CHAPTER 11
CHEMICAL BONDS: THE FORMATION OF COMPOUNDS FROM ATOMS
SOLUTIONS TO REVIEW QUESTIONS

1. smallest Cl, Mg, Na, K, Rb largest.

2. More energy is required for neon because it has a very stable outer electron structure consisting of an octet of electrons in filled orbitals (noble gas electron structure). Sodium, an alkali metal, has a relatively unstable outer electron structure with a single electron in an unfilled orbital. The sodium electron is also farther away from the nucleus and is shielded by more inner electrons than are neon outer electrons.

3. When a third electron is removed from beryllium, it must come from a very stable electron structure corresponding to that of the noble gas, helium. In addition, the third electron must be removed from a +2 beryllium ion, which increases the difficulty of removing it.

4. The first ionization energy decreases from top to bottom because in the successive alkali metals, the outermost electron is farther away from the nucleus and is more shielded from the positive nucleus by additional electron energy levels.

5. The first ionization energy decreases from top to bottom because the outermost electrons in the successive noble gases are farther away from the nucleus and are more shielded by additional inner electron energy levels.

6. Helium has a much higher first ionization energy than does hydrogen because helium has a very stable outer electron structure, consisting of a filled principal energy level.

7. After the first electron is removed from an atom of lithium, much more energy would be required to remove a second electron since that one would come from a noble gas electron configuration, a filled principal energy level.

8. (a) Li > Be (d) Cl > F
    (b) Rb > K (e) Se > O
    (c) Al > P (f) As > Kr

9. The first element in each group has the smallest radius.

10. Atomic size increases down a column since each successive element has an additional energy level which contains electrons located farther from the nucleus.

11. By losing one electron, a potassium atom acquires a noble gas structure and becomes a $K^+$ ion. To become a $K^{2+}$ ion requires the loss of a second electron and breaking into the noble gas structure of the $K^+$ ion. This requires too much energy to generally occur.

12. An aluminum ion has a $+3$ charge because it has lost 3 electrons in acquiring a noble gas electron structure.
13. A Lewis structure is a representation of the bonding in a compound. Valence electrons are responsible for bonding, therefore they are the only electrons that need to be shown in a Lewis structure.

14. Group 1A 2A 3A 4A 5A 6A 7A
   E·   E·   E·   E·   E·   E·   E·

15. Lewis structure:
   Cs·  Ba·  Tl·  Pb·  Po·  At·  Rn·

Each of these is a representative element and has the same number of electrons in its outer energy level as its periodic group.

16. Valence electrons are the electrons found in the outermost s and p energy levels of an atom.

17. Metals are less electronegative than nonmetals. Therefore, metals lose electrons more easily than nonmetals. So, metals will transfer electrons to nonmetals leaving the metals with a positive charge and the nonmetals with a negative charge.

18. The noble gases are the most stable of all the elements because they have a complete octet (8 electrons) in their valence level. When the elements in Groups 1A, 2A, 6A, and 7A form ions, they do so to establish a stable electron structure.
   (a) The elements in Group 1A lose an electron (obtain a positive charge) in order to achieve a noble gas electron configuration.
   (b) The elements in Group 2A lose electrons (obtain a positive charge) in order to achieve a noble gas electron configuration.
   (c) The elements in Group 6A gain electrons (obtain a negative charge) in order to achieve a noble gas electron configuration.
   (d) The elements in Group 7A gain an electron (obtain a negative charge) in order to achieve a noble gas electron configuration.

19. All compounds must be neutrally charged. This means the overall charge on an ionic compound must be zero.

20. The alkaline earth elements.

21. Mg₂(PO₄)₂, Be₁(PO₄)₂, Sr₃(PO₄)₂, and Ba₁(PO₄)₂. Note that the basic formula is the same for all of the elements in the same family when they form an ionic compound with a phosphate ion.

22. When magnesium loses two electrons it will achieve the same electron configuration as the noble gas neon. Elements tend to gain or lose electrons to achieve a noble gas electron configuration.

