6. (a) NF$_3$ is nitrogen trifluoride.  
(b) HI is hydrogen iodide.  
(c) BBr$_3$ is boron tribromide.  
(d) C$_6$H$_{14}$ is hexane.  
7. (a) C$_8$H$_{18}$ is octane.  
(b) P$_2$S$_3$ is diphosphorus trisulfide.  
(c) OF$_2$ is oxygen difluoride.  
(d) XeF$_4$ is xenon tetrafluoride.  
8. (a) Sulfur trioxide is SO$_3$.  
(b) Dinitrogen pentoxide N$_2$O$_5$.  
(c) Phosphorus pentachloride is PCl$_5$.  
(d) Silicon tetrachloride is SiCl$_4$.  
(e) Diboron trioxide is B$_2$O$_3$.  

20. A general rule for the charge on a metal cation: the group number represents the number of electrons lost. Hence, the group number will be the cation’s positive charge. 

(a) Lithium (Group 1A)  
(b) Strontium (Group 2A)  
(c) Aluminum (Group 3A)  
(d) Calcium (Group 2A)  
(e) Zinc (Group 2B)  

21. For nonmetal elements in Groups 5A-7A, the electrons gained by an atom to form a stable anion are calculated using the formula: 8 – (group number). That means the (group number) – 8 is the negative charge of the anion. 

(a) nitrogen (Group 5A) 5 – 8 = -3  
(b) sulfur (Group 6A) 6 – 8 = -2  
(c) chlorine (Group 7A) 7 – 8 = -1  
(d) iodine (Group 7A) 7 – 8 = -1  
(e) phosphorus (Group 5A) 5 – 8 = -3  

22. A general rule for the charge on a metal cation: the group number represents the number of electrons lost. Hence, the group number will be the cation’s positive charge. For nonmetal elements in Groups 5A-7A, the electrons gained by an atom to form a stable anion are calculated using the formula: 8 – (group number). That means the (group number) – 8 is the negative charge of the anion. 

Barium (Group 2A) has a +2 charge. Bromine (Group 7A) has a –1 charge. So, the ions are Ba$^{2+}$ and Br$^–$.  

28. (a) Calcium ion (from Group 2A) is Ca$^{2+}$. Oxide ion (from Group 6A) is O$^{2–}$.  

Ca$_2$O is not a neutral combination of these two ions. The proper formula would be CaO. [net charge = +2 + (–2) = 0]  
(b) Strontium ion (Group 2A) is Sr$^{2+}$. Chloride ion (from Group 7A) is Cl$^–$.  

SrCl$_2$ is the proper neutral combination of these two ions.  

[net charge = +2 + 2 × (–1) = 0]  
(c) Iron ion (from the transition elements) is Fe$^{3+}$ or Fe$^{2+}$. Oxide ion (from Group 6A) is O$^{2–}$. Fe$_2$O$_3$ is not a neutral combination of these ions. The proper possible formulas would be FeO [net charge = +2 + (–2) = 0] or  

Fe$_2$O$_3$ [net charge = 2 × (+3) + 3 × (–2) = 0]
(d) Potassium ion (from Group 1A) is $K^+$. Oxide ion (from Group 6A) is $O^{2-}$.

$K_2O$ is the proper neutral combination of these two ions.

[net charge = $2 \times (+1) + (-2) = 0$]

30. (a) $Ca(CH_3CO_2)_2$ has one ion of calcium ($Ca^{2+}$) and two ions of acetate ($CH_3CO_2^-$ also written: $CH_3COO^-$).

(b) $Co_2(SO_4)_3$ has two ions of cobalt(III) ($Co^{3+}$) and three ions of sulfate ($SO_4^{2-}$).

(c) $Al(OH)_3$ has one ion of aluminum ($Al^{3+}$) and three ions of hydroxide ($OH^-$).

(d) $(NH_4)_2CO_3$ has two ions of ammonium ($NH_4^+$) and one ion of carbonate ($CO_3^{2-}$).

