Form a bond: Shown is the energy change as $\text{H}_2$ forms a covalent bond. Covalent bond: Two orbitals overlap and contain two shared electrons of opposite spin in the overlap region. When the covalent bond is formed the overall energy of the system decreases as shown in the graph below. For $\text{H}_2$ the bond energy is $-436 \text{kJ/mol}$ and the bond length is $74 \text{pm}$.

Electronegativity: The degree to which an atom can attract an $e^-$ in a bond. The greater the electronegativity (EN), the greater the "pull" an atom has on an $e^-$ pair in a bond.

Trend: Increasing toward F (up and to the right).

Types of Bonds

Nonpolar Covalent: equal sharing of $e^-$ between two identical nonmetal elements (e.g., $\text{O}_2$, and $\text{Cl}_2$).

Electronegativity difference ($\Delta \text{EN}$) is zero.

$\Delta \text{EN} = |\text{EN}_1 - \text{EN}_2|$ with $\text{EN}_1$ and $\text{EN}_2$ representing the EN for the 2 atoms in the bond

Polar Covalent: unequal sharing of $e^-$ between two nonmetals (e.g., HF and CO).

Electronegativity difference ($\Delta \text{EN}$) is small.

Ionic: metal + nonmetal, $e^-$ are transferred from the metal to the nonmetal (e.g., $\text{NaCl}$).

Electronegativity difference ($\Delta \text{EN}$) is large.

Bond Polarity: Greater $\Delta \text{EN} \rightarrow$ the more polar, the more ionic, and the less covalent the bond is.

$\Delta \text{EN} \uparrow \Rightarrow \text{bond polarity} \uparrow; \text{bond ionic character} \uparrow; \text{bond covalent character} \downarrow$

Polar bond: A bond in which $e^-$ are pulled to one side of the molecule with a slight positive charge residing on one side ($\delta^+$) and a slight negative charge ($\delta^-$) residing on the other side of the molecule. The arrow points toward the negative side of the bond. ($\rightarrow$).

A dipole moment is a measure of bond polarity and occurs when there is charge separation.

Example 1: Which bond is most polar? H–F, N–F, or O–F

Answer 1: H–F bond is most polar. {The greater the $\Delta \text{EN}$ the more polar (or less covalent or more ionic) the bond is. Larger $\Delta \text{EN}$ values occur when the atoms of the bond are further apart on the Periodic Table. H is a “tricky” element because its EN (2.1) value lands between the values of B (1.9) and C (2.4); hence, think of H being placed between B and C as far as its EN value is concerned.}
LEWIS DOT STRUCTURES (LDS): Use valence e\(^-\) to create covalent bonding pictures (this is one method; there are other methods that can be used to build Lewis dot structures).

1. Count valence e\(^-\) in molecule. Add valence e\(^-\) from each atom, only s and p electrons past the last noble gas; do not include d-electrons.
2. Draw a skeleton structure; keep symmetric if possible. Usually (not always) the central atom in the Lewis dot structure is the first atom in the chemical formula.
3. Build octets (8 e\(^-\)) around each atom except H (2 e\(^-\) for H).
4. Count #e\(^-\) drawn; if it equals the #valence e\(^-\) determined in Step 1, Lewis dot structure is done.
5. If #e\(^-\) count is too large: Go to Step 2 and re-draw skeleton structure, add 1 double bond, and continue with the steps. If #e\(^-\) count is still too large, repeat and add another double bond or triple bond.

If #e\(^-\) count is too small: Add e\(^-\) pairs to the central atom.

Other info: F, Cl, Br, and I can have only 1 bond unless they're the central atom; H can only have 1 bond.

LEWIS DOT STRUCTURES: Categories of compounds to draw LDS for (see the next page for examples):

