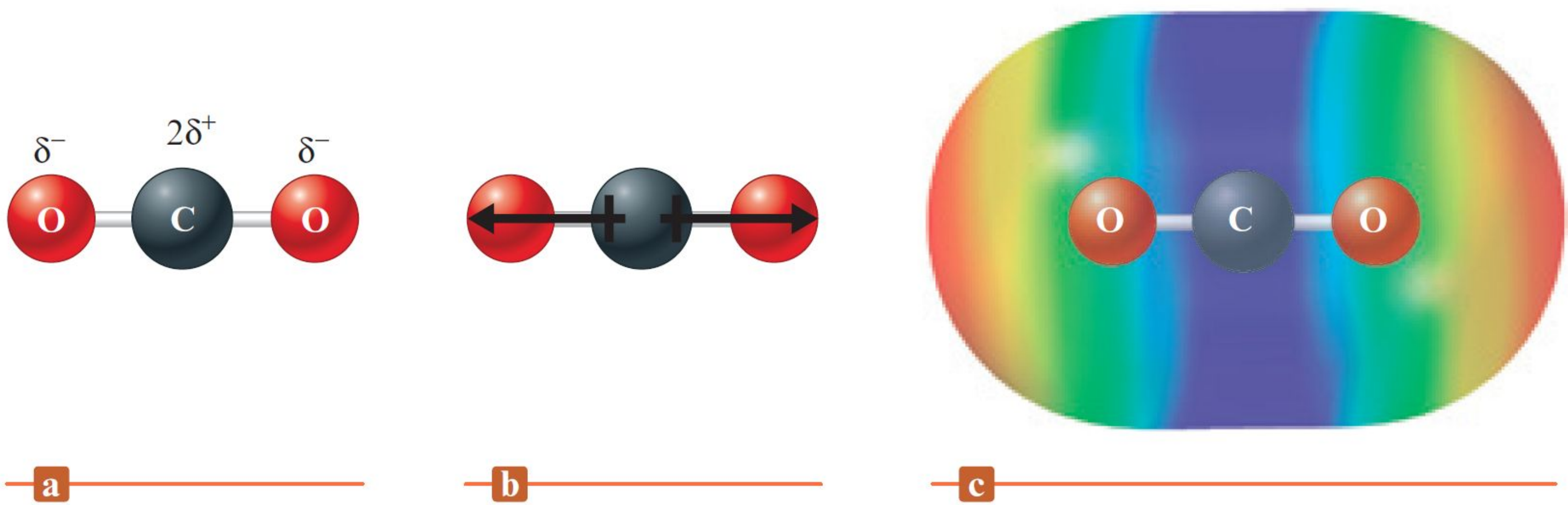


Figure 8.7 - CO_2 Molecule

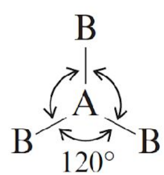
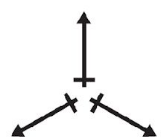
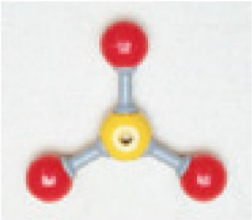
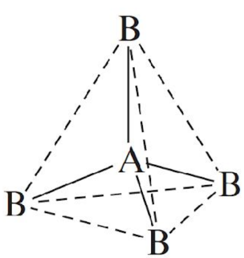
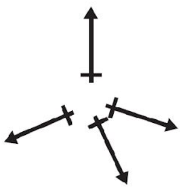


a — Charge distribution

b — The molecule has no dipole moment as the opposed polarities cancel out

c — Electrostatic potential diagram

Table 8.2 - Molecules with Polar Bonds but No Resulting Dipole Moment

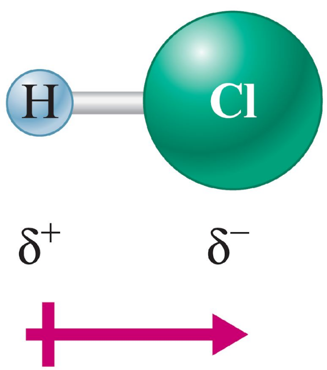
Type	General Example	Cancellation of Polar Bonds	Specific Example	Ball-and-Stick Model
Linear molecules with two identical bonds	B—A—B		CO_2	
Planar molecules with three identical bonds 120 degrees apart			SO_3	
Tetrahedral molecules with four identical bonds 109.5 degrees apart			CCl_4	

