Unit 6 – Covalent Compounds

Early Booklet E.C.:	**/ 2**
Unit 6 Hwk. Pts:	**/ 38**
Unit 6 Lab Pts:	**/ 64**
Late, Incomplete, No Work, No Units Fees?	Y / N

Learning Targets for Unit 6

1.1 I can apply the octet rule to atoms that form covalent bonds
1.2 I can describe the formation of single, double and triple covalent bonds
1.3 I can contrast sigma and pi bonds
1.4 I can relate the strength of a covalent bond to its bond length and bond dissociation and energy
1.5 I can translate molecular formulas into binary molecular compound names
1.6 I can name acidic solutions
1.7 I can list the basic steps used to draw Lewis structures
1.8 I can explain why resonance occurs, and identify resonance structures
1.9 I can identify three exceptions to the octet rule, and name molecules in which these exceptions occur
1.10 I can summarize the VSEPR bonding theory
1.11 I can predict the shape of and the bond angles in a molecule
1.12 I can define hybridization
1.13 I can describe how electronegativity is used to determine bond type
1.14 I can compare and contrast polar and nonpolar covalent bonds and polar and nonpolar molecules
1.15 I can generalize about the characteristics of covalently bonded compounds

Unit Vocabulary for Unit 6

Covalent bond	Endothermic reaction	Exothermic reaction	Lewis structure
Molecule	Pi bond	Sigma bond	Oxyacid
Coordinate covalent bond	Resonance		Hybridization
VSEPR model	Polar covalent bond		

Completed 6.1 Pts.:	/ 9
Late, Incomplete, No work, No Units Fee:	- 1 - 2 - 3
Final Score:	/ 9

6.1 Problems – Covalent Bonds

Go to hut-lhansen.weebly.com for comprehensive lesson notes.

Objectives:

A. Distinguish between single, double, and triple bonds.
B. Draw Lewis Structures for molecules.
C. Predict whether a reaction releases or absorbs energy.

Synopsis

Molecules are compounds where electrons in covalent bonds are shared between atoms, rather than ionic bonds, where attractions between ions come from positive and negative ions attracting each other due to electron imbalance. Depending on an atom's valence electron count, atoms form single, double, and triple bonds to satisfy the octet rule.

Chemical reactions involve transfers of energy. A reaction could give off energy, or absorb it, depending on what is reacting.

Terms

Covalent Bonds form in non-ionic compounds where electrons are shared. This is different from ionic compounds, where electrons are lost or gained by elements. Atoms form covalent bonds to gain stability, and follow the octet rule (see 5.1) – atoms having eight electrons in their outer energy level are chemically stable, meaning that they don't react more.

Molecule: a group of covalently-bonded elements – generally non-metals.

Single Bonds: called sigma (σ) bonds. Two electrons form a link between atoms.

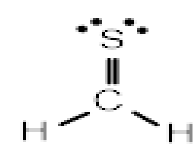
Double Bonds: called pi (σ) bonds. Four electrons link atoms.

Triple Bonds: called pi (σ) bonds. Six electrons link atoms.

Diatomic Elements form covalent bonds with themselves, and is important in chemical reactions. Mnemonic to remember them: if an element ends in "gen" or "ine". Only seven: Hydrogen = H_2 Nitrogen = N_2 Oxygen = O_2 Fluorine = F_2 Chlorine = Cl_2 Bromine = Br_2 Iodine = I_2

1. Guided Example: Covalent Bonds

Circle the single bonds, box the double bonds, triangle the lone pairs of electrons of the



molecule.

Lewis Structures show bonds in molecules, as well as placement of atoms with respect to each other. They also show non-bonding electrons, which pair up but don't link to other atoms.

Bond Dissociation Energy: Energy needed to break chemical bonds. The shorter a bond, the stronger it is. The more bonds present between two atoms, the shorter the bond.

Endothermic Reaction: more energy required to break bonds than is released during bond formation.

Exothermic Reaction: more energy released during bond formation than needed for bond breaking.

Simple Lewis Structure Process – simple molecules only!

1. Draw each atom of your molecule, with singleton atoms from the formula in the middle.
2. Draw the valence electrons of each atom – remember not to pair electrons until you have to.
3. Link lone electrons to spaces where 'missing' electrons are. Each atom should then have eight electrons: either from its original valence count, or by sharing them. Exceptions: hydrogen can NEVER have more than two electrons, boron never has more than six.
4. Redraw your molecule, with the central atom linked to terminal atoms and lone pairs symmetrically.

