Questions 1 through 29 are multiple choice questions. Questions 30-34 are free response questions.

1. Which pair of atoms should form the most polar bond?
   (a) F and B  
   (b) C and O  
   (c) F and O  
   (d) N and F  
   (e) B and N

   Bond polarity increases with an increasing difference in electronegativity of the bonded atoms. F has the greatest electronegativity and electronegativity generally increases across a period - so B would have the smallest electronegativity. Electronegativity differences are greatest for elements that are farther apart on the periodic table.

2. Which pair of ions should form the ionic lattice with the highest energy?
   (a) Na⁺ and Br⁻  
   (b) Li⁺ and F⁻  
   (c) Cs⁺ and F⁻  
   (d) Li⁺ and O²⁻  
   (e) K⁺ and F⁻

   Lattice energy is largest for ions of small size and large charge. Charge generally has greater effect on lattice energy than does size. Lithium ion is the smallest cation listed and oxygen is small and has the largest negative charge.

3. The electrostatic force of attraction is greatest in which compound?
   (a) BaO  
   (b) MgO  
   (c) CaS  
   (d) MgS  
   (e) CaO

   A strong electrostatic force of attraction between two ions is dependent on a high charge and small size. All of the choices are compounds containing 2+ and 2- ions. But Mg²⁺ is the smallest cation and O²⁻ is the smallest anion.

4. Which molecule has the weakest bond?
   (a) CO  
   (b) O₂  
   (c) NO  
   (d) N₂  
   (e) Cl₂

   The strength of covalent bonds increases with increasing multiplicity (more bonds) and decreasing length. Triple bonds are stronger than double bonds which are stronger than single bonds. Cl₂ has a single bond. O₂ and NO have double bonds. CO and N₂ have triple bonds.

5. For which species are octet resonance structures necessary and sufficient to describe the bonding satisfactorily?
   (a) BCl₃  
   (b) SO₂  
   (c) CO₂  
   (d) BeF₂  
   (e) SO₄²⁻

   There is really only one structure for BCl₃ and BeF₂. SO₄⁻ has several resonance forms but they are identical - just rotate the structure by 90°. CO₂ has multiple forms [O≡C=O: and : O=C≡C] but they are truly valid and can be eliminated since the formal charges on the carbon atoms are +1 and -1. Leaving SO₂ as the correct choice. O=S=O and O-S=O and O=S-O.

6. How are the bonding pairs arranged in the best Lewis structure for ozone?
   (a) O−O−O  
   (b) O=O−O  
   (c) O≡O−O  
   (d) O≡O=O  
   (e) O=O=O

   Oxygen must obey the octet rule so the structure is O=O−O.

7. Which species has the shortest bond length?
   (a) CN⁻  
   (b) O₂  
   (c) SO₂  
   (d) SO₃  
   (e) CO₂

   Triple bonds are shorter and stronger than double bonds which are shorter and stronger than single bonds. The Lewis structure for CN⁻ has a triple bond. Only double bonds exist in O₂, SO₂, SO₃, and CO₂.

8. Which species requires the least amount of energy to remove an electron from the outermost energy level?
   (a) Na⁺  
   (b) Ne  
   (c) F⁻  
   (d) O²⁻  
   (e) Mg²⁺

   All five species are isoelectronic. Each of the species is shielded by two inner core electrons. Oxide has the fewest number of protons and the smallest effective nuclear charge. (The least protons are pulling on the most electrons.)
9. Which species has a valid non-octet Lewis structure?
(a) GeCl₄    (b) SiF₄    (c) NH₄⁺    (d) SeCl₄    (e) CO₃²⁻

10. How does the electronegativity (EN) of an element differ from its electron affinity (EA)?
(a) EN and EA are the same; they both measure the same characteristic.
(b) EA measures the energy released when an isolated atom/ion gains an electron; EN is a measure of the ability of an atom in a molecule to attract electrons to itself.
(c) EN values of the neutral atoms are just the negative of EA values of neutral atoms.
(d) EN measures the energy released when an isolated atom/ion gains an electron; EA is a measure of the ability of an atom in a molecule to attract electrons to itself.