33. (a) Nickel(II) nitrate $Ni(NO_3)_2$

(b) sodium bicarbonate $NaHCO_3$

(c) Lithium hypochlorite $LiClO$

(d) magnesium chlorate $Mg(ClO_3)_2$

(e) Calcium sulfite $CaSO_3$

34. To tell if a compound is ionic or not, look for metals and nonmetals together, or common cations and anions. If a compound contains only nonmetals or metalloids and nonmetals, it is probably not ionic.

(a) CF$_4$ contains only nonmetals. Not ionic.

(b) SrBr$_2$ has a metal and nonmetal together. Ionic.

(c) Co(NO$_3$)$_3$ has a metal and nonmetals together. Ionic.

(d) SiO$_2$ contains a metalloid and a nonmetal. Not ionic.

(e) KCN has a metal and a common diatomic ion ($CN^-$) together. Ionic.

(f) SCl$_2$ contains only nonmetals. Not ionic.

39. (a) $Ca(CH_3CO_2)_2$ is calcium acetate.

(b) $Co_2(SO_4)_3$ is cobalt(III) sulfate.

(c) $Al(OH)_3$ is aluminum hydroxide.

54. Define the problem: Determine the number of moles in a given mass of a compound.

Develop a plan: Adapt the method described in the answer to Question 51 to calculate the molar mass for the compound, then use the molar mass as a conversion factor between grams and moles.

Execute the plan:

(a) Molar mass $CH_3OH = (12.0107 \text{ g}) + 4 \times (1.0079 \text{ g}) + (15.9994 \text{ g}) = 32.0417 \text{ g/mol}$

$$1.00 \text{ g } CH_3OH \times \frac{1 \text{ mol } CH_3OH}{32.0417 \text{ g } CH_3OH} = 0.0312 \text{ mol } CH_3OH$$

(b) Molar mass $Cl_2CO = 2 \times (35.453 \text{ g}) + (12.0107 \text{ g}) + (15.9994 \text{ g}) = 98.916 \text{ g/mol}$

$$1.00 \text{ g } Cl_2CO \times \frac{1 \text{ mol } Cl_2CO}{98.916 \text{ g } Cl_2CO} = 0.0101 \text{ mol } Cl_2CO$$

(c) Molar mass $NH_4NO_3 = 2 \times (14.0067 \text{ g}) + 4 \times (1.0079 \text{ g}) + 3 \times (15.9994 \text{ g}) = 80.043 \text{ g/mol}$

$$1.00 \text{ g } NH_4NO_3 \times \frac{1 \text{ mol } NH_4NO_3}{80.043 \text{ g } NH_4NO_3} = 0.0125 \text{ mol } NH_4NO_3$$
(d) Molar mass MgSO₄·7H₂O

\[ \text{Molar mass } = (24.305 \text{ g}) + (32.065 \text{ g}) + 11 \times (15.9994 \text{ g}) + 14 \times (1.0079 \text{ g}) = 246.474 \text{ g/mol} \]

\[ 1.00 \text{ g MgSO}_4 \cdot 7\text{H}_2\text{O} \times \frac{1 \text{ mol MgSO}_4 \cdot 7\text{H}_2\text{O}}{246.474 \text{ g MgSO}_4 \cdot 7\text{H}_2\text{O}} = 0.00406 \text{ mol MgSO}_4 \cdot 7\text{H}_2\text{O} \]

(e) Molar mass AgCH₃CO₂

\[ \text{Molar mass } = (107.8682 \text{ g}) + 2 \times (12.0107 \text{ g}) + 3 \times (1.0079 \text{ g}) + 2 \times (15.9994 \text{ g}) = 166.9121 \text{ g/mol} \]

\[ 1.00 \text{ g AgCH}_3\text{CO}_2 \times \frac{1 \text{ mol AgCH}_3\text{CO}_2}{166.9121 \text{ g AgCH}_3\text{CO}_2} = 0.00599 \text{ mol AgCH}_3\text{CO}_2 \]

**Check your answers:** The quantity in moles is always going to be smaller than the mass in grams. These numbers look right.