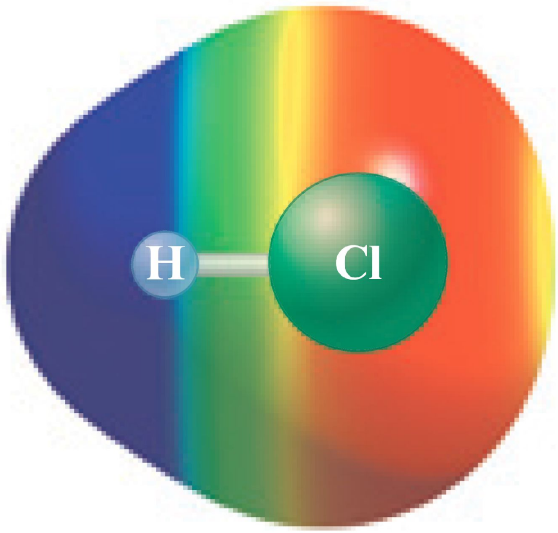
Answer the following in your notes, compare with a partner

- For each of the following molecules, show the direction of the bond polarities and indicate which ones have a dipole moment
  - HCl
  - Cl_2
  - SO_3

Example 8.2 - Solution (Continued 1)

- The HCl molecule has a dipole moment



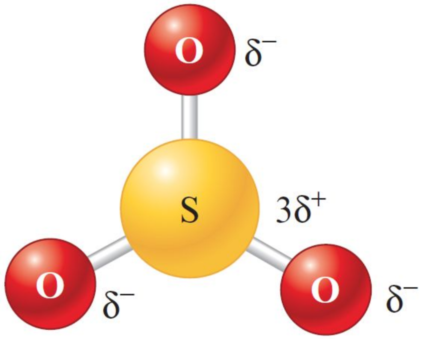


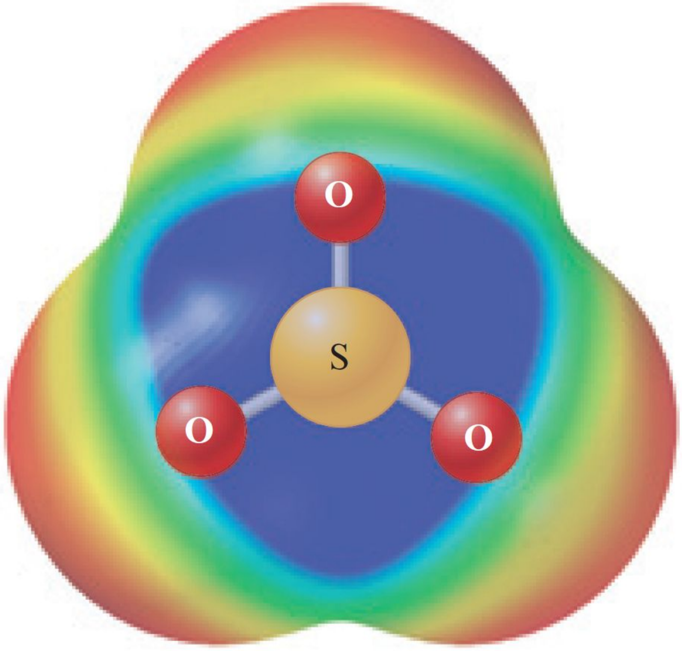
Example 8.2 - Solution (Continued 2)

- Cl_2 molecule
  - The two chlorine atoms share the electrons equally
  - No bond polarity occurs
  - The Cl_2 molecule has no dipole moment

Example 8.2 - Solution (Continued 4)

- The molecule has no dipole moment
  - Symmetrically arranged bonds cancel





Answer the following in your notes, compare with a partner

- Arrange the following ions in order of decreasing size
  - Se^{2-}, Br^{-}, Rb^{+}, Sr^{2+}

