

## Chapter 7

### Chemical Bonding

## Chapter 7 Outline

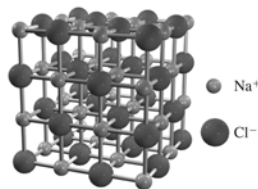
- I. Bond Formation
- II. Lewis Structures
- III. Formal Charge
- IV. Resonance
- V. Polarity

# I. Bond Formation

- Helpful tool: *Lewis Formulas*
- Goal of most chemical bonds: *Give an atom a noble gas electron configuration*

Two ways of doing this:

1. **Transfer** electrons from one atom to another (IONIC BONDING)



2. **Share** electrons between atoms (COVALENT BONDING)

	Single Bond	Double Bond	Triple Bond
Bond length (Å)	1.54 Å	1.34 Å	1.21 Å
Bond energy (kJ/mol)	346	602	835

Shorter Bonds

Stronger Bonds

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### Covalent Bonding: A Closer Look

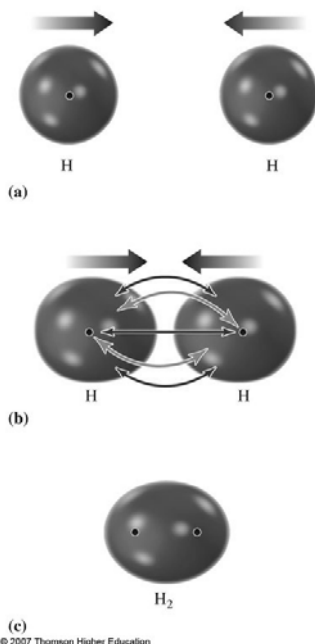


Fig. 7-3, p. 259

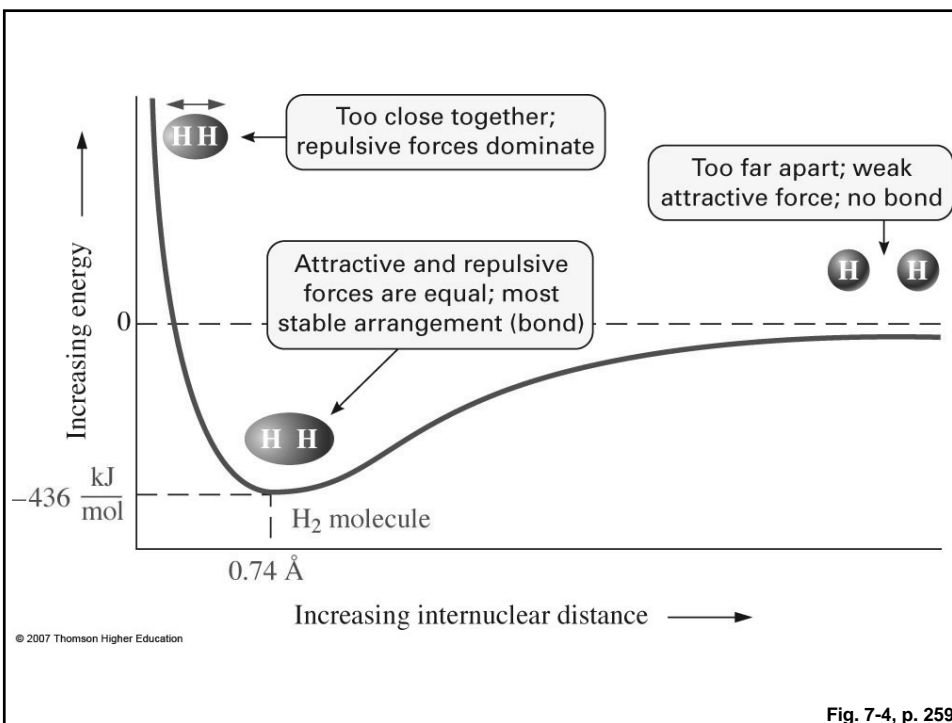


Fig. 7-4, p. 259

## II. Lewis Structures

- *Lewis Structures*

Lewis structures show where the valence electrons are located in a molecule with covalent bonds

- *Octet Rule*

Atoms in molecules and polyatomic ions tend to be surrounded by eight electrons

- *The best Lewis structures are formed simply by combining unpaired electrons from Lewis symbols (not always possible, however)*

### General Steps in Writing Lewis Structures

**1) Draw a skeletal structure of the molecule.**

Which atom is the central atom?

