

CHAPTER 3 STOICHIOMETRY OF FORMULAS AND EQUATIONS

END-OF-CHAPTER PROBLEMS

- 3.1 Plan: The atomic mass of an element expressed in amu is numerically the same as the mass of 1 mole of the element expressed in grams. We know the moles of each element and have to find the mass (in g). To convert moles of element to grams of element, multiply the number of moles by the molar mass of the element.

Solution:

Al 26.98 amu $\equiv$ 26.98 g/mol Al

$$\text{Mass Al (g)} = (3 \text{ mol Al}) \left(\frac{26.98 \text{ g Al}}{1 \text{ mol Al}} \right) = \mathbf{80.94 \text{ g Al}}$$

Cl 35.45 amu $\equiv$ 35.45 g/mol Cl

$$\text{Mass Cl (g)} = (2 \text{ mol Cl}) \left(\frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}} \right) = \mathbf{70.90 \text{ g Cl}}$$

- 3.2 Plan: The molecular formula of sucrose tells us that 1 mole of sucrose contains 12 moles of carbon atoms. Multiply the moles of sucrose by 12 to obtain moles of carbon atoms; multiply the moles of carbon atoms by Avogadro's number to convert from moles to atoms.

Solution:

$$\text{a) Moles of C atoms} = (1 \text{ mol C}_{12}\text{H}_{22}\text{O}_{11}) \left(\frac{12 \text{ mol C}}{1 \text{ mol C}_{12}\text{H}_{22}\text{O}_{11}} \right) = \mathbf{12 \text{ mol C}}$$

$$\text{b) C atoms} = (2 \text{ mol C}_{12}\text{H}_{22}\text{O}_{11}) \left(\frac{12 \text{ mol C}}{1 \text{ mol C}_{12}\text{H}_{22}\text{O}_{11}} \right) \left(\frac{6.022 \times 10^{23} \text{ C atoms}}{1 \text{ mol C}} \right) = \mathbf{1.445 \times 10^{25} \text{ C atoms}}$$

- 3.3 Plan: Review the list of elements that exist as diatomic or polyatomic molecules.

Solution:

"1 mol of chlorine" could be interpreted as a mole of chlorine atoms or a mole of chlorine molecules, Cl₂. Specify which to avoid confusion. The same problem is possible with other diatomic or polyatomic molecules, e.g., F₂, Br₂, I₂, H₂, O₂, N₂, S₈, and P₄. For these elements, as for chlorine, it is not clear if atoms or molecules are being discussed.

- 3.4 The molecular mass is the sum of the atomic masses of the atoms or ions in a molecule. The molar mass is the mass of 1 mole of a chemical entity. Both will have the same numeric value for a given chemical substance but molecular mass will have the units of amu and molar mass will have the units of g/mol.

- 3.5 A mole of a particular substance represents a fixed number of chemical entities and has a fixed mass. Therefore the mole gives us an easy way to determine the number of particles (atoms, molecules, etc.) in a sample by weighing it. The mole maintains the same mass relationship between macroscopic samples as exist between individual chemical entities. It relates the number of chemical entities (atoms, molecules, ions, electrons) to the mass.

- 3.6 Plan: The relative atomic masses of each element can be found by counting the number of atoms of each element and comparing the overall masses of the two samples.

Solution:

a) Balance A: The element on the **left** (green) has the higher molar mass because only 5 green balls are necessary to counterbalance the mass of 6 yellow balls. Since the green ball is heavier, its atomic mass is larger, and therefore its molar mass is larger. Balance B: The element on the **right** (blue) has the higher molar mass since 3 blue balls are heavier than 6 red balls. Since the blue ball is heavier, its atomic mass is larger, and therefore its

molar mass is larger. Balance C: The element on the **left** (orange) has the higher molar mass because 5 orange balls are heavier than 5 purple balls. Since the orange ball is heavier, its atomic mass is larger, and therefore its molar mass is larger. Balance D: The element on the **left** (gray) has the higher molar mass because only 5 gray balls are necessary to counterbalance the mass of 7 red balls. Since the gray ball is heavier, its atomic mass is larger, and therefore its molar mass is larger.

b) The elements on the **right** in Balances A, C, and D have more atoms per gram. The element on the right in each of these balances is lighter. Because these elements are lighter, more atoms are required to make 1 g. In Balance B, the element on the **left** is lighter and would therefore require more atoms to make 1 g.

c) The elements on the **left** in Balances A, C, and D have fewer atoms per gram. Atoms of these elements are heavier, and it takes fewer balls to make 1 g. In Balance B, the element on the **right** is heavier and therefore has fewer atoms per gram.

d) **Neither** element on any of the balances has more atoms per mole. Both the left and right elements have the same number of atoms per mole. The number of atoms per mole (6.022×10^{23}) is constant and so is the same for every element.

- 3.7 Plan: Locate each of the elements on the periodic table and record its atomic mass. The atomic mass of the element multiplied by the number of atoms present in the formula gives the mass of that element in one mole of the substance. The molar mass is the sum of the masses of the elements in the substance expressed in g/mol.

Solution:

- a) $M = (1 \times M \text{ of Sr}) + (2 \times M \text{ of O}) + (2 \times M \text{ of H})$
 $= (1 \times 87.62 \text{ g/mol Sr}) + (2 \times 16.00 \text{ g/mol O}) + (2 \times 1.008 \text{ g/mol H})$
 $= \mathbf{121.64 \text{ g/mol of Sr(OH)}_2}$
- b) $M = (2 \times M \text{ of N}) + (3 \times M \text{ of O})$
 $= (2 \times 14.01 \text{ g/mol N}) + (3 \times 16.00 \text{ g/mol O})$
 $= \mathbf{76.02 \text{ g/mol of N}_2\text{O}_3}$
- c) $M = (1 \times M \text{ of Na}) + (1 \times M \text{ of Cl}) + (3 \times M \text{ of O})$
 $= (1 \times 22.99 \text{ g/mol Na}) + (1 \times 35.45 \text{ g/mol Cl}) + (3 \times 16.00 \text{ g/mol O})$
 $= \mathbf{106.44 \text{ g/mol of NaClO}_3}$
- d) $M = (2 \times M \text{ of Cr}) + (3 \times M \text{ of O})$
 $= (2 \times 52.00 \text{ g/mol Cr}) + (3 \times 16.00 \text{ g/mol O})$
 $= \mathbf{152.00 \text{ g/mol of Cr}_2\text{O}_3}$

- 3.8 Plan: Locate each of the elements on the periodic table and record its atomic mass. The atomic mass of the element multiplied by the number of atoms present in the formula gives the mass of that element in one mole of the substance. The molar mass is the sum of the masses of the elements in the substance expressed in g/mol.

Solution:

- a) $M = (3 \times M \text{ of N}) + (12 \times M \text{ of H}) + (1 \times M \text{ of P}) + (4 \times M \text{ of O})$
 $= (3 \times 14.01 \text{ g/mol N}) + (12 \times 1.008 \text{ g/mol H}) + (1 \times 30.97 \text{ g/mol P}) + (4 \times 16.00 \text{ g/mol O})$
 $= \mathbf{149.10 \text{ g/mol of (NH}_4)_3\text{PO}_4}$
- b) $M = (1 \times M \text{ of C}) + (2 \times M \text{ of H}) + (2 \times M \text{ of Cl})$
 $= (1 \times 12.01 \text{ g/mol C}) + (2 \times 1.008 \text{ g/mol H}) + (2 \times 35.45 \text{ g/mol Cl})$
 $= \mathbf{84.93 \text{ g/mol of CH}_2\text{Cl}_2}$
- c) $M = (1 \times M \text{ of Cu}) + (1 \times M \text{ of S}) + (9 \times M \text{ of O}) + (10 \times M \text{ of H})$
 $= (1 \times 63.55 \text{ g/mol Cu}) + (1 \times 32.07 \text{ g/mol S}) + (9 \times 16.00 \text{ g/mol O}) + (10 \times 1.008 \text{ g/mol H})$
 $= \mathbf{249.70 \text{ g/mol of CuSO}_4 \cdot 5\text{H}_2\text{O}}$
- d) $M = (1 \times M \text{ of Br}) + (3 \times M \text{ of F})$
 $= (1 \times 79.90 \text{ g/mol Br}) + (3 \times 19.00 \text{ g/mol F})$
 $= \mathbf{136.90 \text{ g/mol of BrF}_3}$

- 3.9 Plan: Locate each of the elements on the periodic table and record its atomic mass. The atomic mass of the element multiplied by the number of atoms present in the formula gives the mass of that element in one mole of the substance. The molar mass is the sum of the masses of the elements in the substance expressed in g/mol.

Solution:

- a) $M = (1 \times M \text{ of Sn}) + (1 \times M \text{ of O})$
 $= (1 \times 118.7 \text{ g/mol Sn}) + (1 \times 16.00 \text{ g/mol O})$
 $= \mathbf{134.7 \text{ g/mol of SnO}}$
- b) $M = (1 \times M \text{ of Ba}) + (2 \times M \text{ of F})$
 $= (1 \times 137.3 \text{ g/mol Ba}) + (2 \times 19.00 \text{ g/mol F})$
 $= \mathbf{175.3 \text{ g/mol of BaF}_2}$
- c) $M = (2 \times M \text{ of Al}) + (3 \times M \text{ of S}) + (12 \times M \text{ of O})$
 $= (2 \times 26.98 \text{ g/mol Al}) + (3 \times 32.07 \text{ g/mol S}) + (12 \times 16.00 \text{ g/mol O})$
 $= \mathbf{342.17 \text{ g/mol of Al}_2(\text{SO}_4)_3}$
- d) $M = (1 \times M \text{ of Mn}) + (2 \times M \text{ of Cl})$
 $= (1 \times 54.94 \text{ g/mol Mn}) + (2 \times 35.45 \text{ g/mol Cl})$
 $= \mathbf{125.84 \text{ g/mol of MnCl}_2}$

- 3.10 Plan: Locate each of the elements on the periodic table and record its atomic mass. The atomic mass of the element multiplied by the number of atoms present in the formula gives the mass of that element in one mole of the substance. The molar mass is the sum of the masses of the elements in the substance expressed in g/mol.

Solution:

- a) $M = (2 \times M \text{ of N}) + (4 \times M \text{ of O})$
 $= (2 \times 14.01 \text{ g/mol N}) + (4 \times 16.00 \text{ g/mol O})$
 $= \mathbf{92.02 \text{ g/mol of N}_2\text{O}_4}$
- b) $M = (4 \times M \text{ of C}) + (10 \times M \text{ of H}) + (1 \times M \text{ of O})$
 $= (4 \times 12.01 \text{ g/mol C}) + (10 \times 1.008 \text{ g/mol H}) + (1 \times 16.00 \text{ g/mol O})$
 $= \mathbf{74.12 \text{ g/mol of C}_4\text{H}_{10}\text{O}}$
- c) $M = (1 \times M \text{ of Mg}) + (1 \times M \text{ of S}) + (11 \times M \text{ of O}) + (14 \times M \text{ of H})$
 $= (1 \times 24.31 \text{ g/mol Mg}) + (1 \times 32.07 \text{ g/mol S}) + (11 \times 16.00 \text{ g/mol O}) + (14 \times 1.008 \text{ g/mol H})$
 $= \mathbf{246.49 \text{ g/mol of MgSO}_4 \cdot 7\text{H}_2\text{O}}$
- d) $M = (1 \times M \text{ of Ca}) + (4 \times M \text{ of C}) + (6 \times M \text{ of H}) + (4 \times M \text{ of O})$
 $= (1 \times 40.08 \text{ g/mol Ca}) + (4 \times 12.01 \text{ g/mol C}) + (6 \times 1.008 \text{ g/mol H}) + (4 \times 16.00 \text{ g/mol O})$
 $= \mathbf{158.17 \text{ g/mol of Ca(C}_2\text{H}_3\text{O}_2)_2}$

- 3.11 Plan: Determine the molar mass of each substance, then perform the appropriate molar conversions. To find the mass in part a), multiply the number of moles by the molar mass of the substance. In part b), first convert mass of compound to moles of compound by dividing by the molar mass of the compound. The molecular formula of the compound tells us that 1 mole of compound contains 6 moles of oxygen atoms; use the 1:6 ratio to convert moles of compound to moles of oxygen atoms. In part c), convert mass of compound to moles of compound by dividing by the molar mass of the compound. Since 1 mole of compound contains 6 moles of oxygen atoms, multiply the moles of compound by 6 to obtain moles of oxygen atoms; then multiply by Avogadro's number to obtain the number of oxygen atoms.

Solution:

- a) $M \text{ of KMnO}_4 = (1 \times M \text{ of K}) + (1 \times M \text{ of Mn}) + (4 \times M \text{ of O})$
 $= (1 \times 39.10 \text{ g/mol K}) + (1 \times 54.94 \text{ g/mol Mn}) + (4 \times 16.00 \text{ g/mol O}) = 158.04 \text{ g/mol of KMnO}_4$
$$\text{Mass of KMnO}_4 = (0.68 \text{ mol KMnO}_4) \left(\frac{158.04 \text{ g KMnO}_4}{1 \text{ mol KMnO}_4} \right) = 107.467 = \mathbf{1.1 \times 10^2 \text{ g KMnO}_4}$$
- b) $M \text{ of Ba(NO}_3)_2 = (1 \times M \text{ of Ba}) + (2 \times M \text{ of N}) + (6 \times M \text{ of O})$
 $= (1 \times 137.3 \text{ g/mol Ba}) + (2 \times 14.01 \text{ g/mol N}) + (6 \times 16.00 \text{ g/mol O}) = 261.3 \text{ g/mol Ba(NO}_3)_2$
$$\text{Moles of Ba(NO}_3)_2 = (8.18 \text{ g Ba(NO}_3)_2) \left(\frac{1 \text{ mol Ba(NO}_3)_2}{261.3 \text{ g Ba(NO}_3)_2} \right) = 0.031305 \text{ mol Ba(NO}_3)_2$$

$$\text{Moles of O atoms} = (0.031305 \text{ mol Ba(NO}_3)_2) \left(\frac{6 \text{ mol O atoms}}{1 \text{ mol Ba(NO}_3)_2} \right) = 0.18783 = \mathbf{0.188 \text{ mol O atoms}}$$
- c) $M \text{ of CaSO}_4 \cdot 2\text{H}_2\text{O} = (1 \times M \text{ of Ca}) + (1 \times M \text{ of S}) + (6 \times M \text{ of O}) + (4 \times M \text{ of H})$
 $= (1 \times 40.08 \text{ g/mol Ca}) + (1 \times 32.07 \text{ g/mol S}) + (6 \times 16.00 \text{ g/mol O}) + (4 \times 1.008 \text{ g/mol H})$
 $= \mathbf{172.18 \text{ g/mol}}$

(Note that the waters of hydration are included in the molar mass.)

$$\text{Moles of CaSO}_4 \cdot 2\text{H}_2\text{O} = \left(7.3 \times 10^{-3} \text{ g CaSO}_4 \cdot 2\text{H}_2\text{O}\right) \left(\frac{1 \text{ mol CaSO}_4 \cdot 2\text{H}_2\text{O}}{172.18 \text{ g CaSO}_4 \cdot 2\text{H}_2\text{O}}\right) = 4.239749 \times 10^{-5} \text{ mol}$$

$$\begin{aligned} \text{Moles of O atoms} &= \left(4.239749 \times 10^{-5} \text{ mol CaSO}_4 \cdot 2\text{H}_2\text{O}\right) \left(\frac{6 \text{ mol O atoms}}{1 \text{ mol CaSO}_4 \cdot 2\text{H}_2\text{O}}\right) \\ &= 2.54385 \times 10^{-5} \text{ mol O atoms} \end{aligned}$$

$$\begin{aligned} \text{Number of O atoms} &= \left(2.54385 \times 10^{-5} \text{ mol O atoms}\right) \left(\frac{6.022 \times 10^{23} \text{ O atoms}}{1 \text{ mol O atoms}}\right) \\ &= 1.5319 \times 10^{20} = \mathbf{1.5 \times 10^{20} \text{ O atoms}} \end{aligned}$$

3.12 Plan: Determine the molar mass of each substance, then perform the appropriate molar conversions.

To find the mass in part a), divide the number of molecules by Avogadro's number to find moles of compound and then multiply the mole amount by the molar mass in grams; convert from mass in g to mass in kg. In part b), first convert mass of compound to moles of compound by dividing by the molar mass of the compound. The molecular formula of the compound tells us that 1 mole of compound contains 2 moles of chlorine atoms; use the 1:2 ratio to convert moles of compound to moles of chlorine atoms. In part c), convert mass of compound to moles of compound by dividing by the molar mass of the compound. Since 1 mole of compound contains 2 moles of H^- ions, multiply the moles of compound by 2 to obtain moles of H^- ions; then multiply by Avogadro's number to obtain the number of H^- ions.

Solution:

$$\begin{aligned} \text{a) } \mathcal{M} \text{ of NO}_2 &= (1 \times \mathcal{M} \text{ of N}) + (2 \times \mathcal{M} \text{ of O}) \\ &= (1 \times 14.01 \text{ g/mol N}) + (2 \times 16.00 \text{ g/mol O}) = 46.01 \text{ g/mol of NO}_2 \end{aligned}$$

$$\text{Moles of NO}_2 = \left(4.6 \times 10^{21} \text{ molecules NO}_2\right) \left(\frac{1 \text{ mol NO}_2}{6.022 \times 10^{23} \text{ molecules NO}_2}\right) = 7.63866 \times 10^{-3} \text{ mol NO}_2$$

$$\text{Mass (kg) of NO}_2 = \left(7.63866 \times 10^{-3} \text{ mol NO}_2\right) \left(\frac{46.01 \text{ g NO}_2}{1 \text{ mol NO}_2}\right) \left(\frac{1 \text{ kg}}{10^3 \text{ g}}\right) = 3.51455 \times 10^{-4} = \mathbf{3.5 \times 10^{-4} \text{ kg NO}_2}$$

$$\begin{aligned} \text{b) } \mathcal{M} \text{ of C}_2\text{H}_4\text{Cl}_2 &= (2 \times \mathcal{M} \text{ of C}) + (4 \times \mathcal{M} \text{ of H}) + (2 \times \mathcal{M} \text{ of Cl}) \\ &= (2 \times 12.01 \text{ g/mol C}) + (4 \times 1.008 \text{ g/mol H}) + (2 \times 35.45 \text{ g/mol Cl}) = 98.95 \text{ g/mol of C}_2\text{H}_4\text{Cl}_2 \end{aligned}$$

$$\text{Moles of C}_2\text{H}_4\text{Cl}_2 = \left(0.0615 \text{ g C}_2\text{H}_4\text{Cl}_2\right) \left(\frac{1 \text{ mol C}_2\text{H}_4\text{Cl}_2}{98.95 \text{ g C}_2\text{H}_4\text{Cl}_2}\right) = 6.21526 \times 10^{-4} \text{ mol C}_2\text{H}_4\text{Cl}_2$$

$$\begin{aligned} \text{Moles of Cl atoms} &= \left(6.21526 \times 10^{-4} \text{ mol C}_2\text{H}_4\text{Cl}_2\right) \left(\frac{2 \text{ mol Cl atoms}}{1 \text{ mol C}_2\text{H}_4\text{Cl}_2}\right) = 1.2431 \times 10^{-3} \\ &= \mathbf{1.24 \times 10^{-3} \text{ mol Cl atoms}} \end{aligned}$$

$$\text{c) } \mathcal{M} \text{ of SrH}_2 = (1 \times \mathcal{M} \text{ of Sr}) + (2 \times \mathcal{M} \text{ of H}) = (1 \times 87.62 \text{ g/mol Sr}) + (2 \times 1.008 \text{ g/mol H}) = 89.64 \text{ g/mol of SrH}_2$$

$$\text{Moles of SrH}_2 = \left(5.82 \text{ g SrH}_2\right) \left(\frac{1 \text{ mol SrH}_2}{89.64 \text{ g SrH}_2}\right) = 0.0649264 \text{ mol SrH}_2$$

$$\text{Moles of H}^- \text{ ions} = \left(0.0649264 \text{ mol SrH}_2\right) \left(\frac{2 \text{ mol H}^-}{1 \text{ mol SrH}_2}\right) = 0.1298528 \text{ mol H}^- \text{ ions}$$

$$\text{Number of H}^- \text{ ions} = \left(0.1298528 \text{ mol H}^- \text{ ions}\right) \left(\frac{6.022 \times 10^{23} \text{ H}^- \text{ ions}}{1 \text{ mol H}^-}\right) = 7.81974 \times 10^{22} = \mathbf{7.82 \times 10^{22} \text{ H}^- \text{ ions}}$$

3.13 Plan: Determine the molar mass of each substance, then perform the appropriate molar conversions. To find the mass in part a), multiply the number of moles by the molar mass of the substance. In part b), first convert the mass of compound in kg to mass in g and divide by the molar mass of the compound to find moles of compound. In part c), convert mass of compound in mg to mass in g and divide by the molar mass of the compound to find

moles of compound. Since 1 mole of compound contains 2 moles of nitrogen atoms, multiply the moles of compound by 2 to obtain moles of nitrogen atoms; then multiply by Avogadro's number to obtain the number of nitrogen atoms.

