1. Define Lewis symbol and what do they specifically show about elements?

A Lewis symbol is an element's symbol that has the s and p block electrons drawn in as dots. d and f can also be drawn but is not common.

2. Define octet rule

Any s and p block element wants to gain or lose enough e- to get to a noble gas configuration which is a more stable energy configuration.

3. Define ionic bonding.

The transfer of electrons between two elements usually a metal and nonmetal forming a cation and anion.

4. Describe the movement of electrons when ionic bonds are formed.

The e- are completely transferred between two nuclei.

5. Define lattice energy.

The energy required to completely separate one mole of a solid ionic compound into its gaseous ions ex: \( NaCl(s) \rightarrow Na^+(g) + Cl^-(g) \)

6. Is the process of separating ions endothermic or exothermic?

Endothermic

7. What is the equation for the electrostatic potential energy of two interacting charged particles?

\[ E_{el} = \frac{kQ_1Q_2}{d} \]

Where \( Q_1 \) and \( Q_2 \) are the charges, \( d \) is the distance between their centers in meters.

\[ k = \frac{1}{4\pi\varepsilon_0} \approx 8.99 \times 10^9 \text{ N m}^2\text{C}^{-2} \]

8. Without consulting table 8.2, arrange the ionic compounds NaF, CsI, and CaO in order of increasing lattice energy.

\( CaO > NaF > CsI \)

9. What is \( k \) from the previous equation represent?

\[ k = \frac{\varepsilon_0}{4\pi} \frac{J}{m^2} \]
### Table 8.2 Lattice Energies for Some Ionic Compounds

<table>
<thead>
<tr>
<th>Compound</th>
<th>Lattice Energy (kJ/mol)</th>
<th>Compound</th>
<th>Lattice Energy (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>LiF</td>
<td>1030</td>
<td>MgCl₂</td>
<td>2326</td>
</tr>
<tr>
<td>LiCl</td>
<td>834</td>
<td>SrCl₂</td>
<td>2127</td>
</tr>
<tr>
<td>LiI</td>
<td>730</td>
<td>MgO</td>
<td>3795</td>
</tr>
<tr>
<td>NaF</td>
<td>910</td>
<td>CaO</td>
<td>3414</td>
</tr>
<tr>
<td>NaCl</td>
<td>788</td>
<td>SrO</td>
<td>3217</td>
</tr>
<tr>
<td>NaBr</td>
<td>732</td>
<td>ScN</td>
<td>7547</td>
</tr>
<tr>
<td>NaI</td>
<td>682</td>
<td>KF</td>
<td></td>
</tr>
<tr>
<td>KCl</td>
<td>701</td>
<td>KCl</td>
<td></td>
</tr>
<tr>
<td>KBBr</td>
<td>671</td>
<td>CaCl₂</td>
<td></td>
</tr>
<tr>
<td>CaI</td>
<td>657</td>
<td>CaI</td>
<td></td>
</tr>
</tbody>
</table>

10. How do s-block ions form?
   
   They lose 1-2 e⁻

11. How do f-block ions form?
   
   They gain e⁻

12. Predict the ions generally formed by Sr, S, and Al.
   
   \[
   +2 \quad -2 \quad +3
   \]

13. How do transition metals form ions?
   
   They lose e⁻ in a fashion that either half empties an orbital or completely empties one orbital. s-block e⁻ can be moved to accomplish this.

14. How are covalent bonds formed?
   
   e⁻ are shared between 2 nuclei.

15. What are Lewis Structures (Lewis Dot Structures)?
   
   Lewis dot shows the valence e⁻ of 2 or more interacting nuclei and how the centers (nuclai) are moving e⁻ to form stable e⁻ configuration.

   
   Nonbonding electrons that fill the octet but do not interact with other nuclei. They still cause repulsion of other lone pairs and bonds.

17. Define nonbonding pairs.
   
   They are also known as lone pairs.

