Unit 2—Molecular and Ionic Compound Structure and Properties
Ch 8, 9, 12

Bonding and octet rule *ch 8 introduction and 8.1*

What happens to the electrons in:

An ionic bond

A covalent bond

A metallic bond

What do Lewis symbols represent?

What is the octet rule?

Practice 1

1. Draw the Lewis symbols for arsenic, calcium, krypton, and iodine.

2. Indicate if the above atoms will gain or lose electrons and how many electrons to satisfy the octet rule.

Ionic bonding *ch 8.2, some of 12.5*

Which atom will give up electrons? What common property do these atoms have?

Which atom will gain electrons? What common property do these atoms have?

Practice 2

1. Use Lewis symbols to show the reaction between calcium and sulfur forming an ionic compound.

Why do ionic compounds form crystal structures?

What properties to ionic compounds have?

Define lattice energy.

Lattice energy is exothermic/endothermic (circle one). What does this mean?

Potential energy equation: $\boxed{E_{el} = \frac{Kq_1q_2}{r}}$

What commonalities do ionic compounds with higher lattice energies have?

Why does Na only form a +1 charge and not +2 or +3?

Metallic bonding *some of ch 12.4 and 12.3 (metallic solids, structures of metallic solids, alloys, electron-sea model)*

What is a "sea of electrons?"

What are some common properties of metals?

How does the sea of electrons contribute to the properties of metals above?

What is an alloy?

Define substitutional alloy and give one example.

Define interstitial alloy and give one example.

Practice 3

1. Is steel (an alloy of iron and carbon) a substitutional alloy or an interstitial alloy? How can you tell?

2. Is bronze (an alloy of copper and tin) a substitutional alloy or an interstitial alloy? How can you tell?

Covalent bonding *ch 8.3*

What are some properties common to covalent/molecular compounds?

Define bonding pair of electrons and lone pair of electrons.

Drawing Lewis dot structures, *ch 8.3, 8.5*

1. Count the number of total valence electrons in molecule, determine # electron pairs
2. Write the atoms in the order they'll go
    a. Atom with lowest IE in center
    b. H never in center
3. Make a bond (2 e^-) between each atom
4. Complete octet with more bonds or lone pairs of e^-

Draw dot structures for:

i. Nitrogen trifluoride

ii. Carbon dioxide

iii. Carbonate ion

iv. Hypochlorite

v. Water ion

vi. Nitrogen gas

Incomplete octet: H wants ___ electrons, Be can have ___ electrons, B is stable with ___ or ___ electrons

Practice 4

1. Draw dot structures for:
    a. Ammonium ion
    b. Carbon monoxide
    c. Oxygen gas
    d. Boron tribromide
    e. Phosphorus trichloride
    f. Hydrochloric acid
    g. Carbon tetrafluoride
    h. Dihydrogen monosulfide

Resonance and formal charges, *ch 8.6*

Draw the TWO dot structures for ozone, O_3:

What are resonance structures?

Draw the three resonance structures for nitrate ion:

Practice 5

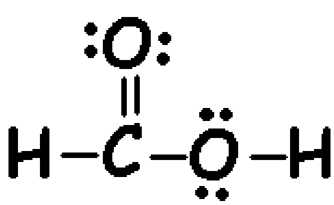
1. Draw all resonance structures for carbonate ion.

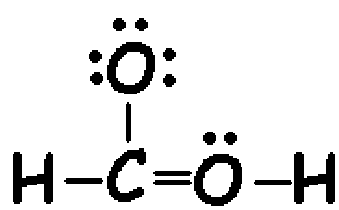
All resonance structures contribute to actual electron arrangement (but maybe not equally). Formal charge can help you evaluate dot structures to determine the most important structure(s) to actual arrangement of electrons in molecule.

Formal charge = # valence e^- on atom - # unshared e^- - ½ # bonding e^-

- For neutral molecules, the sum of all formal charges = _______
- For ions, sum of formal charges = _________
- As low a formal charge as possible
- Negative formal charges should be on electronegative elements

Determine formal charges on each atom:





Practice 6

1. Draw all possible resonance structures for nitrogen monoxide.
2. Use formal charges to determine the structure that contributes most to the actual arrangement of the e^- in the molecule.

Exceptions to the octet rule *ch 8.7*

Odd number of valence electrons

Practice 7

Draw the structures (including resonance if applicable) for:

ClO_2 NO_2

Less than an octet

Which elements do not require a full octet of valence electrons when bonding?

Why?

Practice 8

Draw the structures (including resonance if applicable) for:

$BeCl_2$ BCl_3 AlH_3

More than an octet (expanded octet)

Which elements can have more than an octet when bonding?

Why?

Practice 9

Draw the structures (including resonance if applicable) for:

SF_4 BrF_3 SF_6

SbF_5 ICl_2^- SO_2

Bond polarity and electronegativity *ch 8.4*

Definitions:

Bond polarity –

Nonpolar covalent bond –

Polar covalent bond –

Electronegativity –

What is the periodic trend for electronegativity? Explain.

Which element is the most electronegative? Least electronegative?

Draw the dot structure of CO twice. Indicate the polarity using δ+/ δ- and $+\!\!\longrightarrow$

What is a dipole?

What is a dipole moment?

What factors affect the dipole moment of a bond?

Strengths of covalent bonds *ch 8.8*

What is bond enthalpy?

Why do some bonds have higher enthalpies? Which types of bonds are these?

What does a higher bond enthalpy mean?

Molecular shapes and VSEPR *ch 9.1 and 9.2*

What does VSEPR stand for?

What does it mean?

Definitions:

Bonding pair of electrons –

Nonbonding pair/lone pair –

Electron domain –

Electron domain geometry –

Molecular geometry –

Molecule polarity *ch 9.3*

When determining whether a molecule is polar or nonpolar, you need to consider the ________________ and ________________________.

Symmetrical molecules are polar/nonpolar (circle one) because:

Asymmetrical molecules are polar/nonpolar (circle one) because:

For our purposes, the electronegativities of C and H are similar and C—H is considered a nonpolar bond.

Draw the dot structure, determine the molecular geometry and polarity of:

NCl_3 BF_3

Practice 10

Are the following molecules from *practice 9* polar or nonpolar?

1. SF_4
2. BrF_3
3. SF_6
4. SbF_5
5. ICl_2^-
6. SO_2