Solution:

$$\begin{aligned}\text{a) } \mathcal{M} \text{ of MnSO}_4 &= (1 \times \mathcal{M} \text{ of Mn}) + (1 \times \mathcal{M} \text{ of S}) + (4 \times \mathcal{M} \text{ of O}) \\ &= (1 \times 54.94 \text{ g/mol Mn}) + (1 \times 32.07 \text{ g/mol S}) + (4 \times 16.00 \text{ g/mol O}) = 151.01 \text{ g/mol of MnSO}_4\end{aligned}$$

$$\text{Mass (g) of MnSO}_4 = \left(6.44 \times 10^{-2} \text{ mol MnSO}_4\right) \left(\frac{151.01 \text{ g MnSO}_4}{1 \text{ mol MnSO}_4}\right) = 9.725044 = \mathbf{9.73 \text{ g MnSO}_4}$$

$$\begin{aligned}\text{b) } \mathcal{M} \text{ of Fe(ClO}_4)_3 &= (1 \times \mathcal{M} \text{ of Fe}) + (3 \times \mathcal{M} \text{ of Cl}) + (12 \times \mathcal{M} \text{ of O}) \\ &= (1 \times 55.85 \text{ g/mol Fe}) + (3 \times 35.45 \text{ g/mol S}) + (12 \times 16.00 \text{ g/mol O}) \\ &= 354.20 \text{ g/mol of Fe(ClO}_4)_3\end{aligned}$$

$$\text{Mass (g) of Fe(ClO}_4)_3 = \left(15.8 \text{ kg Fe(ClO}_4)_3\right) \left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right) = 1.58 \times 10^4 \text{ kg Fe(ClO}_4)_3$$

$$\text{Moles of Fe(ClO}_4)_3 = \left(1.58 \times 10^4 \text{ g Fe(ClO}_4)_3\right) \left(\frac{1 \text{ mol Fe(ClO}_4)_3}{354.20 \text{ g Fe(ClO}_4)_3}\right) = 44.6076 = \mathbf{44.6 \text{ mol Fe(ClO}_4)_3}$$

$$\begin{aligned}\text{c) } \mathcal{M} \text{ of NH}_4\text{NO}_2 &= (2 \times \mathcal{M} \text{ of N}) + (4 \times \mathcal{M} \text{ of H}) + (2 \times \mathcal{M} \text{ of O}) \\ &= (2 \times 14.01 \text{ g/mol N}) + (4 \times 1.008 \text{ g/mol H}) + (2 \times 16.00 \text{ g/mol O}) = 64.05 \text{ g/mol NH}_4\text{NO}_2\end{aligned}$$

$$\text{Mass (g) of NH}_4\text{NO}_2 = \left(92.6 \text{ mg NH}_4\text{NO}_2\right) \left(\frac{10^{-3} \text{ g}}{1 \text{ mg}}\right) = 0.0926 \text{ g NH}_4\text{NO}_2$$

$$\text{Moles of NH}_4\text{NO}_2 = \left(0.0926 \text{ g NH}_4\text{NO}_2\right) \left(\frac{1 \text{ mol NH}_4\text{NO}_2}{64.05 \text{ g NH}_4\text{NO}_2}\right) = 1.44575 \times 10^{-3} \text{ mol NH}_4\text{NO}_2$$

$$\text{Moles of N atoms} = \left(1.44575 \times 10^{-3} \text{ mol NH}_4\text{NO}_2\right) \left(\frac{2 \text{ mol N atoms}}{1 \text{ mol NH}_4\text{NO}_2}\right) = 2.8915 \times 10^{-3} \text{ mol N atoms}$$

$$\begin{aligned}\text{Number of N atoms} &= \left(2.8915 \times 10^{-3} \text{ mol N atoms}\right) \left(\frac{6.022 \times 10^{23} \text{ N atoms}}{1 \text{ mol N atoms}}\right) \\ &= 1.74126 \times 10^{21} = \mathbf{1.74 \times 10^{21} \text{ N atoms}}\end{aligned}$$

- 3.14 Plan: Determine the molar mass of each substance, then perform the appropriate molar conversions. In part a), divide the mass by the molar mass of the compound to find moles of compound. Since 1 mole of compound contains 3 moles of ions (1 mole of Sr^{2+} and 2 moles of F^-), multiply the moles of compound by 3 to obtain moles of ions and then multiply by Avogadro's number to obtain the number of ions. In part b), multiply the number of moles by the molar mass of the substance to find the mass in g and then convert to kg. In part c), divide the number of formula units by Avogadro's number to find moles; multiply the number of moles by the molar mass to obtain the mass in g and then convert to mg.

Solution:

$$\begin{aligned}\text{a) } \mathcal{M} \text{ of SrF}_2 &= (1 \times \mathcal{M} \text{ of Sr}) + (2 \times \mathcal{M} \text{ of F}) \\ &= (1 \times 87.62 \text{ g/mol Sr}) + (2 \times 19.00 \text{ g/mol F}) = 125.62 \text{ g/mol of SrF}_2\end{aligned}$$

$$\text{Moles of SrF}_2 = \left(38.1 \text{ g SrF}_2\right) \left(\frac{1 \text{ mol SrF}_2}{125.62 \text{ g SrF}_2}\right) = 0.303296 \text{ mol SrF}_2$$

$$\text{Moles of ions} = \left(0.303296 \text{ mol SrF}_2\right) \left(\frac{3 \text{ mol ions}}{1 \text{ mol SrF}_2}\right) = 0.909888 \text{ mol ions}$$

$$\text{Number of ions} = \left(0.909888 \text{ mol ions}\right) \left(\frac{6.022 \times 10^{23} \text{ ions}}{1 \text{ mol ions}}\right) = 5.47935 \times 10^{23} = \mathbf{5.48 \times 10^{23} \text{ ions}}$$

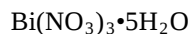
$$\begin{aligned}\text{b) } \mathcal{M} \text{ of CuCl}_2 \cdot 2\text{H}_2\text{O} &= (1 \times \mathcal{M} \text{ of Cu}) + (2 \times \mathcal{M} \text{ of Cl}) + (4 \times \mathcal{M} \text{ of H}) + (2 \times \mathcal{M} \text{ of O}) \\ &= (1 \times 63.55 \text{ g/mol Cu}) + (2 \times 35.45 \text{ g/mol Cl}) + (4 \times 1.008 \text{ g/mol H}) + (2 \times 16.00 \text{ g/mol O}) \\ &= 170.48 \text{ g/mol of CuCl}_2 \cdot 2\text{H}_2\text{O}\end{aligned}$$

(Note that the waters of hydration are included in the molar mass.)

$$\text{Mass (g) of CuCl}_2 \cdot 2\text{H}_2\text{O} = (3.58 \text{ mol CuCl}_2 \cdot 2\text{H}_2\text{O}) \left(\frac{170.48 \text{ g CuCl}_2 \cdot 2\text{H}_2\text{O}}{1 \text{ mol CuCl}_2 \cdot 2\text{H}_2\text{O}} \right) = 610.32 \text{ g CuCl}_2 \cdot 2\text{H}_2\text{O}$$

$$\text{Mass (kg) of CuCl}_2 \cdot 2\text{H}_2\text{O} = (610.32 \text{ g CuCl}_2 \cdot 2\text{H}_2\text{O}) \left(\frac{1 \text{ kg}}{10^3 \text{ g}} \right) = 0.61032 = \mathbf{0.610 \text{ kg CuCl}_2 \cdot 2\text{H}_2\text{O}}$$

$$\begin{aligned} \text{c) } \mathcal{M} \text{ of Bi(NO}_3)_3 \cdot 5\text{H}_2\text{O} &= (1 \times \mathcal{M} \text{ of Bi}) + (3 \times \mathcal{M} \text{ of N}) + (10 \times \mathcal{M} \text{ of H}) + (14 \times \mathcal{M} \text{ of O}) \\ &= (1 \times 209.0 \text{ g/mol Bi}) + (3 \times 14.01 \text{ g/mol N}) + (10 \times 1.008 \text{ g/mol H}) \\ &\quad + (14 \times 16.00 \text{ g/mol O}) = 485.11 \text{ g/mol of} \end{aligned}$$



(Note that the waters of hydration are included in the molar mass.)

$$\text{Moles of Bi(NO}_3)_3 \cdot 5\text{H}_2\text{O} = (2.88 \times 10^{22} \text{ FU}) \left(\frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ FU}} \right) = 0.047825 \text{ mol Bi(NO}_3)_3 \cdot 5\text{H}_2\text{O}$$

$$\text{Mass (g) of Bi(NO}_3)_3 \cdot 5\text{H}_2\text{O} = (0.047825 \text{ mol Bi(NO}_3)_3 \cdot 5\text{H}_2\text{O}) \left(\frac{485.1 \text{ g Bi(NO}_3)_3 \cdot 5\text{H}_2\text{O}}{1 \text{ mol Bi(NO}_3)_3 \cdot 5\text{H}_2\text{O}} \right) = 23.1999 \text{ g}$$

$$\begin{aligned} \text{Mass (mg) of Bi(NO}_3)_3 \cdot 5\text{H}_2\text{O} &= (23.1999 \text{ g Bi(NO}_3)_3 \cdot 5\text{H}_2\text{O}) \left(\frac{1 \text{ mg}}{10^{-3} \text{ g}} \right) \\ &= 23199.9 = \mathbf{2.32 \times 10^4 \text{ mg Bi(NO}_3)_3 \cdot 5\text{H}_2\text{O}} \end{aligned}$$

- 3.15 Plan: The formula of each compound must be determined from its name. The molar mass for each formula comes from the formula and atomic masses from the periodic table. Determine the molar mass of each substance, then perform the appropriate molar conversions. In part a), multiply the moles by the molar mass of the compound to find the mass of the sample. In part b), divide the number of molecules by Avogadro's number to find moles; multiply the number of moles by the molar mass to obtain the mass. In part c), divide the mass by the molar mass to find moles of compound and multiply moles by Avogadro's number to find the number of formula units. In part d), use the fact that each formula unit contains 1 Na ion, 1 perchlorate ion, 1 Cl atom, and 4 O atoms.

Solution:

a) Carbonate is a polyatomic anion with the formula, CO_3^{2-} . Copper(I) indicates Cu^+ . The correct formula for this ionic compound is Cu_2CO_3 .

$$\begin{aligned} \mathcal{M} \text{ of Cu}_2\text{CO}_3 &= (2 \times \mathcal{M} \text{ of Cu}) + (1 \times \mathcal{M} \text{ of C}) + (3 \times \mathcal{M} \text{ of O}) \\ &= (2 \times 63.55 \text{ g/mol Cu}) + (1 \times 12.01 \text{ g/mol C}) + (3 \times 16.00 \text{ g/mol O}) = 187.11 \text{ g/mol of Cu}_2\text{CO}_3 \end{aligned}$$

$$\text{Mass (g) of Cu}_2\text{CO}_3 = (8.35 \text{ mol Cu}_2\text{CO}_3) \left(\frac{187.11 \text{ g Cu}_2\text{CO}_3}{1 \text{ mol Cu}_2\text{CO}_3} \right) = 1562.4 = \mathbf{1.56 \times 10^3 \text{ g Cu}_2\text{CO}_3}$$

b) Dinitrogen pentoxide has the formula N_2O_5 . Di- indicates 2 N atoms and penta- indicates 5 O atoms.

$$\begin{aligned} \mathcal{M} \text{ of N}_2\text{O}_5 &= (2 \times \mathcal{M} \text{ of N}) + (5 \times \mathcal{M} \text{ of O}) \\ &= (2 \times 14.01 \text{ g/mol N}) + (5 \times 16.00 \text{ g/mol O}) = 108.02 \text{ g/mol of N}_2\text{O}_5 \end{aligned}$$

$$\text{Moles of N}_2\text{O}_5 = (4.04 \times 10^{20} \text{ N}_2\text{O}_5 \text{ molecules}) \left(\frac{1 \text{ mol N}_2\text{O}_5}{6.022 \times 10^{23} \text{ N}_2\text{O}_5 \text{ molecules}} \right) = 6.7087 \times 10^{-4} \text{ mol N}_2\text{O}_5$$

$$\text{Mass (g) of N}_2\text{O}_5 = (6.7087 \times 10^{-4} \text{ mol N}_2\text{O}_5) \left(\frac{108.02 \text{ g N}_2\text{O}_5}{1 \text{ mol N}_2\text{O}_5} \right) = 0.072467 = \mathbf{0.0725 \text{ g N}_2\text{O}_5}$$

c) The correct formula for this ionic compound is NaClO_4 ; Na has a charge of +1 (Group 1 ion) and the perchlorate ion is ClO_4^- .

$$\begin{aligned} \mathcal{M} \text{ of NaClO}_4 &= (1 \times \mathcal{M} \text{ of Na}) + (1 \times \mathcal{M} \text{ of Cl}) + (4 \times \mathcal{M} \text{ of O}) \\ &= (1 \times 22.99 \text{ g/mol Na}) + (1 \times 35.45 \text{ g/mol Cl}) + (4 \times 16.00 \text{ g/mol O}) = 122.44 \text{ g/mol of NaClO}_4 \end{aligned}$$

$$\text{Moles of NaClO}_4 = (78.9 \text{ g NaClO}_4) \left(\frac{1 \text{ mol NaClO}_4}{122.44 \text{ g NaClO}_4} \right) = 0.644397 = \mathbf{0.644 \text{ mol NaClO}_4}$$

FU = formula units

$$\begin{aligned}\text{FU of NaClO}_4 &= (0.644397 \text{ mol NaClO}_4) \left(\frac{6.022 \times 10^{23} \text{ FU NaClO}_4}{1 \text{ mol NaClO}_4} \right) \\ &= 3.88056 \times 10^{23} = \mathbf{3.88 \times 10^{23} \text{ FU NaClO}_4}\end{aligned}$$

$$\text{d) Number of Na}^+ \text{ ions} = (3.88056 \times 10^{23} \text{ FU NaClO}_4) \left(\frac{1 \text{ Na}^+ \text{ ion}}{1 \text{ FU NaClO}_4} \right) = \mathbf{3.88 \times 10^{23} \text{ Na}^+ \text{ ions}}$$

$$\text{Number of ClO}_4^- \text{ ions} = (3.88056 \times 10^{23} \text{ FU NaClO}_4) \left(\frac{1 \text{ ClO}_4^- \text{ ion}}{1 \text{ FU NaClO}_4} \right) = \mathbf{3.88 \times 10^{23} \text{ ClO}_4^- \text{ ions}}$$

$$\text{Number of Cl atoms} = (3.88056 \times 10^{23} \text{ FU NaClO}_4) \left(\frac{1 \text{ Cl atom}}{1 \text{ FU NaClO}_4} \right) = \mathbf{3.88 \times 10^{23} \text{ Cl atoms}}$$

$$\text{Number of O atoms} = (3.88056 \times 10^{23} \text{ FU NaClO}_4) \left(\frac{4 \text{ O atoms}}{1 \text{ FU NaClO}_4} \right) = \mathbf{1.55 \times 10^{24} \text{ O atoms}}$$

- 3.16 Plan: The formula of each compound must be determined from its name. The molar mass for each formula comes from the formula and atomic masses from the periodic table. Determine the molar mass of each substance, then perform the appropriate molar conversions. In part a), multiply the moles by the molar mass of the compound to find the mass of the sample. In part b), divide the number of molecules by Avogadro's number to find moles; multiply the number of moles by the molar mass to obtain the mass. In part c), divide the mass by the molar mass to find moles of compound and multiply moles by Avogadro's number to find the number of formula units. In part d), use the fact that each formula unit contains 2 Li ions, 1 sulfate ion, 1 S atom, and 4 O atoms.

Solution:

a) Sulfate is a polyatomic anion with the formula, SO_4^{2-} . Chromium(III) indicates Cr^{3+} . Decahydrate indicates 10 water molecules ("waters of hydration"). The correct formula for this ionic compound is $\text{Cr}_2(\text{SO}_4)_3 \cdot 10\text{H}_2\text{O}$.

$$\begin{aligned}\mathcal{M} \text{ of } \text{Cr}_2(\text{SO}_4)_3 \cdot 10\text{H}_2\text{O} &= (2 \times \mathcal{M} \text{ of Cr}) + (3 \times \mathcal{M} \text{ of S}) + (22 \times \mathcal{M} \text{ of O}) + (20 \times \mathcal{M} \text{ of H}) \\ &= (2 \times 52.00 \text{ g/mol Cr}) + (3 \times 32.07 \text{ g/mol S}) + (22 \times 16.00 \text{ g/mol O}) + (20 \times 1.008 \text{ g/mol H}) \\ &= 572.4 \text{ g/mol of } \text{Cr}_2(\text{SO}_4)_3 \cdot 10\text{H}_2\text{O}\end{aligned}$$

$$\begin{aligned}\text{Mass (g) of } \text{Cr}_2(\text{SO}_4)_3 \cdot 10\text{H}_2\text{O} &= (8.42 \text{ mol } \text{Cr}_2(\text{SO}_4)_3 \cdot 10\text{H}_2\text{O}) \left(\frac{572.4 \text{ g}}{\text{mol}} \right) \\ &= 4819.608 = \mathbf{4.82 \times 10^3 \text{ g } \text{Cr}_2(\text{SO}_4)_3 \cdot 10\text{H}_2\text{O}}\end{aligned}$$

b) Dichlorine heptaoxide has the formula Cl_2O_7 . Di- indicates 2 Cl atoms and hepta- indicates 7 O atoms.

$$\begin{aligned}\mathcal{M} \text{ of } \text{Cl}_2\text{O}_7 &= (2 \times \mathcal{M} \text{ of Cl}) + (7 \times \mathcal{M} \text{ of O}) \\ &= (2 \times 35.45 \text{ g/mol Cl}) + (7 \times 16.00 \text{ g/mol O}) = 182.9 \text{ g/mol of } \text{Cl}_2\text{O}_7\end{aligned}$$

$$\text{Moles of } \text{Cl}_2\text{O}_7 = (1.83 \times 10^{24} \text{ molecules } \text{Cl}_2\text{O}_7) \left(\frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molecules}} \right) = 3.038858 \text{ mol } \text{Cl}_2\text{O}_7$$

$$\text{Mass (g) of } \text{Cl}_2\text{O}_7 = (3.038858 \text{ mol } \text{Cl}_2\text{O}_7) \left(\frac{182.9 \text{ g } \text{Cl}_2\text{O}_7}{1 \text{ mol}} \right) = 555.807 = \mathbf{5.56 \times 10^2 \text{ g } \text{Cl}_2\text{O}_7}$$

c) The correct formula for this ionic compound is Li_2SO_4 ; Li has a charge of +1 (Group 1 ion) and the sulfate ion is SO_4^{2-} .

$$\begin{aligned}\mathcal{M} \text{ of } \text{Li}_2\text{SO}_4 &= (2 \times \mathcal{M} \text{ of Li}) + (1 \times \mathcal{M} \text{ of S}) + (4 \times \mathcal{M} \text{ of O}) \\ &= (2 \times 6.941 \text{ g/mol Li}) + (1 \times 32.07 \text{ g/mol S}) + (4 \times 16.00 \text{ g/mol O}) = 109.95 \text{ g/mol of } \text{Li}_2\text{SO}_4\end{aligned}$$

$$\text{Moles of } \text{Li}_2\text{SO}_4 = (6.2 \text{ g } \text{Li}_2\text{SO}_4) \left(\frac{1 \text{ mol } \text{Li}_2\text{SO}_4}{109.95 \text{ g } \text{Li}_2\text{SO}_4} \right) = 0.056389 = \mathbf{0.056 \text{ mol } \text{Li}_2\text{SO}_4}$$

$$\text{FU of } \text{Li}_2\text{SO}_4 = (0.056389 \text{ mol } \text{Li}_2\text{SO}_4) \left(\frac{6.022 \times 10^{23} \text{ FU}}{1 \text{ mol } \text{Li}_2\text{SO}_4} \right) = 3.3957 \times 10^{22} = \mathbf{3.4 \times 10^{22} \text{ FU } \text{Li}_2\text{SO}_4}$$

$$\text{d) Number of Li}^+ \text{ ions} = \left(3.3957 \times 10^{22} \text{ FU Li}_2\text{SO}_4 \right) \left(\frac{2 \text{ Li}^+ \text{ ions}}{1 \text{ FU Li}_2\text{SO}_4} \right) = 6.7914 \times 10^{22} = \mathbf{6.8 \times 10^{22} \text{ Li}^+ \text{ ions}}$$

$$\text{Number of SO}_4^{2-} \text{ ions} = \left(3.3957 \times 10^{22} \text{ FU Li}_2\text{SO}_4 \right) \left(\frac{1 \text{ SO}_4^{2-} \text{ ion}}{1 \text{ FU Li}_2\text{SO}_4} \right) = 3.3957 \times 10^{22} = \mathbf{3.4 \times 10^{22} \text{ SO}_4^{2-} \text{ ions}}$$

$$\text{Number of S atoms} = \left(3.3957 \times 10^{22} \text{ FU Li}_2\text{SO}_4 \right) \left(\frac{1 \text{ S atom}}{1 \text{ FU Li}_2\text{SO}_4} \right) = 3.3957 \times 10^{22} = \mathbf{3.4 \times 10^{22} \text{ S atoms}}$$

$$\text{Number of O atoms} = \left(3.3957 \times 10^{22} \text{ FU Li}_2\text{SO}_4 \right) \left(\frac{4 \text{ O atoms}}{1 \text{ FU Li}_2\text{SO}_4} \right) = 1.3583 \times 10^{23} = \mathbf{1.4 \times 10^{23} \text{ O atoms}}$$

- 3.17 Plan: Determine the formula and the molar mass of each compound. The formula gives the relative number of moles of each element present. Multiply the number of moles of each element by its molar mass to find the total mass of element in 1 mole of compound. Mass percent = $\frac{\text{total mass of element}}{\text{molar mass of compound}}(100)$.