1. "Normal" compounds: follows the octet rule; e.g., NCl\(_3\)
2. Charged species: add charge to #valence e\(^-\) for anions, e.g., for SO\(_4\)\(^{2-}\) add 2 to valence e\(^-\) count
   \((1(6) + 4(6) + 2 = 32e^-)\); subtract charge from #valence e\(^-\) for cations, e.g., NH\(_4^+\) (1(5) + 4(1) - 1 = 8e^-)
3. Contains multiple bonds; i.e., double or triple bonds; e.g., O\(_2\), N\(_2\), etc.
4. Resonance structures: typically, a double bond that moves to an equivalent location, e.g., CO\(_3\)\(^{2-}\), C\(_6\)H\(_6\)
5. Central atom greater than 8e\(^-\): central atom must be in 3rd period or later of Periodic Table, e.g., SF\(_6\)
6. Central atom with less than 8e\(^-\); often B and Be compounds; e.g., BF\(_3\); in this example B has 6e\(^-\); B and Be usually do have 8e\(^-\) (e.g., BF\(_4^-\)) but they can have less than 8
7. Oxyacids = H\(_n\)XO\(_m\): skeleton structure: X in middle; H’s bonded to O’s, e.g., H\(_2\)SO\(_4\), HNO\(_3\)
8. Individual atoms
9. Organic molecules (e.g., CH\(_4\), C\(_2\)H\(_2\), C\(_2\)H\(_4\), C\(_3\)H\(_6\), C\(_3\)H\(_8\), CH\(_3\)COOH; H\(_2\)CO; NH\(_2\)(CO)NH\(_2\), etc.)
10. Radicals: have an odd #e\(^-\), e.g., NO (1(5) +1(6) = 11e^-)

RESONANCE: Sometimes multiple bonds can be placed in several, often equivalent, locations. If this occurs, you must draw all structures (called resonance structures) for the complete and correct LDS (e.g., there are two LDS for benzene, C\(_6\)H\(_6\) and both must be drawn).

BOND ORDER (BO):

Single bond = sharing of 1 e\(^-\) pair; BO = 1
Double bond = sharing of 2 e\(^-\) pairs; BO = 2
Triple bond = sharing of 3 e\(^-\) pairs; BO = 3

BO with resonance structures (one method): Bond Order = \#bonds / \#locations bonds are distributed

BOND ORDER↑ or Atomic Radii↓ ⇒ bond strength↑, bond length↓, bond dissociation energy↑
### Step-by-Step Examples: Lewis dot structure (LDS) (organic molecules not shown)

<table>
<thead>
<tr>
<th><strong>Category</strong></th>
<th><strong>Step 1: count valence electrons</strong></th>
<th><strong>Step 2: draw skeleton structure</strong></th>
<th><strong>Step 3: build octets</strong></th>
<th><strong>Step 4: check LDS for #e^–</strong></th>
<th><strong>ANSWERS</strong></th>
</tr>
</thead>
<tbody>
<tr>
<td>1. NCl&lt;sub&gt;3&lt;/sub&gt; (*&quot;normal&quot;)</td>
<td>1(N) + 3(Cl) = 1(5) + 3(7) = <strong>26e^–</strong></td>
<td><img src="image" alt="NCl3 skeletal structure" /></td>
<td><img src="image" alt="NCl3 octet structure" /></td>
<td><strong>26e^– → done! ✓</strong></td>
<td><img src="image" alt="NCl3 LDS" /></td>
</tr>
<tr>
<td>2. SO&lt;sub&gt;4&lt;/sub&gt;&lt;sup&gt;2-&lt;/sup&gt; (charged)</td>
<td>1(S) + 4(O) + 2 = 1(6) + 4(6) + 2 = <strong>32e^–</strong></td>
<td><img src="image" alt="SO42- skeletal structure" /></td>
<td><img src="image" alt="SO42- octet structure" /></td>
<td><strong>32e^– → done! ✓</strong></td>
<td><img src="image" alt="SO42- LDS" /></td>
</tr>
<tr>
<td>3. N&lt;sub&gt;2&lt;/sub&gt; (multiple bond)</td>
<td>2(N) = 2(5) = <strong>10e^–</strong></td>
<td><img src="image" alt="N2 skeletal structure" /></td>
<td><img src="image" alt="N2 octet structure" /></td>
<td><strong>10e^– → done! ✓</strong></td>
<td><img src="image" alt="N2 LDS" /></td>
</tr>
<tr>
<td>4. O&lt;sub&gt;3&lt;/sub&gt; (resonance)</td>
<td>3(O) = 3(6) = <strong>18e^–</strong></td>
<td><img src="image" alt="O3 skeletal structure" /></td>
<td><img src="image" alt="O3 octet structure" /></td>
<td><strong>18e^– → done! ✓</strong></td>
<td><img src="image" alt="O3 LDS" /></td>
</tr>
<tr>
<td>5. SF&lt;sub&gt;4&lt;/sub&gt; (more than an octet)</td>
<td>1(S) + 4(F) = 1(6) + 4(7) = <strong>34e^–</strong></td>
<td><img src="image" alt="SF4 skeletal structure" /></td>
<td><img src="image" alt="SF4 octet structure" /></td>
<td><strong>34e^– → done! ✓</strong></td>
<td><img src="image" alt="SF4 LDS" /></td>
</tr>
<tr>
<td>6. BCl&lt;sub&gt;3&lt;/sub&gt; (less than an octet)</td>
<td>1(B) + 3(Cl) = 1(3) + 3(7) = <strong>24e^–</strong></td>
<td><img src="image" alt="BCl3 skeletal structure" /></td>
<td><img src="image" alt="BCl3 octet structure" /></td>
<td><strong>26e^– → too many e^– → add a multiple bond</strong></td>
<td><img src="image" alt="BCl3 LDS" /></td>
</tr>
<tr>
<td>7. H&lt;sub&gt;3&lt;/sub&gt;PO&lt;sub&gt;4&lt;/sub&gt; (oxyacid)</td>
<td>1(P) + 4(O) + 3(H) = 1(5) + 4(6) + 3(1) = <strong>32e^–</strong></td>
<td><img src="image" alt="H3PO4 skeletal structure" /></td>
<td><img src="image" alt="H3PO4 octet structure" /></td>
<td><strong>32e^– → done! ✓</strong></td>
<td><img src="image" alt="H3PO4 LDS" /></td>
</tr>
<tr>
<td>8. N (atom)</td>
<td>1(N) = 1(5) = <strong>5e^–</strong></td>
<td><img src="image" alt="N skeletal structure" /></td>
<td><img src="image" alt="N octet structure" /></td>
<td><strong>5e^– → done! ✓</strong></td>
<td><img src="image" alt="N LDS" /></td>
</tr>
<tr>
<td>9. Organic molecules (not shown)</td>
<td><img src="image" alt="Organic molecules" /></td>
<td><img src="image" alt="Organic molecules" /></td>
<td><img src="image" alt="Organic molecules" /></td>
<td><img src="image" alt="Organic molecules" /></td>
<td><img src="image" alt="Organic molecules" /></td>
</tr>
</tbody>
</table>