4. Guided Example
Draw the Lewis Structure of water: H_2O

3. **Guided Example:**
ammonia: NH_3

2. Guided Example:
methane: CH_4

5. How is a covalent bond different from an ionic bond?

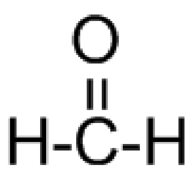
6. Draw the electron dot structure for the following elements, then predict the number of covalent bonds needed for each of these elements to satisfy the octet rule.

N   As

Br   Se

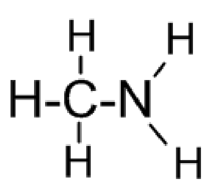
7. Circle and label the single (sigma) and multiple (pi) bonds in each of the molecules:

a.



b.

8. Compare the following two examples of carbon-nitrogen bonds. In which bond is the carbon-nitrogen bond shorter? In which is it stronger? What makes a bond stronger?



9. Draw Lewis structures for the following:
   b. PH_3 b. H_2S c. F_2

Reflections. What parts of this section made sense? ______________________________
What interested you most? ______________________________
What parts need more explanation? ______________________________

Completed 6.2 Pts.:	/ 6
Late, Incomplete, No work, No Units Fee:	- 1 - 2 - 3
Final Score:	/ 6

6.2 Problems – Naming Molecules

Go to hut-lhansen.weebly.com for comprehensive lesson notes.

Objectives:

A. Name binary molecules from formulas, and writing formulas from names.
B. Name binary acids and oxyacids.

Synopsis

The names of binary compounds are based on the numbers of atoms of each element. Naming certain types of acids also is covered.

Terms

Binary Molecule Compounds: Covalently bonded compound formed from ratios of two different elements.

Acid: Ionic compound that releases hydrogen ions (H^+) in aqueous solutions.

Binary Acid: Acid formed from hydrogen and another element.

Oxyacids: Contain hydrogen and an oxyanion.

Naming Binary Compounds Process

1. Name 1st element. *Ex: $SeCl_2$* *Ex: selenium*
2. Name 2nd element, with -IDE ending. *Ex: chloride*
3. Use prefixes (Resource 4) to tell *how many* of each element there are. *Ex: selenium dichloride*
   Note: 1st element never gets "mono-"
4. Use the reverse process to determine formulas from names.

1. Guided Binary Compounds Naming Practice

1. H_2O 2. CO_2 3. CCl_4 4. H_2O

Naming Binary Acids Process

1. Write the name in ionic format *Ex:* HCl: hydrogen chloride
2. Remove *gen* and *ide* endings. *Ex:* HCl: hydro- chlor-
3. Combine word parts, and add the ending: *ic acid*: *Ex:* hydrochloric acid
4. Check you acid name: Resources Page 5.

2. Guided Binary Acids Naming Practice

1. HBr 2. HI 3. H_2S

Naming Oxyacids Process

1. Write the name in ionic format. Examples: HNO_2 = hydrogen nitrite HNO_3 = hydrogen nitrate
2. Remove the word 'hydrogen', then
   A. Replace -*ite* ending with *–ous acid* nitrous acid.
   B. Replace -*ate* ending with *–ic acid* nitric acid.
3. Check you acid name: Resources Page 5.

3. Guided Oxycids Naming Practice

1. HClO 2. $HClO_2$ 3. $HClO_3$ 4. $HClO_4$

4. Name each binary molecular compound:
   a. NF_3
   b. NO
   c. SO_3
   d. SiF_4

5. Write the formula for each binary molecule:
   a. sulfur difluoride
   b. selenium dibromide
   c. carbon tetrafluoride
   d. hydrobromic acid

6. Name each acid:
   a. H_2SO_3
   b. H_2SO_4
   c. H_3PO_4

Reflections. What parts of this section made sense? ______________________________
What interested you most? ______________________________
What parts need more explanation? ______________________________

6.3 Problems – Molecular Structures

Go to hut-lhansen.weebly.com for comprehensive lesson notes.

Completed 6.3 Pts.: / 8
Late, Incomplete, No work, No Units Fee: - 1 - 2 - 3
Final Score: / 8

Objectives:

A. Draw Lewis structures for more all molecules.
B. Recognize when multiple Lewis structures for the same molecule are possible.

Synopsis

This subunit deals with more complex Lewis structures, including those with resonance.