Solution:

a) Ammonium bicarbonate is an ionic compound consisting of ammonium ions, NH_4^+ and bicarbonate ions, HCO_3^- . The formula of the compound is NH_4HCO_3 .

$$\begin{aligned} \mathcal{M} \text{ of } \text{NH}_4\text{HCO}_3 &= (1 \times \mathcal{M} \text{ of N}) + (5 \times \mathcal{M} \text{ of H}) + (1 \times \mathcal{M} \text{ of C}) + (3 \times \mathcal{M} \text{ of O}) \\ &= (1 \times 14.01 \text{ g/mol N}) + (5 \times 1.008 \text{ g/mol H}) + (1 \times 12.01 \text{ g/mol C}) + (3 \times 16.00 \text{ g/mol O}) \\ &= 79.06 \text{ g/mol of } \text{NH}_4\text{HCO}_3 \end{aligned}$$

There are 5 moles of H in 1 mole of NH_4HCO_3 .

$$\text{Mass (g) of H} = (5 \text{ mol H}) \left(\frac{1.008 \text{ g H}}{1 \text{ mol H}} \right) = 5.040 \text{ g H}$$

$$\text{Mass percent} = \frac{\text{total mass H}}{\text{molar mass of compound}}(100) = \frac{5.040 \text{ g H}}{79.06 \text{ g } \text{NH}_4\text{HCO}_3}(100) = 6.374905 = \mathbf{6.375\% \text{ H}}$$

b) Sodium dihydrogen phosphate heptahydrate is a salt that consists of sodium ions, Na^+ , dihydrogen phosphate ions, H_2PO_4^- , and seven waters of hydration. The formula is $\text{NaH}_2\text{PO}_4 \cdot 7\text{H}_2\text{O}$. Note that the waters of hydration are included in the molar mass.

$$\begin{aligned} \mathcal{M} \text{ of } \text{NaH}_2\text{PO}_4 \cdot 7\text{H}_2\text{O} &= (1 \times \mathcal{M} \text{ of Na}) + (16 \times \mathcal{M} \text{ of H}) + (1 \times \mathcal{M} \text{ of P}) + (11 \times \mathcal{M} \text{ of O}) \\ &= (1 \times 22.99 \text{ g/mol Na}) + (16 \times 1.008 \text{ g/mol H}) + (1 \times 30.97 \text{ g/mol P}) + (11 \times 16.00 \text{ g/mol O}) \\ &= 246.09 \text{ g/mol } \text{NaH}_2\text{PO}_4 \cdot 7\text{H}_2\text{O} \end{aligned}$$

There are 11 moles of O in 1 mole of $\text{NaH}_2\text{PO}_4 \cdot 7\text{H}_2\text{O}$.

$$\text{Mass (g) of O} = (11 \text{ mol O}) \left(\frac{16.00 \text{ g O}}{1 \text{ mol O}} \right) = 176.00 \text{ g O}$$

$$\begin{aligned} \text{Mass percent} &= \frac{\text{total mass O}}{\text{molar mass of compound}}(100) = \frac{176.00 \text{ g O}}{246.09 \text{ g } \text{NaH}_2\text{PO}_4 \cdot 7\text{H}_2\text{O}}(100) \\ &= 71.51855 = \mathbf{71.52\% \text{ O}} \end{aligned}$$

- 3.18 Plan: Determine the formula and the molar mass of each compound. The formula gives the relative number of moles of each element present. Multiply the number of moles of each element by its molar mass to find the total mass of element in 1 mole of compound. Mass percent = $\frac{\text{total mass of element}}{\text{molar mass of compound}}(100)$.

Solution:

a) Strontium periodate is an ionic compound consisting of strontium ions, Sr^{2+} and periodate ions, IO_4^- . The formula of the compound is $\text{Sr}(\text{IO}_4)_2$.

$$\begin{aligned} \mathcal{M} \text{ of } \text{Sr}(\text{IO}_4)_2 &= (1 \times \mathcal{M} \text{ of Sr}) + (2 \times \mathcal{M} \text{ of I}) + (8 \times \mathcal{M} \text{ of O}) \\ &= (1 \times 87.62 \text{ g/mol Sr}) + (2 \times 126.9 \text{ g/mol I}) + (8 \times 16.00 \text{ g/mol O}) \\ &= 469.4 \text{ g/mol of } \text{Sr}(\text{IO}_4)_2 \end{aligned}$$

There are 2 moles of I in 1 mole of $\text{Sr}(\text{IO}_4)_2$.

$$\text{Mass (g) of I} = (2 \text{ mol I}) \left(\frac{126.9 \text{ g I}}{1 \text{ mol I}} \right) = 253.8 \text{ g I}$$

$$\text{Mass percent} = \frac{\text{total mass I}}{\text{molar mass of compound}} (100) = \frac{253.8 \text{ g I}}{469.4 \text{ g Sr(IO}_4)_2} (100) = 54.0690 = \mathbf{54.07\% \text{ I}}$$

b) Potassium permanganate is an ionic compound consisting of potassium ions, K^+ and permanganate ions, MnO_4^- . The formula of the compound is KMnO_4 .

$$\begin{aligned} \mathcal{M} \text{ of } \text{KMnO}_4 &= (1 \times \mathcal{M} \text{ of K}) + (1 \times \mathcal{M} \text{ of Mn}) + (4 \times \mathcal{M} \text{ of O}) \\ &= (1 \times 39.10 \text{ g/mol K}) + (1 \times 54.94 \text{ g/mol Mn}) + (4 \times 16.00 \text{ g/mol O}) \\ &= 158.04 \text{ g/mol of } \text{KMnO}_4 \end{aligned}$$

There is 1 mole of Mn in 1 mole of KMnO_4 .

$$\text{Mass (g) of Mn} = (1 \text{ mol Mn}) \left(\frac{54.94 \text{ g Mn}}{1 \text{ mol Mn}} \right) = 54.94 \text{ g Mn}$$

$$\text{Mass percent} = \frac{\text{total mass Mn}}{\text{molar mass of compound}} (100) = \frac{54.94 \text{ g Mn}}{158.04 \text{ g } \text{KMnO}_4} (100) = 34.76335 = \mathbf{34.76\% \text{ Mn}}$$

- 3.19 Plan: Determine the formula of cisplatin from the figure, and then calculate the molar mass from the formula. Divide the mass given by the molar mass to find moles of cisplatin. Since 1 mole of cisplatin contains 6 moles of hydrogen atoms, multiply the moles given by 6 to obtain moles of hydrogen and then multiply by Avogadro's number to obtain the number of atoms.

Solution:

The formula for cisplatin is $\text{Pt}(\text{Cl})_2(\text{NH}_3)_2$.

$$\begin{aligned} \mathcal{M} \text{ of } \text{Pt}(\text{Cl})_2(\text{NH}_3)_2 &= (1 \times \mathcal{M} \text{ of Pt}) + (2 \times \mathcal{M} \text{ of Cl}) + (2 \times \mathcal{M} \text{ of N}) + (6 \times \mathcal{M} \text{ of H}) \\ &= (1 \times 195.1 \text{ g/mol Pt}) + (2 \times 35.45 \text{ g/mol Cl}) + (2 \times 14.01 \text{ g/mol N}) + (6 \times 1.008 \text{ g/mol H}) \\ &= 300.1 \text{ g/mol of } \text{Pt}(\text{Cl})_2(\text{NH}_3)_2 \end{aligned}$$

$$\text{a) Moles of cisplatin} = (285.3 \text{ g cisplatin}) \left(\frac{1 \text{ mol cisplatin}}{300.1 \text{ g cisplatin}} \right) = 0.9506831 = \mathbf{0.9507 \text{ mol cisplatin}}$$

$$\text{b) Moles of H atoms} = (0.98 \text{ mol cisplatin}) \left(\frac{6 \text{ mol H}}{1 \text{ mol cisplatin}} \right) = 5.88 \text{ mol H atoms}$$

$$\text{Number of H atoms} = (5.88 \text{ mol H atoms}) \left(\frac{6.022 \times 10^{23} \text{ H atoms}}{1 \text{ mol H atoms}} \right) = 3.540936 \times 10^{24} = \mathbf{3.5 \times 10^{24} \text{ H atoms}}$$

- 3.20 Plan: Determine the molar mass of propane. Divide the given mass by the molar mass to find the moles. Since each mole of propane contains 3 moles of carbon, multiply the moles of propane by 3 to obtain moles of C atoms. Multiply the moles of C by its molar mass to obtain mass of carbon.

Solution:

a) The formula of propane is C_3H_8 .

$$\mathcal{M} \text{ of } \text{C}_3\text{H}_8 = (3 \times \mathcal{M} \text{ of C}) + (8 \times \mathcal{M} \text{ of H}) = (3 \times 12.01 \text{ g/mol C}) + (8 \times 1.008 \text{ g/mol H}) = 44.09 \text{ g/mol}$$

$$\text{Moles of } \text{C}_3\text{H}_8 = (85.5 \text{ g } \text{C}_3\text{H}_8) \left(\frac{1 \text{ mol } \text{C}_3\text{H}_8}{44.09 \text{ g } \text{C}_3\text{H}_8} \right) = 1.939215 = \mathbf{1.94 \text{ mol } \text{C}_3\text{H}_8}$$

$$\text{b) Moles of C} = (1.939215 \text{ mol } \text{C}_3\text{H}_8) \left(\frac{3 \text{ mol C}}{1 \text{ mol } \text{C}_3\text{H}_8} \right) = 5.817645 \text{ mol C}$$

$$\text{Mass (g) of C} = (5.817645 \text{ mol C}) \left(\frac{12.01 \text{ g C}}{1 \text{ mol C}} \right) = 69.86992 = \mathbf{69.9 \text{ g C}}$$

- 3.21 Plan: Determine the formula and the molar mass of each compound. The formula gives the relative number of moles of nitrogen present. Multiply the number of moles of nitrogen by its molar mass to find the total mass of nitrogen in 1 mole of compound. Divide the total mass of nitrogen by the molar mass of compound and multiply by 100 to determine mass percent. $\text{Mass percent} = \frac{(\text{mol N}) \times (\text{molar mass N})}{\text{molar mass of compound}} (100)$. Then rank the values in order of decreasing mass percent N.

Solution:

Name	Formula	Molar Mass (g/mol)
Potassium nitrate	KNO_3	101.11
Ammonium nitrate	NH_4NO_3	80.05
Ammonium sulfate	$(\text{NH}_4)_2\text{SO}_4$	132.15
Urea	$\text{CO}(\text{NH}_2)_2$	60.06

$$\text{Mass \% N in potassium nitrate} = \frac{(1 \text{ mol N})(14.01 \text{ g/mol N})}{101.11 \text{ g/mol}} \times 100 = 13.856196 = \mathbf{13.86\% \text{ N}}$$

$$\text{Mass \% N in ammonium nitrate} = \frac{(2 \text{ mol N})(14.01 \text{ g/mol N})}{80.05 \text{ g/mol}} \times 100 = 35.003123 = \mathbf{35.00\% \text{ N}}$$

$$\text{Mass \% N in ammonium sulfate} = \frac{(2 \text{ mol N})(14.01 \text{ g/mol N})}{132.15 \text{ g/mol}} \times 100 = 21.20318 = \mathbf{21.20\% \text{ N}}$$

$$\text{Mass \% N in urea} = \frac{(2 \text{ mol N})(14.01 \text{ g/mol N})}{60.06 \text{ g/mol}} \times 100 = 46.6533 = \mathbf{46.65\% \text{ N}}$$

Rank is $\text{CO}(\text{NH}_2)_2 > \text{NH}_4\text{NO}_3 > (\text{NH}_4)_2\text{SO}_4 > \text{KNO}_3$

- 3.22 Plan: The volume must be converted from cubic feet to cubic centimeters. The volume and the density will give the mass of galena which is then divided by molar mass to obtain moles. Part b) requires a conversion from cubic decimeters to cubic centimeters. The density allows a change from volume in cubic centimeters to mass which is then divided by the molar mass to obtain moles; the amount in moles is multiplied by Avogadro's number to obtain formula units of PbS which is also the number of Pb atoms due to the 1:1 PbS:Pb mole ratio.

Solution:

Lead(II) sulfide is composed of Pb^{2+} and S^{2-} ions and has a formula of PbS.

$$\mathcal{M} \text{ of PbS} = (1 \times \mathcal{M} \text{ of Pb}) + (1 \times \mathcal{M} \text{ of S}) = (1 \times 207.2 \text{ g/mol Pb}) + (1 \times 32.07 \text{ g/mol S}) = 239.3 \text{ g/mol}$$

$$\text{a) Volume (cm}^3\text{)} = (1.00 \text{ ft}^3 \text{ PbS}) \left(\frac{(12 \text{ in})^3}{(1 \text{ ft})^3} \right) \left(\frac{(2.54 \text{ cm})^3}{(1 \text{ in})^3} \right) = 28316.85 \text{ cm}^3$$

$$\text{Mass (g) of PbS} = (28316.85 \text{ cm}^3 \text{ PbS}) \left(\frac{7.46 \text{ g PbS}}{1 \text{ cm}^3} \right) = 211243.7 \text{ g PbS}$$

$$\text{Moles of PbS} = (211243.7 \text{ g PbS}) \left(\frac{1 \text{ mol PbS}}{239.3 \text{ g PbS}} \right) = 882.7568 = \mathbf{883 \text{ mol PbS}}$$

$$\text{b) Volume (cm}^3\text{)} = (1.00 \text{ dm}^3 \text{ PbS}) \left(\frac{(0.1 \text{ m})^3}{(1 \text{ dm})^3} \right) \left(\frac{(1 \text{ cm})^3}{(10^{-2} \text{ m})^3} \right) = 1.00 \times 10^3 \text{ cm}^3$$

$$\text{Mass (g) of PbS} = (1.00 \times 10^3 \text{ cm}^3 \text{ PbS}) \left(\frac{7.46 \text{ g PbS}}{1 \text{ cm}^3} \right) = 7460 \text{ g PbS}$$

$$\text{Moles of PbS} = (7460 \text{ g PbS}) \left(\frac{1 \text{ mol PbS}}{239.3 \text{ g PbS}} \right) = 31.17426 \text{ mol PbS}$$

$$\text{Moles of Pb} = (31.17426 \text{ mol PbS}) \left(\frac{1 \text{ mol Pb}}{1 \text{ mol PbS}} \right) = 31.17426 \text{ mol Pb}$$

Number of lead atoms =

$$(31.17426 \text{ mol Pb}) \left(\frac{6.022 \times 10^{23} \text{ Pb atoms}}{1 \text{ mol Pb}} \right) = 1.87731 \times 10^{25} = \mathbf{1.88 \times 10^{25} \text{ Pb atoms}}$$

- 3.23 Plan: If the molecular formula for hemoglobin (Hb) were known, the number of Fe^{2+} ions in a molecule of hemoglobin could be calculated. It is possible to calculate the mass of iron from the percentage of iron and the molar mass of the compound. Assuming you have 1 mole of hemoglobin, take 0.33% of its molar mass as the mass of Fe in that 1 mole. Divide the mass of Fe by its molar mass to find moles of Fe in 1 mole of hemoglobin which is also the number of ions in 1 molecule.

Solution:

$$\text{Mass of Fe} = \left(\frac{0.33\% \text{ Fe}}{100\% \text{ Hb}} \right) \left(\frac{6.8 \times 10^4 \text{ g}}{\text{mol}} \right) = 224.4 \text{ g Fe}$$

$$\text{Moles of Fe} = (224.4 \text{ g Fe}) \left(\frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} \right) = 4.0179 = 4.0 \text{ mol Fe}^{2+}/\text{mol Hb}$$

Thus, there are **4 Fe^{2+} /molecule Hb**.

- 3.24 Plan: Remember that the molecular formula tells the *actual* number of moles of each element in one mole of compound.

Solution:

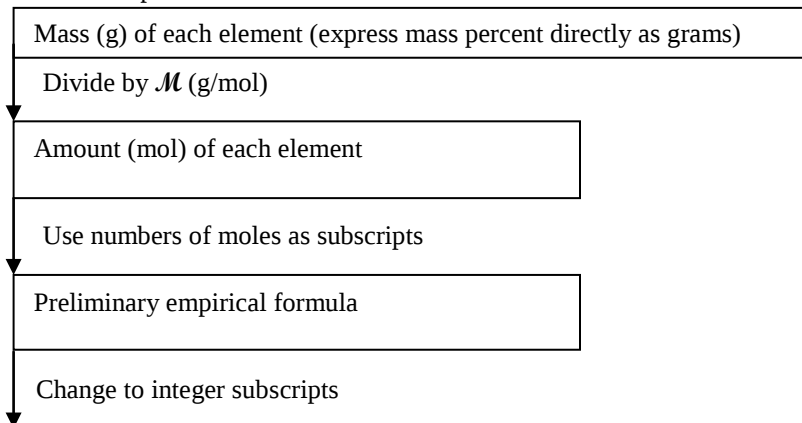
a) No, this information does not allow you to obtain the molecular formula. You can obtain the empirical formula from the number of moles of each type of atom in a compound, but not the molecular formula.

b) Yes, you can obtain the molecular formula from the mass percentages and the total number of atoms.

Plan:

- 1) Assume a 100.0 g sample and convert masses (from the mass % of each element) to moles using molar mass.
- 2) Identify the element with the lowest number of moles and use this number to divide into the number of moles for each element. You now have at least one elemental mole ratio (the one with the smallest number of moles) equal to 1.00 and the remaining mole ratios that are larger than one.
- 3) Examine the numbers to determine if they are whole numbers. If not, multiply each number by a whole-number factor to get whole numbers for each element. You will have to use some judgment to decide when to round. Write the empirical formula using these whole numbers.
- 4) Check the total number of atoms in the empirical formula. If it equals the total number of atoms given then the empirical formula is also the molecular formula. If not, then divide the total number of atoms given by the total number of atoms in the empirical formula. This should give a whole number. Multiply the number of atoms of each element in the empirical formula by this whole number to get the molecular formula. If you do not get a whole number when you divide, return to step 3 and revise how you multiplied and rounded to get whole numbers for each element.

Roadmap:



Empirical formula

Divide total number of atoms in molecule by the number of atoms in the empirical formula and multiply the empirical formula by that factor

Molecular formula

c) Yes, you can determine the molecular formula from the mass percent and the number of atoms of one element in a compound. Plan:

1) Follow steps 1–3 in part b).

2) Compare the number of atoms given for the one element to the number in the empirical formula. Determine the factor the number in the empirical formula must be multiplied by to obtain the given number of atoms for that element. Multiply the empirical formula by this number to get the molecular formula.

Roadmap:

(Same first three steps as in b).

Empirical formula

Divide the number of atoms of the one element in the molecule by the number of atoms of that element in the empirical formula and multiply the empirical formula by that factor

Molecular formula

d) No, the mass % will only lead to the empirical formula.

e) Yes, a structural formula shows all the atoms in the compound. Plan: Count the number of atoms of each type of element and record as the number for the molecular formula.

Roadmap:

Structural formula

Count the number of atoms of each element and use these numbers as subscripts

Molecular formula

3.25 Plan: Examine the number of atoms of each type in the compound. Divide all atom numbers by the common factor that results in the lowest whole-number values. Add the molar masses of the atoms to obtain the empirical formula mass.

Solution:

a) C_2H_4 has a ratio of 2 carbon atoms to 4 hydrogen atoms, or 2:4. This ratio can be reduced to 1:2, so that the empirical formula is **CH_2** . The empirical formula mass is $12.01 \text{ g/mol C} + 2(1.008 \text{ g/mol H}) = \mathbf{14.03 \text{ g/mol}}$.

b) The ratio of atoms is 2:6:2, or 1:3:1. The empirical formula is **CH_3O** and its empirical formula mass is $12.01 \text{ g/mol C} + 3(1.008 \text{ g/mol H}) + 16.00 \text{ g/mol O} = \mathbf{31.03 \text{ g/mol}}$.

c) Since, the ratio of elements cannot be further reduced, the molecular formula and empirical formula are the same, **N_2O_5** . The formula mass is $2(14.01 \text{ g/mol N}) + 5(16.00 \text{ g/mol O}) = \mathbf{108.02 \text{ g/mol}}$.

d) The ratio of elements is 3 atoms of barium to 2 atoms of phosphorus to 8 atoms of oxygen, or 3:2:8. This ratio cannot be further reduced, so the empirical formula is also **$Ba_3(PO_4)_2$** , with a formula mass of $3(137.3 \text{ g/mol Ba}) + 2(30.97 \text{ g/mol P}) + 8(16.00 \text{ g/mol O}) = \mathbf{601.8 \text{ g/mol}}$.

e) The ratio of atoms is 4:16, or 1:4. The empirical formula is **TeI_4** , and the formula mass is $127.6 \text{ g/mol Te} + 4(126.9 \text{ g/mol I}) = \mathbf{635.2 \text{ g/mol}}$.

- 3.26 **Plan:** Examine the number of atoms of each type in the compound. Divide all atom numbers by the common factor that results in the lowest whole-number values. Add the molar masses of the atoms to obtain the empirical formula mass.

Solution:

- a) C_4H_8 has a ratio of 4 carbon atoms to 8 hydrogen atoms, or 4:8. This ratio can be reduced to 1:2, so that the empirical formula is CH_2 . The empirical formula mass is $12.01 \text{ g/mol C} + 2(1.008 \text{ g/mol H}) = \mathbf{14.03 \text{ g/mol}}$.
 b) $C_3H_6O_3$ has a ratio of atoms of 3:6:3, or 1:2:1. The empirical formula is CH_2O and its empirical formula mass is $12.01 \text{ g/mol C} + 2(1.008 \text{ g/mol H}) + 16.00 \text{ g/mol O} = \mathbf{30.03 \text{ g/mol}}$.
 c) P_4O_{10} has a ratio of 4 P atoms to 10 O atoms, or 4:10. This ratio can be reduced to 2:5, so that the empirical formula is P_2O_5 . The empirical formula mass is $2(30.97 \text{ g/mol P}) + 5(16.00 \text{ g/mol O}) = \mathbf{141.94 \text{ g/mol}}$.
 d) $Ga_2(SO_4)_3$ has a ratio of 2 atoms of gallium to 3 atoms of sulfur to 12 atoms of oxygen, or 2:3:12. This ratio cannot be further reduced, so the empirical formula is also $Ga_2(SO_4)_3$, with a formula mass of $2(69.72 \text{ g/mol Ga}) + 3(32.07 \text{ g/mol S}) + 12(16.00 \text{ g/mol O}) = \mathbf{427.6 \text{ g/mol}}$.
 e) Al_2Br_6 has a ratio of atoms of 2:6, or 1:3. The empirical formula is $AlBr_3$, and the formula mass is $26.98 \text{ g/mol Al} + 3(79.90 \text{ g/mol Br}) = \mathbf{266.7 \text{ g/mol}}$.

- 3.27 **Plan:** Determine the molar mass of each empirical formula. The subscripts in the molecular formula are whole-number multiples of the subscripts in the empirical formula. To find this whole number, divide the molar mass of the compound by its empirical formula mass. Multiply each subscript in the empirical formula by the whole number.

Solution:

Only approximate whole-number values are needed.

- a) CH_2 has empirical mass equal to $12.01 \text{ g/mol C} + 2(1.008 \text{ g/mol H}) = 14.03 \text{ g/mol}$

$$\text{Whole-number multiple} = \frac{\text{molar mass of compound}}{\text{empirical formula mass}} = \left(\frac{42.08 \text{ g/mol}}{14.03 \text{ g/mol}} \right) = 3$$

Multiplying the subscripts in CH_2 by 3 gives C_3H_6 .