11. Which of these molecules is not polar?
(a) H₂O    (b) CO₂    (c) NO₂    (d) SO₂    (e) NH₃

H₂O, NO₂, and SO₂ are all nonlinear and have a net dipole moment. NH₃ is trigonal pyramidal and each of the bond dipoles point toward the nitrogen. While CO₂ has bond dipoles, the molecule is linear and the bond dipoles point is opposite directions, thereby canceling each other out, leaving no net dipole moment.

12. Which species contains a central atom with sp² hybridization?
(a) C₂H₂    (b) SO₃²⁻    (c) O₃    (d) BrI₃    (e) NH₃

A trigonal planar electron-domain geometry exhibited by O₃ is characteristic of the sp² hybrid set. C₂H₂ is linear and sp hybridized. SO₃²⁻ is tetrahedral and sp³ hybridized. BrI₃ is trigonal bipyramidal and sp³d hybridized. NH₃ is trigonal pyramidal and an sp³ hybrid.

13. For ClF₃, the electron domain geometry of Cl and the molecular geometry are, respectively . . .
(a) trigonal planar and trigonal planar.
(b) trigonal planar and trigonal bipyramidal.
(c) trigonal bipyramidal and trigonal planar.
(d) trigonal bipyramidal and T-shaped.

14. The size of the H-N-H bond angles of the following species increases in which order?
(a) NH₃ < NH₄⁺ < NH₂⁻
(b) NH₃ < NH₂⁻ < NH₄⁺
(c) NH₂⁻ < NH₃ < NH₄⁺
(d) NH₂⁻ < NH₄⁺ < NH₃
(e) NH₄⁺ < NH₃ < NH₂⁻

All three species have a tetrahedral electron-domain geometry. NH₄⁺ has no lone pairs so the four bonded pairs are all equally far apart (109.5°). NH₃ has one lone pair (which repels more than the bonded pairs) which pushes the three bonded pairs slightly closer making the bond angle slightly smaller (107°). NH₂⁻ has two lone pairs (which repel even more) and push the remaining two bonded pairs even closer and make the bond angle even smaller (104.5°).
15. What is the molecular geometry and polarity of the BF₃ molecule?
   (a) trigonal pyramidal and polar  (b) trigonal pyramidal and nonpolar
   (c) trigonal planar and polar   (d) trigonal planar and nonpolar
   (e) T-shaped and polar

   BF₃ has an incomplete octet and does not satisfy the octet rule. It is a trigonal planar
   and has bond angles of 120°. The three bonds have bond dipoles but the bond dipoles are
   all oriented at the corners of an equilateral triangle and the bond dipoles cancel each
   other out, leaving the molecule with a net zero bond dipole moment.

16. In which species is the F-X-F bond angle the smallest?
   (a) NF₃   (b) BF₃   (c) CF₄   (d) BrF₃   (e) OF₂

   \[ \begin{array}{ccc}
   & F & \\
   & N & F \\
   F & & F \\
   \end{array} \quad \begin{array}{ccc}
   & F & \\
   & B & F \\
   F & & F \\
   \end{array} \quad \begin{array}{ccc}
   & F & \\
   & C & F \\
   F & & F \\
   \end{array} \quad \begin{array}{ccc}
   & F & \\
   & Br & F \\
   F & & F \\
   \end{array} \quad \begin{array}{ccc}
   & O & H \\
   & H & \\
   \end{array} \]

   \[ \sim 107° \quad 120° \quad 109.5° \quad 90° \quad \sim 104.5° \]

17. Which set does not contain a linear species?
   (a) CO₂, SO₂, NO₂   (b) H₂O, HCN, BeI₂
   (c) OCN⁻, C₂H₂, OF₂   (d) I⁻, BrF₃, SCN⁻
   (e) H₂S, ClO₂⁻, NH₂⁻

   The following are all linear: CO₂, HCN, OCN⁻, I⁻, and SCN⁻. (Note that HCN, OCN⁻, and
   SCN⁻ are isoelectronic species and have the same Lewis structure and geometry.)