18. Predict the Lewis Structure of when fluorine and nitrogen bond and draw the lewis structure.
   
   \[
   \overset{5}{\text{F}} \quad \overset{3}{\text{N}} \quad : \text{F} \equiv \text{N}: 
   \]

\[ \text{\underline{single}} \quad \text{\underline{double}} \quad \text{\underline{triple}} \]

20. Why are double and triple bonds necessary?

\[ \text{to get all } 8e^- \text{ needed for octet} \]

21. Define bond polarity

\[ \text{the measure of how equally or unequally the } e^- \text{ are being shared in any covalent bond} \]

22. Define nonpolar covalent bond

\[ \text{the } e^- \text{ are shared equally} \]

23. Define polar covalent bond

\[ \text{our nuclei exerts more pull on the shared } e^- \text{ than the other causing unequal sharing} \]

24. Define electronegativity and electron affinity

\[ \text{the ability of an atom to attract } e^- \text{ to itself} \quad \text{the energy change that occurs when an } e^- \text{ is added to a neutral atom or ion} \]

25. What are the approximate ranges for nonpolar, polar, and ionic bonds.

\[ 0 \text{ to } 0.4 \quad 0.5 \text{ to } 1.5 \quad >1.5 \]

26. What does }^\dagger_\ddagger \text{ mean?

\[ \text{partial positive } \quad \text{partial negative} \]

27. Draw what the three different bonds would look like from an electron perspective.

\[ \text{nonpolar} \quad \text{polar} \quad \text{ionic} \]

28. In each case, which bond is more polar? (a) B–Cl or C–Cl, (b) P–F or P–Cl. Indicate in each case which atom has the partial negative charge.

\[ \text{Cl} \quad \text{F} \]

29. What is a polar molecule and nonpolar molecule?

\[ \text{a molecule that has a dipole moment} \quad \text{no net dipole moment due to polar bonds being present} \]

30. Define dipole.

\[ \text{two electrical charges of equal magnitude but opposite sign are separated by a distance} \]

31. Define dipole moment and give the equation for it.

\[ \mu = Qr \quad \text{the product of } Q \text{ and } r \text{ gives the magnitude of the dipole moment} \]
32. The bond length in the HCl molecule is 1.27 Å. (a) Calculate the dipole moment, in debyes, that results if the charges on the H and Cl atoms were 1+ and 1- respectively. (b) The experimentally measured dipole moment of HCl(g) is 1.08 D. What magnitude of charge, in units of e, on the H and Cl atoms leads to this dipole moment?

A) \( e = 1.60 \times 10^{-19} \text{ C} \) \( r = 1.27 \text{ Å} \) \( \mu = Qr = \frac{(1.08 \times 10^{-10} \text{ C})}{(1.60 \times 10^{-19} \text{ C})} \times (1.27 \times 10^{-10} \text{ m}) = 0.68 \text{ D} \)

B) \( Q = \frac{2.84 \times 10^{-20} \text{ C}}{2.17 \times 10^{-10} \text{ m}} = 0.13 \text{ e} \)

33. Write the procedure for drawing Lewis Structures.
1) Sum the valence e- from all atoms, taking into account overall charge
2) Write the symbols for the atoms and show which are connected
3) Complete the octet around all atoms bonded to central atom
4) Place the rest of e- on central atom
5) If there are not enough e- to give central atom try multiple bonds

34. Draw the Lewis structure of phosphorus trichloride.

\[ \vdots \leftarrow P \rightarrow \vdots \]

35. Draw the Lewis structure for HCN

\[ \text{H} - \text{C} = \text{N} \]

36. Draw the Lewis structure of bromate ion.

\[ \text{BrO}_3^{-} \]

\[ \left[ \begin{array}{c} \text{Br}^- \end{array} \right] \]

37. Define formal charge and give the equation for formal charge.

The charge the atom would have if each bonding e- pair in the molecule were shared equally between 2 atoms.

\( F.C. = \text{valence e} - \frac{1}{2} (\text{bonding e} - \text{shared e}) \)

38. Three possible Lewis structures for the thiocyanate ion are

\[ [\text{CN}]=\text{S} \]

a) Determine the formal charges in each structure

b) Based on the formal charges, which Lewis structure is the dominate one?

2) The Lewis structure in which any negative charges reside on the most electronegative atoms generally more dominant

39. Define resonance structure.

In cases where 2 or more Lewis structures are equally good descriptions of a single molecule, the R.S. are “averaged” to give a more accurate description of the molecule.