Lattice Energy Calculations

- Represented by a modified form of Coulomb's law

$$\text{Lattice energy} = k\left(\frac{Q_1 Q_2}{\text{r}}\right)$$

  - *k* - Proportionality constant
    - Depends on the structure of the solid and the electronic configurations of the ions
  - Q_1 and Q_2 - Charges on the ions
  - *r* - Shortest distance between the centers of the anions and the cations

Bond Energy

- E added to the system to break bonds
  - Endothermic
- E released when bonds are formed
  - Exothermic

Answer in your notes, compare with a partner

- Use the bond energies listed in Table 8.4, and calculate ΔH for the reaction of methane with chlorine and fluorine to give Freon-12 (CF_2Cl_2)

$$CH_4(g) + 2Cl_2(g) + 2F_2(g) \longrightarrow CF_2Cl_2(g) + 2HF(g) + 2HCl(g)$$

Interactive Example 8.5 - Solution

- Combine energy changes to calculate ΔH

  $$\Delta H = \text{E to break bonds} - \text{E released when bonds form}$$

  - The minus sign gives the correct sign to the energy terms for the exothermic processes

Interactive Example 8.5 - Solution (Continued 1)

- Reactant bonds broken

$$CH_4\text{: 4 mol C—H} \quad 4\ \cancel{\text{mol}} \times \frac{413\text{ kJ}}{\cancel{\text{mol}}} = 1652\text{ kJ}$$

$$2Cl_2\text{: 2 mol Cl—Cl} \quad 2\ \cancel{\text{mol}} \times \frac{239\text{ kJ}}{\cancel{\text{mol}}} = 478\text{ kJ}$$

$$2F_2\text{: 2 mol F—F} \quad 2\ \cancel{\text{mol}} \times \frac{154\text{ kJ}}{\cancel{\text{mol}}} = \underline{308\text{ kJ}}$$

$$\text{Total energy requied} = 2438\text{ kJ}$$

Interactive Example 8.5 - Solution (Continued 2)

- Product bonds formed

$$CF_2Cl_2\text{: 2 mol C—F} \quad 2\ \cancel{\text{mol}} \times \frac{485\ \text{kJ}}{\cancel{\text{mol}}} = 970\ \text{kJ}$$

and

$$\text{2 mol C—Cl} \quad 2\ \cancel{\text{mol}} \times \frac{339\ \text{kJ}}{\cancel{\text{mol}}} = 678\ \text{kJ}$$

$$\text{2HF: 2 mol H—F} \quad 2\ \cancel{\text{mol}} \times \frac{565\ \text{kJ}}{\cancel{\text{mol}}} = 1130\ \text{kJ}$$

$$\text{2HCl: 2 mol H—Cl} \quad 2\ \cancel{\text{mol}} \times \frac{427\ \text{kJ}}{\cancel{\text{mol}}} = \underline{854\ \text{kJ}}$$

$$\text{Total energy released} = 3632\ \text{kJ}$$

Interactive Example 8.5 - Solution (Continued 3)

- Calculating ΔH

$$\Delta H = \text{energy required to break bonds} - \text{energy released when bonds form}$$
$$= 2438 \text{ kJ} - 3632 \text{ kJ}$$
$$= -1194 \text{ kJ}$$

Problem Solving Strategy - Steps for Writing Lewis Structures

1. Sum the valence e- from all atoms
2. Use a pair of e- to form a bond between each pair of atoms
3. Arrange the remaining e- to satisfy the duet rule for hydrogen and the octet rule for other elements

Drawing the Lewis Structure of Water

- Sum the valence e- for H_2O
  - 1 + 1 + 6 = 8 valence e-

$$\begin{array}{ccc} \uparrow & \uparrow & \uparrow \\ \mathrm{H} & \mathrm{H} & \mathrm{O} \end{array}$$