**Notes:**
- **Step 1:** Count valence electrons.
- **Step 2:** Draw skeleton structure.
- **Step 3:** Build octets.
- **Step 4:** Check LDS.
- **ANSWERS:** Include charge and brackets when necessary.
- **Examples:** Sometimes one structure is better than another based on formal charge (FC) and electron negativity (EN).
FORMAL CHARGE (FC): The "charge" on an atom as determined from the LDS with bonded e⁻ shared equally

\[ \text{Formal Charge} = \#\text{valence e}^- \text{ (from the Periodic Table)} - (\#\text{e}^- \text{ around atom shown in LDS}) \]

\[ \Sigma \text{Formal Charges} = \text{Total Charge on Molecule} \]

**Example 2:** Assign formal charges to each atom in NCO⁻ (use this LDS): \[ \overset{\cdot}{:}N=\overset{\cdot}{C}=\overset{\cdot}{O}^- \]

**Answer 2:**

<table>
<thead>
<tr>
<th>Structure</th>
<th>FC (N):</th>
<th>FC (C):</th>
<th>FC (O):</th>
</tr>
</thead>
<tbody>
<tr>
<td>I</td>
<td>5 \cdot 6 = -1</td>
<td>4 \cdot 4 = 0</td>
<td>6 \cdot 6 = 0</td>
</tr>
<tr>
<td>II</td>
<td>5 \cdot 6 = -1</td>
<td>4 \cdot 4 = 0</td>
<td>6 \cdot 6 = 0</td>
</tr>
<tr>
<td>III</td>
<td>5 \cdot 6 = -1</td>
<td>4 \cdot 4 = 0</td>
<td>6 \cdot 6 = 0</td>
</tr>
<tr>
<td>IV</td>
<td>5 \cdot 6 = -1</td>
<td>4 \cdot 4 = 0</td>
<td>6 \cdot 6 = 0</td>
</tr>
</tbody>
</table>

Replace bonds with e⁻ to see #e⁻ around atom in LDS; e⁻ in bonds are split between adjoining atoms; lone pairs of e⁻ stay with atom.