23. The term molecule is used to describe covalent compounds that have a distinct set of atoms in their structure. Ionic compounds are composed of collections of ions which come together in ratios that balance their charges, but there are no independent molecules formed.

24. A single covalent bond is composed of two electrons. A maximum of three covalent bonds may be formed between any two atoms.
25. A covalent bond is formed by the overlap of the orbitals on individual atoms. There is a sharing of electrons in the region where the orbitals overlap.

26. In a Lewis structure the dash represents a two electron covalent bond.

27. Not all molecules that contain polar bonds are polar due to dipoles that cancel each other by acting in equal and opposite direction.

28. The more electronegative atom in the bond between two atoms will more strongly attract electrons so it will have a partial negative charge (\(\delta^-\)). The less electronegative atom will have a partial positive charge (\(\delta^+\)) because the bonding pair of electrons has been pulled away by the more electronegative atom.

29. (a) Elements with the highest electronegativities are found in the upper right hand corner of the periodic table.
   (b) Elements with the lowest electronegativities are found in the lower left of the periodic table.

30. A Lewis structure is a visual representation of the arrangement of atoms and electrons in a molecule or an ion. It shows how the atoms in a molecule are bonded together.

31. The dots in a Lewis structure represent electrons. The lines represent bonding pairs of electrons.

32. There are times when two or more Lewis structures can be drawn for a single skeleton structure. These different structures all exist and are called resonance structures.

33. Resonance structures

34. The Lewis structure for an ion should be drawn inside of a set of square brackets and the charge of the entire ion specified at the upper right hand outer corner of the brackets. An example is \([\text{Na}^+]\).

35. \(\text{Na}_2\text{SO}_4\) and \(\text{CaCO}_3\) are two of many possible examples of compounds with both ionic and covalent bonds.

36. The electron pair arrangement is the arrangement of both bonding and nonbonding electrons in a Lewis structure. The molecular shape of a molecule is the three dimensional arrangement of its atoms in space.
SOLUTIONS TO EXERCISES

1. Drawing 1 K⁺ ion     Drawing 2 K atom
Both particles have the same number of protons, so K⁺ with one fewer electron is smaller.

2. Drawing 1 Cl atom    Drawing 2 Cl⁻ ion
Both particles have the same number of protons, so Cl⁻ with one more electron is larger.

3. (a) A calcium atom is larger because it has electrons in the 4th shell, while a calcium ion does not. In addition, Ca²⁺ ion has 20 protons and 18 electrons, creating a charge imbalance and drawing the electrons in towards the positive nucleus.
(b) A chloride ion is larger because it has one more electron than a chlorine atom, in its outer shell. Also, the ion has 17 protons and 18 electrons, creating a charge imbalance, resulting in a lessening of the attraction of the electrons towards the nucleus.
(c) A magnesium ion is larger than an aluminum ion. Both ions will have 10 electrons in their electron shells, but the aluminium ion will have a greater charge imbalance since it has 13 protons and the magnesium ion has 12 protons. The charge imbalance draws the electrons in more closely to the nucleus.
(d) A sodium atom is larger than a silicon atom. Sodium and silicon are both in period 3. Going across a period, the radii of atoms decrease.
(e) A bromide ion is larger than a potassium ion. The bromide ion has 35 protons and 36 electrons, creating a charge imbalance that results in a lessening of the attraction of the electrons towards the nucleus. The potassium ion has 19 protons and 18 electrons, creating a charge imbalance that results in the electrons being drawn more closely to the nucleus.