18. The hybrid orbitals of nitrogen in N₂O₄ are . . .
   (a) sp   (b) sp²   (c) sp³   (d) dsp³   (e) d²sp³

   Each sp² hybrid set is characteristic of trigonal planar geometry.

19. How many sigma and how many pi bonds are in CH₂=CHCH₂C=OCH₃?
   (a) 5 sigma and 2 pi   (b) 8 sigma and 4 pi
   (c) 11 sigma and 2 pi   (d) 11 sigma and 4 pi
   (e) 13 sigma and 2 pi

   Each single bond is a sigma bond and each double bond consists of one sigma bond and
   one pi bond. There are 8 C-H sigma bonds, 4 C-C sigma bonds, and one C-O sigma bond.
   There is one C-C pi bond and one C-O pi bond.

20. What is the best estimate of the H-O-H bond angle in H₃O⁺?
   (a) 109.5°   (b) 107°   (c) 104.5°   (d) 116°   (e) 120°

   The electron domain geometry of hydronium is tetrahedral. However, from VSEPR, there
   is one lone pair that exerts a greater repulsive force than the bonded pairs of electrons.
So as the bonds are pushed slightly together, there is a small decrease in the bond angle. But not as much push as if there were two lone pairs. Hydronium is isoelectronic with ammonia.

21. The bond between carbon and hydrogen is one of the most important types of bonds in chemistry. The length of a H-C bond is approximately 1.1 Å. Based on this distance and the differences in electronegativity, do you expect the dipole moment of an individual H-C bond to be larger or smaller than that of the H-I bond?
   (a) H-C has the same bond dipole moment as H-I, because the magnitude of \( Q \) and \( r \) in \( \mu = Qr \) are about the same for both.
   (b) H-C has the larger bond dipole moment, because \( Q \) and \( r \) are larger in H-C than in H-I.
   (c) H-I has the smaller dipole moment, because \( Q \) is about the same for both, but \( r \) is larger in H-C than in H-I.
   (d) H-I has the larger dipole moment because \( Q \) is about the same for both, but \( r \) is larger in H-I than in H-C.

22. Suppose a Lewis structure for a neutral fluorine-containing molecule results in a formal charge on the fluorine atom of +1. What conclusion can you draw?
   (a) The structure actually represents an ion.
   (b) The F atom in the structure must have four covalent bonds attached to it.
   (c) There must be another F atom in the structure carrying a −1 formal charge, since F is the most electronegative element and it should a negative formal charge.
   (d) There must be a better Lewis structure, since F is the most electronegative element and it should a negative formal charge.

23. In the same sense that the O-O bonds in ozone are described as “one-and-a-half” bonds, how would you describe the N-O bonds in NO\(_3^-\)?
   (a) \( 1 \frac{1}{5} \)
   (b) \( 1\frac{1}{4} \)
   (c) \( 1\frac{1}{3} \)
   (d) \( 1\frac{1}{2} \)

24. Which of the Lewis structures for NO (shown right) is dominant based on analysis of the formal charge.
   (a) Structure I, because all atoms have zero formal charge.
   (b) Structure I, because N and O have equal but opposite formal charges.
   (c) Structure II, because all atoms have zero formal charge.
   (d) Structure II, because N should not have a formal charge.

25. Based on bond enthalpies, which do you expect to be more reactive, oxygen, O\(_2\), or hydrogen peroxide, H\(_2\)O\(_2\)?
   (a) O\(_2\) is more reactive, because the O=O bond enthalpy is less than that of the O-O bond enthalpy in H\(_2\)O\(_2\).
   (b) O\(_2\) is more reactive, because the O=O bond enthalpy is greater than that of the O-O bond enthalpy in H\(_2\)O\(_2\).
   (c) H\(_2\)O\(_2\) is more reactive, because the O-O enthalpy is less than that of the O=O bond enthalpy in O\(_2\).
   (d) H\(_2\)O\(_2\) is more reactive, because the O-O bond enthalpy is greater than that of the O=O bond enthalpy in O\(_2\).