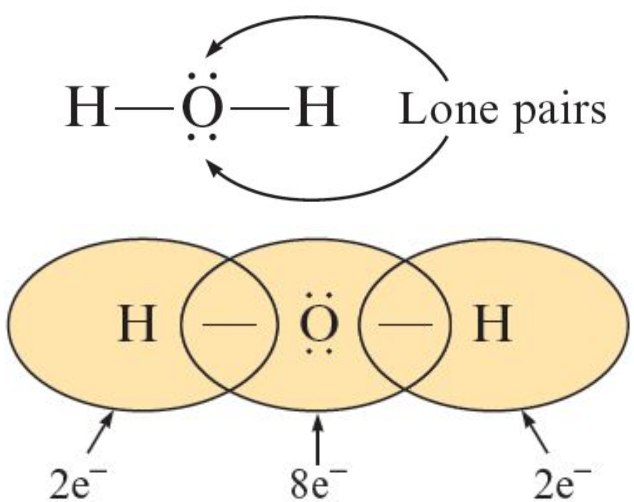
- Draw the O—H single bonds

H—O—H

  - A line is used to indicate each pair of bonding e-

Drawing the Lewis Structure of Water (Continued)

- Distribute the remaining e- to achieve a noble gas configuration for each atom
  - Dots represent lone e- pairs

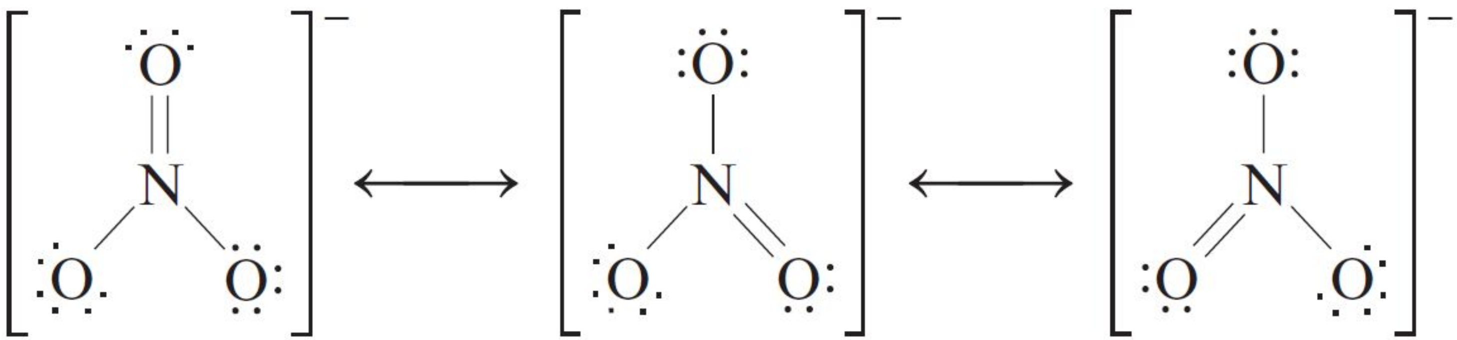


Answer in your notes, compare with a partner

- Give the Lewis structure for each of the following
  - HF
  - N_2
  - NH_3
  - NO^+

Resonance

- Example - Nitrate ion
  - Has three valid Lewis structures



  - The most accurate structure is obtained when the three structures are superimposed

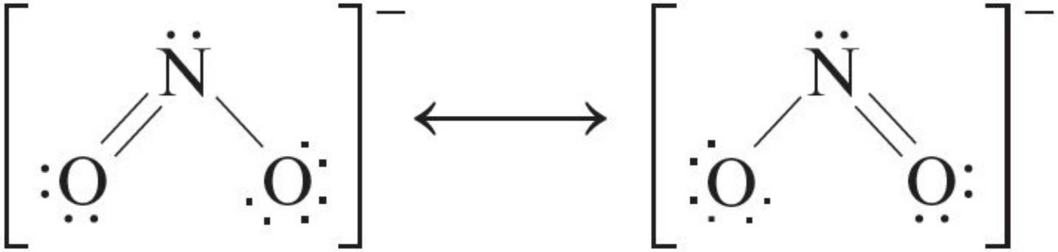
Section 8.12
Resonance

Answer in your notes, compare with a partner

- Describe the electron arrangement in the nitrite anion (NO_2^-) using the localized electron model

Example 8.9 - Solution

- NO_2^- possesses 18 valence electrons
  - 5 + 2(6) + 1 = 18
- Indicating the single bonds gives the structure O—N—O
- The remaining 14 electrons (18 – 4) can be used to produce these structures