4. (a) Fe²⁺ has one electron more than Fe³⁺, so it will have a larger radius.
(b) A potassium atom is larger than a potassium ion. They both have 19 protons. The potassium atom has 19 electrons. The potassium ion has 18 electrons, creating a charge imbalance that results in the electrons being drawn more closely to the nucleus.
(c) A chloride ion is larger than a sodium ion. The chloride ion has 17 protons and 18 electrons, creating a charge imbalance that results in a lessening of the attraction of the electrons towards the nucleus. The sodium ion has 11 protons and 10 electrons, creating a charge imbalance that results in the electrons being drawn more closely to the nucleus.
(d) A strontium atom is larger than an iodine atom. Strontium and iodine are both in period 5. Going across a period, the radii of atoms decrease.
(e) A rubidium ion is larger than a strontium ion. Both ions will have 36 electrons in their electron shells, but the strontium ion will have a greater charge imbalance since it has 38 protons and the rubidium ion has 37 protons. The charge imbalance draws the electrons in more closely to the nucleus.

5. (a) H O       (d) Pb S
(b) Rb Cl      (e) P F
(c) H N       (f) H C
6. 
(a) $\text{H}^+$ $\text{Cl}^-$  
(b) $\text{S}^2-$ $\text{O}^+$  
(c) $\text{C}^+$ $\text{Cl}^-$  
(d) $\text{I}^-$ $\text{Br}^+$  
(e) $\text{Cs}^+$ $\text{I}^-$  
(f) $\text{O}^-$ $\text{F}^+$

7. (a) covalent  
(b) ionic  
(c) ionic  
(d) covalent

8. (a) ionic  
(b) covalent  
(c) covalent  
(d) covalent

9. (a) $\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-  $  
(b) $\text{Br} + \text{e}^- \rightarrow \text{Br}^-$

10. (a) $\text{K} \rightarrow \text{K}^+ + \text{e}^-  $  
(b) $\text{S} + 2\text{e}^- \rightarrow \text{S}^{2-}$

11. (a) $\text{Li} \rightarrow \text{Li}_2\text{O}  $ 
(b) $\text{K} \rightarrow \text{K}_3\text{N}  $ 

12. (a) $\text{K} \rightarrow \text{K}_2\text{S}  $ 
(b) $\text{Ca}_3\text{N}_2  $ 

13. (a) $\text{Se} \text{(6)}  $ 
(b) $\text{P} \text{(5)}  $ 
(c) $\text{Br} \text{(7)}  $ 
(d) $\text{Mg} \text{(2)}  $ 
(e) $\text{He} \text{(2)}  $ 
(f) $\text{As} \text{(5)}  $ 

14. (a) $\text{Pb} \text{(4)}  $ 
(b) $\text{Li} \text{(1)}  $ 
(c) $\text{O} \text{(6)}  $ 
(d) $\text{Cs} \text{(1)}  $ 
(e) $\text{Ga} \text{(3)}  $ 
(f) $\text{Ar} \text{(8)}  $ 

15. Noble gas structures:
(a) potassium atom, lose 1 e$^-$  
(b) aluminum ion, none  
(c) bromine atom, gain 1 e$^-$  
(d) selenium atom, gain 2 e$^-$

16. (a) sulfur atom, gain 2 e$^-$  
(b) calcium atom, lose 2 e$^-$  
(c) nitrogen atom, gain 3e$^-$  
(d) iodide ion, none

17. (a) ionic, $\text{NaCl}  $ 
(b) molecular, $\text{CH}_4  $ 
(c) ionic, $\text{MgBr}_2  $ 
(d) molecular, $\text{Br}_2  $ 
(e) molecular, $\text{CO}_2  $ 

18. (a) Two oxygen atoms will form a nonpolar covalent compound. The formula is $\text{O}_2$.  
(b) Hydrogen and bromine will form a polar covalent compound. The formula is $\text{HBr}$. 

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(c) Oxygen and two hydrogen atoms will form a polar covalent compound. The formula is H₂O.
(d) Two iodine atoms will form a nonpolar covalent compound, the formula is I₂.