Terms

Five ways to portray molecules Ex. = phosphorus trihydride

1. Molecular (chemical) formula: PH_3
2. Lewis structure:
3. Structural formula – no lone electrons:
4. Ball and stick model:
5. Space filling model:

Resonance: When a double or triple bond is present, more than one correct Lewis structure can be made.

Sub-Octets: Some compounds are stable with less than an octet.

Expanded Octets: Some compounds are stable with more than an octet.

Coordinate Covalent Bonds: Lone pairs of electrons are donated completely to an atom.

Lewis Structure Process – Complex Molecules

1. Place atoms roughly:
   A. Central atom is a formula's single atom.
   B. Hydrogen is <u>always</u> a terminal atom.
2. Add up <u>all</u> valence electrons, add or subtract charge electrons   Ex: NO_3^- has 1 more e^-.
3. Determine e^- pairs: divide valence e^- number by two.
4. Draw single bond b/w central and terminal atoms.
5. Count and place remaining e^- pairs:
   A. Place on terminal atoms to satisfy octet rule.
   B. Place final pairs on central atom.
   C. If you are making the structure of an ion, put it in brackets with its charge superscripted.
6. Ask: does the central atom obey octet rule?
   If not, make double or triple bonds <u>using lone pairs</u> from terminal atoms to the central atom.

1. Lewis Structure Guided Example Phosphate Ion (PO_4^{3-}) How many V. E.? How many pairs?	2. Lewis Structure Guided Example Sulfur Dioxide (SO_2) (resonance) How many V. E.? How many pairs?

3. Draw the Lewis structures of sulfur monoxide (SO), sulfur trioxide (SO_3) (this has resonance structures), and the sulfate ion (SO_4^{2-}).

4. Draw the shorthand resonance structure for the polyatomic carbonate ion CO_3^{2-}.

Draw Lewis structures for these, each of which has a central atom that is an exception to the octet rule:

5. PCl_3
6. BF_3
7. ClF_5
8. XeF_4

<u>Reflections.</u> What parts of this section made sense? ______________________________

What interested you most? ______________________________

What parts need more explanation? ______________________________

Completed 6.4 Pts.: / 8
Late, Incomplete, No work, No Units Fee: - 1 - 2 - 3
Final Score: / 8

6.4 Problems – Molecular Shapes

Go to hut-lhansen.weebly.com for comprehensive lesson notes.

Objectives:

A. Students will be able to predict molecular shapes based on chemical formulas.

Synopsis

Molecules have unique shapes, based on two factors: how many atoms there are, and how many lone pairs of electrons there are.

Terms

VSEPR Model: Valence Shell Electron Pair Repulsion model predicts shapes of molecules. Shape minimizes repulsion of shared/unshared electron pairs.

Terminal Atom: atom connected to one other atom in a molecule.

Central Atom: atom in the middle of a molecule.

Lone Pair: pair of electrons not involved in a bond

Bond Angle: angle between a central atom and its terminal atoms.

Hybrid: two things combine forming something with characteristics of both. Hybridized orbitals (s,p, and d) mix, & form new, symmetrical orbitals. The hybridization of the central atom determines shape.

Expanded Octet: Central atoms can have more than 8 electrons: D orbitals are involved, and leads to sp^3d or sp^3d^2 hybridization.

Determining Hybridization & Molecular Shape Process

0. Make a Lewis Structure.
1. Add lone pairs & terminal atoms. (Number = 1 - 6)
2. Determine hybridization of central atom using the chart on Resources 6.
3. Determine the shape using Resources Pages 6 and 7: match lone pair count, terminal atom count, and hybridization column by column until there's a match.

1. Guided Practice: Shape of Ammonia – NH_3

Lewis structure:

How many terminal atoms?

How many lone pairs?

What is the hybridization?

What shape does it have?

What is the bond angle of the central atom?

2. Guided Practice: Shape of sulfur hexachloride – SCl_{36}

Lewis structure:

How many terminal atoms?

How many lone pairs?

What is the hybridization?

What shape does it have?

What is the bond angle of the central atom?

Make Lewis structures for the following, and determine the molecular shape. What is the bond angle between terminal atoms and the central atom?

3. SCl_2
4. HCl
5. SiH_4
6. XeF_2
7. NH_2Cl (Nitrogen is central atom)
8. BF_3

Reflections. What parts of this section made sense? ______________________________

What interested you most? ______________________________

What parts need more explanation? ______________________________

Completed 6.5 Pts.: / 7
Late, Incomplete, No work, No Units Fee: - 1 - 2 - 3
Final Score: / 7

6.5 Problems – Electronegativity & Polarity

Go to hut-lhansen.weebly.com for comprehensive lesson notes.