40. The O-O bonds in ozone are often described as “one and a half” bonds. Is this description consistent with the idea of resonance?

\[ \sqrt{2} \] \( O=\ddot{O}=-\dot{O} \) or \( \dot{O}=-\ddot{O}=\ddot{O} \)

\[ \text{drawn} \rightarrow O=O=O \]
41. Which is predicted to have the shorter sulfur-oxygen bonds SO₃ or SO₃²⁻?

\[ \text{SO}_3 \quad \overset{\text{short}}{\leftrightarrow} \quad \text{SO}_3^2^- \]

* double bonds are shorter than single

42. Which of these statements about resonance is true? (a) When you draw resonance structures, it is permissible to alter the way atoms are connected. (b) The nitrate ion has one long N-O bond and two short N-O bonds. (c) "Resonance" refers to the idea that molecules are resonating rapidly between different bonding patterns. (d) The cyanide ion has only one dominant resonance structure. (e) All of the above are true.

A) F
B) F they are anomol
C) T
D) T H-C=N:
E) F

43. How is resonance in benzene special?

\( \text{Benzene has 3 bonds that can "re-rule" positions} \)

44. What are the 3 expectations to the octet rule?
1) Molecules and polyatomic ions containing an odd number of e⁻.
2) Molecules and polyatomic ions in which an atom has fewer than an octet of valence e⁻.
3) Molecules and polyatomic ions in which an atom has more than an octet of valence e⁻.

45. Define hypervalent

molecules and ions with more than an octet of electrons around the central atom.

46. Draw the Lewis Structure for ICl₄⁻:

\[ \text{I} \quad \text{Cl} \quad \text{Cl} \quad \text{Cl} \quad \text{I} \]

47. Define bond enthalpy for covalent bonds

bond enthalpy change, \( \Delta H \), for the breaking of a particular bond in one mole of a gas

48. Define atomization

a process in which a molecule is decomposed into its atoms.

\[ \text{C}_2\text{H}_4(g) \rightarrow \text{C} \quad (g) + 2 \text{H}_2(g) \quad \Delta H = 1080 \text{ kJ} \]
49. How is the $\Delta H_{\text{rxn}}$ calculated using the table above?

$$\Delta H_{\text{rxn}} = \sum \text{(bond enthalpies of bonds broken)} - \sum \text{(bond enthalpies of bonds formed)}$$

Don't forget to keep track of coefficients.

50. How are bond enthalpy and bond length related?

As bond length $\downarrow$, bond enthalpy $\uparrow$

<table>
<thead>
<tr>
<th>Bond</th>
<th>Bond Length (Å)</th>
<th>Bond</th>
<th>Bond Length (Å)</th>
</tr>
</thead>
<tbody>
<tr>
<td>C—H</td>
<td>1.13</td>
<td>N—N</td>
<td>1.47</td>
</tr>
<tr>
<td>C—C</td>
<td>1.54</td>
<td>N=N</td>
<td>1.24</td>
</tr>
<tr>
<td>C=N</td>
<td>1.34</td>
<td>N=O</td>
<td>1.10</td>
</tr>
<tr>
<td>C—N</td>
<td>1.20</td>
<td>N=O</td>
<td>1.36</td>
</tr>
<tr>
<td>C≡N</td>
<td>1.38</td>
<td>N≡O</td>
<td>1.22</td>
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<tr>
<td>C≡N</td>
<td>1.16</td>
<td>O=O</td>
<td>1.48</td>
</tr>
<tr>
<td>C≡O</td>
<td>1.43</td>
<td>O=O</td>
<td>1.21</td>
</tr>
<tr>
<td>C≡O</td>
<td>1.23</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
51. Phosgene, a substance used in poisonous gas warfare during World War I, is so named because it was first prepared by the action of sunlight on a mixture of carbon monoxide and chlorine gases. Its name comes from the Greek words *phos* (light) and *genes* (born of). Phosgene has the following elemental composition: 12.14% C, 16.17% O, and 71.69% Cl by mass. Its molar mass is 98.9 g/mol. (a) Determine the molecular formula of this compound. (b) Draw three Lewis structures for the molecule that satisfy the octet rule for each atom. (The Cl and O atoms bond to C.) (c) Using formal charges, determine which Lewis structure is the dominant one. (d) Using average bond enthalpies, estimate $\Delta H$ for the formation of gaseous phosgene from CO(g) and Cl₂(g).

\[ \text{COCl} \]

52. (a) True or false: An element’s number of valence electrons is the same as its atomic number. (b) How many valence electrons does a nitrogen atom possess? (c) An atom has the electron configuration $1s^22s^22p^63s^23p^2$. How many valence electrons does the atom have?