19. (a) NaH, Na₂O  (c) AlH₃, Al₂O₃
   (b) CaH₂, CaO   (d) SnH₄, SnO₂

20. (a) SbH₃, Sb₂O₃  (c) HCl, Cl₂O₇
   (b) H₂Se, SeO₃   (d) CH₄, CO₂

21. Li₂SO₄  lithium sulfate  K₂SO₄  potassium sulfate
   Rb₂SO₄  rubidium sulfate  Cs₂SO₄  cesium sulfate
   Fr₂SO₄  francium sulfate

22. BeBr₂  beryllium bromide  BaBr₂  barium bromide
   MgBr₂  magnesium bromide  RaBr₂  radium bromide
   SrBr₂  strontium bromide

23. Lewis structures:
   (a) Na⁺  (b) Br⁻  (c) O²⁻

24. (a) Ga³⁺  (b) [Ga]³⁺  (c) [Ca]²⁺

25. (a) covalent  (c) ionic
   (b) ionic  (d) covalent

26. (a) covalent  (c) covalent
   (b) ionic  (d) covalent

27. (a) covalent  (b) covalent  (c) covalent

28. (a) covalent  (b) covalent  (c) covalent

29. (a) H::H  (b) :N::N:  (c) :Cl::Cl:

30. (a) :O::O:  (b) :Br::Br:  (c) :I::I:

31. (a) :Cl::N::Cl:  (c) H::H
   (b) H::C::O::H  (d) [Na]⁺\left[ \text{ion} \right]

32. (a) :S::H  (c) H::N::H
   (b) :S::C::S:  (d) \left[ \text{ion} \right]^⁺\left[ \text{ion} \right]⁻
33. (a) [Ba]$^{2+}$  (d) $\text{C} \equiv \text{N} \equiv$  
(b) [Al]$^{3+}$  (e) $\text{O} - \text{C} \equiv \text{O} -$  
(c) $\text{O} : \text{S} : \text{O} : \text{O} ::$  
34. (a) $\text{I} -$  (d) $\text{O} - \text{Cl} - \text{O} -$  
(b) $\text{S} ^{2-}$  (e) $\text{O} - \text{N} : \text{O} : \text{O} -$  
(c) $\text{O} : \text{C} : \text{O} : \text{O} ::$  
35. (a) CH$_3$Cl, polar  (c) OF$_2$, polar  
(b) Cl$_2$, nonpolar  (d) PBr$_3$, polar  
36. (a) H$_2$, nonpolar  (c) CH$_3$OH, polar  
(b) NI$_3$, polar  (d) CS$_2$, nonpolar  
37. (a) 4 electron pairs, tetrahedral  
(b) 4 electron pairs, tetrahedral  
(c) 3 electron pairs, trigonal planar  
38. (a) 3 electron pairs, trigonal planar  
(b) 4 electron pairs, tetrahedral  
(c) 4 electron pairs, tetrahedral  
39. (a) tetrahedral  (b) trigonal pyramidal  (c) tetrahedral  
40. (a) tetrahedral  (b) trigonal pyramidal  (c) tetrahedral  
41. (a) tetrahedral  (b) trigonal pyramidal  (c) bent  
42. (a) tetrahedral  (b) bent  (c) bent  
43. Oxygen  
44. Potassium
45. Atom 1 Argon
    Atom 2 Sodium
    Atom 3 Cesium

46. Drawing 1 Sr$^{2+}$ All of these particles are isoelectronic, meaning they have the same number of electrons. The more protons in their nucleus, the smaller the particle.
    Drawing 2 Rb$^{+}$
    Drawing 3 Kr
    Drawing 4 Br$^{-}$
    Drawing 5 Se$^{2-}$

47. hydrazine N$_2$H$_4$ 14 e$^-$ $\text{N}=\text{N}$
    hydrozoic acid HN$_3$ 16 e$^-$ H$=\text{N}=\text{N}=\text{N}$

48. (a) NO$_2^-$ 18 e$^-$ bent
    (b) SO$_4^{2-}$ 32 e$^-$ tetrahedral
    (c) SOCl$_2$ 26 e$^-$ trigonal pyramidal
    (d) Cl$_2$O 20 e$^-$ bent