Objectives:

A. Predict whether a molecule is polar based on its shape and composition.
B. Describe properties of polar and non-polar molecules.

Synopsis

Different molecular properties arise when bonds between atoms are not totally ionic or covalent.

Terms

Electronegativity: atom's ability to attract electrons in a bond. defines bond type. Resources Page 5.

Note: in formulas, the least electronegative element is written first, followed by the most. Ex: HCl

Polar Covalent Bond: connection between atoms where electrons are not shared equally.

Dipole: a molecule with charged ends: partial charges are present on a molecule due to lopsided electron distribution.

The charged ends of the molecule are labeled with δ^+ if it's positive or δ^- if it's negative (δ is Greek small delta). A dipole is denoted with an arrow pointing towards the negatively charged end.

1. Indicating Polarity Guided Practice

Draw HCl, showing positive and negative ends.

Or:

Properties of Covalent Compounds:

A. Polar compounds dissolve in polar solvents (acetone, water, ammonia – etc.).
B. Non-polar compounds dissolve in non-polar solvents (vegetable oil, gasoline, mineral oil – etc.).
C. Dissimilar compounds tend not to dissolve in each other (oil and water).
D. Covalent compounds have lower melting/boiling temperatures than ionic compounds
E. They have some intermolecular forces:
    I. Dipole-dipole: attraction between positive and negative ends of molecules.
    II. Hydrogen bond: a dipole-dipole force where the positive end is a hydrogen atom; the negative end is either fluorine, oxygen, or nitrogen

Determining Polarity Process – Two Criteria

1. Determine absolute (size only – disregard plus or minus) electronegativity difference between central and individual terminal atoms. It must be between 0.4 and 1.7 to be polar.
2. Draw Lewis structure and determine shape.
3. Molecule must be asymmetric to be polar, either by:
    A. shape: linear diatomic, bent, trigonal pyramidal, t-shaped, seesaw, or square pyramidal.
    B. composition: different terminal atoms.

3. Polarity Guided Practice: Is HOBr Polar?

Electronegativity differences: H vs O =

Br vs O =

Lewis structure:

2. Polarity Guided Practice: Is SO_3^{2-} Polar?

Electronegativity differences: H vs O =

Br vs O =

Lewis structure:

Which atom of the pair is more negatively charged based on electronegativity values (Use resources Page 5)?

4. C-H   C-N   C-S   C-O

5. Rank the following bonds in order of increasing polarity (Use resources Page 5):

C-H   N-H   Si-H   O-H   Cl-H   H-H.

Determine if these molecules are polar. Draw Lewis structures to support your answers.

6. H_3O^+ PCl_5

7. H_2S ClO_3

Reflections. What parts of this section made sense? ____________________

What interested you most? ____________________

What parts need more explanation? ____________________

Chemistry	6.1 Lab - Covalent (Molecular) Compounds Lab				
Name:					Correction Credit: Half
Lab Points:	Missed:	Late, No Units, No Work Fee:	First Score:	Corrections:	Final Score:
12		-1 -2 -3			

Theory:

Covalent compounds often are substances made of nonmetals: the elements hydrogen, boron, carbon, nitrogen, oxygen, fluorine, silicon, phosphorus, sulfur, chlorine, arsenic, selenium, bromine, iodine.

In this lab you will be making three covalent compounds: oxygen, carbon dioxide, and hydrogen. Each compound has a specific test that can be done to identify it. Be sure to write complete sentences for your observations.

Safety:

You will be using candles to ignite several wooden splints.

Goggles, long pants, and close-toed shoes must be worn during the entire lab.

Equipment and Chemicals:

Tea Candle – Clamp it to the ring stand.
10 mL graduated cylinder
Three small test tubes
Three large test tubes
Wooden test tube rack
1.0 M hydrochloric acid (HCl)
Magnesium ribbon (Mg)
Marble chips ($CaCO_3$)
Hydrogen peroxide (H_2O_2)
Manganese (IV) oxide (MnO_2)
Wooden splints
Lighter (teacher Use Only)
Metal spatula

Procedure 1: Oxygen

The chemical reaction to make oxygen is as follows:

$$2\ H_2O_2 + MnO_2 \rightarrow O_2 + 2H_2O + MnO_2$$

The MnO_2 does not react in this reaction: it serves as a catalyst – a chemical that promotes a reaction but is not consumed.

Oxygen is needed for combustion, and makes a long **glowing** splint burst into flames.