**Lewis dot resonance structure that contributes the most based on better Formal Charges:**

1. Has the smaller (preferably zeros) formal charges.
2. Has more zeros for formal charges.
3. If there is a choice between 2 structure with same set of formal charges and same number of zeros, the more prevalent structure has the more negative formal charges on the more electronegative atom.

**Example 3:** Of the four ClO₃⁻ structures shown below which contributes the most based on formal charges? \( EN(Cl) = 2.7; EN(O) = 3.4 \)

**Answer 3: Step 1:** Assign FC to all atoms in all structures.

<table>
<thead>
<tr>
<th>Structure</th>
<th>FC (Cl):</th>
<th>FC (O):</th>
</tr>
</thead>
<tbody>
<tr>
<td>I</td>
<td>7 - 5 = +2</td>
<td>6 - 7 = -1</td>
</tr>
<tr>
<td>II</td>
<td>7 - 6 = +1</td>
<td>6 - 7 = -1</td>
</tr>
<tr>
<td>III</td>
<td>7 - 7 = 0</td>
<td>6 - 6 = 0</td>
</tr>
<tr>
<td>IV</td>
<td>7 - 8 = -1</td>
<td>6 - 7 = -1</td>
</tr>
</tbody>
</table>

**Step 2:** Eliminate structures based on Rule 1, then Rule 2, then Rule 3 above to find the structure that contributes the most based on better formal charges.

1. Eliminate Structure I because of the larger FC on the Cl (Rule 1: smaller FC are better). This is the only structure with a +2 FC.
2. Eliminate Structure II because it only has 1 zero FC while Structures III and IV each have 3 zeros for FC (Rule 2: more zeros are better).
3. Eliminate Structure IV because the more EN atom, O (EN(Cl) = 2.7 versus EN(O) = 3.4), should have the more negative FC (Rule 3).

4. Better structure based on FC.

**Rule 1:** smaller FC better.

**Rule 2:** more zeros for FC are better. Rule 3: the more EN atom gets the more negative FC. Structure III contributes more based on FC even though it violates the octet rule.
**BOND ENTHALPIES (D):** The energy needed to break or make a bond; break bonds: endothermic; form bonds: exothermic; can be used to estimate \( \Delta H_{\text{rxn}} \) (more accurate if chemicals are in the gas phase)

\[
\Delta_f H = \Sigma D_{\text{bonds broken}} - \Sigma D_{\text{bonds formed}}
\]

**Example 4:** What is the enthalpy of reaction of \( \text{CH}_4(g) \) combusting?

\[
\text{CH}_4(g) + 2\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(g)
\]

**Answer 4:**

**Step 1:** Draw LDS of each molecule to determine whether bonds are single, double, etc.

**Step 2:** Determine bonds broken (reactants) and formed (products); keep in mind stoichiometric coefficients

- **Bonds broken:** 4 C–H bonds; 2 O=O bonds (each \( \text{O}_2 \) molecule has 1 O=O bond and there are 2 \( \text{O}_2 \) molecules so there are 2 O=O bonds)
- **Bonds formed:** 2 C=O bonds; 4 O–H bonds (each \( \text{H}_2\text{O} \) has 2 O–H bonds and there are 2 \( \text{H}_2\text{O} \) molecules so there are 4 O–H bonds)

**Step 3:** Calculate \( \Delta H_{\text{rxn}} \)

\[
\Delta H_{\text{rxn}} = [4(D_{\text{C-H}}) + 2(D_{\text{O=O}})] - [2(D_{\text{C=O}}) + 4(D_{\text{O-H}})] ; \text{the needed bond enthalpy values (D) would be provided in a table for the next step}
\]

\[
\Delta H_{\text{rxn}} = [4(416) + 2(498)] - [(2(803) + 4(467))] = -814 \text{kJ/mol}
\]

1. Predict the order of **increasing** electronegativity in each of the following groups of elements:
   I. C, N, O
   II. S, Se, Cl
   III. Si, Ge, Sn
   IV. Tl, S, Ge

2. Based on the electronegativity differences, choose the molecule that is most likely to have **ionic bonds**?
   a. BaF\(_2\)
   b. Cl\(_2\)
   c. NH\(_3\)
   d. NO\(_3\)
   e. CH\(_4\)

3. In each set of bonds, determine which bond is the **most polar** bond?
   I. O—F, Cl—F, F—F, N—F, C—F
   II. Si—Cl, P—Cl, S—Cl
   III. Se—F, Se—Cl, Se—Br