- central atom is often the element farthest to the left and/or lowest in the periodic table.
- central atom is often the element listed first in the formula
- central atom is often the element in the compound for which there is only one atom.
- hydrogen is *never* the central atom.

**2) Count the total number of valence electrons in all the atoms. Make sure to add/subtract electrons if the molecule is a polyatomic ion.**

**3) Draw single bonds between the central atom and the surrounding atoms.**

**4) Fill the octets of the atoms bonded to the central atom.**

- Remember that hydrogen only needs 2 electrons in its outer shell, not 8!

**5) Put leftover electrons on the central atom as lone pairs.\***

**6) If central atom doesn't have an octet, make double or triple bonds.\*\***

\* Atoms in and beyond the third period may have an expanded octet (i.e., more than 8 electrons in their outer shell)!

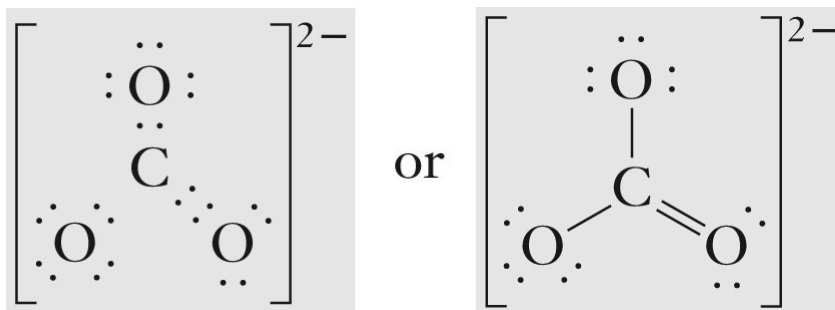
\*\* Remember that atoms in Groups IIA and IIIA don't need an octet.

### III. Formal Charge

#### *Formal Charge -*

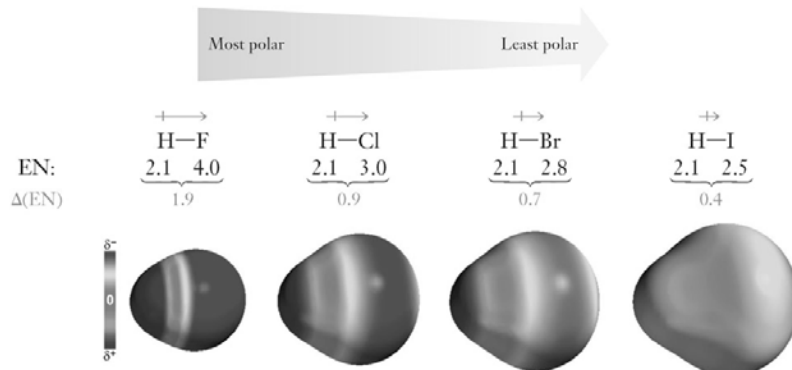
- Compares number of valence electrons for an atom in a Lewis structure with the free atom (group #)
- Sum of formal charges = charge on molecule
- *Trick:* Give all lone pairs to the atom. Give half of the bond pair electrons to each atom in the bond.
- Always try to minimize the non-zero formal charge within a molecule
- When possible, always place the negative formal charge on the more electronegative element

### IV. Resonance



# V. Polarity

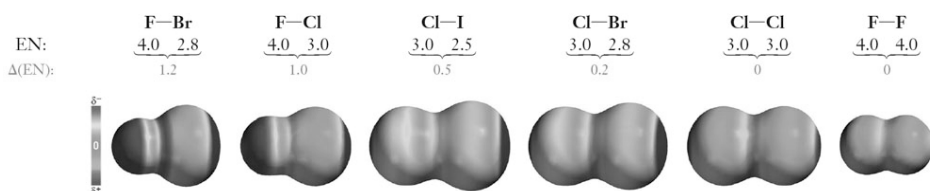
## Compare $\text{Cl}_2$ and $\text{IBr}$



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*When is a bond truly covalent?*

When the electronegativities of the atoms are the same.



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