Oxidation Numbers: Rules

1) The oxidation number of the atoms in any free, uncombined element, is zero
2) The sum of the oxidation numbers of all atoms in a compound is zero
3) The sum of the oxidation numbers of all atoms in an ion is equal to the charge of the ion
4) The oxidation number of fluorine in all its compounds is –1
5) The oxidation number of other halogens in their compounds is usually –1

Oxidation Numbers: Rules

6) The oxidation number of hydrogen is +1 when it is combined with more electronegative elements (most nonmetals) and –1 when it is combined with more electropositive elements (metals)

7) The oxidation number of Group 1A elements is always +1 and the oxidation number of Group 2A elements is always +2

8) The oxidation number of oxygen in most compounds is –2

9) Oxidation numbers for other elements are usually determined by the number of electrons they need to gain or lose to attain the electron configuration of a noble gas

Ionic Bonding

$$Na - e^- \rightarrow Na^+$$

$$Cl + e^- \rightarrow Cl^-$$

$$Na + Cl \rightarrow Na^+ + Cl^-$$

- Na^+ cations and Cl^- anions are electrostatically attracted to each other resulting in an extended ionic lattice
- We say that Na^+ and Cl^- ions are held together by ionic bonding

F_2 Molecule

- This bond is called a nonpolar covalent bond
- It is characterized by the symmetrical charge distribution

HF Molecule

- F is more electronegative than H
- In this molecule the electron pair will be shifted towards the F atom

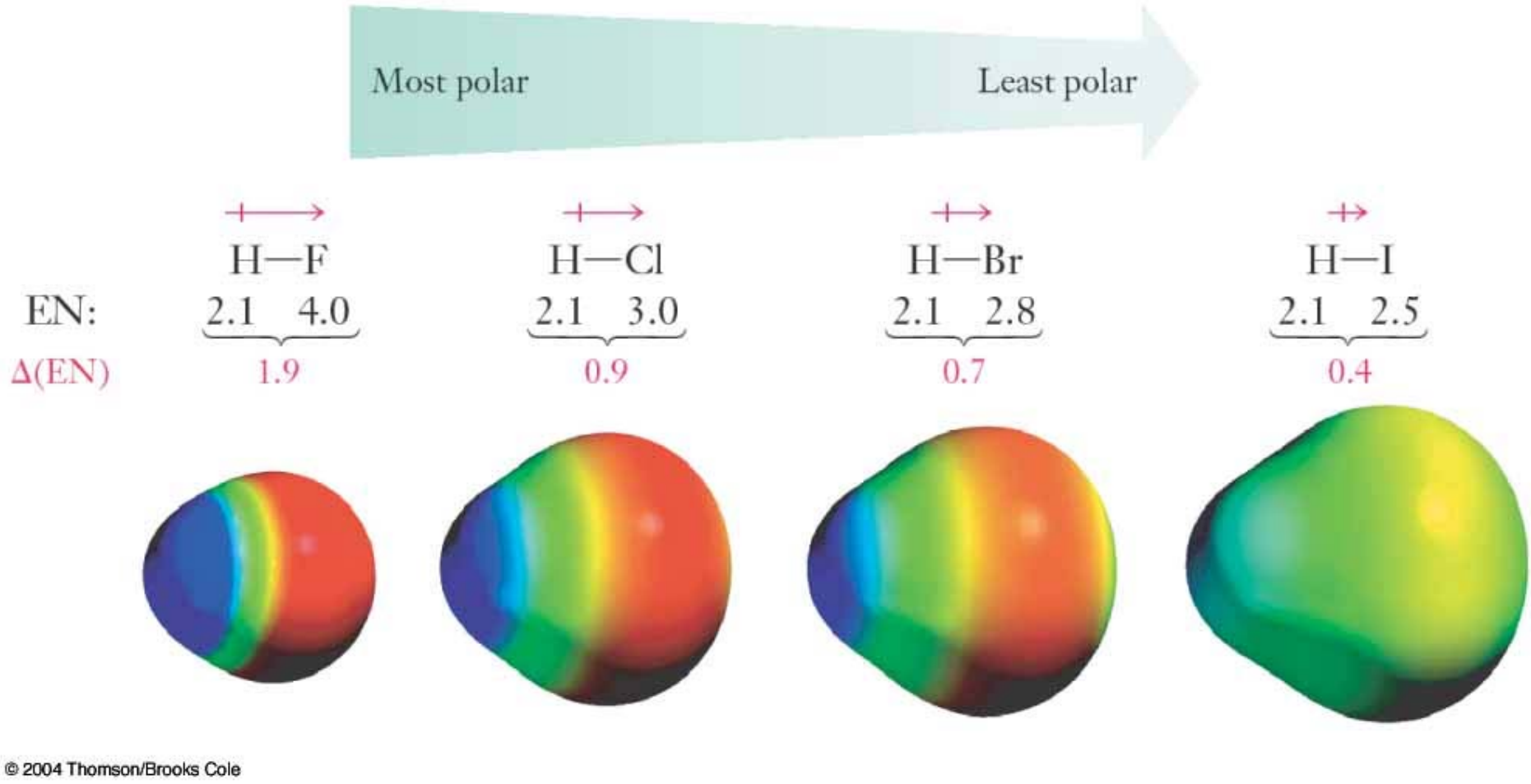
- This bond is called a polar covalent bond
- The charge distribution is not symmetrical