**Percent Composition**

65. **Define the problem:** Given the formula of a compound, determine the molar mass, and the mass percent of each element.

**Develop a plan:** Calculate the mass of each element in one mole of compound, while calculating the molar mass of the compound. Divide the calculated mass of the element by the molar mass of the compound and multiply by 100 % to get mass percent. To get the last element’s mass percent, subtract the other percentages from 100 %.

**Execute the plan:**

(a) Mass of Pb per mole of PbS = 207.2 g Pb

Mass of S per mole of PbS = 32.065 g S

Molar mass PbS = (207.2 g) + (32.065 g) = 239.3 g/mol PbS

\[ \% \text{ Pb} = \frac{\text{mass of Pb per mol PbS}}{\text{mass of PbS per mol PbS}} \times 100 \% = \frac{207.2 \text{ g Pb}}{239.3 \text{ g PbS}} \times 100 \% = 86.60 \% \text{ Pb in PbS} \]

\[ \% \text{ S} = 100 \% - 86.60 \% \text{ Pb} = 13.40 \% \text{ S in PbS} \]

(b) Mass of C per mole of C₂H₆ = 2 × (12.0107 g) = 24.0214 g C

Mass of H per mole of C₂H₆ = 6 × (1.0079 g) = 6.0474 g H

Molar mass C₂H₆ = (24.0214 g) + (6.0474 g) = 30.0688 g/mol C₂H₆

\[ \% \text{ C} = \frac{\text{mass of C per mol C}_2\text{H}_6}{\text{mass of C}_2\text{H}_6 / \text{mol C}_2\text{H}_6} \times 100 \% = \frac{24.0214 \text{ g C}}{30.0688 \text{ g C}_2\text{H}_6} \times 100 \% = 79.8881 \% \text{ C in C}_2\text{H}_6 \]

\[ \% \text{ H} = 100 \% - 79.8881 \% \text{ C} = 20.1119 \% \text{ H in C}_2\text{H}_6 \]

(c) Mass of C per mole of CH₃CO₂H = 2 × (12.0107 g) = 24.0214 g C

Mass of H per mole of CH₃CO₂H = 4 × (1.0079 g) = 4.0316 g H

Mass of O per mole of CH₃CO₂H = 2 × (15.9994 g) = 31.9988 g O

Molar mass CH₃CO₂H = (24.0214 g) + (4.0316 g) + (31.9988 g) = 60.0518 g/mol CH₃CO₂H