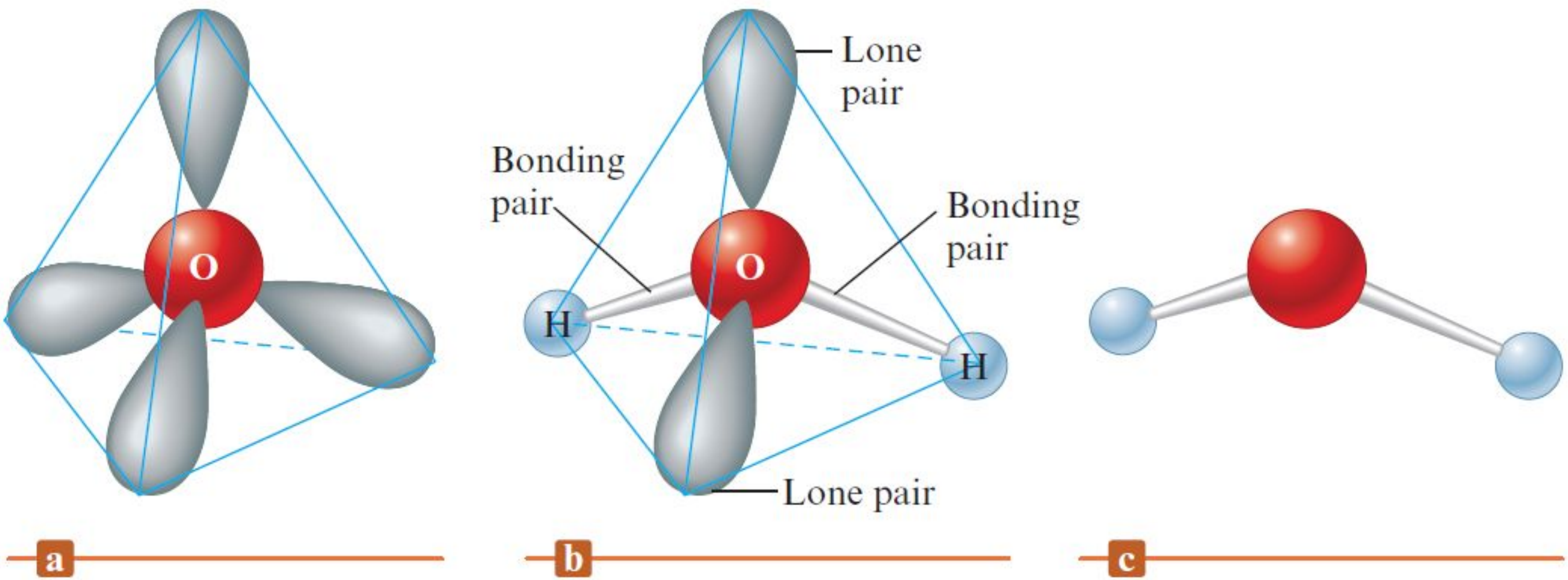
Rules Governing Formal Charge

- To calculate the formal charge on an atom:
  - Take the sum of the lone pair e- and one-half the shared e-
  - Subtract the number of assigned e-from the number of valence e- on the free, neutral atom to obtain the formal charge

Answer in your notes, compare with a partner

- Describe the molecular structure of the water molecule

Example 8.11 - Solution (Continued)

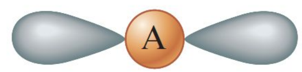
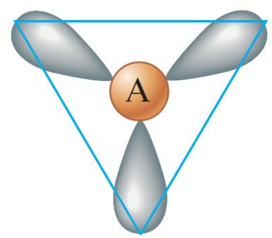
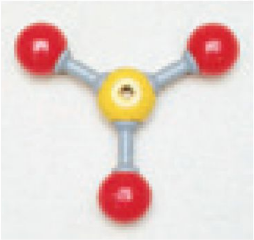
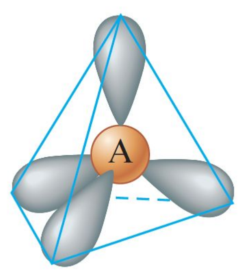


a The tetrahedral arrangement of the electron pairs around oxygen in the water molecule