26. Using average bond energies, calculate \( \Delta H_{\text{rxn}} \) for the combustion of acetylene.

\[
2 \ H \equiv C \equiv C \equiv H \ + \ 5 \ O=O \ \rightarrow \ 4 \ O=O \equiv C \equiv O \ + \ 2 \ H \equiv O \equiv H
\]

   (a) +1359 kJ 
   (b) −1359 kJ 
   (c) +2428 kJ 
   (d) −2428 kJ 
   (e) −510 kJ
27. It might seem that a square-planar geometry of four electron domains around a central atom would be more favorable than a tetrahedral arrangement. Rationalize why the tetrahedron is preferred, based on the angles between electron domains.

(a) Bond angles are not determined by a particular arrangement of electron domains.
(b) A tetrahedral arrangement of electron domains results in greater electron repulsions and a less favorable geometry than electron domains in a square planar geometry.
(c) A tetrahedral arrangement of electron domains results in smaller electron repulsions and a more favorable geometry than electron domains in a square planar geometry.
(d) A tetrahedral arrangement of electron domains results in no greater electron repulsions and no greater favorable geometry than electron domains in a square planar geometry.

28. In an $sp^2$ hybridized atom, what is the orientation of the unhybridized $p$ atomic orbital relative to the three $sp^2$ hybridized orbitals?

(a) The unhybridized $p$ orbital is 180° from the plane of the $sp^2$ orbitals.
(b) The unhybridized $p$ orbital is coplanar with the plane of the $sp^2$ orbitals.
(c) The unhybridized $p$ orbital is 109.5° from the plane of the $sp^2$ orbitals.
(d) The unhybridized $p$ orbital is perpendicular to the plane of the $sp^2$ orbitals.

29. The molecule called diazine has the formula N$_2$H$_2$ and a structure shown to the right. Do you expect diazine to be a linear molecule (all four atoms on the same line)? If not, do you expect the molecule to be planar (all four atoms in the same plane)?

(a) The molecule is both linear and planar.
(b) The molecule is not linear, but it is planar.
(c) The molecule is linear, but it is not planar.
(d) The molecule is neither linear nor planar.

30. Consider the following chemical species: the nitrogen molecule, the nitrite ion, and the nitrate ion.

a. Write the chemical formulas for each of the species and identify the oxidation number of the nitrogen atom in each formula.

   - nitrogen molecule = N$_2$, nitrite ion = NO$_2^-$, and nitrate ion = NO$_3^-$

b. Draw Lewis structure for each of the species. Where appropriate, draw resonance structures for each.

   - nitrogen molecule: \( \cdot\text{N} ≡ \text{N} \cdot \)
   - nitrite ion: \[ \begin{array}{c}
   \text{N} \\
   \text{O} \\
   \text{O}
   \end{array} \] 
   - nitrate ion: \[ \begin{array}{c}
   \text{N} \\
   \text{O} \\
   \text{O}
   \end{array} \]

   Multiple bonds are shorter than single bonds. Nitrogen has a triple bond - the shortest of all. Nitrite has 2 bonds (one double and one single) and each is considered equal - so each has a bond strength equivalent to 1.5 bonds. Nitrate has 3 bonds (one double and one single) and each is considered equal - so each has a bond strength equivalent to 1.333 bonds. The double bond character is greater in the nitrite ion than it is in the nitrate ion - so the nitrite ion has a shorter bond length and the nitrate ion has the longest bond length.

d. Write a balanced net ionic equation for the reaction of nitrogen dioxide with water. Comment on the molecular and/or ionic species that are formed.

   \[ 2 \text{NO}_2 (g) + \text{H}_2\text{O} (l) \rightarrow \text{H}^+ (aq) + \text{NO}_3^- (aq) + \text{HNO}_2 (aq) \]

   Nitric acid is a strong acid and ionizes completely. Nitrous acid is weak, so it is written in its molecular form.
e. Draw each of the resonance structures for nitrogen monoxide and assign formal charges to each structure.

\[ \text{NO} \rightarrow \begin{array}{c}
\text{:N} \rightarrow \text{O} \\
\text{+1} \rightarrow -1
\end{array} \]

31. Carbon dioxide gas is bubbled into water.
   a. Write and balance a chemical equation to describe the process.
      \[ \text{CO}_2 (g) + \text{H}_2\text{O} (l) \rightarrow \text{H}_2\text{CO}_3 (aq) \]
      It’s a weak acid so it is written in molecular form.
   b. Draw the Lewis structures of the reactants and products. Include any valid resonance structures.