1. Pour 10.0 mL of the hydrogen peroxide into a small test tube in your rack.
2. Put the small tube in your rack and use the spatula to put a small amount of manganese dioxide into the tube.
3. Cover the small tube with a large tube and observe the reaction for four minutes, or until the bubbles stop. Record your observations here (2 pts). Does it heat up?

4. Ignite a wooden splint with the candle, let it burn for a few seconds, then blow it out (it must be **glowing**). Immediately put it in the **small** tube all the way to the top of the liquid, but don't submerge it. Record what happens (1 pt).

5. Draw the Lewis structure of O_2 (1 pt).

Procedure 2: Carbon Dioxide

The chemical reaction to make carbon dioxide is as follows:

$$2\ HCl + CaCO_3 \rightarrow CaCl_2 + H_2O + CO_2$$

Carbon dioxide suffocates flames, and will do so to your burning splint.

1. Pour 5.0 mL of the hydrochloric acid (HCl) into a small test tube in your rack.
2. Drop 3 marble chips into it.
3. Cover the small tube with a large tube and allow the reaction to go for three minutes. Record your observations here (2 pts). Does it heat up?

4. Ignite a wooden splint with the candle, then put it in the small tube without submerging it. Record what happens (1 pt).

5. Draw the Lewis structure of CO_2 (1 pt) – carbon is in the middle.

Procedure 3: Hydrogen

The chemical reaction to make hydrogen is as follows:

$$2\ HCl + Mg \rightarrow H_2 + MgCl_2$$

Hydrogen burns during "The Barking Dog Test" to make a popping sound.

1. Pour 5.0 mL of the hydrochloric acid (HCl) into a small test tube in your rack.
2. Put one 1.5 cm piece of magnesium ribbon in the HCl tube.
3. Cover the small tube with a large tube and allow the reaction to finish. Record your observations here (1 pt). Does it heat up?

4. Making sure that the large tube remains upside down, ignite a wooden splint with the candle, and put it in the smaller (lower) tube first. Then, rapidly reignite your splint if necessary, and raise it into the larger (upper) tube. Record what happens for both tubes (2 pts).

5. Draw the Lewis structure of H_2 (1 pt).

Clean up:

Bring your test tubes to the front and pour them into the waste beaker.
Wash the rest of your dishes and clean your lab station.

Chemistry	6.2 Lab - Lewis Structures				
Name:					Correction Credit: Half
Lab Points:	Missed:	Late, No Units, No Work Fee:	First Score:	Corrections:	Final Score:
16		-1 -2 -3 -4			

Part 0. Draw Lewis dot diagrams of these elements or ions. 3 points total.

Na Na	Cl Cl	O O	I I	Al Al
K^+ K	F^- F	S^{2-} S	Br^- Br	Ca^{2+} Ca

Part 1. Draw Lewis structures (½ point) and name (½ point) the following molecules. Include all non-bonding (lone) electron pairs in your diagrams.

1. SO Name ___________________________

2. CCl_4 Name ___________________________

3. CS_2 Name ___________________________

4. SeF_2 Name ___________________________

5. BF_4^- Name: tetrafluoroborate ion

Part 2. (2 pts each) Draw resonance structures for the following molecules or ions. Be sure to include all non-bonding (lone) electron pairs in your diagrams. Use resonance shorthand if you want to.

6. ozone - O_3 (Two structures – Hint: This molecule does <u>NOT</u> form a triangular ring)

7. nitrite ion - NO_2^- (Two structures for this ion)

8. formate ion – HCO_2^- (Carbon is the central atom - two structures for this ion)

Chemistry	6.3 Lab - Molecular Shapes Activity				
Name:					Correction Credit: Full
Lab Points:	Missed:	Late, No Units, No Work Fee:	First Score:	Corrections:	Final Score:
20		-1 -2 -3 -4 -5			

In your lab groups, draw Lewis structures for the following. Determine hybridization of the central atom and write the shape of the molecule. Use the modeling kit to build selected molecules, and show me for approval. Two points each.

We will discuss this next class.

Keep the kits together please!

Suggested Atomic Model Colors:

Short (single) Bonds: 6
Long (double) Bonds: 4
White = Hydrogen: 4
Red = Oxygen: 4

Black = Carbon: 1
Blue = Nitrogen: 1
Yellow = Sulfur: 1
Green = Halogen:6

Tan = Phosphorus, Boron: 1
Silver = Metals:1

A Note on Bonds: Short gray – single; long gray = double/triple; purple = lone pairs.