- b) NH_2 has empirical mass equal to $14.01 \text{ g/mol N} + 2(1.008 \text{ g/mol H}) = 16.03 \text{ g/mol}$

$$\text{Whole-number multiple} = \frac{\text{molar mass of compound}}{\text{empirical formula mass}} = \left(\frac{32.05 \text{ g/mol}}{16.03 \text{ g/mol}} \right) = 2$$

Multiplying the subscripts in NH_2 by 2 gives N_2H_4 .

- c) NO_2 has empirical mass equal to $14.01 \text{ g/mol N} + 2(16.00 \text{ g/mol O}) = 46.01 \text{ g/mol}$

$$\text{Whole-number multiple} = \frac{\text{molar mass of compound}}{\text{empirical formula mass}} = \left(\frac{92.02 \text{ g/mol}}{46.01 \text{ g/mol}} \right) = 2$$

Multiplying the subscripts in NO_2 by 2 gives N_2O_4 .

- d) CHN has empirical mass equal to $12.01 \text{ g/mol C} + 1.008 \text{ g/mol H} + 14.01 \text{ g/mol N} = 27.03 \text{ g/mol}$

$$\text{Whole-number multiple} = \frac{\text{molar mass of compound}}{\text{empirical formula mass}} = \left(\frac{135.14 \text{ g/mol}}{27.03 \text{ g/mol}} \right) = 5$$

Multiplying the subscripts in CHN by 5 gives $C_5H_5N_5$.

- 3.28 **Plan:** Determine the molar mass of each empirical formula. The subscripts in the molecular formula are whole-number multiples of the subscripts in the empirical formula. To find this whole number, divide the molar mass of the compound by its empirical formula mass. Multiply each subscript in the empirical formula by the whole number.

Solution:

Only approximate whole-number values are needed.

- a) CH has empirical mass equal to $12.01 \text{ g/mol C} + 1.008 \text{ g/mol H} = 13.02 \text{ g/mol}$

$$\text{Whole-number multiple} = \frac{\text{molar mass of compound}}{\text{empirical formula mass}} = \left(\frac{78.11 \text{ g/mol}}{13.02 \text{ g/mol}} \right) = 6$$

Multiplying the subscripts in CH by 6 gives C_6H_6 .

b) $\text{C}_3\text{H}_6\text{O}_2$ has empirical mass equal to $3(12.01 \text{ g/mol C}) + 6(1.008 \text{ g/mol H}) + 2(16.00 \text{ g/mol O}) = 74.08 \text{ g/mol}$

$$\text{Whole-number multiple} = \frac{\text{molar mass of compound}}{\text{empirical formula mass}} = \left(\frac{74.08 \text{ g/mol}}{74.08 \text{ g/mol}} \right) = 1$$

Multiplying the subscripts in $\text{C}_3\text{H}_6\text{O}_2$ by 1 gives **$\text{C}_3\text{H}_6\text{O}_2$** .

c) HgCl has empirical mass equal to $200.6 \text{ g/mol Hg} + 35.45 \text{ g/mol Cl} = 236.0 \text{ g/mol}$

$$\text{Whole-number multiple} = \frac{\text{molar mass of compound}}{\text{empirical formula mass}} = \left(\frac{472.1 \text{ g/mol}}{236.0 \text{ g/mol}} \right) = 2$$

Multiplying the subscripts in HgCl by 2 gives **Hg_2Cl_2** .

d) $\text{C}_7\text{H}_4\text{O}_2$ has empirical mass equal to $7(12.01 \text{ g/mol C}) + 4(1.008 \text{ g/mol H}) + 2(16.00 \text{ g/mol O}) = 120.10 \text{ g/mol}$

$$\text{Whole-number multiple} = \frac{\text{molar mass of compound}}{\text{empirical formula mass}} = \left(\frac{240.20 \text{ g/mol}}{120.10 \text{ g/mol}} \right) = 2$$

Multiplying the subscripts in $\text{C}_7\text{H}_4\text{O}_2$ by 2 gives **$\text{C}_{14}\text{H}_8\text{O}_4$** .

- 3.29 **Plan:** The empirical formula is the smallest whole-number ratio of the atoms or moles in a formula. All data must be converted to moles of an element by dividing mass by the molar mass. Divide each mole number by the smallest mole number to convert the mole ratios to whole numbers.

Solution:

a) 0.063 mol Cl and 0.22 mol O: preliminary formula is $\text{Cl}_{0.063}\text{O}_{0.22}$

Converting to integer subscripts (dividing all by the smallest subscript):

$$\frac{\text{Cl}_{0.063}\text{O}_{0.22}}{0.063 \quad 0.063} \rightarrow \text{Cl}_1\text{O}_{3.5}$$

The formula is $\text{Cl}_1\text{O}_{3.5}$, which in whole numbers (x 2) is **Cl_2O_7** .

b) Find moles of elements by dividing by molar mass:

$$\text{Moles of Si} = (2.45 \text{ g Si}) \left(\frac{1 \text{ mol Si}}{28.09 \text{ g Si}} \right) = 0.08722 \text{ mol Si}$$

$$\text{Moles of Cl} = (12.4 \text{ g Cl}) \left(\frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} \right) = 0.349788 \text{ mol Cl}$$

Preliminary formula is $\text{Si}_{0.08722}\text{Cl}_{0.349788}$

Converting to integer subscripts (dividing all by the smallest subscript):

$$\frac{\text{Si}_{0.08722}\text{Cl}_{0.349788}}{0.08722 \quad 0.349788} \rightarrow \text{Si}_1\text{Cl}_4$$

The empirical formula is **SiCl_4** .

c) Assume a 100 g sample and convert the masses to moles by dividing by the molar mass:

$$\text{Moles of C} = (100 \text{ g}) \left(\frac{27.3 \text{ parts C by mass}}{100 \text{ parts by mass}} \right) \left(\frac{1 \text{ mol C}}{12.01 \text{ g C}} \right) = 2.2731 \text{ mol C}$$

$$\text{Moles of O} = (100 \text{ g}) \left(\frac{72.7 \text{ parts O by mass}}{100 \text{ parts by mass}} \right) \left(\frac{1 \text{ mol O}}{16.00 \text{ g O}} \right) = 4.5438 \text{ mol O}$$

Preliminary formula is $\text{C}_{2.2731}\text{O}_{4.5438}$

Converting to integer subscripts (dividing all by the smallest subscript):

$$\frac{\text{C}_{2.2731}\text{O}_{4.5438}}{2.2731 \quad 2.2731} \rightarrow \text{C}_1\text{O}_2$$

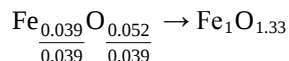
The empirical formula is **CO_2** .

- 3.30 **Plan:** The empirical formula is the smallest whole-number ratio of the atoms or moles in a formula. All data must be converted to moles of an element by dividing mass by the molar mass. Divide each mole number by the smallest mole number to convert the mole ratios to whole numbers.

Solution:

a) 0.039 mol Fe and 0.052 mol O: preliminary formula is $\text{Fe}_{0.039}\text{O}_{0.052}$

Converting to integer subscripts (dividing all by the smallest subscript):



The formula is $\text{Fe}_1\text{O}_{1.33}$, which in whole numbers (x 3) is **Fe_3O_4** .

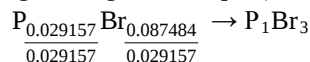
b) Find moles of elements by dividing by molar mass:

$$\text{Moles of P} = (0.903 \text{ g P}) \left(\frac{1 \text{ mol P}}{30.97 \text{ g P}} \right) = 0.029157 \text{ mol P}$$

$$\text{Moles of Br} = (6.99 \text{ g Br}) \left(\frac{1 \text{ mol Br}}{79.90 \text{ g Br}} \right) = 0.087484 \text{ mol Br}$$

Preliminary formula is $\text{P}_{0.029157}\text{Br}_{0.087484}$

Converting to integer subscripts (dividing all by the smallest subscript):



The empirical formula is **PBr_3** .

c) Assume a 100 g sample and convert the masses to moles by dividing by the molar mass:

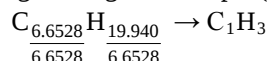
79.9% C and $100 - 79.9 = 20.1\%$ H

$$\text{Moles of C} = (100 \text{ g}) \left(\frac{79.9 \text{ parts C by mass}}{100 \text{ parts by mass}} \right) \left(\frac{1 \text{ mol C}}{12.01 \text{ g C}} \right) = 6.6528 \text{ mol C}$$

$$\text{Moles of H} = (100 \text{ g}) \left(\frac{20.1 \text{ parts H by mass}}{100 \text{ parts by mass}} \right) \left(\frac{1 \text{ mol H}}{1.008 \text{ g H}} \right) = 19.940 \text{ mol H}$$

Preliminary formula is $\text{C}_{6.6528}\text{H}_{19.940}$

Converting to integer subscripts (dividing all by the smallest subscript):



The empirical formula is **CH_3** .

- 3.31 **Plan:** The moles of the metal are known, and the moles of fluorine atoms may be found in part a) from the M:F mole ratio in the compound formula. In part b), convert moles of F atoms to mass and subtract the mass of F from the mass of MF_2 to find the mass of M. In part c), divide the mass of M by moles of M to determine the molar mass of M which can be used to identify the element.

Solution:

a) Determine the moles of fluorine.

$$\text{Moles of F} = (0.600 \text{ mol M}) \left(\frac{2 \text{ mol F}}{1 \text{ mol M}} \right) = \mathbf{1.20 \text{ mol F}}$$

b) Determine the mass of M.

$$\text{Mass of F} = (1.20 \text{ mol F}) \left(\frac{19.00 \text{ g F}}{1 \text{ mol F}} \right) = 22.8 \text{ g F}$$

$$\text{Mass (g) of M} = \text{MF}_2(\text{g}) - \text{F}(\text{g}) = 46.8 \text{ g} - 22.8 \text{ g} = \mathbf{24.0 \text{ g M}}$$

c) The molar mass is needed to identify the element.

$$\text{Molar mass of M} = \frac{24.0 \text{ g M}}{0.600 \text{ mol M}} = 40.0 \text{ g/mol}$$

The metal with the closest molar mass to 40.0 g/mol is **calcium**.

- 3.32 **Plan:** The moles of the metal oxide are known, and the moles of oxygen atoms may be found in part a) from the compound:oxygen mole ratio in the compound formula. In part b), convert moles of O atoms to mass and subtract the mass of O from the mass of M_2O_3 to find the mass of M. In part c), find moles of M from the compound:M mole ratio and divide the mass of M by moles of M to determine the molar mass of M which can be used to identify the element.

Solution:

a) Determine the moles of oxygen.

$$\text{Moles of O} = (0.370 \text{ mol M}_2\text{O}_3) \left(\frac{3 \text{ mol O}}{1 \text{ mol M}_2\text{O}_3} \right) = \mathbf{1.11 \text{ mol O}}$$

b) Determine the mass of M.

$$\text{Mass of O} = (1.11 \text{ mol O}) \left(\frac{16.00 \text{ g O}}{1 \text{ mol O}} \right) = 17.76 \text{ g O}$$

$$\text{Mass(g) of M} = \text{M}_2\text{O}_3(\text{g}) - \text{O(g)} = 55.4 \text{ g (M + O)} - 17.76 = 37.64 = \mathbf{37.6 \text{ g M}}$$

c) First, the number of moles of M must be calculated.

$$\text{Moles M} = (0.370 \text{ mol M}_2\text{O}_3) \left(\frac{2 \text{ mol M}}{1 \text{ mol M}_2\text{O}_3} \right) = 0.740 \text{ mol M}$$

The molar mass is needed to identify the element.

$$\text{Molar mass of M} = \frac{37.6 \text{ g M}}{0.740 \text{ mol M}} = 50.86 \text{ g/mol}$$

The metal with the closest molar mass to 50.9 g/mol is **vanadium**.

- 3.33 Plan: The empirical formula is the smallest whole-number ratio of the atoms or moles in a formula. Assume 100 grams of cortisol so the percentages are numerically equivalent to the masses of each element. Convert each of the masses to moles by dividing by the molar mass of each element involved. Divide each mole number by the smallest mole number to convert the mole ratios to whole numbers. The subscripts in the molecular formula are whole-number multiples of the subscripts in the empirical formula. To find this whole number, divide the molar mass of the compound by its empirical formula mass. Multiply each subscript in the empirical formula by the whole number.

Solution:

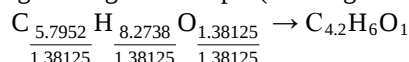
$$\text{Moles of C} = (69.6 \text{ g C}) \left(\frac{1 \text{ mol C}}{12.01 \text{ g C}} \right) = 5.7952 \text{ mol C}$$

$$\text{Moles of H} = (8.34 \text{ g H}) \left(\frac{1 \text{ mol H}}{1.008 \text{ g H}} \right) = 8.2738 \text{ mol H}$$

$$\text{Moles of O} = (22.1 \text{ g O}) \left(\frac{1 \text{ mol O}}{16.00 \text{ g O}} \right) = 1.38125 \text{ mol O}$$

Preliminary formula is $\text{C}_{5.7952}\text{H}_{8.2738}\text{O}_{1.38125}$

Converting to integer subscripts (dividing all by the smallest subscript):



The carbon value is not close enough to a whole number to round the value. The smallest number that 4.20 may be multiplied by to get close to a whole number is 5. (You may wish to prove this to yourself.) All three ratios need to be multiplied by five: $5(\text{C}_{4.2}\text{H}_6\text{O}_1) = \text{C}_{21}\text{H}_{30}\text{O}_5$.

The empirical formula mass is $= 21(12.01 \text{ g/mol C}) + 30(1.008 \text{ g/mol H}) + 5(16.00 \text{ g/mol O}) = 362.45 \text{ g/mol}$

$$\text{Whole-number multiple} = \frac{\text{molar mass of compound}}{\text{empirical formula mass}} = \left(\frac{362.47 \text{ g/mol}}{362.45 \text{ g/mol}} \right) = 1$$

The empirical formula mass and the molar mass given are the same, so the empirical and the molecular formulas are the same. The molecular formula is **C₂₁H₃₀O₅**.

- 3.34 Plan: In combustion analysis, finding the moles of carbon and hydrogen is relatively simple because all of the carbon present in the sample is found in the carbon of CO₂, and all of the hydrogen present in the sample is found in the hydrogen of H₂O. Convert the mass of CO₂ to moles and use the ratio between CO₂ and C to find the moles and mass of C present. Do the same to find the moles and mass of H from H₂O. The moles of oxygen are more difficult to find, because additional O₂ was added to cause the combustion reaction. Subtracting the masses of C and H from the mass of the sample gives the mass of O. Convert the mass of O to moles of O. Take the moles of C, H, and O and divide by the smallest value to convert to whole numbers to get the empirical formula.

Determine the empirical formula mass and compare it to the molar mass given in the problem to see how the empirical and molecular formulas are related. Finally, determine the molecular formula.

Solution:

$$\text{Moles of C} = (0.449 \text{ g CO}_2) \left(\frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \right) \left(\frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \right) = 0.010202 \text{ mol C}$$

$$\text{Mass (g) of C} = (0.010202 \text{ mol C}) \left(\frac{12.01 \text{ g C}}{1 \text{ mol C}} \right) = 0.122526 \text{ g C}$$

$$\text{Moles of H} = (0.184 \text{ g H}_2\text{O}) \left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) \left(\frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \right) = 0.020422 \text{ mol H}$$

$$\text{Mass (g) of H} = (0.020422 \text{ mol H}) \left(\frac{1.008 \text{ g H}}{1 \text{ mol H}} \right) = 0.020585 \text{ g H}$$

$$\begin{aligned} \text{Mass (g) of O} &= \text{Sample mass} - (\text{mass of C} + \text{mass of H}) \\ &= 0.1595 \text{ g} - (0.122526 \text{ g C} + 0.020585 \text{ g H}) = 0.016389 \text{ g O} \end{aligned}$$

$$\text{Moles of O} = (0.016389 \text{ g O}) \left(\frac{1 \text{ mol O}}{16.00 \text{ g O}} \right) = 0.0010243 \text{ mol O}$$

$$\text{Preliminary formula} = \text{C}_{0.010202} \text{H}_{0.020422} \text{O}_{0.0010243}$$

Converting to integer subscripts (dividing all by the smallest subscript):

$$\frac{\text{C}_{0.010202}}{0.0010243} \frac{\text{H}_{0.020422}}{0.0010243} \frac{\text{O}_{0.0010243}}{0.0010243} \rightarrow \text{C}_{10} \text{H}_{20} \text{O}_1$$

$$\text{Empirical formula} = \text{C}_{10} \text{H}_{20} \text{O}$$

$$\text{Empirical formula mass} = 10(12.01 \text{ g/mol C}) + 20(1.008 \text{ g/mol H}) + 1(16.00 \text{ g/mol O}) = 156.26 \text{ g/mol}$$

The empirical formula mass is the same as the given molar mass so the empirical and molecular formulas are the same. The molecular formula is **C₁₀H₂₀O**.

- 3.35 Students I and II are incorrect. Both students changed a given formula. Only coefficients should be changed when balancing; subscripts cannot be changed. Student I failed to identify the product correctly, writing AlCl₂ instead of AlCl₃. Student II used atomic chlorine instead of molecular chlorine as a reactant. Student III followed the correct process, changing only coefficients.

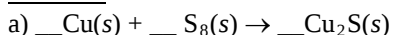
- 3.36 Plan: Examine the diagram and label each formula. We will use A for red atoms and B for green atoms.

Solution:

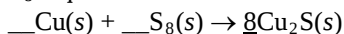
The reaction shows A₂ and B₂ diatomic molecules forming AB molecules. Equal numbers of A₂ and B₂ combine to give twice as many molecules of AB. Thus, the reaction is A₂ + B₂ → 2 AB. This is the balanced equation in **b**.

- 3.37 Plan: Balancing is a trial-and-error procedure. Balance one element at a time, placing coefficients where needed to have the same number of atoms of a particular element on each side of the equation. The smallest whole-number coefficients should be used.

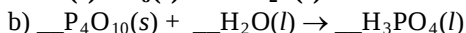
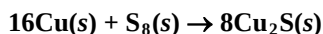
Solution:



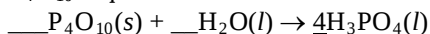
Balance the S first, because there is an obvious deficiency of S on the right side of the equation. The 8 S atoms in S₈ require the coefficient 8 in front of Cu₂S:



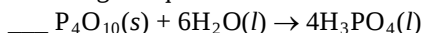
Then balance the Cu. The 16 Cu atoms in Cu₂S require the coefficient 16 in front of Cu:



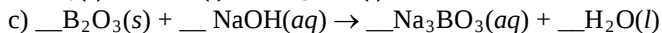
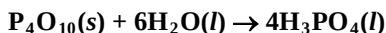
Balance the P first, because there is an obvious deficiency of P on the right side of the equation. The 4 P atoms in P₄O₁₀ require a coefficient of 4 in front of H₃PO₄:



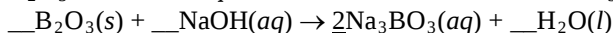
Balance the H next, because H is present in only one reactant and only one product. The 12 H atoms in 4H₃PO₄ on the right require a coefficient of 6 in front of H₂O:



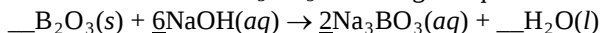
Balance the O last, because it appears in both reactants and is harder to balance. There are 16 O atoms on each side:



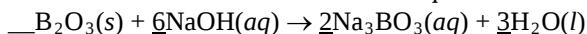
Balance oxygen last because it is present in more than one place on each side of the reaction. The 2 B atoms in B_2O_3 on the left require a coefficient of 2 in front of Na_3BO_3 on the right:



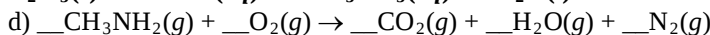
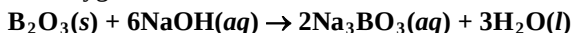
The 6 Na atoms in $2Na_3BO_3$ on the right require a coefficient of 6 in front of NaOH on the left:



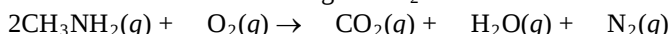
The 6 H atoms in $6NaOH$ on the left require a coefficient of 3 in front of H_2O on the right:



The oxygen is now balanced with 9 O atoms on each side:



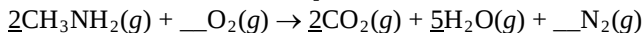
There are 2 N atoms on the right in N_2 so a coefficient of 2 is required in front of CH_3NH_2 on the left:



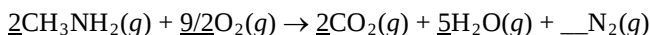
There are now 10 H atoms in $2CH_3NH_2$ on the left so a coefficient of 5 is required in front of H_2O on the right:



The 2 C atoms on the left require a coefficient of 2 in front of CO_2 on the right:



The 9 O atoms on the right (4 O atoms in $2CO_2$ plus 5 in $5H_2O$) require a coefficient of 9/2 in front of O_2 on the left:

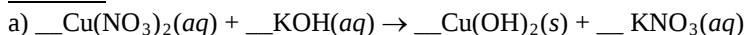


Multiply all coefficients by 2 to obtain whole numbers:

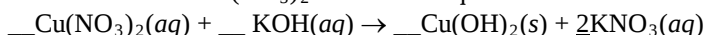


- 3.38 Plan: Balancing is a trial-and-error procedure. Balance one element at a time, placing coefficients where needed to have the same number of atoms of a particular element on each side of the equation. The smallest whole-number coefficients should be used.