49. (a) C$_2$H$_6$ 14 e$^-$ $\text{H} \text{C} = \text{C} \text{H}$
    (b) C$_2$H$_4$ 12 e$^-$ $\text{H} = \text{C} = \text{C} = \text{H}$
    (c) C$_2$H$_2$ 10 e$^-$ $\text{H} = \text{C} = \text{C} = \text{H}$

50. (a) Be (b) He (c) K (d) F (e) Fr (f) Ne

51. (a) Cl (b) O (c) Ca

52. The noble gases already have a full outer electron configuration and therefore there is no need for them to attract electrons.
53. Lithium has a $+1$ charge after the first electron is removed. It takes more energy to overcome that charge and to remove another electron than to remove a single electron from an uncharged He atom.

54. Yes. Each of these elements have an $ns^1$ electron and they could lose that electron in the same way elements in Group 1A do. They would then form $+1$ ions and ionic compounds such as CuCl, AgCl, and AuCl.

55. SnBr$_2$, GeBr$_2$.

56. The bond between sodium and chlorine is ionic. An electron has been transferred from a sodium atom to a chlorine atom. The substance is composed of ions not molecules. Use of the word molecule implies covalent bonding.

57. A covalent bond results from the sharing of a pair of electrons between two atoms, while an ionic bond involves the transfer of one or more electrons from one atom to another.

58. This structure shown is incorrect since the bond is ionic. It should be represented as:

$$[\text{Na}^+]\cdot[\text{Cl}^-]$$

59. The four most electronegative elements are F, O, N, Cl.

60. highest F, O, S, H, Mg, Cs lowest.

61. It is possible for a molecule to be nonpolar even though it contains polar bonds. If the molecule is symmetrical, the polarities of the bonds will cancel (in a manner similar to a positive and negative number of the same size) resulting in a nonpolar molecule. An example is CO$_2$ which is linear and nonpolar.

62. Both molecules contain polar bonds. CO$_2$ is symmetrical about the C atom, so the polarities cancel. In CO, there is only one polar bond, therefore the molecule is polar.

63. (a) NO < CO < NaO (b) GeO < SiO < CO (c) BSe < BS < BO

64. (a) 109.5° (actual angle closer to 105°) (b) 109.5° (actual angle closer to 107°) (c) 109.5° (d) 109.5°

65. (a) Both use the p orbitals for bonding. B uses one s and two p orbitals while N uses one s and three p orbitals for bonding.

(b) BF$_3$ is trigonal planar while NF$_3$ is trigonal pyramidal.

(c) BF$_3$ has no lone pairs

NF$_3$ has one lone pair

(d) BF$_3$ has 3 very polar covalent bonds. NF$_3$ has 3 polar covalent bonds

66. Fluorine’s electronegativity is greater than any other element. Ionic bonds form between atoms of widely different electronegativities. Therefore, Fr–F, Cs–F, Rb–F, or K–F would be ionic substances with the greatest electronegativity difference.
67. Each element in a particular column has the same number of valence electrons and therefore the same Lewis structure.

68. \[
\begin{align*}
S & \quad \frac{1.40 \text{ g}}{32.07 \text{ g/mol}} = 0.0437 \text{ mol} \\
& \quad \frac{0.0437}{0.0437} = 1.00 \\
O & \quad \frac{2.10 \text{ g}}{16.00 \text{ g/mol}} = 0.131 \text{ mol} \\
& \quad \frac{0.131}{0.0437} = 3.00
\end{align*}
\]
Empirical formula is \(\text{SO}_3\)

69. We need to know the molecular formula before we can draw the Lewis structure. From the data, determine the empirical and then the molecular formula.

\[
\begin{align*}
C & \quad \frac{14.5 \text{ g}}{12.01 \text{ g/mol}} = 1.21 \text{ mol} \\
& \quad \frac{1.21}{1.21} = 1.00 \\
\text{Cl} & \quad \frac{85.5 \text{ g}}{35.45 \text{ g/mol}} = 2.41 \text{ mol} \\
& \quad \frac{2.41}{1.21} = 2.01
\end{align*}
\]