4. (Practice for Lewis dot structures) Draw Lewis dot structures for each of the following molecules.
   a. PH\(_3\)
   b. CHCl\(_3\)
   c. NH\(_4^+\)
   d. BF\(_4^-\)
   e. SeF\(_2\)
   f. SO\(_4^{2-}\)
   g. ClO\(_4^-\)
   h. N\(_2\)
   i. HCN
   j. PF\(_5\)
   k. BrF\(_3\)
   l. BCl\(_3\)
   m. XeF\(_4\)
   n. SeF\(_4\)
   o. SO\(_3\)
   p. HNO\(_3\)
   q. SCN\(^-\)
   r. O\(_3\)
   s. N\(_3^-\)
   t. H\(_2\)SO\(_4\)
   u. PF\(_6^-\)
   v. N\(_2\)
   w. CO
   x. N\(_2\)H\(_2\)
   y. O\(_2\)
   z. CH\(_2\)CH\(_2\)
   aa. CO\(_3^{2-}\)
   bb. CH\(_3\)(CO)CH\(_3\)
   cc. NO
   dd. B

5. Which of the following chemicals **does not obey the octet rule**?
   a. CH\(_4\)
   b. SiH\(_4\)
   c. SeF\(_4\)
   d. BH\(_4^-\)
   e. IF\(_4^{+3}\)

6. Which molecule has the **strongest** carbon–oxygen bond?
   a. CO\(_3^{2-}\)
   b. CO
   c. CO\(_2\)
   d. H\(_2\)CO (C is central atom)
   e. They are all the same strength.
7. I. What is the formal charge on the S in SO$_3^{2-}$?
   a. -2  
   b. -1  
   c. 0  
   d. 1  
   e. 2

II. What is the formal charge on the N in NO$_2^-$?
   a. -2  
   b. -1  
   c. 0  
   d. 1  
   e. 2

8. Given the 3 resonance drawings of N$_2$O, which structure would contribute the most based on formal charges?
   I. [Image]
   II. [Image]
   III. [Image]
   a. I  
   b. II  
   c. III  
   d. I and III  
   e. All equally important.

9. Consider the Lewis structures of the following molecules. Which molecules list below would include resonance in their Lewis structures?
   PF$_5$  
   HNO$_3$  
   SO$_2$  
   F$_2$O
   a. PF$_5$, HNO$_3$  
   b. HNO$_3$, SO$_2$  
   c. SO$_2$, F$_2$O  
   d. PF$_5$, SO$_2$  
   e. HNO$_3$, F$_2$O

10. Which compound follows the octet rule?
    a. NO$_2$  
    b. ClO$_3$  
    c. NO  
    d. BF$_3$  
    e. None obey the octet rule.

11. Which of the following statements is incorrect?
    a. Ionic bonding results from the complete transfer of electrons from one atom to another.
    b. Dipoles are the result of unequal electron distribution in a molecule.
    c. The electrons in a polar bond are found nearer to the more electronegative element.
    d. A molecule with very polar bonds can be nonpolar.
    e. Linear molecules cannot have a net dipole moment.

12. Calculate the change in enthalpy, $\Delta H_{\text{rxn}}$, for the reaction shown given the bond enthalpies below. (Hint: Draw Lewis dot structures for each chemical in the reaction; note that the C atoms are bonded together in the product molecule.)

$\text{CH}_3$–$\text{CH}=\text{CH}_2$ + HBr $\rightarrow$ $\text{CH}_3$–$\text{CHBr}$–$\text{CH}_3$

<table>
<thead>
<tr>
<th>Bond</th>
<th>Bond enthalpy (kJ/mol)</th>
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<th>Bond</th>
<th>Bond enthalpy (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>C–C</td>
<td>356</td>
<td>Cl–Cl</td>
<td>242</td>
<td>C–H</td>
<td>416</td>
</tr>
<tr>
<td>C=C</td>
<td>598</td>
<td>H–Br</td>
<td>366</td>
<td>O–H</td>
<td>467</td>
</tr>
<tr>
<td>C–Br</td>
<td>813</td>
<td>H–H</td>
<td>436</td>
<td>C–Br</td>
<td>285</td>
</tr>
</tbody>
</table>

13. What is the bond enthalpy of a carbon-chlorine bond (C–Cl) in CHCl$_2$–CHCl$_2$ given the $\Delta H_{\text{rxn}}$ and bond enthalpies below? (Hint: Draw Lewis dot structures for each chemical in the reaction; note that the C atoms are bonded together in the product molecule.)