1. CO
Hybridization: _____
Shape: ___________

2. CH_2S (C is central atom)
Hybridization: _____
Shape: ___________

3. CH_3Cl (C is central atom)
Hybridization: _____
Shape: ___________
Build this molecule. Approval: _____

4. NH_2Cl (N is central atom)
Hybridization: _____
Shape: ___________

5. H_2S
Hybridization: _____
Shape: ___________

6. CO_2
Hybridization: _____
Shape: ___________
Build this molecule. Approval: _____

7. BH_3
Hybridization: _____
Shape: ___________
Approval: _____

8. PH_3Cl_2 (P is the central atom)
Hybridization: _____
Shape: ___________
Build this molecule. Approval: _____

9. SH_2Cl_2 (S is the central atom)
Hybridization: _____
Shape: ___________

10. TeF_6
Hybridization: _____
Shape: ___________
Approval: _____

Chemistry	6.4 Lab - Polar Molecules Activity				
Name:					Correction Credit: Full
Hwk. Points:	Missed:	Late, No Units, No Work Fee:	First Score:	Corrections:	Final Score:
16		-1 -2 -3 -4			

Draw Lewis structures for the following, including non-bonding pairs. Write down the shape of the molecule, and circle whether it is polar or not. Use the modeling kit to build selected molecules, and show me for approval. Two points each.

We will correct this in class next time we meet.

Suggested Atomic Model Colors:

Short (single) Bonds: 6
Long (double) Bonds: 4
White = Hydrogen: 4
Red = Oxygen: 4

Black = Carbon: 1
Blue = Nitrogen: 1
Yellow = Sulfur: 1
Green = Halogen:6

Tan = Phosphorus, Boron: 1
Silver = Metals:1

1. NF_3
Shape: ______________
Electronegativity difference: ________
Polar or nonpolar?

2. BF_3
Shape: ______________
Electronegativity difference: ________
Polar or nonpolar?
Build this molecule. Approval: _____

3. HBr
Shape: ______________
Electronegativity difference: ________
Polar or nonpolar?

4. NH_4^+
Shape: ______________
Electronegativity difference: ________
Polar or nonpolar?

5. COF_2 (C is central) (Note: fluorine only can have single bonds)
Shape: ______________
2 Electronegativity differences: ________
Polar or nonpolar?
Build this molecule. Approval: _____

6. I_3^- (Tri-iodide ion – Hint: not triangular)
Shape: ______________
Electronegativity difference: ________
Polar or nonpolar?

7. AsO_4^{3-}
Shape: ______________
Electronegativity difference: ________
Polar or nonpolar?
Approval: _____

8. $CHBr_3$ (Carbon is central)
Shape: ______________
2 Electronegativity differences: ________
Polar or nonpolar?
Build this molecule. Approval: _____

Completed Points: / 10
Late/Inc. Fee: -1 -2 - 3
Final Score: / 10

Unit 6 Test Review – Covalent Compounds

This serves as test preparation for the Unit 6 exam. Points earned are based on completion, and we will go over any questions you have during the review. Additionally, you may be called upon to present a selection of these problems to the class.

1. What is a sigma bond? What is a pi bond?

2. How many covalent bonds could the following elements have: carbon, oxygen, nitrogen, and phosphorus?

3. How do two atoms form covalent bonds?

4. How do covalent bonds differ from ionic bonds?

5. What are the names of the following binary compounds:

   CF_4 SO SO_2 ClF_3 N_2O_2?

6. Write the formulas for the following compounds:

   dichlorine monoxide selenium diiodide.

7. Write the formulas of the following acids:

hydrochloric hydroiodic sulfurous nitrous acetic.

8. Draw the Lewis structures for the following molecules:

   a. PBr_3
   b. HCN (carbon in middle)

9. Draw resonance structures for SO_3.

10. What is the shape of the carbonate ion? Draw a structure to answer this.

11. What is the shape of the cyanide ion? Draw a structure to answer this.

12. What is the shape of the CSO molecule? Sulfur is the central atom, and draw a structure to answer this.

13. Based on electronegativity values, which of the atoms in the following bonded pairs is more positively charged:

C-O	N-O	H-O	C-H	N-H

14. Rank the following bonds in order of increasing polarity:

C-Cl	C-O	H-F	C-H

15. Is $SiCl_4$ polar? Draw its Lewis structure to determine your answer.

16. Is NF_3 polar? Draw its Lewis structure to determine your answer.