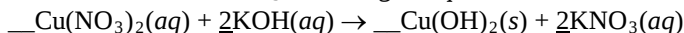
Solution:



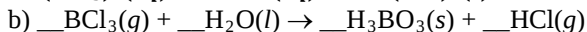
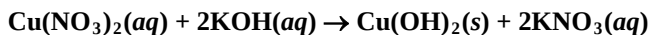
The 2 N atoms in $Cu(NO_3)_2$ on the left require a coefficient of 2 in front of KNO_3 on the right:



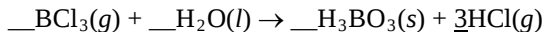
The 2 K atoms in $2KNO_3$ on the right require a coefficient of 2 in front of KOH on the left:



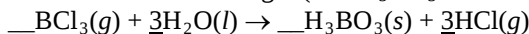
There are 8 O atoms and 2 H atoms on each side:



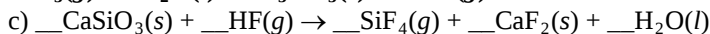
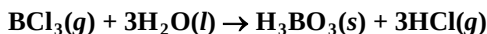
The 3 Cl atoms in BCl_3 on the left require a coefficient of 3 in front of HCl on the right:



The 6 H atoms on the right (3 in H_3BO_3 and 3 in HCl) require a coefficient of 3 in front of H_2O on the left:



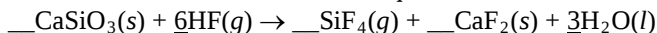
There are 3 O atoms and 1 B atom on each side:



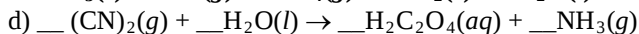
The 6 F atoms on the right (4 in SiF_4 and 2 in CaF_2) require a coefficient of 6 in front of HF on the left:



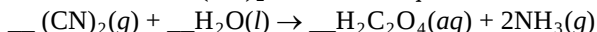
The 6 H atoms in $6HF$ on the left require a coefficient of 3 in front of H_2O on the right:



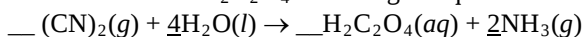
There are 1 Ca atom, 1 Si atom, and 3 O atoms on each side:



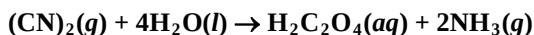
The 2 N atoms in $(CN)_2$ on the left requires a coefficient of 2 in front of NH_3 on the left:



The 4 O atoms in $\text{H}_2\text{C}_2\text{O}_4$ on the right requires a coefficient of 4 in front of H_2O on the right:



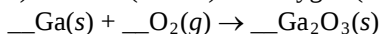
There are 2 C atoms and 8 H atoms on each side:



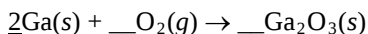
- 3.39 **Plan:** The names must first be converted to chemical formulas. Balancing is a trial-and-error procedure. Balance one element at a time, placing coefficients where needed to have the same number of atoms of a particular element on each side of the equation. The smallest whole-number coefficients should be used. Remember that oxygen is diatomic.

Solution:

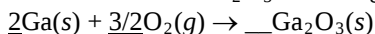
a) Gallium (a solid) and oxygen (a gas) are reactants and solid gallium(III) oxide is the only product:



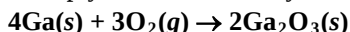
A coefficient of 2 in front of Ga on the left is needed to balance the 2 Ga atoms in Ga_2O_3 :



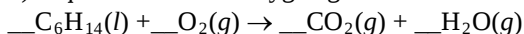
The 3 O atoms in Ga_2O_3 on the right require a coefficient of 3/2 in front of O_2 on the left:



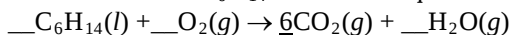
Multiply all coefficients by 2 to obtain whole numbers:



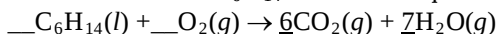
b) Liquid hexane and oxygen gas are the reactants while carbon dioxide gas and gaseous water are the products:



The 6 C atoms in C_6H_{14} on the left require a coefficient of 6 in front of CO_2 on the right:

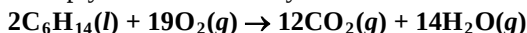


The 14 H atoms in C_6H_{14} on the left require a coefficient of 7 in front of H_2O on the right:

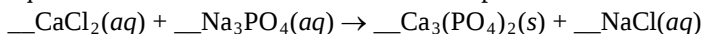


The 19 O atoms on the right (12 in 6CO_2 and 7 in $7\text{H}_2\text{O}$) require a coefficient of 19/2 in front of O_2 on the left:

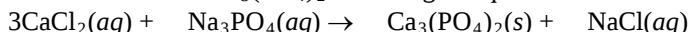
Multiply all coefficients by 2 to obtain whole numbers:



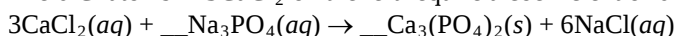
c) Aqueous solutions of calcium chloride and sodium phosphate are the reactants; solid calcium phosphate and an aqueous solution of sodium chloride are the products:



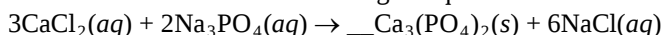
The 3 Ca atoms in $\text{Ca}_3(\text{PO}_4)_2$ on the right require a coefficient of 3 in front of CaCl_2 on the left:



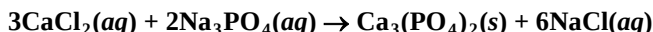
The 6 Cl atoms in 3CaCl_2 on the left require a coefficient of 6 in front of NaCl on the right:



The 6 Na atoms in 6NaCl on the right require a coefficient of 2 in front of Na_3PO_4 on the left:



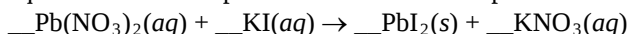
There are now 2 P atoms on each side:



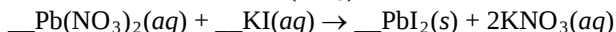
- 3.40 **Plan:** The names must first be converted to chemical formulas. Balancing is a trial-and-error procedure. Balance one element at a time, placing coefficients where needed to have the same number of atoms of a particular element on each side of the equation. The smallest whole-number coefficients should be used. Remember that oxygen is diatomic.

Solution:

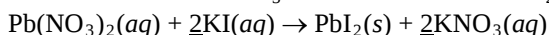
a) Aqueous solutions of lead(II) nitrate and potassium iodide are the reactants; solid lead(II) iodide and an aqueous solution of potassium nitrate are the products:



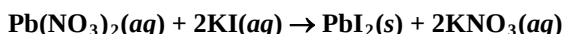
There are 2 N atoms in $\text{Pb}(\text{NO}_3)_2$ on the left so a coefficient of 2 is required in front of KNO_3 on the right:



The 2 K atoms in 2KNO_3 and the 2 I atoms in PbI_2 on the right require a coefficient of 2 in front of KI on the left:



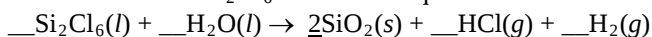
There are now 6 O atoms on each side:



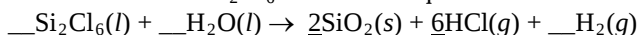
b) Liquid disilicon hexachloride and water are the reactants and solid silicon dioxide, hydrogen chloride gas and hydrogen gas are the products:



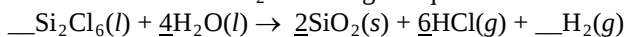
The 2 Si atoms in Si_2Cl_6 on the left require a coefficient of 2 in front of SiO_2 on the right:



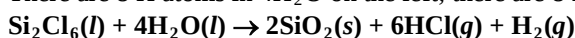
The 6 Cl atoms in Si_2Cl_6 on the left require a coefficient of 6 in front of HCl on the right:



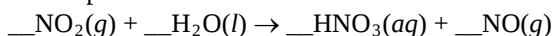
The 4 O atoms in 2SiO_2 on the right require a coefficient of 4 in front of H_2O on the left.



There are 8 H atoms in $4\text{H}_2\text{O}$ on the left; there are 8 H atoms on the right (6 in 6HCl and 2 in H_2):

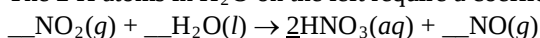


c) Nitrogen dioxide and water are the reactants and an aqueous solution of nitric acid and nitrogen monoxide gas are the products:

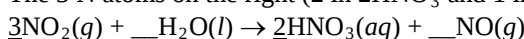


Start with hydrogen it occurs in only one reactant and one product:

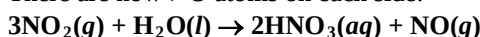
The 2 H atoms in H_2O on the left require a coefficient of 2 in front of HNO_3 on the right:



The 3 N atoms on the right (2 in 2HNO_3 and 1 in NO) require a coefficient of 3 in front of NO_2 on the left;



There are now 7 O atoms on each side:



- 3.41 Plan: Write a balanced chemical reaction to obtain the mole ratio between the reactants. Compare the number of particles of each reactant with the mole ratio to find the limiting reactant. Use the limiting reactant to calculate the number of product molecules that will form.

Solution:

a) The reaction is $\text{A}_2 + \text{B}_2 \rightarrow \text{AB}_3$ or $\text{A}_2 + 3\text{B}_2 \rightarrow 2\text{AB}_3$. The mole ratio between A_2 and B_2 is 1:3. Three times as many B_2 molecules are required as you have of A_2 molecules. With 3 A_2 molecules present, $3 \times 3 = 9$ B_2 molecules would be required. Since you have only 6 B_2 molecules, **B_2 is the limiting reagent.**

b) The balanced equation shows that 2AB_3 molecules are produced for every 3 B_2 molecules that react. Use the 3:2 mole ratio between the limiting reactant, B_2 , and AB_3 :

$$\text{Number of molecules of product} = (6 \text{ B}_2 \text{ molecules}) \left(\frac{2 \text{ AB}_3 \text{ molecules}}{3 \text{ B}_2 \text{ molecules}} \right) = 4 \text{ AB}_3 \text{ molecules}$$

- 3.42 Plan: Convert the kilograms of oxygen to grams of oxygen and then moles of oxygen by dividing by its molar mass. Use the moles of oxygen and the mole ratio from the balanced chemical equation to determine the moles of KNO_3 required. Multiply the moles of KNO_3 by its molar mass to obtain the mass in grams.

Solution:

$$\text{a) Mass (g) of O}_2 = (56.6 \text{ kg O}_2) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) = 5.66 \times 10^4 \text{ g O}_2$$

$$\text{Moles of O}_2 = (5.66 \times 10^4 \text{ g O}_2) \left(\frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \right) = 1.76875 \times 10^3 \text{ mol O}_2$$

$$\text{Moles of KNO}_3 = (1.76875 \text{ mol O}_2) \left(\frac{4 \text{ mol KNO}_3}{5 \text{ mol O}_2} \right) = 1415 = 1.42 \times 10^3 \text{ mol KNO}_3$$

$$\text{b) Mass (g) of KNO}_3 = (1415 \text{ mol KNO}_3) \left(\frac{101.11 \text{ g KNO}_3}{1 \text{ mol KNO}_3} \right) = 143070.65 = 1.43 \times 10^5 \text{ g KNO}_3$$

Combining all steps gives:

$$\begin{aligned} \text{Mass (g) of KNO}_3 &= (56.6 \text{ kg O}_2) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left(\frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \right) \left(\frac{4 \text{ mol KNO}_3}{5 \text{ mol O}_2} \right) \left(\frac{101.11 \text{ g KNO}_3}{1 \text{ mol KNO}_3} \right) \\ &= 143070.65 = 1.43 \times 10^5 \text{ g KNO}_3 \end{aligned}$$

- 3.43 Plan: Convert mass of Cr_2S_3 to moles by dividing by its molar mass. Use the mole ratio between Cr_2S_3 and Cr_2O_3 from the balanced chemical equation to determine the moles of Cr_2O_3 required. Multiply the moles of Cr_2O_3 by its molar mass to obtain the mass in grams.

Solution:

$$\text{a) Moles of } \text{Cr}_2\text{S}_3 = (421 \text{ g } \text{Cr}_2\text{S}_3) \left(\frac{1 \text{ mol } \text{Cr}_2\text{S}_3}{200.21 \text{ g } \text{Cr}_2\text{S}_3} \right) = 2.102792 \text{ mol } \text{Cr}_2\text{S}_3$$

$$\text{Moles of } \text{Cr}_2\text{O}_3 = (2.102792 \text{ mol } \text{Cr}_2\text{S}_3) \left(\frac{1 \text{ mol } \text{Cr}_2\text{O}_3}{1 \text{ mol } \text{Cr}_2\text{S}_3} \right) = 2.102792 = \mathbf{2.10 \text{ mol } \text{Cr}_2\text{O}_3}$$

$$\text{b) Mass (g) of } \text{Cr}_2\text{O}_3 = (2.102792 \text{ mol } \text{Cr}_2\text{O}_3) \left(\frac{152.00 \text{ g } \text{Cr}_2\text{O}_3}{1 \text{ mol } \text{Cr}_2\text{O}_3} \right) = 319.624 = \mathbf{3.20 \times 10^2 \text{ g } \text{Cr}_2\text{O}_3}$$

Combining all steps gives:

$$\begin{aligned} \text{Mass (g) of } \text{Cr}_2\text{O}_3 &= (421 \text{ g } \text{Cr}_2\text{S}_3) \left(\frac{1 \text{ mol } \text{Cr}_2\text{S}_3}{200.21 \text{ g } \text{Cr}_2\text{S}_3} \right) \left(\frac{1 \text{ mol } \text{Cr}_2\text{O}_3}{1 \text{ mol } \text{Cr}_2\text{S}_3} \right) \left(\frac{152.00 \text{ g } \text{Cr}_2\text{O}_3}{1 \text{ mol } \text{Cr}_2\text{O}_3} \right) \\ &= 319.624 = \mathbf{3.20 \times 10^2 \text{ g } \text{Cr}_2\text{O}_3} \end{aligned}$$

- 3.44 Plan: First, balance the equation. Convert the grams of diborane to moles of diborane by dividing by its molar mass. Use mole ratios from the balanced chemical equation to determine the moles of the products. Multiply the mole amount of each product by its molar mass to obtain mass in grams.

Solution:

The balanced equation is: $\text{B}_2\text{H}_6(\text{g}) + 6\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{H}_3\text{BO}_3(\text{s}) + 6\text{H}_2(\text{g})$.

$$\text{Moles of } \text{B}_2\text{H}_6 = (43.82 \text{ g } \text{B}_2\text{H}_6) \left(\frac{1 \text{ mol } \text{B}_2\text{H}_6}{27.67 \text{ g } \text{B}_2\text{H}_6} \right) = 1.583665 \text{ mol } \text{B}_2\text{H}_6$$

$$\text{Moles of } \text{H}_3\text{BO}_3 = (1.583665 \text{ mol } \text{B}_2\text{H}_6) \left(\frac{2 \text{ mol } \text{H}_3\text{BO}_3}{1 \text{ mol } \text{B}_2\text{H}_6} \right) = 3.16733 \text{ mol } \text{H}_3\text{BO}_3$$

$$\text{Mass (g) of } \text{H}_3\text{BO}_3 = (3.16733 \text{ mol } \text{H}_3\text{BO}_3) \left(\frac{61.83 \text{ g } \text{H}_3\text{BO}_3}{1 \text{ mol } \text{H}_3\text{BO}_3} \right) = 195.83597 = \mathbf{195.8 \text{ g } \text{H}_3\text{BO}_3}$$

Combining all steps gives:

$$\begin{aligned} \text{Mass (g) of } \text{H}_3\text{BO}_3 &= (43.82 \text{ g } \text{B}_2\text{H}_6) \left(\frac{1 \text{ mol } \text{B}_2\text{H}_6}{27.67 \text{ g } \text{B}_2\text{H}_6} \right) \left(\frac{2 \text{ mol } \text{H}_3\text{BO}_3}{1 \text{ mol } \text{B}_2\text{H}_6} \right) \left(\frac{61.83 \text{ g } \text{H}_3\text{BO}_3}{1 \text{ mol } \text{H}_3\text{BO}_3} \right) \\ &= 195.83597 = \mathbf{195.8 \text{ g } \text{H}_3\text{BO}_3} \end{aligned}$$

$$\text{Moles of } \text{H}_2 = (1.583665 \text{ mol } \text{B}_2\text{H}_6) \left(\frac{6 \text{ mol } \text{H}_2}{1 \text{ mol } \text{B}_2\text{H}_6} \right) = 9.50199 \text{ mol } \text{H}_2$$

$$\text{Mass (g) of } \text{H}_2 = (9.50199 \text{ mol } \text{H}_2) \left(\frac{2.016 \text{ g } \text{H}_2}{1 \text{ mol } \text{H}_2} \right) = 19.15901 \text{ g } \text{H}_2 = \mathbf{19.16 \text{ g } \text{H}_2}$$

Combining all steps gives:

$$\text{Mass (g) of } \text{H}_2 = (43.82 \text{ g } \text{B}_2\text{H}_6) \left(\frac{1 \text{ mol } \text{B}_2\text{H}_6}{27.67 \text{ g } \text{B}_2\text{H}_6} \right) \left(\frac{6 \text{ mol } \text{H}_2}{1 \text{ mol } \text{B}_2\text{H}_6} \right) \left(\frac{2.016 \text{ g } \text{H}_2}{1 \text{ mol } \text{H}_2} \right) = 19.15601 = \mathbf{19.16 \text{ g } \text{H}_2}$$

- 3.45 Plan: First, balance the equation. Convert the grams of silver sulfide to moles of silver sulfide by dividing by its molar mass. Use mole ratios from the balanced chemical equation to determine the moles of the products. Multiply the mole amount of each product by its molar mass to obtain mass in grams.

Solution:

First, balance the equation: $\text{Ag}_2\text{S}(\text{s}) + 2 \text{HCl}(\text{aq}) \rightarrow 2 \text{AgCl}(\text{s}) + \text{H}_2\text{S}(\text{g})$

$$\text{Moles of Ag}_2\text{S} = (174 \text{ g Ag}_2\text{S}) \left(\frac{1 \text{ mol Ag}_2\text{S}}{247.9 \text{ g Ag}_2\text{S}} \right) = 0.7018959 \text{ mol Ag}_2\text{S}$$

$$\text{Moles of AgCl} = (0.7018959 \text{ mol Ag}_2\text{S}) \left(\frac{2 \text{ mol AgCl}}{1 \text{ mol Ag}_2\text{S}} \right) = 1.403792 \text{ mol AgCl}$$

$$\text{Mass (g) of AgCl} = (1.403792 \text{ mol Ag}_2\text{S}) \left(\frac{143.4 \text{ g AgCl}}{1 \text{ mol AgCl}} \right) = 201.304 = \mathbf{201 \text{ g AgCl}}$$

Combining all steps gives:

$$\text{Mass (g) AgCl} = (174 \text{ g Ag}_2\text{S}) \left(\frac{1 \text{ mol Ag}_2\text{S}}{247.9 \text{ g Ag}_2\text{S}} \right) \left(\frac{2 \text{ mol AgCl}}{1 \text{ mol Ag}_2\text{S}} \right) \left(\frac{143.4 \text{ g AgCl}}{1 \text{ mol AgCl}} \right) = 201.304 = \mathbf{201 \text{ g AgCl}}$$

$$\text{Moles of H}_2\text{S} = (0.7018959 \text{ mol Ag}_2\text{S}) \left(\frac{1 \text{ mol H}_2\text{S}}{1 \text{ mol Ag}_2\text{S}} \right) = 0.7018959 \text{ mol H}_2\text{S}$$

$$\text{Mass (g) of H}_2\text{S} = 0.7018959 \text{ mol H}_2\text{S} \left(\frac{34.09 \text{ g H}_2\text{S}}{1 \text{ mol H}_2\text{S}} \right) = 23.9276 = \mathbf{23.9 \text{ g H}_2\text{S}}$$

Combining all steps gives:

$$\text{Mass (g) of H}_2\text{S} = 174 \text{ g Ag}_2\text{S} \left(\frac{1 \text{ mol Ag}_2\text{S}}{247.9 \text{ g Ag}_2\text{S}} \right) \left(\frac{1 \text{ mol H}_2\text{S}}{1 \text{ mol Ag}_2\text{S}} \right) \left(\frac{34.09 \text{ g H}_2\text{S}}{1 \text{ mol H}_2\text{S}} \right) = 23.9276 = \mathbf{23.9 \text{ g H}_2\text{S}}$$

- 3.46 Plan: Write the balanced equation by first writing the formulas for the reactants and products. Convert the mass of phosphorus to moles by dividing by the molar mass, use the mole ratio between phosphorus and chlorine from the balanced chemical equation to obtain moles of chlorine, and finally divide the moles of chlorine by its molar mass to obtain amount in grams.

Solution:

Reactants: formula for phosphorus is given as P₄ and formula for chlorine gas is Cl₂ (chlorine occurs as a diatomic molecule). Product: formula for phosphorus pentachloride (the name indicates one phosphorus atom and five chlorine atoms) is PCl₅.

Equation: P₄ + Cl₂ → PCl₅

Balancing the equation: P₄ + 10Cl₂ → 4PCl₅

$$\text{Moles of P}_4 = (455 \text{ g P}_4) \left(\frac{1 \text{ mol P}_4}{123.88 \text{ g P}_4} \right) = 3.67291 \text{ mol P}_4$$

$$\text{Moles of Cl}_2 = (3.67291 \text{ mol P}_4) \left(\frac{10 \text{ mol Cl}_2}{1 \text{ mol P}_4} \right) = 36.7291 \text{ mol Cl}_2$$

$$\text{Mass (g) of Cl}_2 = (36.7291 \text{ mol Cl}_2) \left(\frac{70.90 \text{ g Cl}_2}{1 \text{ mol Cl}_2} \right) = 2604.09 = \mathbf{2.60 \times 10^3 \text{ g Cl}_2}$$

Combining all steps gives:

$$\text{Mass (g) of Cl}_2 = (455 \text{ g P}_4) \left(\frac{1 \text{ mol P}_4}{123.88 \text{ g P}_4} \right) \left(\frac{10 \text{ mol Cl}_2}{1 \text{ mol P}_4} \right) \left(\frac{70.90 \text{ g Cl}_2}{1 \text{ mol Cl}_2} \right) = 2604.09267 = \mathbf{2.60 \times 10^3 \text{ g Cl}_2}$$

- 3.47 Plan: Write the balanced equation by first writing the formulas for the reactants and products. Convert the mass of sulfur to moles by dividing by the molar mass, use the mole ratio between sulfur and fluorine from the balanced chemical equation to obtain moles of fluorine, and finally divide the moles of fluorine by its molar mass to obtain amount in grams.

Solution:

Reactants: formula for sulfur is given as S₈ and formula for fluorine gas is F₂ (fluorine occurs as a diatomic molecule). Product: formula for sulfur hexafluoride (the name indicates one sulfur atom and six fluoride atoms) is SF₆.