A) 7  
B) 5

53. (a) True or false: The hydrogen atom is most stable when it has a full octet of electrons. (b) How many electrons must a sulfur atom gain to achieve an octet in its valence shell? (c) If an atom has the electron configuration $1s^22s^22p^3$, how many electrons must it gain to achieve an octet?

A) F  
B) 2

54. Consider the element silicon, Si. (a) Write its electron configuration. (b) How many valence electrons does a silicon atom have? (c) Which subshells hold the valence electrons?

A) $[Ne]3s^23p^2$  
B) 4

55. (a) Write the electron configuration for the element titanium, Ti. How many valence electrons does this atom possess? (b) Hafnium, Hf, is also found in group 4B. Write the electron configuration for Hf. (c) Ti and Hf behave as though they possess the same number of valence electrons. Which of the subshells in the electron configuration of Hf behave as valence orbitals? Which behave as core orbitals?

A) $[Ar]4s^23d^2$  
B) $[Xe]6s^25p^2$

56. Write the Lewis symbol for atoms of each of the following elements: (a) Al, (b) Br, (c) Ar, (d) Sr.

A) $\text{Al}^+$  
B) $\text{Br}^-$  
C) $\text{Ar}^-$  
D) $\text{Sr}^+$
57. (a) Using Lewis symbols, diagram the reaction between magnesium and oxygen atoms to give the ionic substance MgO. (b) How many electrons are transferred? (c) Which atom loses electrons in the reaction?

\[ \text{Mg}^{2+} + 2\text{O}^2- \rightarrow \text{MgO} \]

58. Predict the chemical formula of the ionic compound formed between the following pairs of elements: (a) Al and F, (b) K and S, (c) Y and O, (d) Mg and N.

\[ \text{A) AlF}_3 \quad \text{B) K}_2\text{S} \quad \text{C) Y}_2\text{O} \quad \text{D) Mg}_3\text{N}_2 \]

59. Write the electron configuration for each of the following ions, and determine which ones possess noble-gas configurations: (a) Sr\(^{2+}\), (b) Ti\(^{2+}\), (c) Se\(^{2-}\), (d) Ni\(^{2+}\), (e) Br\(^{+}\), (f) Mn\(^{3+}\).

\[ \text{A) [Ar]} \quad \text{B) [Ar]} \quad \text{C) [Ar]} \quad \text{D) [Ar]} \quad \text{E) [Ar]} \quad \text{F) [Ar]} \]

60. (a) Is lattice energy usually endothermic or exothermic? (b) Write the chemical equation that represents the process of lattice energy for the case of NaCl. (c) Would you expect salts like NaCl, which have singly-charged ions, to have larger or smaller lattice energies compared to salts like CaO which are composed of doubly-charged ions?

\[ \text{A) endothermic} \quad \text{B) NaCl(s) \rightarrow Na^+(g) + Cl^-(g)} \quad \text{C) CaO will have a higher lattice energy compared to the singly-charged ions.} \]

61. The ionic substances NaF, CaO, and ScN are isoelectronic (they have the same number of electrons). Examine the lattice energies for these substances in Table 8.2. Make a graph of lattice energy on the vertical axis versus the charge on the cation on the horizontal axis. (a) What is the slope of the line? (b) Make a graph of lattice energy on the vertical axis versus the square of the cation charge on the horizontal axis. What is the slope of this line? (c) Compare how well the data points fall on a line for the graphs in (a) and (b). Which trend is more linear, lattice energy versus cation charge or lattice energy versus cation charge squared? (d) Predict the lattice energy for the compound TiC, if we consider the carbon to have a 4- charge.

\[ \text{A) use excel} \quad \text{B) use excel} \quad \text{C) The data points are closer to the line of the charge squared and it is more linear than the other} \]

\[ \text{A) } 3.5 \times 10^{-3} \text{kJ/mole} \quad \text{B) } 8.2 \times 10^{-1} \text{kJ/mole} \]

62. Consider the ionic compounds KF, NaCl, NaBr, and LiCl. (a) Use ionic radii (Figure 7.8) to estimate the cation–anion distance for each compound. (b) Based on your answer to part (a), arrange these four compounds in order of decreasing lattice energy. (c) Check your predictions in part (b) with the experimental values of lattice energy from Table 8.2. Are the predictions from ionic radii correct?