b Two of the electron pairs are shared between oxygen and the hydrogen atoms and two are lone pairs

c The V-shaped molecular structure of the water molecule

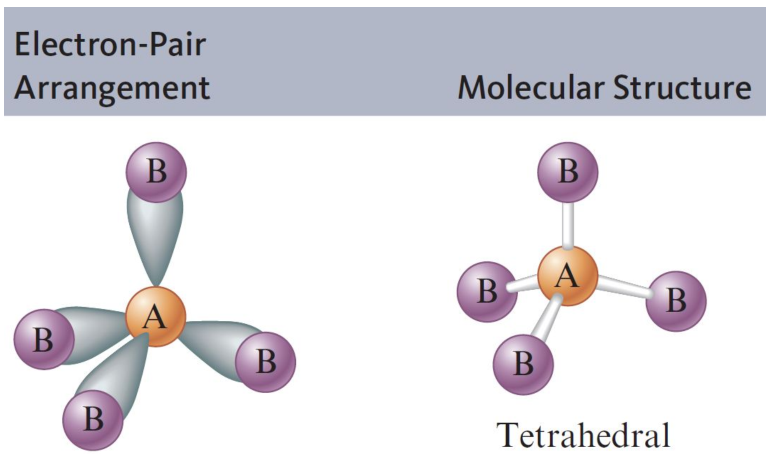
Table 8.6 - Arrangements of Electron Pairs around an Atom Yielding Minimum Repulsion

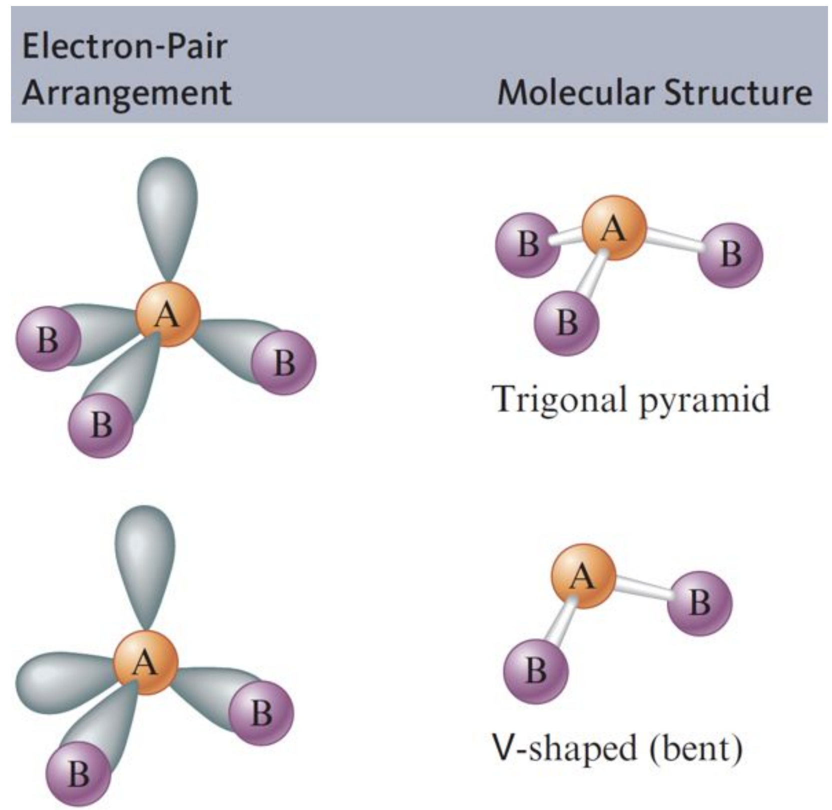
Number of Electron Pairs	Arrangement of Electron Pairs		Example
2	Linear		
3	Trigonal planar		
4	Tetrahedral		

Discuss with partner, find another group and decide on an answer

- You and a friend are studying for a chemistry exam
  - What if your friend tells you that all molecules with polar bonds are polar molecules?
    - How would you explain to your friend that this is not correct?
    - Provide two examples to support your answer

Table 8.7 - Structures of Molecules with Four Electron Pairs around the Central Atom