Equation: S₈ + F₂ → SF₆

Balancing the equation: $\text{S}_8(\text{s}) + 24\text{F}_2(\text{g}) \rightarrow 8\text{SF}_6(\text{s})$

$$\text{Moles of S}_8 = (17.8 \text{ g S}_8) \left(\frac{1 \text{ mol S}_8}{256.56 \text{ g S}_8} \right) = 0.0693795 \text{ mol S}_8$$

$$\text{Moles of F}_2 = (0.0693795 \text{ mol S}_8) \left(\frac{24 \text{ mol F}_2}{1 \text{ mol S}_8} \right) = 1.665108 \text{ mol F}_2$$

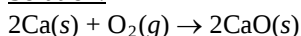
$$\text{Mass (g) of F}_2 = (1.665108 \text{ mol F}_2) \left(\frac{38.00 \text{ g F}_2}{1 \text{ mol F}_2} \right) = 63.274 = \mathbf{63.3 \text{ g F}_2}$$

Combining all steps gives:

$$\text{Mass (g) of F}_2 = (17.8 \text{ g S}_8) \left(\frac{1 \text{ mol S}_8}{256.56 \text{ g S}_8} \right) \left(\frac{24 \text{ mol F}_2}{1 \text{ mol S}_8} \right) \left(\frac{38.00 \text{ g F}_2}{1 \text{ mol F}_2} \right) = 63.27409 = \mathbf{63.3 \text{ g F}_2}$$

- 3.48 Plan: Convert the given mass of each reactant to moles by dividing by the molar mass of that reactant. Use the mole ratio from the balanced chemical equation to find the moles of CaO formed from each reactant, assuming an excess of the other reactant. The reactant that produces fewer moles of CaO is the limiting reactant. Convert the moles of CaO obtained from the limiting reactant to grams using the molar mass.

Solution:



$$\text{a) Moles of Ca} = (4.20 \text{ g Ca}) \left(\frac{1 \text{ mol Ca}}{40.08 \text{ g Ca}} \right) = 0.104790 \text{ mol Ca}$$

$$\text{Moles of CaO from Ca} = (0.104790 \text{ mol Ca}) \left(\frac{2 \text{ mol CaO}}{2 \text{ mol Ca}} \right) = 0.104790 = \mathbf{0.105 \text{ mol CaO}}$$

$$\text{b) Moles of O}_2 = (2.80 \text{ g O}_2) \left(\frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \right) = 0.0875 \text{ mol O}_2$$

$$\text{Moles of CaO from O}_2 = (0.0875 \text{ mol O}_2) \left(\frac{2 \text{ mol CaO}}{1 \text{ mol O}_2} \right) = 0.17500 = \mathbf{0.175 \text{ mol CaO}}$$

c) **Calcium** is the limiting reactant since it will form less calcium oxide.

d) The mass of CaO formed is determined by the limiting reactant, Ca.

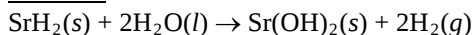
$$\text{Mass (g) of CaO} = (0.104790 \text{ mol CaO}) \left(\frac{56.08 \text{ g CaO}}{1 \text{ mol CaO}} \right) = 5.8766 = \mathbf{5.88 \text{ g CaO}}$$

Combining all steps gives:

$$\text{Mass (g) of CaO} = (4.20 \text{ g Ca}) \left(\frac{1 \text{ mol Ca}}{40.08 \text{ g Ca}} \right) \left(\frac{2 \text{ mol CaO}}{2 \text{ mol Ca}} \right) \left(\frac{56.08 \text{ g CaO}}{1 \text{ mol CaO}} \right) = 5.8766 = \mathbf{5.88 \text{ g CaO}}$$

- 3.49 Plan: Convert the given mass of each reactant to moles by dividing by the molar mass of that reactant. Use the mole ratio from the balanced chemical equation to find the moles of H₂ formed from each reactant, assuming an excess of the other reactant. The reactant that produces fewer moles of H₂ is the limiting reactant. Convert the moles of H₂ obtained from the limiting reactant to grams using the molar mass.

Solution:



$$\text{a) Moles of SrH}_2 = (5.70 \text{ g SrH}_2) \left(\frac{1 \text{ mol SrH}_2}{89.64 \text{ g SrH}_2} \right) = 0.0635877 \text{ mol SrH}_2$$

$$\text{Moles of H}_2 \text{ from SrH}_2 = (0.0635877 \text{ mol SrH}_2) \left(\frac{2 \text{ mol H}_2}{1 \text{ mol SrH}_2} \right) = 0.127175 = \mathbf{0.127 \text{ mol H}_2}$$

$$\text{b) Mass (g) of H}_2\text{O} = (4.75 \text{ g H}_2\text{O}) \left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) = 0.263596 \text{ mol H}_2\text{O}$$

$$\text{Moles of H}_2 \text{ from H}_2\text{O} = (0.263596 \text{ mol H}_2\text{O}) \left(\frac{2 \text{ mol H}_2}{2 \text{ mol H}_2\text{O}} \right) = 0.263596 = \mathbf{0.264 \text{ mol H}_2}$$

c) **SrH₂** is the limiting reagent since it will yield fewer moles of hydrogen gas.

d) The mass of H₂ formed is determined by the limiting reactant, SrH₂.

$$\text{Mass (g) of H}_2 = (0.127175 \text{ mol H}_2) \left(\frac{2.016 \text{ g H}_2}{1 \text{ mol H}_2} \right) = 0.256385 = \mathbf{0.256 \text{ g H}_2}$$

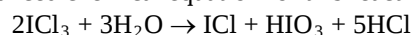
Combining all steps gives:

$$\text{Mass (g) of H}_2 = (5.70 \text{ g SrH}_2) \left(\frac{1 \text{ mol SrH}_2}{89.64 \text{ g SrH}_2} \right) \left(\frac{2 \text{ mol H}_2}{1 \text{ mol SrH}_2} \right) \left(\frac{2.016 \text{ g H}_2}{1 \text{ mol H}_2} \right) = 0.256385 = \mathbf{0.256 \text{ g H}_2}$$

- 3.50 Plan: First, balance the chemical equation. To determine which reactant is limiting, calculate the amount of HIO₃ formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the amount of HIO₃ formed and the amount of the excess reactant that reacts. The difference between the amount of excess reactant that reacts and the initial amount of reactant supplied gives the amount of excess reactant remaining.

Solution:

The balanced chemical equation for this reaction is:



Hint: Balance the equation by starting with oxygen. The other elements are in multiple reactants and/or products and are harder to balance initially.

Finding the moles of HIO₃ from the moles of ICl₃ (if H₂O is limiting):

$$\text{Moles of ICl}_3 = (635 \text{ g ICl}_3) \left(\frac{1 \text{ mol ICl}_3}{233.2 \text{ g ICl}_3} \right) = 2.722985 \text{ mol ICl}_3$$

$$\text{Moles of HIO}_3 \text{ from ICl}_3 = (2.722985 \text{ mol ICl}_3) \left(\frac{1 \text{ mol HIO}_3}{2 \text{ mol ICl}_3} \right) = 1.361492 = 1.36 \text{ mol HIO}_3$$

Finding the moles of HIO₃ from the moles of H₂O (if ICl₃ is limiting):

$$\text{Moles of H}_2\text{O} = (118.5 \text{ g H}_2\text{O}) \left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) = 6.57603 \text{ mol H}_2\text{O}$$

$$\text{Moles HIO}_3 \text{ from H}_2\text{O} = (6.57603 \text{ mol H}_2\text{O}) \left(\frac{1 \text{ mol HIO}_3}{3 \text{ mol H}_2\text{O}} \right) = 2.19201 = 2.19 \text{ mol HIO}_3$$

ICl₃ is the limiting reagent and will produce **1.36 mol HIO₃**.

$$\text{Mass (g) of HIO}_3 = (1.361492 \text{ mol HIO}_3) \left(\frac{175.9 \text{ g HIO}_3}{1 \text{ mol HIO}_3} \right) = 239.486 = \mathbf{239 \text{ g HIO}_3}$$

Combining all steps gives:

$$\text{Mass (g) of HIO}_3 = (635 \text{ g ICl}_3) \left(\frac{1 \text{ mol ICl}_3}{233.2 \text{ g ICl}_3} \right) \left(\frac{1 \text{ mol HIO}_3}{2 \text{ mol ICl}_3} \right) \left(\frac{175.9 \text{ g HIO}_3}{1 \text{ mol HIO}_3} \right) = 239.486 = \mathbf{239 \text{ g HIO}_3}$$

The remaining mass of the excess reagent can be calculated from the amount of H₂O combining with the limiting reagent.

$$\text{Moles of H}_2\text{O required to react with 635 g ICl}_3 = (2.722985 \text{ mol ICl}_3) \left(\frac{3 \text{ mol H}_2\text{O}}{2 \text{ mol ICl}_3} \right) = 4.0844775 \text{ mol H}_2\text{O}$$

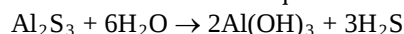
$$\text{Mass (g) of H}_2\text{O required to react with 635 g ICl}_3 = (4.0844775 \text{ mol H}_2\text{O}) \left(\frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \right) \\ = 73.6023 = 73.6 \text{ g H}_2\text{O reacted}$$

$$\text{Remaining H}_2\text{O} = 118.5 \text{ g} - 73.6 \text{ g} = \mathbf{44.9 \text{ g H}_2\text{O}}$$

- 3.51 **Plan:** First, balance the chemical equation. To determine which reactant is limiting, calculate the amount of H₂S formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the amount of H₂S formed and the amount of the excess reactant that reacts. The difference between the amount of excess reactant that reacts and the initial amount of reactant supplied gives the amount of excess reactant remaining.

Solution:

The balanced chemical equation for this reaction is:



Finding the moles of H₂S from the moles of Al₂S₃ (if H₂O is limiting):

$$\text{Moles of Al}_2\text{S}_3 = (158 \text{ g Al}_2\text{S}_3) \left(\frac{1 \text{ mol Al}_2\text{S}_3}{150.17 \text{ g Al}_2\text{S}_3} \right) = 1.05214 \text{ mol Al}_2\text{S}_3$$

$$\text{Moles of H}_2\text{S from Al}_2\text{S}_3 = (1.05214 \text{ mol Al}_2\text{S}_3) \left(\frac{3 \text{ mol H}_2\text{S}}{1 \text{ mol Al}_2\text{S}_3} \right) = 3.15642 = 3.16 \text{ mol H}_2\text{S}$$

Finding the moles of H₂S from the moles of H₂O (if Al₂S₃ is limiting):

$$\text{Moles of H}_2\text{O} = (131 \text{ g H}_2\text{O}) \left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) = 7.26970 \text{ mol H}_2\text{O}$$

$$\text{Moles of H}_2\text{S from H}_2\text{O} = (7.26970 \text{ mol H}_2\text{O}) \left(\frac{3 \text{ mol H}_2\text{S}}{6 \text{ mol H}_2\text{O}} \right) = 3.63485 = 3.63 \text{ mol H}_2\text{S}$$

Al₂S₃ is the limiting reagent and **3.16 mol of H₂S** will form.

$$\text{Mass (g) of H}_2\text{S} = (3.15642 \text{ mol H}_2\text{S}) \left(\frac{34.09 \text{ g H}_2\text{S}}{1 \text{ mol H}_2\text{S}} \right) = 107.602 = \mathbf{108 \text{ g H}_2\text{S}}$$

Combining all steps gives:

$$\text{Grams H}_2\text{S} = (158 \text{ g Al}_2\text{S}_3) \left(\frac{1 \text{ mol Al}_2\text{S}_3}{150.17 \text{ g Al}_2\text{S}_3} \right) \left(\frac{3 \text{ mol H}_2\text{S}}{1 \text{ mol Al}_2\text{S}_3} \right) \left(\frac{34.09 \text{ g H}_2\text{S}}{1 \text{ mol H}_2\text{S}} \right) = 107.602 = \mathbf{108 \text{ g H}_2\text{S}}$$

The remaining mass of the excess reagent can be calculated from the amount of H₂O combining with the limiting reagent.

$$\text{Moles of H}_2\text{O required to react with 158 g of Al}_2\text{S}_3 = (1.05214 \text{ mol Al}_2\text{S}_3) \left(\frac{6 \text{ mol H}_2\text{O}}{1 \text{ mol Al}_2\text{S}_3} \right) = 6.31284 \text{ mol H}_2\text{O}$$

$$\text{Mass (g) of H}_2\text{O required to react with 158 g of Al}_2\text{S}_3 = (6.31284 \text{ mol H}_2\text{O}) \left(\frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \right) = 113.757 \text{ g H}_2\text{O}$$

$$\text{Remaining H}_2\text{O} = 131 \text{ g H}_2\text{O} - 113.757 \text{ g H}_2\text{O} = 17.243 = \mathbf{17 \text{ g H}_2\text{O}}$$

- 3.52 **Plan:** Write the balanced equation; the formula for carbon is C, the formula for oxygen is O₂, and the formula for carbon dioxide is CO₂. To determine which reactant is limiting, calculate the amount of CO₂ formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the amount of CO₂ formed and the amount of the excess reactant that reacts. The difference between the amount of excess reactant that reacts and the initial amount of reactant supplied gives the amount of excess reactant remaining.

Solution:

The balanced equation is: C(s) + O₂(g) → CO₂(g)

Finding the moles of CO₂ from the moles of carbon (if O₂ is limiting):

$$\text{Moles of CO}_2 \text{ from C} = (0.100 \text{ mol C}) \left(\frac{1 \text{ mol CO}_2}{1 \text{ mol C}} \right) = 0.100 \text{ mol CO}_2$$

Finding the moles of CO₂ from the moles of oxygen (if C is limiting):

$$\text{Moles of O}_2 = (8.00 \text{ g O}_2) \left(\frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \right) = 0.250 \text{ mol O}_2$$

$$\text{Moles of CO}_2 \text{ from O}_2 = (0.250 \text{ mol O}_2) \left(\frac{1 \text{ mol CO}_2}{1 \text{ mol O}_2} \right) = 0.25000 = 0.250 \text{ mol CO}_2$$

Carbon is the limiting reactant and will be used to determine the amount of CO₂ that will form.

$$\text{Mass (g) of CO}_2 = (0.100 \text{ mol CO}_2) \left(\frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} \right) = 4.401 = \mathbf{4.40 \text{ g CO}_2}$$

Since carbon is limiting, the **O₂ is in excess**. The amount remaining depends on how much combines with the limiting reagent.

$$\text{Moles of O}_2 \text{ required to react with 0.100 mol of C} = (0.100 \text{ mol C}) \left(\frac{1 \text{ mol O}_2}{1 \text{ mol C}} \right) = 0.100 \text{ mol O}_2$$

$$\text{Mass (g) of O}_2 \text{ required to react with 0.100 mol of C} = (0.100 \text{ mol O}_2) \left(\frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} \right) = 3.20 \text{ g O}_2$$

$$\text{Remaining O}_2 = 8.00 \text{ g} - 3.20 \text{ g} = \mathbf{4.80 \text{ g O}_2}$$

- 3.53 Plan: Write the balanced equation; the formula for hydrogen is H₂, the formula for oxygen is O₂, and the formula for water is H₂O. To determine which reactant is limiting, calculate the amount of H₂O formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the amount of H₂O formed and the amount of the excess reactant that reacts. The difference between the amount of excess reactant that reacts and the initial amount of reactant supplied gives the amount of excess reactant remaining.

Solution:

The balanced equation is: $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l})$

Finding the moles of H₂O from the moles of hydrogen (if O₂ is limiting):

$$\text{Moles of H}_2 = (0.0375 \text{ g H}_2) \left(\frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} \right) = 0.01860 \text{ mol H}_2$$

$$\text{Moles of H}_2\text{O from H}_2 = (0.01860 \text{ mol H}_2) \left(\frac{2 \text{ mol H}_2\text{O}}{2 \text{ mol H}_2} \right) = 0.01860 = 0.0186 \text{ mol H}_2\text{O}$$

Finding the moles of H₂O from the moles of oxygen (if H₂ is limiting):

$$\text{Mole of H}_2\text{O from O}_2 = (0.0185 \text{ mol O}_2) \left(\frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol O}_2} \right) = 0.0370 \text{ mol H}_2\text{O}$$

The hydrogen is the limiting reactant, and will be used to determine the amount of water that will form.

$$\text{Mass (g) of H}_2\text{O} = (0.01860 \text{ mol H}_2\text{O}) \left(\frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \right) = 0.335172 = \mathbf{0.335 \text{ g H}_2\text{O}}$$

Since the hydrogen is limiting; the **oxygen must be the excess reactant**. The amount of excess reactant is determined from the limiting reactant.

$$\text{Moles of O}_2 \text{ required to react with 0.0375 g of H}_2 = (0.01860 \text{ mol H}_2) \left(\frac{1 \text{ mol O}_2}{2 \text{ mol H}_2} \right) = 0.00930 \text{ mol O}_2$$

$$\text{Mass (g) of O}_2 \text{ required to react with 0.0375 g of H}_2 = (0.00930 \text{ mol O}_2) \left(\frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} \right) = 0.2976 \text{ g O}_2$$

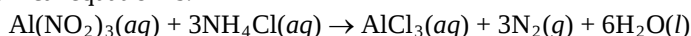
$$\text{Mass of O}_2 \text{ supplied} = (0.0185 \text{ mol O}_2) \left(\frac{32.00 \text{ mol O}_2}{1 \text{ mol O}_2} \right) = 0.5920 \text{ g O}_2$$

$$\text{Remaining O}_2 = 0.5920 \text{ g} - 0.2976 \text{ g} = 0.2944 = \mathbf{0.294 \text{ g O}_2}$$

- 3.54 Plan: The question asks for the mass of each substance present at the end of the reaction. "Substance" refers to both reactants and products. Solve this problem using multiple steps. Recognizing that this is a limiting reactant problem, first write a balanced chemical equation. To determine which reactant is limiting, calculate the amount of any product formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Any product can be used to predict the limiting reactant; in this case, AlCl_3 is used. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the amount of both products formed and the amount of the excess reactant that reacts. The difference between the amount of excess reactant that reacts and the initial amount of reactant supplied gives the amount of excess reactant remaining.

Solution:

The balanced chemical equation is:



Now determine the limiting reagent. We will use the moles of AlCl_3 produced to determine which is limiting. Finding the moles of AlCl_3 from the moles of $\text{Al}(\text{NO}_2)_3$ (if NH_4Cl is limiting):

$$\text{Moles of Al}(\text{NO}_2)_3 = (72.5 \text{ g Al}(\text{NO}_2)_3) \left(\frac{1 \text{ mol Al}(\text{NO}_2)_3}{165.01 \text{ g Al}(\text{NO}_2)_3} \right) = 0.439367 \text{ mol Al}(\text{NO}_2)_3$$

$$\text{Moles of AlCl}_3 \text{ from Al}(\text{NO}_2)_3 = (0.439367 \text{ mol Al}(\text{NO}_2)_3) \left(\frac{1 \text{ mol AlCl}_3}{1 \text{ mol Al}(\text{NO}_2)_3} \right) = 0.439367 = 0.439 \text{ mol AlCl}_3$$

Finding the moles of AlCl_3 from the moles of NH_4Cl (if $\text{Al}(\text{NO}_2)_3$ is limiting):

$$\text{Moles of NH}_4\text{Cl} = (58.6 \text{ g NH}_4\text{Cl}) \left(\frac{1 \text{ mol NH}_4\text{Cl}}{53.49 \text{ g NH}_4\text{Cl}} \right) = 1.09553 \text{ mol NH}_4\text{Cl}$$

$$\text{Moles of AlCl}_3 \text{ from NH}_4\text{Cl} = (1.09553 \text{ mol NH}_4\text{Cl}) \left(\frac{1 \text{ mol AlCl}_3}{3 \text{ mol NH}_4\text{Cl}} \right) = 0.365177 = 0.365 \text{ mol AlCl}_3$$

Ammonium chloride is the limiting reactant, and it is used for all subsequent calculations.

Mass of substances after the reaction:

$\text{Al}(\text{NO}_2)_3$:

Mass (g) of $\text{Al}(\text{NO}_2)_3$ (the excess reactant) required to react with 58.6 g of NH_4Cl =

$$(1.09553 \text{ mol NH}_4\text{Cl}) \left(\frac{1 \text{ mol Al}(\text{NO}_2)_3}{3 \text{ mol NH}_4\text{Cl}} \right) \left(\frac{165.01 \text{ g Al}(\text{NO}_2)_3}{1 \text{ mol Al}(\text{NO}_2)_3} \right) = 60.2579 = 60.3 \text{ g Al}(\text{NO}_2)_3$$

$\text{Al}(\text{NO}_2)_3$ remaining: $72.5 \text{ g} - 60.3 \text{ g} = \mathbf{12.2 \text{ g Al}(\text{NO}_2)_3}$

NH_4Cl : **None left** since it is the limiting reagent.

AlCl_3 :

$$\text{Mass (g) of AlCl}_3 = (0.365177 \text{ mol AlCl}_3) \left(\frac{133.33 \text{ g AlCl}_3}{1 \text{ mol AlCl}_3} \right) = 48.689 = \mathbf{48.7 \text{ g AlCl}_3}$$

N_2 :

$$\text{Mass (g) of N}_2 = (1.09553 \text{ mol NH}_4\text{Cl}) \left(\frac{3 \text{ mol N}_2}{3 \text{ mol NH}_4\text{Cl}} \right) \left(\frac{28.02 \text{ g N}_2}{1 \text{ mol N}_2} \right) = 30.697 = \mathbf{30.7 \text{ g N}_2}$$

H_2O :

$$\text{Mass (g) of H}_2\text{O} = (1.09553 \text{ mol NH}_4\text{Cl}) \left(\frac{6 \text{ mol H}_2\text{O}}{3 \text{ mol NH}_4\text{Cl}} \right) \left(\frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \right) = 39.483 = \mathbf{39.5 \text{ g H}_2\text{O}}$$

- 3.55 Plan: The question asks for the mass of each substance present at the end of the reaction. "Substance" refers to both reactants and products. Solve this problem using multiple steps. Recognizing that this is a limiting reactant problem, first write a balanced chemical equation. To determine which reactant is limiting, calculate the amount of any product formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Any product can be used to predict the limiting reactant; in this case, CaF_2 is used. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the amount of both products formed and the amount of the excess reactant that reacts. The difference between the amount of excess reactant that reacts and the initial amount of reactant supplied gives the amount of excess reactant remaining.