\[ \text{A) } \text{K}^+ < \text{Na}^+ < \text{Li}^+ \quad \text{B) LiCl > KF > NaCl > NaBr} \]

\[ \text{C) The predictions from ionic radii are incorrect} \]

\[ \text{K}^+ : 80 \text{pm} \quad \text{Na}^+ : 102 \text{pm} \quad \text{Li}^+ : 102 \text{pm} \quad \text{Na}^+ : 2.48 \text{pm} \quad \text{K}^+ : 2.57 \text{pm} \quad \text{Li}^+ : 1.5 \text{pm} \quad \text{Cl}^− : 180 \text{pm} \quad \text{Br}^− : 188 \text{pm} \quad \text{Cl}^− : 180 \text{pm} \quad \text{O}^2− : 138 \text{pm} \]
63. Energy is required to remove two electrons from Ca to form Ca^{2+}, and energy is required to add two electrons to O to form O^{2-}. Yet CaO is stable relative to the free elements. Which statement is the best explanation? (a) The lattice energy of CaO is large enough to overcome these processes. (b) CaO is a covalent compound, and these processes are irrelevant. (c) CaO has a higher molar mass than either Ca or O. (d) The enthalpy of formation of CaO is small. (e) CaO is stable to atmospheric conditions.

A is the best.

64. List the individual steps used in constructing a Born–Haber cycle for the formation of BaI₂ from the elements. Which of the steps would you expect to be exothermic?

1) Use Aₚ, C to find ΔH°
2) Use Aₚ, C to find ΔH° for each ion in ionic form (e.g., Na⁺(g) → Na⁺(l))
3) Use ΔG° at 298 K and 278 K to find ΔS° for each ion in ionic form (e.g., Na⁺(g) → Na⁺(l))

65. Use data from Appendix C, Figure 7.10, and Figure 7.12 to calculate the lattice energy of RbCl.

1) Use Born–Haber cycle
2) Rb(g) + Cl(g) → RbCl(s)
3) \( \frac{1}{2} \text{Cl}_{(g)} + e^- \rightarrow \text{Cl}_{(g)} \)
4) \( -349.8 \text{kJ/mol} \)

66. (a) State whether the bonding in each compound is likely to be covalent or not: (i) iron, (ii) sodium chloride, (iii) water, (iv) oxygen, (v) argon. (b) A substance XY, formed from two different elements, boils at -33 °C. Is XY likely to be a covalent or an ionic substance?

(a) I not 2 not 3 cov - (c) cov 5 not

(b) Covalent - because covalent compounds boil at low temps.

67. Using Lewis symbols and Lewis structures, diagram the formation of SiCl₄ from Si and Cl atoms, showing valence-shell electrons. (a) How many valence electrons does Si have initially? (b) How many valence electrons does each Cl have initially? (c) How many valence electrons surround the Si in the SiCl₄ molecule? (d) How many valence electrons surround each Cl in the SiCl₄ molecule? (e) How many bonding pairs of electrons are in the SiCl₄ molecule?

(a) 4
(b) 7
(c) 1 bonding pair

68. (a) Construct a Lewis structure for O₂ in which each atom achieves an octet of electrons. (b) How many bonding electrons are in the structure? (c) Would you expect the O–O bond in O₂ to be shorter or longer than the O–O bond in compounds that contain an O–O single bond? Explain.

(a) \( \approx \)
(b) 1 bonding pair
(c) Single bond is longer than double, not shared means shorter bond.
69. Which of the following statements about electronegativity is false? (a) Electronegativity is the ability of an atom in a molecule to attract electron density toward itself. (b) Electronegativity is the same thing as electron affinity. (c) The numerical values for electronegativity have no units. (d) Fluorine is the most electronegative element. (e) Cesium is the least electronegative element.

Only true because Br is unable to be measured.

70. Using only the periodic table as your guide, select the most electronegative atom in each of the following sets: (a) Na, Mg, K, Ca; (b) P, S, As, Se; (c) Be, B, C, Si; (d) Zn, Ge, Ga, As.

71. Which of the following bonds are polar? (a) B-F, (b) O-Cl, (c) Se-O, (d) H-I. Which is the more electronegative atom in each polar bond?