32. Consider the chemical species IF\(_5\) and IF\(_4^+\).
   a. Draw the Lewis structure and make a rough three-dimensional sketch of each of these species.

   b. Identify the orbital hybridization, the electron domain geometry, and the molecular geometry of each species.
      IF\(_4^+\) is sp\(^3\)d hybridized; its electron domain geometry is trigonal bipyramidal; and its molecular geometry is seesaw.
      IF\(_5\) is sp\(^3\)d\(^2\) hybridized; its electron domain geometry is octahedral; and its molecular geometry is square pyramidal.
c. Identify the approximate bond angle of each species.
   \( \text{IF}_4^+ \): equatorial bond angle is slightly less than 120° due to a larger repulsive force exerted by the equatorial lone pair and the axial bond angle is 180°.
   \( \text{IF}_5^2- \): all bond angles are ~90°

   d. Predict which, if any, is a polar species. Justify your answer.
   Both species are polar. Each I-F bond is polar due to a large difference in electronegativity between the two atoms. The three-dimensional arrangement of bond dipoles in both species contribute to a net molecular dipole.

   e. Predict the most probable oxidation state of the iodine atom in each species. Give an example of another chemical species having the same oxidation number as \( \text{IF}_4^+ \).
   Both species have iodine with an oxidation state of +5. Fluorine is always -1 since it is the most negative, leaving iodine as a +5 in both species.

   f. Would you expect that the conversion from \( \text{IF}_5^2- \) to \( \text{IF}_4^+ + \text{F}^- \) to be exothermic or endothermic? Explain.
   The reaction \( \text{IF}_5^2- \rightarrow \text{IF}_4^+ + \text{F}^- \) is most likely endothermic. The reaction requires that an I-F bond be broken. Bond breaking is always endothermic.

33. Consider each of these molecules: \( \text{C}_3\text{H}_4 \), \( \text{C}_3\text{H}_6 \), and \( \text{C}_3\text{H}_8 \).
   a. Draw the Lewis structure for each molecule and identify the orbital hybridization of each carbon atom.
   \begin{align*}
   \text{CH}_3\text{C}≡\text{CH} & \quad \text{sp}^3, \text{sp}, \text{sp} \quad \text{sp}^3, \text{sp}^2, \text{sp}^2 \quad \text{all sp}^3 \\
   \text{CH}_3\text{CH}=\text{CH}_2 & \quad \text{sp}^3, \text{sp}, \text{sp} \quad \text{sp}^3, \text{sp}, \text{sp} \quad \text{all sp}^3 \\
   \text{CH}_3\text{CH}_2\text{CH}_3 & \quad \text{sp}^3, \text{sp}, \text{sp} \quad \text{sp}^3, \text{sp}, \text{sp} \quad \text{all sp}^3
   \end{align*}
   b. Specify the geometry of each central carbon atom.
   All sp hybridized carbons are linear, all sp\(^2\) carbons are trigonal planar, and all sp\(^3\) carbons are tetrahedral.
   c. Write a balanced chemical equation for the complete combustion of each molecule.
   \( \text{C}_3\text{H}_4 \) (g) + 4 \( \text{O}_2 \) (g) \( \rightarrow \) 3 \( \text{CO}_2 \) (g) + 2 \( \text{H}_2\text{O} \) (l)
   2 \( \text{C}_3\text{H}_6 \) (g) + 9 \( \text{O}_2 \) (g) \( \rightarrow \) 6 \( \text{CO}_2 \) (g) + 6 \( \text{H}_2\text{O} \) (l) or \( \text{C}_3\text{H}_6 + 9/2 \text{O}_2 \rightarrow \text{3 CO}_2 + 3 \text{H}_2\text{O} \)
   \( \text{C}_3\text{H}_8 \) (g) + 5 \( \text{O}_2 \) (g) \( \rightarrow \) 3 \( \text{CO}_2 \) (g) + 4 \( \text{H}_2\text{O} \) (l)
   d. Use the following bond enthalpies to determine the heat of combustion of each molecule. Specify your answer in \( \text{kJ/mol} \).