Solution:

The balanced chemical equation is:



Now determine the limiting reagent. We will use the moles of CaF_2 produced to determine which is limiting.

Finding the moles of CaF_2 from the moles of $\text{Ca}(\text{NO}_3)_2$ (if NH_4F is limiting):

$$\text{Moles of } \text{Ca}(\text{NO}_3)_2 = (16.8 \text{ g } \text{Ca}(\text{NO}_3)_2) \left(\frac{1 \text{ mol } \text{Ca}(\text{NO}_3)_2}{164.10 \text{ g } \text{Ca}(\text{NO}_3)_2} \right) = 0.1023766 \text{ mol } \text{Ca}(\text{NO}_3)_2$$

$$\begin{aligned} \text{Moles of } \text{CaF}_2 \text{ from } \text{Ca}(\text{NO}_3)_2 &= (0.1023766 \text{ mol } \text{Ca}(\text{NO}_3)_2) \left(\frac{1 \text{ mol } \text{CaF}_2}{1 \text{ mol } \text{Ca}(\text{NO}_3)_2} \right) \\ &= 0.1023766 = 0.102 \text{ mol } \text{CaF}_2 \end{aligned}$$

Finding the moles of CaF_2 from the moles of NH_4F (if $\text{Ca}(\text{NO}_3)_2$ is limiting):

$$\text{Moles of } \text{NH}_4\text{F} = (17.50 \text{ g } \text{NH}_4\text{F}) \left(\frac{1 \text{ mol } \text{NH}_4\text{F}}{37.04 \text{ g } \text{NH}_4\text{F}} \right) = 0.47246 \text{ mol } \text{NH}_4\text{F}$$

$$\text{Moles of } \text{CaF}_2 \text{ from } \text{NH}_4\text{F} = (0.47246 \text{ mol } \text{NH}_4\text{F}) \left(\frac{1 \text{ mol } \text{CaF}_2}{2 \text{ mol } \text{NH}_4\text{F}} \right) = 0.23623 = 0.236 \text{ mol } \text{CaF}_2$$

Calcium nitrate is the limiting reactant, and it is used for all subsequent calculations

Mass of substances after the reaction:

$\text{Ca}(\text{NO}_3)_2$: **None** (It is the limiting reactant.)

NH_4F :

Mass (g) of NH_4F (the excess reactant) required to react with 16.8 g of $\text{Ca}(\text{NO}_3)_2$ =

$$(0.1023766 \text{ mol } \text{Ca}(\text{NO}_3)_2) \left(\frac{2 \text{ mol } \text{NH}_4\text{F}}{1 \text{ mol } \text{Ca}(\text{NO}_3)_2} \right) \left(\frac{37.04 \text{ g } \text{NH}_4\text{F}}{1 \text{ mol } \text{NH}_4\text{F}} \right) = 7.58406 \text{ g } \text{Ca}(\text{NO}_3)_2$$

NH_4F remaining: $17.50 \text{ g} - 7.58 \text{ g} = 9.9159 = \mathbf{9.92 \text{ g } \text{NH}_4\text{F}}$

CaF_2 :

$$\text{Mass (g) of } \text{CaF}_2 = (0.1023766 \text{ mol } \text{Ca}(\text{NO}_3)_2) \left(\frac{1 \text{ mol } \text{CaF}_2}{1 \text{ mol } \text{Ca}(\text{NO}_3)_2} \right) \left(\frac{78.08 \text{ g } \text{CaF}_2}{1 \text{ mol } \text{CaF}_2} \right) = 7.99356 = \mathbf{7.99 \text{ g } \text{CaF}_2}$$

N_2O :

$$\text{Mass (g) of } \text{N}_2\text{O} = (0.1023766 \text{ mol } \text{Ca}(\text{NO}_3)_2) \left(\frac{2 \text{ mol } \text{N}_2\text{O}}{1 \text{ mol } \text{Ca}(\text{NO}_3)_2} \right) \left(\frac{44.02 \text{ g } \text{N}_2\text{O}}{1 \text{ mol } \text{N}_2\text{O}} \right) = 9.0132 = \mathbf{9.01 \text{ g } \text{N}_2\text{O}}$$

H_2O :

$$\text{Mass (g) of } \text{H}_2\text{O} = (0.1023766 \text{ mol } \text{Ca}(\text{NO}_3)_2) \left(\frac{4 \text{ mol } \text{H}_2\text{O}}{1 \text{ mol } \text{Ca}(\text{NO}_3)_2} \right) \left(\frac{18.02 \text{ g } \text{H}_2\text{O}}{1 \text{ mol } \text{H}_2\text{O}} \right) = 7.3793 = \mathbf{7.38 \text{ g } \text{H}_2\text{O}}$$

- 3.56 Plan: Express the yield of each step as a fraction of 1.00; multiply the fraction of the first step by that of the second step and then multiply by 100 to get the overall percent yield.

Solution:

$$73\% = 0.73; 68\% = 0.68$$

$$(0.73 \times 0.68) \times 100 = 49.64 = \mathbf{50. \%}$$

- 3.57 Plan: Express the yield of each step as a fraction of 1.00; multiply the fraction of the first step by that of the second step and then multiply by 100 to get the overall percent yield.

Solution:

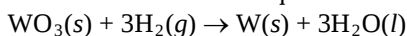
$$48\% = 0.48; 73\% = 0.73$$

$$(0.48 \times 0.73) \times 100 = 35.04 = \mathbf{35\%}$$

- 3.58 Plan: Write and balance the chemical equation using the formulas of the substances. Determine the theoretical yield of the reaction from the mass of tungsten(VI) oxide. To do that, convert the mass of tungsten(VI) oxide to moles by dividing by its molar mass and then use the mole ratio between tungsten(VI) oxide and water to determine the moles and then mass of water that should be produced. Use the density of water to determine the actual yield of water in grams. The actual yield divided by the theoretical yield just calculated (with the result multiplied by 100%) gives the percent yield.

Solution:

The balanced chemical equation is:



Determining the theoretical yield of H_2O :

$$\text{Moles of } \text{WO}_3 = (45.5 \text{ g } \text{WO}_3) \left(\frac{1 \text{ mol } \text{WO}_3}{231.9 \text{ g } \text{WO}_3} \right) = 0.1962053 \text{ mol } \text{WO}_3$$

$$\text{Mass (g) of } \text{H}_2\text{O (theoretical yield)} = (0.1962053 \text{ mol } \text{WO}_3) \left(\frac{3 \text{ mol } \text{H}_2\text{O}}{1 \text{ mol } \text{WO}_3} \right) \left(\frac{18.02 \text{ g } \text{H}_2\text{O}}{1 \text{ mol } \text{H}_2\text{O}} \right) = 10.60686 \text{ g } \text{H}_2\text{O}$$

Determining the actual yield of H_2O :

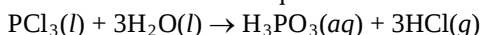
$$\text{Mass (g) of } \text{H}_2\text{O (actual yield)} = (9.60 \text{ mL } \text{H}_2\text{O}) \left(\frac{1.00 \text{ g } \text{H}_2\text{O}}{1 \text{ mL } \text{H}_2\text{O}} \right) = 9.60 \text{ g } \text{H}_2\text{O}$$

$$\% \text{ yield} = \left(\frac{\text{actual Yield}}{\text{theoretical Yield}} \right) \times 100\% = \left(\frac{9.60 \text{ g } \text{H}_2\text{O}}{10.60686 \text{ g } \text{H}_2\text{O}} \right) \times 100\% = 90.5075 = \mathbf{90.5\%}$$

- 3.59 Plan: Write and balance the chemical equation using the formulas of the substances. Determine the theoretical yield of the reaction from the mass of phosphorus trichloride. To do that, convert the mass of phosphorus trichloride to moles by dividing by its molar mass and then use the mole ratio between phosphorus trichloride and HCl to determine the moles and then mass of HCl that should be produced. The actual yield of the HCl is given. The actual yield divided by the theoretical yield just calculated (with the result multiplied by 100%) gives the percent yield.

Solution:

The balanced chemical equation is:



Determining the theoretical yield of HCl:

$$\text{Moles of } \text{PCl}_3 = (200. \text{ g } \text{PCl}_3) \left(\frac{1 \text{ mol } \text{PCl}_3}{137.32 \text{ g } \text{PCl}_3} \right) = 1.456452 \text{ mol } \text{PCl}_3$$

$$\text{Mass (g) of } \text{HCl (theoretical yield)} = (1.456452 \text{ mol } \text{PCl}_3) \left(\frac{3 \text{ mol } \text{HCl}}{1 \text{ mol } \text{PCl}_3} \right) \left(\frac{36.46 \text{ g } \text{HCl}}{1 \text{ mol } \text{HCl}} \right) = 159.3067 \text{ g } \text{HCl}$$

Actual yield (g) of HCl is given as 128 g HCl.

Calculate the percent yield:

$$\% \text{ yield} = \left(\frac{\text{actual Yield}}{\text{theoretical Yield}} \right) \times 100\% = \left(\frac{128 \text{ g } \text{HCl}}{159.3067 \text{ g } \text{HCl}} \right) \times 100\% = 80.3481586 = \mathbf{80.3\%}$$

- 3.60 Plan: Write the balanced chemical equation. Since quantities of two reactants are given, we must determine which is the limiting reactant. To determine which reactant is limiting, calculate the amount of any product formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Any product can be used to predict the limiting reactant; in this case, CH₃Cl is used. Only 75.0% of the calculated amounts of products actually form, so the actual yield is 75% of the theoretical yield.

Solution:

The balanced equation is: $\text{CH}_4(g) + \text{Cl}_2(g) \rightarrow \text{CH}_3\text{Cl}(g) + \text{HCl}(g)$

Determining the limiting reactant:

Finding the moles of CH₃Cl from the moles of CH₄ (if Cl₂ is limiting):

$$\text{Moles of CH}_4 = (20.5 \text{ g CH}_4) \left(\frac{1 \text{ mol CH}_4}{16.04 \text{ g CH}_4} \right) = 1.278055 \text{ mol CH}_4$$

$$\text{Moles of CH}_3\text{Cl from CH}_4 = (1.278055 \text{ mol CH}_4) \left(\frac{1 \text{ mol CH}_3\text{Cl}}{1 \text{ mol CH}_4} \right) = 1.278055 \text{ mol CH}_3\text{Cl}$$

Finding the moles of CH₃Cl from the moles of Cl₂ (if CH₄ is limiting):

$$\text{Moles of Cl}_2 = (45.0 \text{ g Cl}_2) \left(\frac{1 \text{ mol Cl}_2}{70.90 \text{ g Cl}_2} \right) = 0.634697 \text{ mol Cl}_2$$

$$\text{Moles of CH}_3\text{Cl from Cl}_2 = (0.634697 \text{ mol Cl}_2) \left(\frac{1 \text{ mol CH}_3\text{Cl}}{1 \text{ mol Cl}_2} \right) = 0.634697 \text{ mol CH}_3\text{Cl}$$

Chlorine is the limiting reactant and is used to determine the theoretical yield of CH₃Cl:

$$\text{Mass (g) of CH}_3\text{Cl (theoretical yield)} = (0.634697 \text{ mol CH}_3\text{Cl}) \left(\frac{50.48 \text{ g CH}_3\text{Cl}}{1 \text{ mol CH}_3\text{Cl}} \right) = 32.0395 \text{ g CH}_3\text{Cl}$$

$$\% \text{ yield} = \left(\frac{\text{actual Yield}}{\text{theoretical Yield}} \right) \times 100\%$$

$$\text{Actual yield (g) of CH}_3\text{Cl} = \frac{\% \text{ yield}}{100\%} (\text{theoretical yield}) = \frac{75\%}{100\%} (32.0395 \text{ g CH}_3\text{Cl}) = 24.02962 = \mathbf{24.0 \text{ g CH}_3\text{Cl}}$$

- 3.61 Plan: Write the balanced chemical equation. Since quantities of two reactants are given, we must determine which is the limiting reactant. To determine which reactant is limiting, calculate the amount of product formed from each reactant, assuming an excess of the other reactant. Only 93.0% of the calculated amount of product actually forms, so the actual yield is 93.0% of the theoretical yield.

Solution:

The balanced equation is: $3\text{Ca}(s) + \text{N}_2(g) \rightarrow \text{Ca}_3\text{N}_2(s)$

Determining the limiting reactant:

Finding the moles of Ca₃N₂ from the moles of Ca (if N₂ is limiting):

$$\text{Moles of Ca} = (56.6 \text{ g Ca}) \left(\frac{1 \text{ mol Ca}}{40.08 \text{ g Ca}} \right) = 1.412176 \text{ mol Ca}$$

$$\text{Moles of Ca}_3\text{N}_2 \text{ from Ca} = (1.412176 \text{ mol Ca}) \left(\frac{1 \text{ mol Ca}_3\text{N}_2}{3 \text{ mol Ca}} \right) = 0.470725 \text{ mol Ca}_3\text{N}_2$$

Finding the moles of Ca₃N₂ from the moles of N₂ (if Ca is limiting):

$$\text{Moles of N}_2 = (30.5 \text{ g N}_2) \left(\frac{1 \text{ mol N}_2}{28.02 \text{ g N}_2} \right) = 1.08851 \text{ mol N}_2$$

$$\text{Moles of Ca}_3\text{N}_2 \text{ from N}_2 = (1.08851 \text{ mol N}_2) \left(\frac{1 \text{ mol Ca}_3\text{N}_2}{1 \text{ mol N}_2} \right) = 1.08851 \text{ mol Ca}_3\text{N}_2$$

Ca is the limiting reactant and is used to determine the theoretical yield of Ca₃N₂.

$$\text{Mass (g) of Ca}_3\text{N}_2 \text{ (theoretical yield)} = (0.470725 \text{ mol Ca}_3\text{N}_2) \left(\frac{148.26 \text{ g Ca}_3\text{N}_2}{1 \text{ mol Ca}_3\text{N}_2} \right) = 69.7897 \text{ g Ca}_3\text{N}_2$$

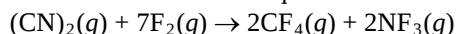
$$\% \text{ yield} = \left(\frac{\text{actual Yield}}{\text{theoretical Yield}} \right) \times 100\%$$

$$\text{Actual yield (g) of Ca}_3\text{N}_2 = \frac{\% \text{ yield}}{100\%} (\text{theoretical yield}) = \frac{93\%}{100\%} (69.7897 \text{ g Ca}_3\text{N}_2) = 64.9044 = \mathbf{64.9 \text{ g Ca}_3\text{N}_2}$$

- 3.62 Plan: Write the balanced equation; the formula for fluorine is F_2 , the formula for carbon tetrafluoride is CF_4 , and the formula for nitrogen trifluoride is NF_3 . To determine which reactant is limiting, calculate the amount of CF_4 formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the mass of CF_4 formed.

Solution:

The balanced chemical equation is:



Determining the limiting reactant:

Finding the moles of CF_4 from the moles of $(\text{CN})_2$ (if F_2 is limiting):

$$\text{Moles of CF}_4 \text{ from } (\text{CN})_2 = (60.0 \text{ g } (\text{CN})_2) \left(\frac{1 \text{ mol } (\text{CN})_2}{52.04 \text{ g } (\text{CN})_2} \right) \left(\frac{2 \text{ mol CF}_4}{1 \text{ mol } (\text{CN})_2} \right) = 2.30592 \text{ mol CF}_4$$

Finding the moles of CF_4 from the moles of F_2 (if $(\text{CN})_2$ is limiting):

$$\text{Moles of CF}_4 \text{ from F}_2 = (60.0 \text{ g F}_2) \left(\frac{1 \text{ mol F}_2}{38.00 \text{ g F}_2} \right) \left(\frac{2 \text{ mol CF}_4}{7 \text{ mol F}_2} \right) = 0.4511278 \text{ mol CF}_4$$

F_2 is the limiting reactant, and will be used to calculate the amount of CF_4 produced.

$$\text{Mass (g) of CF}_4 = (60.0 \text{ g F}_2) \left(\frac{1 \text{ mol F}_2}{38.00 \text{ g F}_2} \right) \left(\frac{2 \text{ mol CF}_4}{7 \text{ mol F}_2} \right) \left(\frac{88.01 \text{ g CF}_4}{1 \text{ mol CF}_4} \right) = 39.70376 = \mathbf{39.7 \text{ g CF}_4}$$

- 3.63 Plan: Write and balance the chemical reaction. Remember that both chlorine and oxygen exist as diatomic molecules. Use the mole ratio between oxygen and dichlorine monoxide to find the moles of dichlorine monoxide that reacted. Multiply the amount in moles by Avogadro's number to convert to number of molecules.

Solution:

a) Both oxygen and chlorine are diatomic. **Scene A** best represents the product mixture as there are O_2 and Cl_2 molecules in Scene A. Scene B shows oxygen and chlorine atoms and Scene C shows atoms and molecules. Oxygen and chlorine atoms are NOT products of this reaction.

b) The balanced reaction is $2\text{Cl}_2\text{O}(\text{g}) \rightarrow 2\text{Cl}_2(\text{g}) + \text{O}_2(\text{g})$.

c) There is a 2:1 mole ratio between Cl_2 and O_2 . In Scene A, there are 6 green molecules and 3 red molecules. Since twice as many Cl_2 molecules are produced as there are O_2 molecules produced, the red molecules are the O_2 molecules.

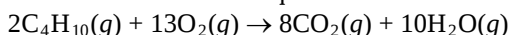
$$\begin{aligned} \text{Moles of Cl}_2\text{O} &= (3 \text{ O}_2 \text{ molecules}) \left(\frac{2 \text{ O atoms}}{1 \text{ O}_2 \text{ molecule}} \right) \left(\frac{0.050 \text{ mol O atoms}}{1 \text{ O atom}} \right) \left(\frac{1 \text{ mol O}_2 \text{ molecules}}{2 \text{ mol O atoms}} \right) \left(\frac{2 \text{ mol Cl}_2\text{O}}{1 \text{ mol O}_2} \right) \\ &= 0.30 \text{ mol Cl}_2\text{O} \end{aligned}$$

$$\begin{aligned} \text{Molecules of Cl}_2\text{O} &= (0.30 \text{ mol Cl}_2\text{O}) \left(\frac{6.022 \times 10^{23} \text{ Cl}_2\text{O molecules}}{1 \text{ mol Cl}_2\text{O}} \right) \\ &= 1.8066 \times 10^{23} = \mathbf{1.8 \times 10^{23} \text{ Cl}_2\text{O molecules}} \end{aligned}$$

- 3.64 Plan: Write a balanced equation. Use the density of butane to convert the given volume of butane to mass and divide by the molar mass of butane to convert mass to moles. Use the mole ratio between butane and oxygen to find the moles and then mass of oxygen required for the reaction. The mole ratio between butane and water is used to find the moles of water produced and the mole ratio between butane and carbon dioxide is used to find the moles of carbon dioxide produced. The total moles of product are multiplied by Avogadro's number to find the number of product molecules.

Solution:

The balanced chemical equation is:



$$\text{a) Moles of C}_4\text{H}_{10} = (5.50 \text{ mL C}_4\text{H}_{10}) \left(\frac{0.579 \text{ g C}_4\text{H}_{10}}{1 \text{ mL C}_4\text{H}_{10}} \right) \left(\frac{1 \text{ mol C}_4\text{H}_{10}}{58.12 \text{ g C}_4\text{H}_{10}} \right) = 0.054792 \text{ mol C}_4\text{H}_{10}$$

$$\text{Mass (g) of O}_2 = (0.054792 \text{ mol C}_4\text{H}_{10}) \left(\frac{13 \text{ mol O}_2}{2 \text{ mol C}_4\text{H}_{10}} \right) \left(\frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} \right) = 11.3967 = \mathbf{11.4 \text{ g O}_2}$$

$$\text{b) Moles of H}_2\text{O} = (0.054792 \text{ mol C}_4\text{H}_{10}) \left(\frac{10 \text{ mol H}_2\text{O}}{2 \text{ mol C}_4\text{H}_{10}} \right) = 0.27396 = \mathbf{0.274 \text{ mol H}_2\text{O}}$$

$$\text{c) Moles of CO}_2 = (0.054792 \text{ mol C}_4\text{H}_{10}) \left(\frac{8 \text{ mol CO}_2}{2 \text{ mol C}_4\text{H}_{10}} \right) = 0.219168 \text{ mol CO}_2$$

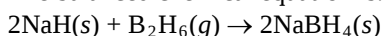
$$\text{Total moles} = 0.27396 \text{ mol H}_2\text{O} + 0.219168 \text{ mol CO}_2 = 0.493128 \text{ mol}$$

$$\text{Total molecules} = (0.493128 \text{ mol}) \left(\frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} \right) = 2.96962 \times 10^{23} = \mathbf{2.97 \times 10^{23} \text{ molecules}}$$

- 3.65 Plan: Write a balanced equation for the reaction. Convert the given mass of each reactant to moles by dividing by the molar mass of that reactant. Use the mole ratio from the balanced chemical equation to find the moles of NaBH₄ formed from each reactant, assuming an excess of the other reactant. The reactant that produces fewer moles of product is the limiting reactant. Convert the moles of NaBH₄ obtained from the limiting reactant to grams using the molar mass. This is the theoretical yield of NaBH₄. Since there is a yield of 88.5%, the amount of NaBH₄ actually obtained will be 88.5% of the theoretical yield.