72. (a) From the data in Table 8.3, calculate the effective charges on the H and Br atoms of the HBr molecule in units of the electronic charge, e. (b) If you were to put HBr under very high pressure, so its bond length decreased significantly, would its dipole moment increase, decrease, or stay the same, if you assume that the effective charges on the atoms do not change?

\[ r = 1.41 \text{ Å}, \quad D = 0.182 \text{ Å}, \quad Q = 1.10 \times 10^{-19} \text{ C}, \quad l_e = 1.10 \times 10^{-19} \text{ C} \]

73. Draw Lewis structures for the following: (a) SiH₄, (b) CO, (c) SF₂, (d) H₂SO₄ (H is bonded to O), (e) ClO₂⁻, (f) NH₂OH.

(a) H-Si-H, (b) C≡O, (c) :F-S-F, (d) H-O-S-OH, (e) O-C≡Cl-O⁻, (f) N-Ȯ-H.
74. Which one of these statements about formal charge is true? (a) Formal charge is the same as oxidation number. (b) To draw the best Lewis structure, you should minimize formal charge. (c) Formal charge takes into account the different electronegativities of the atoms in a molecule. (d) Formal charge is most useful for ionic compounds. (e) Formal charge is used in calculating the dipole moment of a diatomic molecule.

(a) F
(b) T
(c) F (Lewis structures do)
(d) F
(e) F

75. Write Lewis structures that obey the octet rule for each of the following, and assign oxidation numbers and formal charges to each atom: (a) OCS, (b) SOCl₂ (S is the central atom), (c) BrO₃⁻, (d) HClO₂ (H is bonded to O).

(a) \( \begin{align*}
&\text{C} \equiv \text{C} \equiv \text{S} \\
&\text{O} = -1 \\
&\text{O} = 0 \\
&\text{C} = 0 \\
&\text{S} = 0
\end{align*} \)

(b) \( \begin{align*}
&\text{Cl} \equiv \text{S} \equiv \text{Cl} \\
&\text{S} = -2 \\
&\text{O} = -1 \\
&\text{O} = 0
\end{align*} \)

(c) \( \begin{align*}
&\text{Cl} = 0 \\
&\text{O} = -1 \\
&\text{Br} = +1 \\
&\text{Cl} = 0 \\
&\text{Br} = -1
\end{align*} \)

(d) \( \begin{align*}
&\text{Cl} = 0 \\
&\text{O} = 0 \\
&\text{H} = +1
\end{align*} \)

76. (a) Draw the best Lewis structure(s) for the nitrite ion, NO₂⁻. (b) With what allotrope of oxygen is it isoelectronic? (c) What would you predict for the lengths of the bonds in NO₃⁻ relative to N=O single bonds and double bonds?

(a) \( \text{NO}_2^- \) 182°
(b) O₃ is closer
(c) The bonds in NO₂⁻ could be longer than a double and shorter than a single

77. Predict the ordering, from shortest to longest, of the bond lengths in CO, CO₂, and CO₃²⁻.

\( \begin{align*}
&\text{C} \equiv \text{O} \\
&\text{O} = \text{C} = \text{O} \\
&\text{CO} < \text{CO}_3^\text{2-} < \text{CO}_2^\text{2-}
\end{align*} \)

78. (a) Do the C—C bond lengths in benzene alternate short-long- short-long around the ring? Why or why not? (b) Are C—C bond lengths in benzene shorter than C—C single bonds? (c) Are C—C bond lengths in benzene shorter than C—C double bonds?

(a) no because the bonds "tickle" every other time
(b) benzene has shorter bonds
(c) longer than double
79. Indicate whether each statement is true or false: (a) The octet rule is based on the fact that filling in all s and p valence electrons in a shell gives eight electrons. (b) The Si in SiH₄ does not follow the octet rule because hydrogen is in an unusual oxidation state. (c) Boron compounds are frequent exceptions to the octet rule because they have too few electrons surrounding the boron. (d) Compounds in which nitrogen is the central atom are frequent exceptions to the octet rule because they have too many electrons surrounding the nitrogen.  

   A) T  B) F  C) T  D) F

80. Draw the Lewis structures for each of the following ions or molecules. Identify those in which the octet rule is not obeyed; state which atom in each compound does not follow the octet rule; and state, for those atoms, how many electrons surround these atoms: (a) PH₃, (b) AlH₃, (c) N³⁻, (d) CH₂Cl₂, (e) SnF₆.