   \begin{tabular}{|c|c|c|c|c|c|c|c|}
   \hline
   Bond & C-H & C-C & C≡C & C=O & O-H & O_2 & C=O & C-O \\
   Bond Enthalpy (kJ/mol) & 413 & 384 & 614 & 839 & 463 & 495 & 799 & 358 \\
   \hline
   \end{tabular}

   \[ \Delta H_{\text{comb}} = \Sigma \Delta H_{\text{BDE}} \text{ for bonds broken } + \Sigma \Delta H_{\text{BDE}} \text{ for bonds formed} \]

   for \( \text{C}_3\text{H}_4 \):
   \[ \Delta H_{\text{comb}} = 4 \times (495 \text{kJ}) + 4 \times (143 \text{kJ}) + 839 \text{kJ} + 348 \text{kJ} + 6 \times (-799 \text{kJ}) + 4 \times (-463 \text{kJ}) = -1827 \text{ kJ} \]

   for \( \text{C}_3\text{H}_6 \):
   \[ \Delta H_{\text{comb}} = 9/2 \times (495 \text{kJ}) + 6 \times (143 \text{kJ}) + 614 \text{kJ} + 348 \text{kJ} + 6 \times (-799 \text{kJ}) + 6 \times (-463 \text{kJ}) = -1905 \text{ kJ} \]
   This is \( \Delta H_{\text{comb}} \) which is for the combustion of a single mole of hydrocarbon.

   for \( \text{C}_3\text{H}_8 \):
   \[ \Delta H_{\text{comb}} = 5 \times (495 \text{kJ}) + 8 \times (143 \text{kJ}) + 2 \times (348 \text{kJ}) + 6 \times (-799 \text{kJ}) + 8 \times (-463 \text{kJ}) = -2023 \text{ kJ} \]
34. Answer the following questions using the data in the table to the right.

   a. Which binary compound in the table has the highest lattice energy? Explain using Coulomb’s law.

   **Highest lattice energy = ScN.** Energy between charged particles is directly proportional to the charges and Sc is a 3+ and nitride is a 3−; these are the largest charges.

   b. Rank the compounds having the three highest lattice energies, highest to lowest, in column 1 of the table. Use Coulomb’s law to explain the data and ranking.

   **LiF > NaF > KF.** Energy between charges particles is inversely proportional to the distance between them. Li+ is the smallest ion and F− will be closest to it while K+ is the largest so F− will be farthest away.

   c. Predict the approximate lattice energy of RbF. Explain your reasoning. Predict whether CsF will have a larger or smaller lattice energy than RbF. Explain.

   **RbF will have a lattice energy of about 700 kJ/mol, smaller than KF since the Rb+ ion is larger than K+ and the fluoride ion will be at a larger distance so there will be a smaller energy. CsF will have an even smaller value of lattice energy since Cs+ is even larger than Rb+ making the distance even greater and the energy even smaller.**

   d. Which ionic compound in the table has the least ionic character? Explain.

   **CsI is the least polar compound. CsI has the lowest lattice energy and both Cs and I have relatively small and close electronegativities.**

   e. Rank the four sodium compounds in the table according to decreasing lattice energy. Explain the ranking using Coulomb’s law.

   **NaF > NaCl > NaBr > NaI.** The lattice energy ranking follows the inverse order of the size of the anions. Coulomb’s law states that the closer the ions the larger the energy. Fluoride has the smallest size and will be closest to the sodium while I has the largest size and will be farthest from the sodium.