\(\text{CCl}_2\) is the empirical formula

\[
\text{empirical mass} = 1(12.01 \text{ g}) + 2(35.45 \text{ g}) = 82.91 \text{ g}
\]

\[
\frac{166 \text{ g}}{82.91 \text{ g}} = 2.00
\]

Therefore, the molecular formula is \((\text{CCl}_2)_2\) or \(\text{C}_2\text{Cl}_4\)

70. (a) ionic (b) both (c) covalent (d) covalent

71. (a) ionic; sodium phosphide
(b) both; ammonium iodide
(c) covalent; sulfur dioxide
(d) covalent; hydrogen sulfide
(e) both; copper(II) nitrate
(f) ionic; magnesium oxide
72. (a) \[
\begin{align*}
\text{H-} & \text{O-} \text{S-} \text{O-} \text{H}
\end{align*}
\]

(b) \[
\begin{align*}
\text{[Na]}^+ \left[ \begin{array}{c}
\text{O-} \\
\text{N} \\
\text{O-} \\
\end{array} \right]^- 
\end{align*}
\]

(c) \[
\begin{align*}
\text{[K]}^+ \left[ \begin{array}{c}
\text{O-} \\
\text{C} \\
\text{O-} \\
\end{array} \right]^{2-} 
\end{align*}
\]

(d) \[
\begin{align*}
\text{H-C\equivN-} 
\end{align*}
\]

(e) \[
\begin{align*}
\left[ \begin{array}{c}
\text{O:} \\
\text{S:} \\
\text{O:} \\
\end{array} \right]^{2-} & \left[ \begin{array}{c}
\text{O:} \\
\text{S:} \\
\text{O:} \\
\end{array} \right]^{2-} \left[ \begin{array}{c}
\text{Al}^3+ \\
\text{O:} \\
\text{S:} \\
\end{array} \right]^{2-} 
\end{align*}
\]

(f) \[
\begin{align*}
\text{H:} & \text{O-} \\
\text{H-C-C-} & \text{O-} \\
\text{H} & 
\end{align*}
\]

73. \[
(25 \text{ g Li}) \left( \frac{1 \text{ mol}}{6.941 \text{ g}} \right) \left( \frac{520 \text{ kJ}}{\text{mol}} \right) = 1.9 \times 10^3 \text{ kJ}
\]

74. Removing the first electron from 1 mole of sodium atoms requires 496 kJ. To remove a second electron from 1 mole of sodium atoms requires 4,565 kJ. The conversions are:

\[
(15 \text{ mol Na}) \left( \frac{496 \text{ KJ}}{\text{mol}} \right) = 7.4 \times 10^3 \text{ kJ}
\]

\[
(15 \text{ mol Na}) \left( \frac{4565 \text{ KJ}}{\text{mol}} \right) = 6.8 \times 10^4 \text{ kJ}
\]

\[
(7.4 \times 10^3 \text{ kJ}) + (6.8 \times 10^4 \text{ kJ}) = 7.5 \times 10^4 \text{ kJ}
\]

75. \[
\begin{align*}
\text{O-} & \text{S-} \text{O-} \\
\text{H-} & \text{O-} \\
\end{align*}
\]

76. \[
\begin{align*}
\text{C-} & \text{C-} \\
\text{N-} & \\
\end{align*}
\]

Alliin

\[
\begin{align*}
\text{H-C-S-} & \text{C-} \\
\end{align*}
\]

Allicin
77. Tetrahedral electron pair and molecular geometry. Bond angles are ≈ 109.5°.

Tetrahedral electron pair geometry and bent molecular geometry. Bond angles are less than 109.5°.

Trigonal pyramidal electron pair and molecular geometry. Bond angles are 120°.

Linear electron pair and molecular geometry. Bond angles are 180°.

H-C-C-C≡N:

- Chapter 11 -