   A) \( \begin{array}{c} \text{P} \\ \text{H} \end{array} \)  B) \( \begin{array}{c} \text{Al} \\ \text{H} \end{array} \)  C) :\( \text{N} \)\( ^{-2} \)  D) \( \begin{array}{c} \text{C} \\ \text{H} \\ \text{C} \\ \text{Cl} \\ \text{Cl} \\ \text{Cl} \end{array} \)  E) \( \begin{array}{c} \text{Sn} \\ \text{F} \\ \text{F} \\ \text{F} \\ \text{F} \\ \text{F} \end{array} \)

   A) does not follow  B) \text{PH}  C) \text{N} \text{H}  D) \text{H}  E) \text{Sn} \text{F}

81. In the vapor phase, BeCl₂ exists as a discrete molecule. (a) Draw the Lewis structure of this molecule, using only single bonds. Does this Lewis structure satisfy the octet rule? (b) What other resonance structures are possible that satisfy the octet rule? (c) On the basis of the formal charges, which Lewis structure is expected to be dominant for BeCl₂?

   A) \( \text{Be} - \text{Cl} \)  B) \( \text{Cl}^{\ominus} \text{Be} - \text{Cl}^{\dagger} \)  C) \( \text{Be} \) should be \( \text{Be}^{\ominus} \text{Cl}^{\dagger} \)

82. Consider the following statement: “For some molecules and ions, a Lewis structure that satisfies the octet rule does not lead to the lowest formal charges, and a Lewis structure that leads to the lowest formal charges does not satisfy the octet rule.” Illustrate this statement using the hydrogen sulfite ion, HSO₃⁻, as an example (the H atom is bonded to one of the O atoms).

\[
\begin{align*}
\text{HSO}_3^- \quad & \text{HSO}_3^- \quad & \text{HSO}_3^- \\
\text{H} - \text{O} - \text{S} - \text{O} & \quad \text{H} - \text{O} - \text{S} - \text{O} & \quad \text{H} - \text{O} - \text{S} - \text{O} \\
\text{O} & \quad \text{O} & \quad \text{O} \\
\end{align*}
\]
83. Using Table 8.4, estimate $\Delta H$ for each of the following gas phase reactions (note that lone pairs on atoms are not shown):

\[ \begin{align*}
(a) & \quad \text{H}_2\text{O}_2 + \text{H}_2\text{O}_2 \rightarrow 3 \text{H}_2\text{O} \quad \Delta H = -304 \text{ kJ} \\
(b) & \quad \text{H}_2\text{C} + \text{C} = \text{N} \rightarrow \text{H}_2\text{C} + \text{C} = \text{N} \quad \Delta H = -82 \text{ kJ} \\
(c) & \quad 2\text{N} = \text{N} + 3\text{Cl} - \text{Cl} \quad \Delta H = -4(\text{Cl}) \text{ kJ}
\end{align*} \]

see question 49 for equation and use Table 8.4 to calculate.

84. Using Table 8.4, estimate $\Delta H$ for each of the following reactions:

\[ \begin{align*}
& (a) \quad 2\text{CH}_4(g) + \text{O}_2(g) \rightarrow 2\text{CH}_3\text{OH}(g) \quad \Delta H = -521 \\
& (b) \quad \text{H}_2(g) + \text{Br}_2(l) \rightarrow \text{HB}(g) \quad \Delta H = -103 \\
& (c) \quad 2\text{H}_2\text{O}_2(g) \rightarrow 2\text{H}_2(g) + \text{O}_2(g) \quad \Delta H = -205
\end{align*} \]

get above

85. Ammonia is produced directly from nitrogen and hydrogen by using the Haber process, which is perhaps the most widely used industrial chemical reaction on Earth. The chemical reaction is $\text{N}_2(g) + 3\text{H}_2(g) \rightarrow 2\text{NH}_3(g)$ (a) Use Table 8.4 to estimate the enthalpy change for the reaction. Is it exothermic or endothermic? (b) Calculate the enthalpy change as obtained using $\Delta H^\circ$ values.

\[ \begin{align*}
& (a) \quad \text{bonds broken} \quad 3 \text{H-H} \quad 436 \text{ kJ} \\
& \text{1 N=N} \quad 941 \text{ kJ} \\
& \text{bonds formed} \quad 6 \text{N-H} \quad 391 \text{ kJ} \\
& \Delta H_{\text{in}} = 2\text{N} - 2\text{H}_2 = -292 \text{ kJ} \\
& \text{exothermic}
\end{align*} \]

\[ \Delta H_f^\circ = -32 \text{ kJ} \]