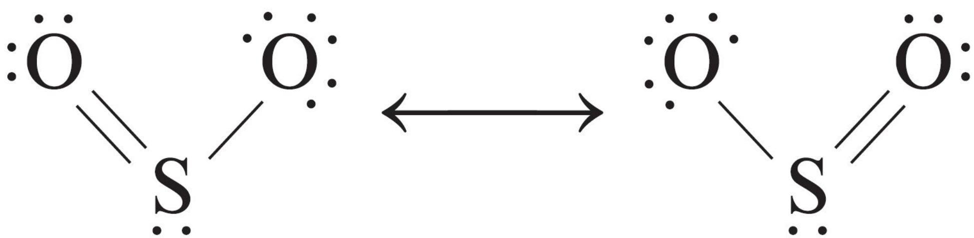


Interactive Example 8.14 - Structures of Molecules with Multiple Bonds

- Predict the molecular structure of the sulfur dioxide molecule
  - Is this molecule expected to have a dipole moment?

Interactive Example 8.14 - Solution

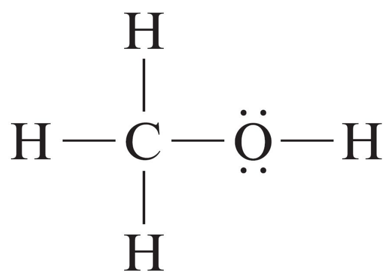
- First, we must determine the Lewis structure for the SO_2 molecule, which has 18 valence electrons
  - Expected resonance structures



- To determine the molecular structure, we must count the electron pairs around the sulfur atom

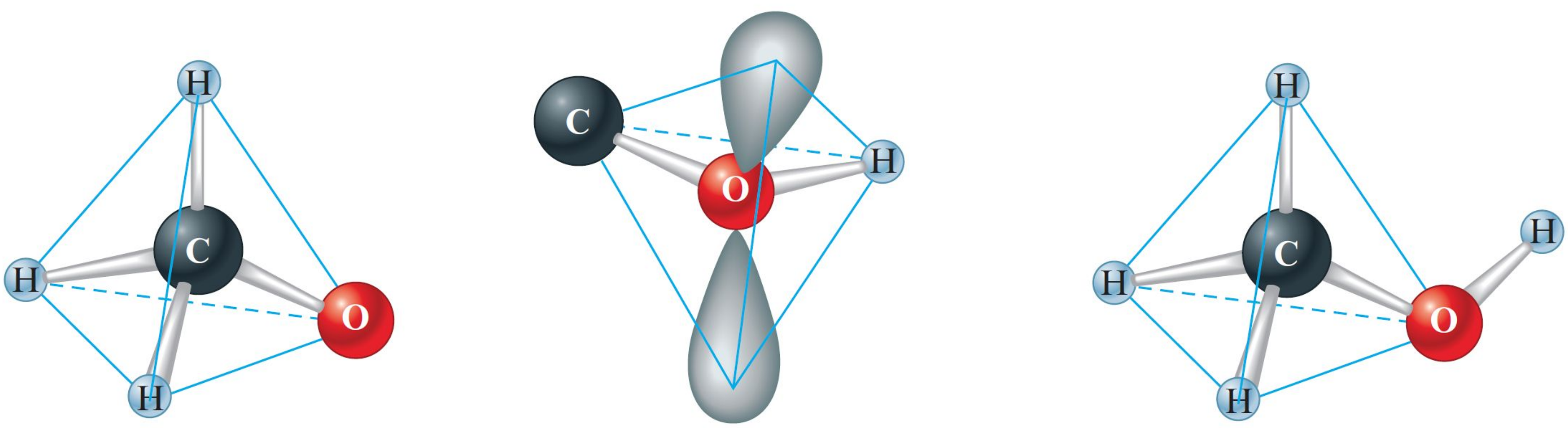
Molecules Containing No Single Atom

- Consider a methanol (CH_3OH) molecule



- The molecular structure can be predicted from the arrangement of pairs around the carbon and oxygen atoms

Figure 8.22 - The Molecular Structure of Methanol



a
The arrangement of electron pairs and atoms around the carbon atom

b
The arrangement of bonding and lone pairs around the oxygen atom

c
The molecular structure