Electron Density Distribution

H —— F

- Blue – low electron density (more positive)
- Red – high electron density (more negative)

Polar Bonds



Polar Molecules

- Polar molecules can be attracted by magnetic and electric fields
- We sometimes represent these molecules as dipoles
- The direction of the dipole is from the positive to the negative pole
- Each dipole is characterized by a dipole moment
- The larger the difference in the electronegativities of the bonded elements, the higher the dipole moment of the molecule

The Continuous Range of Bonding Types

- Covalent and ionic bonding represent two extremes:
  - In pure nonpolar covalent bonds electrons are equally shared by the atoms
  - In pure electrostatic ionic bonds electrons are completely transferred from one atom to the other
- Most compounds fall somewhere between these two extremes

The Continuous Range of Bonding Types

- All bonds have some ionic and some covalent character
  - For example, HI is about 17% ionic and 83% covalent
- As the electronegativity difference increases, the bond becomes
  - more polar
  - less covalent
  - more ionic

Example 1

- Which of these bonds is more polar:
  - N–O
  - C–Cl
  - Na–H
  - Na–Br

Example 2

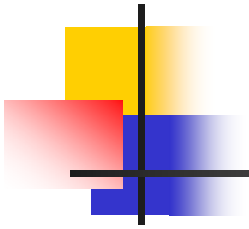
- Which of these bonds is less covalent:
  - Al–I
  - Al–Cl
  - Al–F
  - Al–Br

Example 3

- Which of these bonds has the highest dipole moment:
  - C–B
  - C–C
  - C–N
  - C–O
  - C–F

The Octet Rule

- In most of their compounds, the representative elements achieve noble gas configurations
- Lewis dot formulas are based on the octet rule
- Electrons which are shared among two atoms are called bonding electrons
- Unshared electrons are called lone pairs or nonbonding electrons



H_2O Molecule

NH_3 Molecule

NH_4^+ Ion

- Lewis formulas can also be drawn for polyatomic ions

CO_2 Molecule

N_2 Molecule

Covalent Bonding

- Covalent bonds are formed when atoms share electrons
  - If the atoms share 2 electrons a single covalent bond is formed
  - If the atoms share 4 electrons a double covalent bond is formed
  - If the atoms share 6 electrons a triple covalent bond is formed

The Octet Rule

$$S = N - A$$

- S = total number of electrons **shared** in bonds
- N = total number of electrons **needed** to achieve a noble gas configuration
  - 8 for representative elements
  - 2 for H atoms
- A = total number of electrons **available** in valence shells of the atoms
  - A is equal to the periodic group number for each element
- A-S = number of electrons in **lone** pairs

Examples

- F_2
- H_2O
- CH_4
- CO_2

Examples

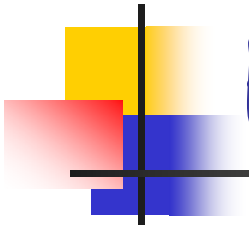
- N_2
- CO
- C_2H_2
- HCN

Examples

- For ions we must adjust the number of electrons available, *A*:
  - Add one e^- to *A* for each negative charge
  - Subtract one e^- from *A* for each positive charge