\[ \% \text{ C} = \frac{\text{mass of C per mol CH}_3\text{CO}_2\text{H}}{\text{mass of CH}_3\text{CO}_2\text{H per mol CH}_3\text{CO}_2\text{H}} \times 100 \% = \frac{24.0214 \text{ g C}}{60.0518 \text{ g CH}_3\text{CO}_2\text{H}} \times 100 \% = 40.0011 \% \text{ C in CH}_3\text{CO}_2\text{H} \]
\[
\%^{\text{H}} = \frac{\text{mass of H}}{\text{mol CH}_3\text{CO}_2\text{H}} \times 100 \%
\]
\[
= \frac{4.0316 \text{ g H}}{60.0518 \text{ g CH}_3\text{CO}_2\text{H}} \times 100 \% = 6.7135 \% \text{ H in CH}_3\text{CO}_2\text{H}
\]
\[
\%^{\text{O}} = 100 \% - 40.0011 \% \text{ C} - 6.7135 \% \text{ H} = 53.2854 \% \text{ O in CH}_3\text{CO}_2\text{H}
\]
\[
(d) \quad \text{Mass of N per mole of NH}_4\text{NO}_3 = 2 \times (14.0067 \text{ g}) = 28.0134 \text{ g N}
\]
\[
\text{Mass of H per mole of NH}_4\text{NO}_3 = 4 \times (1.0079 \text{ g}) = 4.0316 \text{ g H}
\]
\[
\text{Mass of O per mole of NH}_4\text{NO}_3 = 3 \times (15.9994 \text{ g}) = 47.9982 \text{ g O}
\]
\[
\text{Molar mass NH}_4\text{NO}_3 = (28.0134 \text{ g}) + (4.0316 \text{ g}) + (47.9982 \text{ g}) = 80.0432 \text{ g/mol NH}_4\text{NO}_3
\]
\[
\%^{\text{N}} = \frac{\text{mass of N}}{\text{mol NH}_4\text{NO}_3} \times 100 \% = \frac{28.0134 \text{ g N}}{80.0432 \text{ g NH}_4\text{NO}_3} \times 100 \% = 34.9979 \% \text{ N in NH}_4\text{NO}_3
\]
\[
\%^{\text{H}} = \frac{\text{mass of H}}{\text{mol NH}_4\text{NO}_3} \times 100 \% = \frac{4.0316 \text{ g H}}{80.0432 \text{ g NH}_4\text{NO}_3} \times 100 \% = 5.0368 \% \text{ H in NH}_4\text{NO}_3
\]
\[
\%^{\text{O}} = 100 \% - 34.9979 \% \text{ C} - 5.0368 \% \text{ H} = 59.9654 \% \text{ O in NH}_4\text{NO}_3
\]

Note: When masses of different things are used in the same problem, make sure your units clearly specify what each mass refers to.

Check your answers: Calculating the last element’s mass percent using the formula gives the same answer as subtracting the other percentages from 100 %. These answers are consistent.

67. Define the problem: Given the mass percent of one compound, M₂O, containing one known element, O, and one unknown element, M, calculate the percent by mass of another compound, MO.

Develop a plan: Choose a convenient sample mass of M₂O, such as 100.0 g. Find the mass of M and O in the sample, using the given mass percent. Using the molar mass of oxygen as a conversion factor, determine the number of moles of oxygen, then using the formula stoichiometry of M₂O as a conversion factor determine the number of moles of M. Find the molar mass of M by dividing the mass of M by the moles of M. Use the molar mass of M, and the formula stoichiometry of MO, to determine the mass percent of M in MO.

Execute the plan:

73.4 % M in M₂O means that 100.0 grams of M₂O contains 73.4 grams of M.

\[
\text{Mass of O} = 100.0 \text{ g M}_2\text{O} - 73.4 \text{ g M} = 26.6 \text{ g O}
\]

Formula Stoichiometry: 1 mol of M₂O contains 2 mol M and 1 mol O.

\[
\frac{26.6 \text{ g O} \times \frac{1 \text{ mol O}}{15.9994 \text{ g O}} \times \frac{2 \text{ mol M}}{1 \text{ mol O}} = 3.33 \text{ mol M}}
\]

Molar Mass of M = \frac{\text{mass of M in sample}}{\text{mol of M in sample}} = \frac{73.4 \text{ g M}}{3.33 \text{ mol M}} = \frac{22.1 \text{ g}}{\text{mol}}

Molar mass of MO = 22.1 \text{ g} + 15.9994 \text{ g} = 38.07 \text{ g/mol}
\[
\% M = \frac{\text{mass of } M}{\text{mol } MO} \times 100\% = \frac{22.1 \text{ g } M}{38.07 \text{ g } MO} \times 100\% = 58.0\% \text{ M in MO}
\]

Check your answer: It makes sense that the compound with more atoms of M has a higher mass percent of M. The closest element to M’s atomic mass (22.1) is sodium (atomic mass = 22.99). If M is sodium, the two compounds would probably be sodium oxide (Na₂O) and sodium peroxide (Na₂O₂, a compound made up of two Na⁺ ions and one O₂²⁻ ion. The simple ratio of Na and O atoms in this compound is 1:1). The results make sense.