$\text{C}_2\text{H}_2 + 2\text{Cl}_2 \rightarrow \text{CHCl}_2$–$\text{CHCl}_2$  
$\Delta H_{\text{rxn}} = -367$kJ/mol

<table>
<thead>
<tr>
<th>Bond</th>
<th>Bond enthalpy (kJ/mol)</th>
<th>Bond</th>
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<td>813</td>
<td>H–H</td>
<td>436</td>
<td>C–Br</td>
<td></td>
</tr>
</tbody>
</table>

ANSWERS
1. I. C $<$ N $<$ O  
2. Se $<$ S $<$ Cl  
3. Sn $<$ Ge $<$ Si  
4. TI $<$ Ge $<$ S
2. a. the farther the atoms are apart from one another on the Periodic Table the greater the ΔEN; the greater the ΔEN the more ionic/polar the bond will be; Ba and F are farthest apart so they should have the most ionic bond; note that Ba is a metal and F is a nonmetal – the definition used earlier for ionic compound; all the other molecules contain 2 nonmetals – the definition used for molecular compounds.

3. Greatest ΔEN means greatest distance apart on the Periodic Table:
I. C—F II. Si—Cl III. Se—F

4. 

- a.
- b.
- c. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 
- d. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 
- e. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 
- f. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 
- g. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 
- h. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 
- i. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 
- j. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 
- k. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 
- l. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 
- m. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 
- n. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 
- o. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 
- p. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 
- q. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 
- r. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 
- s. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 
- t. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 
- u. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 
- v. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 
- w. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 
- x. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 
- y. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 
- z. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 

5. c. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \]

6. b. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 

- c. \[ \begin{align*} \text{H} & \text{H} \\ \text{H} & \text{H} \end{align*} \] 

greater BO means greater bond strength
7. I. d  
{Draw Lewis dot structure: } \[
\begin{array}{c}
\text{\textcircled{O} :} \\
\text{\textcircled{O} :} \\
\end{array}
\] \[^2 \; ; \text{FC} = \#\text{valence e}^- - \#\text{e}^- \text{ in Lewis dot; } S: 6 - 5 = 1 \]

II. c  
{Draw Lewis dot structure; } \[
\begin{array}{c}
\text{\textcircled{O} :} \\
\text{\textcircled{O} :} \\
\ \text{\textcircled{O} :} \\
\end{array}
\] \[1 \; \; ; \text{FC} = \#\text{valence e}^- - \#\text{e}^- \text{ in Lewis dot; } N: 5 - 5 = 0 \]

8. c  
{eliminate the second diagram because of the larger FC, -2: } \[
\begin{array}{c}
\text{\textcircled{O} :} \\
\text{\textcircled{O} :} \\
\ \text{\textcircled{O} :} \\
\end{array}
\] \[; \text{ between the first and third structures, since O is more EN than N, the O atom should have the more negative FC, hence, eliminate the first structure: }

{the third structure contributes the most because it has the better FC: } \[
\begin{array}{c}
\text{\textcircled{O} :} \\
\text{\textcircled{O} :} \\
\ \text{\textcircled{O} :} \\
\end{array}
\]

9. b  
{Note that the first 3 molecules are radicals (odd #e^-) and hence, cannot follow the octet rule; BF_3 is an exception with only } \[6e^- \text{ around the B; none follow the octet rule } \]

10. e  
{“e” is not correct; for instance, F–Cl is linear and has a dipole}

11. e  
\[\Delta H_{\text{rxn}} = \text{bonds broken} - \text{bonds formed} = [1(C=C) + 1(C–C) + 6(C–H) + 1(H–Br)] - [2(C–C) + 7(C–H) + 1(C–Br)] \]
\[= [1(598) + 1(356) + 6(416) + 1(366)] - [2(356) + 7(416) + 1(285)] = [3816] - [3909] = -93kJ/mol \]

13. 327kJ/mol  
\[\Delta H_{\text{rxn}} = [2(C–H) + 1(C=C) + 2(Cl–Cl)] - [(2(C–H) + 4(C–Cl) + 1(C–C)]; \]
\[x = 327kJ/mol \]