- NH_4^+

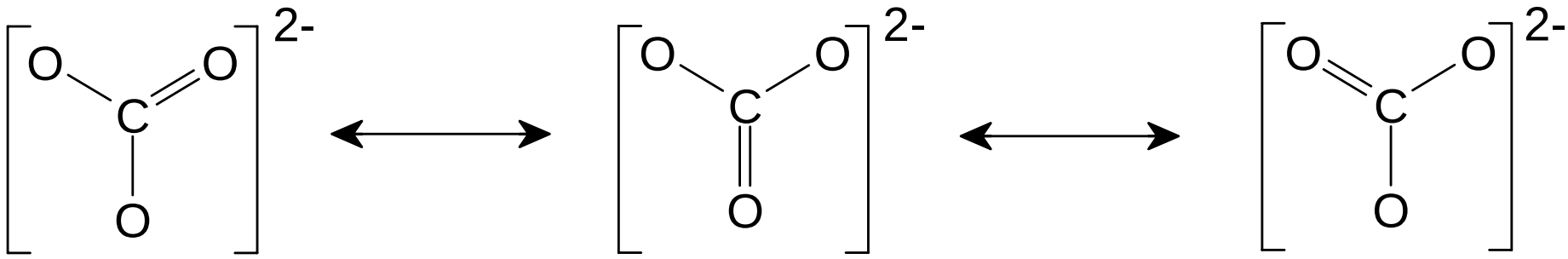
- BF_4^-



Example: CO_3^{2-}

Resonance

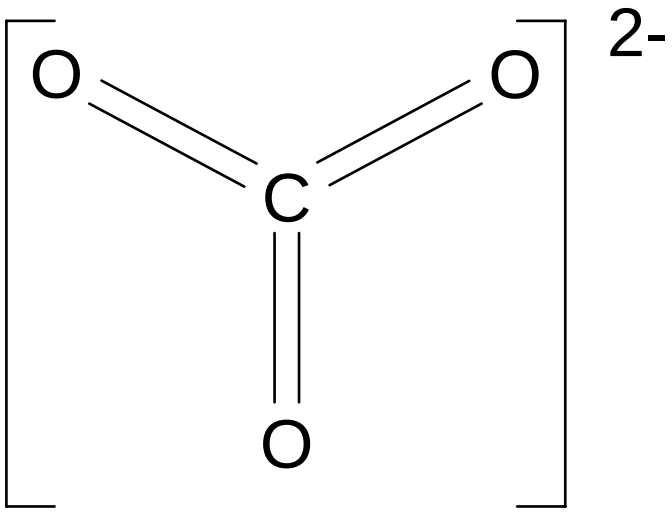
- There are three possible structures for CO_3^{2-}
  - The double bond can be placed in one of three places



- These are called equivalent <u>resonance structures</u>
- The real structure of the CO_3^{2-} anion is an average of these three resonance structures

Resonance

- There are no single or double bonds in CO_3^{2-}
- All three bonds are equivalent
- They are intermediate between the single and double bond



Resonance: Other Examples

- SO_3

Resonance: Other Examples

- NO_3^-

Resonance: Other Examples

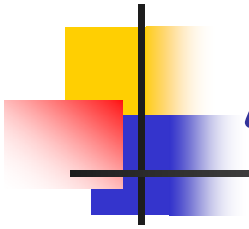
- SO_4^{2-}

Exceptions to the Octet Rule

- In those cases where the octet rule does not apply, the substituents attached to the central atom nearly always attain noble gas configurations
- The central atom does not have a noble gas configuration but may have fewer than 8 or more than 8 electrons

Examples

- BBr_3
- AsF_5



Assignments & Reminders

- Go through the lecture notes
- Read Chapter 7 completely, except for Sections 7-7 & 7-8
- Read Sections 4-5 & 4-6 of Chapter 4
- Homework #4 due by Oct. 16 @ 3 p.m.
- Review Session @ 5:15 p.m. on Sunday