Solution:

The balanced chemical equation is:



Determining the limiting reactant:

Finding the moles of NaBH₄ from the amount of NaH (if B₂H₆ is limiting):

$$\text{Moles of NaBH}_4 \text{ from NaH} = (7.98 \text{ g NaH}) \left(\frac{1 \text{ mol NaH}}{24.00 \text{ g NaH}} \right) \left(\frac{2 \text{ mol NaBH}_4}{2 \text{ mol NaH}} \right) = 0.3325 \text{ mol NaBH}_4$$

Finding the moles of NaBH₄ from the amount of B₂H₆ (if NaH is limiting):

$$\text{Moles of NaBH}_4 \text{ from B}_2\text{H}_6 = (8.16 \text{ g B}_2\text{H}_6) \left(\frac{1 \text{ mol B}_2\text{H}_6}{27.67 \text{ g B}_2\text{H}_6} \right) \left(\frac{2 \text{ mol NaBH}_4}{1 \text{ mol B}_2\text{H}_6} \right) = 0.58981 \text{ mol NaBH}_4$$

NaH is the limiting reactant, and will be used to calculate the theoretical yield of NaBH₄.

$$\text{Mass (g) of NaBH}_4 = (0.3325 \text{ mol NaBH}_4) \left(\frac{37.83 \text{ g NaBH}_4}{1 \text{ mol NaBH}_4} \right) = 12.5785 \text{ g NaBH}_4$$

$$\% \text{ yield} = \left(\frac{\text{actual Yield}}{\text{theoretical Yield}} \right) \times 100\%$$

$$\text{Mass (g) of NaBH}_4 = \left(\frac{\% \text{ yield}}{100\%} \right) (\text{theoretical yield}) = \left(\frac{88.5\%}{100\%} \right) (12.5785 \text{ g NaBH}_4) = 11.13197 = \mathbf{11.1 \text{ g NaBH}_4}$$

Combining all steps gives:

$$\begin{aligned} \text{Mass (g) of NaBH}_4 &= (7.98 \text{ g NaH}) \left(\frac{1 \text{ mol NaH}}{24.00 \text{ g NaH}} \right) \left(\frac{2 \text{ mol NaBH}_4}{2 \text{ mol NaH}} \right) \left(\frac{37.83 \text{ g NaBH}_4}{1 \text{ mol NaBH}_4} \right) \left(\frac{88.5\%}{100\%} \right) \\ &= 11.13197 = \mathbf{11.1 \text{ g NaBH}_4} \end{aligned}$$

- 3.66 Plan: Recall that molarity = moles of solute/volume (L) of solution. Here you can use the number of particles in place of moles of solute.
Solution:
 a) **Solution B** has the highest molarity as it has the largest number of particles, 12, in a volume of 50 mL.
 b) **Solutions A and F** both have 8 particles in a volume of 50 mL and thus the same molarity. **Solutions C, D, and E** all have 4 particles in a volume of 50 mL and thus have the same molarity.
 c) Mixing Solutions A and C results in $8 + 4 = 12$ particles in a volume of 100 mL. That is a **lower molarity** than that of Solution B which has 12 particles in a volume of 50 mL or 24 particles in a volume of 100 mL.
 d) Adding 50 mL to Solution D would result in 4 particles in a total volume of 100 mL; adding 75 mL to Solution F would result in 4 particles in a volume of 100 mL. The molarity of each solution would be the **same**.
 e) Solution A has 8 particles in a volume of 50 mL while Solution E has the equivalent of 4 particles in a volume of 50 mL. The molarity of Solution E is half that of Solution A. Therefore **half of the volume, 12.5 mL**, of Solution E must be evaporated. When 12.5 mL of solvent is evaporated from Solution E, the result will be 2 particles in 12.5 mL or 8 particles in 50 mL as in Solution A.
- 3.67 Plan: The spheres represent particles of solute and the amount of *solute* per given volume of *solution* determines its concentration. Molarity = moles of solute/volume (L) of solution.
Solution:
 a) **Box C** has more solute added because it contains 2 more spheres than Box A contains.
 b) **Box B** has more solvent because solvent molecules have displaced two solute molecules.
 c) **Box C** has a higher molarity, because it has more moles of solute per volume of solution.
 d) **Box B** has a lower concentration (and molarity), because it has fewer moles of solute per volume of solution.
- 3.68 Plan: In all cases, use the known quantities and the definition of molarity $\left(M = \frac{\text{moles solute}}{\text{L of solution}} \right)$ to find the unknown quantity. Volume must be expressed in liters. The molar mass is used to convert moles to grams. The chemical formulas must be written to determine the molar mass. (a) You will need to convert milliliters to liters, multiply by the molarity to find moles, and convert moles to mass in grams. (b) Convert mass of solute to moles and volume from mL to liters. Divide the moles by the volume. (c) Multiply the molarity by the volume.
Solution:
 a) Calculating moles of solute in solution:

$$\text{Moles of Ca(C}_2\text{H}_3\text{O}_2)_2 = (185.8 \text{ mL}) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) \left(\frac{0.267 \text{ mol Ca(C}_2\text{H}_3\text{O}_2)_2}{1 \text{ L}} \right) = 0.0496086 \text{ mol Ca(C}_2\text{H}_3\text{O}_2)_2$$
 Converting from moles of solute to grams:

$$\begin{aligned} \text{Mass (g) of Ca(C}_2\text{H}_3\text{O}_2)_2 &= (0.0496086 \text{ mol Ca(C}_2\text{H}_3\text{O}_2)_2) \left(\frac{158.17 \text{ g Ca(C}_2\text{H}_3\text{O}_2)_2}{1 \text{ mol Ca(C}_2\text{H}_3\text{O}_2)_2} \right) \\ &= 7.84659 = \mathbf{7.85 \text{ g Ca(C}_2\text{H}_3\text{O}_2)_2} \end{aligned}$$
 b) Converting grams of solute to moles:

$$\text{Moles of KI} = (21.1 \text{ g KI}) \left(\frac{1 \text{ mol KI}}{166.0 \text{ g KI}} \right) = 0.127108 \text{ moles KI}$$

$$\text{Volume (L)} = (500. \text{ mL}) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) = 0.500 \text{ L}$$

$$\text{Molarity of KI} = \frac{0.127108 \text{ mol KI}}{0.500 \text{ L}} = 0.254216 = \mathbf{0.254 \text{ M KI}}$$
 c) Moles of NaCN = $(145.6 \text{ L}) \left(\frac{0.850 \text{ mol NaCN}}{1 \text{ L}} \right) = 123.76 = \mathbf{124 \text{ mol NaCN}}$

- 3.69 Plan: In all cases, use the known quantities and the definition of molarity $\left(M = \frac{\text{moles solute}}{\text{L of solution}} \right)$ to find the unknown quantity. Volume must be expressed in liters. The molar mass is used to convert moles to grams. The chemical formulas must be written to determine the molar mass. (a) You will need to convert mass of solute to moles and divide by the molarity to obtain volume in liters, which is then converted to milliliters. (b) Multiply the volume by the molarity to obtain moles of solute. Use Avogadro's number to determine the number of ions present. (c) Divide mmoles by milliliters; molarity may not only be expressed as moles/L, but also as mmoles/mL.

Solution:

a) Converting mass of solute to moles:

$$\text{Moles of KOH} = (8.42 \text{ g KOH}) \left(\frac{1 \text{ mol KOH}}{56.11 \text{ g KOH}} \right) = 0.15006 \text{ mol KOH}$$

$$\text{Volume (L) of KOH solution} = (0.15006 \text{ mol KOH}) \left(\frac{1 \text{ L}}{2.26 \text{ mol}} \right) = 0.066398 \text{ L KOH solution}$$

$$\text{Volume (mL) of KOH solution} = (0.066398 \text{ L KOH}) \left(\frac{1 \text{ L}}{10^{-3} \text{ mL}} \right) = 66.39823 = \mathbf{66.4 \text{ mL KOH solution}}$$

$$\text{b) Moles of CuCl}_2 = (52 \text{ L}) \left(\frac{2.3 \text{ mol CuCl}_2}{\text{L}} \right) = 119.6 \text{ mol CuCl}_2$$

$$\text{Moles of Cu}^{2+} \text{ ions} = (119.6 \text{ mol CuCl}_2) \left(\frac{1 \text{ mol Cu}^{2+}}{1 \text{ mol CuCl}_2} \right) = 119.6 \text{ mol Cu}^{2+} \text{ ions}$$

Converting moles of ions to number of ions:

$$\text{Number of Cu}^{2+} \text{ ions} = (119.6 \text{ mol Cu}^{2+} \text{ ions}) \left(\frac{6.022 \times 10^{23} \text{ Cu}^{2+} \text{ ions}}{1 \text{ mol Cu}^{2+} \text{ ions}} \right) = 7.2023 \times 10^{25} = \mathbf{7.2 \times 10^{25} \text{ Cu}^{2+} \text{ ions}}$$

$$\text{c) } M \text{ glucose} = \left(\frac{135 \text{ mmol glucose}}{275 \text{ mL}} \right) = 0.490909 = \mathbf{0.491 \text{ M glucose}}$$

Note: Since 1 mmol is 10^{-3} mol and 1 mL is 10^{-3} L, we can use these units instead of converting to mol and L since molarity is a ratio of mol/L. Molarity may not only be expressed as moles/L, but also as mmoles/mL.

- 3.70 Plan: These are dilution problems. Dilution problems can be solved by converting to moles and using the new volume; however, it is much easier to use $M_1 V_1 = M_2 V_2$. The dilution equation does not require a volume in liters; it only requires that the volume units match. In part c), it is necessary to find the moles of sodium ions in each separate solution, add these two mole amounts, and divide by the total volume of the two solutions.

Solution:

$$\text{a) } M_1 = 0.250 \text{ M KCl} \quad V_1 = 37.00 \text{ mL} \quad M_2 = ? \quad V_2 = 150.00 \text{ mL}$$

$$M_1 V_1 = M_2 V_2$$

$$M_2 = \frac{M_1 \times V_1}{V_2} = \frac{(0.250 \text{ M})(37.00 \text{ mL})}{150.0 \text{ mL}} = 0.061667 = \mathbf{0.0617 \text{ M KCl}}$$

$$\text{b) } M_1 = 0.0706 \text{ M (NH}_4)_2\text{SO}_4 \quad V_1 = 25.71 \text{ mL} \quad M_2 = ? \quad V_2 = 500.00 \text{ mL}$$

$$M_1 V_1 = M_2 V_2$$

$$M_2 = \frac{M_1 \times V_1}{V_2} = \frac{(0.0706 \text{ M})(25.71 \text{ mL})}{500.0 \text{ mL}} = 0.003630 = \mathbf{0.00363 \text{ M (NH}_4)_2\text{SO}_4}$$

$$\begin{aligned} \text{c) Moles of Na}^+ \text{ from NaCl solution} &= (3.58 \text{ mL}) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) \left(\frac{0.348 \text{ mol NaCl}}{1 \text{ L}} \right) \left(\frac{1 \text{ mol Na}^+}{1 \text{ mol NaCl}} \right) \\ &= 0.00124584 \text{ mol Na}^+ \end{aligned}$$

$$\begin{aligned}\text{Moles of Na}^+ \text{ from Na}_2\text{SO}_4 \text{ solution} &= (500. \text{ mL}) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) \left(\frac{6.81 \times 10^{-2} \text{ mol Na}_2\text{SO}_4}{1 \text{ L}} \right) \left(\frac{2 \text{ mol Na}^+}{1 \text{ mol Na}_2\text{SO}_4} \right) \\ &= 0.0681 \text{ mol Na}^+\end{aligned}$$

$$\text{Total moles of Na}^+ \text{ ions} = 0.00124584 \text{ mol Na}^+ \text{ ions} + 0.0681 \text{ mol Na}^+ \text{ ions} = 0.06934584 \text{ mol Na}^+ \text{ ions}$$

$$\text{Total volume} = 3.58 \text{ mL} + 500. \text{ mL} = 503.58 \text{ mL} = 0.50358 \text{ L}$$

$$\text{Molarity of Na}^+ = \frac{\text{total moles Na}^+ \text{ ions}}{\text{total volume}} = \frac{0.06934584 \text{ mol Na}^+ \text{ ions}}{0.50358 \text{ L}} = 0.1377057 = \mathbf{0.138 \text{ M Na}^+ \text{ ions}}$$

- 3.71 Plan: These are dilution problems. Dilution problems can be solved by converting to moles and using the new volume; however, it is much easier to use $M_1V_1 = M_2V_2$. The dilution equation does not require a volume in liters; it only requires that the volume units match.

Solution:

$$\begin{aligned}\text{a) } M_1 &= 2.050 \text{ M Cu(NO}_3)_2 & V_1 &= ? & M_2 &= 0.8543 \text{ M Cu(NO}_3)_2 & V_2 &= 750.0 \text{ mL} \\ M_1V_1 &= M_2V_2\end{aligned}$$

$$V_1 = \frac{M_2 \times V_2}{M_1} = \frac{(0.8543 \text{ M})(750.0 \text{ mL})}{2.050 \text{ M}} = 312.5488 = \mathbf{312.5 \text{ mL}}$$

$$\text{b) } M_1 = 1.63 \text{ M CaCl}_2 \quad M_1 \text{ Cl}^- = \left(\frac{1.63 \text{ mol CaCl}_2}{1 \text{ L}} \right) \left(\frac{2 \text{ mol Cl}^-}{1 \text{ mol CaCl}_2} \right) = 3.26 \text{ M Cl}^- \text{ ions}$$

$$\begin{aligned}M_1 &= 3.26 \text{ M Cl}^- & V_1 &= ? & M_2 &= 2.86 \times 10^{-2} \text{ M Cl}^- \text{ ions} & V_2 &= 350. \text{ mL} \\ M_1V_1 &= M_2V_2\end{aligned}$$

$$V_1 = \frac{M_2 \times V_2}{M_1} = \frac{(2.86 \times 10^{-2} \text{ M})(350. \text{ mL})}{3.26 \text{ M}} = 3.07055 = \mathbf{3.07 \text{ mL}}$$

$$\begin{aligned}\text{c) } M_1 &= 0.155 \text{ M Li}_2\text{CO}_3 & V_1 &= 18.0 \text{ mL} & M_2 &= 0.0700 \text{ M Li}_2\text{CO}_3 & V_2 &= ? \\ M_1V_1 &= M_2V_2\end{aligned}$$

$$V_2 = \frac{M_1 \times V_1}{M_2} = \frac{(0.155 \text{ M})(18.0 \text{ mL})}{(0.0700 \text{ M})} = 39.8571 = \mathbf{39.9 \text{ mL}}$$

- 3.72 Plan: Use the density of the solution to find the mass of 1 L of solution. Volume in liters must be converted to volume in mL. The 70.0% by mass translates to 70.0 g solute/100 g solution and is used to find the mass of HNO₃ in 1 L of solution. Convert mass of HNO₃ to moles to obtain moles/L, molarity.

Solution:

$$\text{a) Mass (g) of 1 L of solution} = (1 \text{ L solution}) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right) \left(\frac{1.41 \text{ g solution}}{1 \text{ mL}} \right) = 1410 \text{ g solution}$$

$$\text{Mass (g) of HNO}_3 \text{ in 1 L of solution} = (1410 \text{ g solution}) \left(\frac{70.0 \text{ g HNO}_3}{100 \text{ g solution}} \right) = \mathbf{987 \text{ g HNO}_3/\text{L}}$$

$$\text{b) Moles of HNO}_3 = (987 \text{ g HNO}_3) \left(\frac{1 \text{ mol HNO}_3}{63.02 \text{ g HNO}_3} \right) = 15.6617 \text{ mol HNO}_3$$

$$\text{Molarity of HNO}_3 = \left(\frac{15.6617 \text{ mol HNO}_3}{1 \text{ L solution}} \right) = 15.6617 = \mathbf{15.7 \text{ M HNO}_3}$$

- 3.73 Plan: Use the molarity of the solution to find the moles of H₂SO₄ in 1 mL. Convert moles of H₂SO₄ to mass of H₂SO₄, divide that mass by the mass of 1 mL of solution, and multiply by 100 for mass percent. Use the density of the solution to find the mass of 1 mL of solution.

Solution:

$$\text{a) Moles of H}_2\text{SO}_4 \text{ in 1 mL} = \left(\frac{18.3 \text{ mol H}_2\text{SO}_4}{1 \text{ L}} \right) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) = \mathbf{1.83 \times 10^{-2} \text{ mol H}_2\text{SO}_4/\text{mL}}$$

$$\text{b) Mass of H}_2\text{SO}_4 \text{ in 1 mL} = (1.83 \times 10^{-2} \text{ mol H}_2\text{SO}_4) \left(\frac{98.09 \text{ g H}_2\text{SO}_4}{1 \text{ mol H}_2\text{SO}_4} \right) = 1.79505 \text{ g H}_2\text{SO}_4$$

$$\text{Mass of 1 mL of solution} = (1 \text{ mL}) \left(\frac{1.84 \text{ g}}{1 \text{ mL}} \right) = 1.84 \text{ g solution}$$

$$\text{Mass percent} = \frac{\text{mass of H}_2\text{SO}_4}{\text{mass of solution}} (100) = \frac{1.79505 \text{ g H}_2\text{SO}_4}{1.84 \text{ g solution}} (100) = 97.5571 = \mathbf{97.6\% \text{ H}_2\text{SO}_4 \text{ by mass}}$$

- 3.74 Plan: Convert the mass of calcium carbonate to moles, and use the mole ratio in the balanced chemical equation to find the moles of hydrochloric acid required to react with these moles of calcium carbonate. Use the molarity of HCl to find the volume that contains this number of moles.

Solution:



Converting from grams of CaCO_3 to moles:

$$\text{Moles of CaCO}_3 = (16.2 \text{ g CaCO}_3) \left(\frac{1 \text{ mol CaCO}_3}{100.09 \text{ g CaCO}_3} \right) = 0.161854 \text{ mol CaCO}_3$$

Converting from moles of CaCO_3 to moles of HCl:

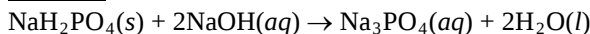
$$\text{Moles of HCl} = (0.161854 \text{ mol CaCO}_3) \left(\frac{2 \text{ mol HCl}}{1 \text{ mol CaCO}_3} \right) = 0.323708 \text{ mol HCl}$$

Converting from moles of HCl to volume:

$$\text{Volume (mL) of HCl} = (0.323708 \text{ mol HCl}) \left(\frac{1 \text{ L}}{0.383 \text{ mol HCl}} \right) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right) = 845.1906 = \mathbf{845 \text{ mL HCl solution}}$$

- 3.75 Plan: Convert the volume of NaOH solution to liters and multiply by the molarity of the solution to obtain moles of NaOH. Use the mole ratio in the balanced chemical equation to find the moles of NaH_2PO_4 required to react with these moles of NaOH. Finally, convert moles of NaH_2PO_4 to moles.

Solution:



$$\text{Volume (L)} = (43.74 \text{ mL}) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) = 0.04374 \text{ L}$$

Finding moles of NaOH:

$$\text{Moles of NaOH} = (0.04374 \text{ L}) \left(\frac{0.285 \text{ mol NaOH}}{1 \text{ L}} \right) = 0.0124659 \text{ mol NaOH}$$

Converting from moles of NaOH to moles of NaH_2PO_4 :

$$\text{Moles of NaH}_2\text{PO}_4 = (0.0124659 \text{ mol NaOH}) \left(\frac{1 \text{ mol NaH}_2\text{PO}_4}{2 \text{ mol NaOH}} \right) = 0.00623295 \text{ mol NaH}_2\text{PO}_4$$

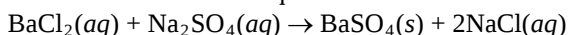
Converting from moles of NaH_2PO_4 to mass:

$$\text{Mass (g) of NaH}_2\text{PO}_4 = (0.00623295 \text{ mol NaH}_2\text{PO}_4) \left(\frac{119.98 \text{ g NaH}_2\text{PO}_4}{1 \text{ mol NaH}_2\text{PO}_4} \right) = 0.747829 = \mathbf{0.748 \text{ g NaH}_2\text{PO}_4}$$

- 3.76 Plan: The first step is to write and balance the chemical equation for the reaction. Multiply the molarity and volume of each of the reactants to determine the moles of each. To determine which reactant is limiting, calculate the amount of barium sulfate formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the mass of barium sulfate formed.

Solution:

The balanced chemical equation is:



$$\text{Moles of BaCl}_2 = (35.0 \text{ mL}) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) \left(\frac{0.160 \text{ mol BaCl}_2}{1 \text{ L}} \right) = 0.00560 \text{ mol BaCl}_2$$

Finding the moles of BaSO₄ from the moles of BaCl₂ (if Na₂SO₄ is limiting):

$$\text{Moles of BaSO}_4 \text{ from BaCl}_2 = (0.00560 \text{ mol BaCl}_2) \left(\frac{1 \text{ mol BaSO}_4}{1 \text{ mol BaCl}_2} \right) = 0.00560 \text{ mol BaSO}_4$$

$$\text{Moles of Na}_2\text{SO}_4 = (58.0 \text{ mL}) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) \left(\frac{0.065 \text{ mol Na}_2\text{SO}_4}{1 \text{ L}} \right) = 0.00377 \text{ mol Na}_2\text{SO}_4$$

Finding the moles of BaSO₄ from the moles of Na₂SO₄ (if BaCl₂ is limiting):

$$\text{Moles BaSO}_4 \text{ from Na}_2\text{SO}_4 = (0.00377 \text{ mol Na}_2\text{SO}_4) \left(\frac{1 \text{ mol BaSO}_4}{1 \text{ mol Na}_2\text{SO}_4} \right) = 0.00377 \text{ mol BaSO}_4$$

Sodium sulfate is the limiting reactant.

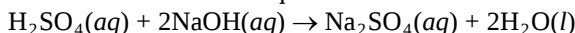
Converting from moles of BaSO₄ to mass:

$$\text{Mass (g) of BaSO}_4 = (0.00377 \text{ mol BaSO}_4) \left(\frac{233.4 \text{ g BaSO}_4}{1 \text{ mol BaSO}_4} \right) = 0.879918 = \mathbf{0.88 \text{ g BaSO}_4}$$

- 3.77 Plan: The first step is to write and balance the chemical equation for the reaction. Use the molarity and volume of each of the reactants to determine the moles of each. To determine which reactant is limiting, calculate the amount of either product formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the amount of the excess reactant that reacts. The difference between the amount of excess reactant that reacts and the initial amount of reactant supplied gives the amount of excess reactant remaining.

Solution:

The balanced chemical equation is:



We can use either product to determine the limiting reactant. We will use sodium sulfate.

$$\text{Moles of H}_2\text{SO}_4 = (350.0 \text{ mL}) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) \left(\frac{0.210 \text{ mol H}_2\text{SO}_4}{1 \text{ L}} \right) = 0.0735 \text{ mol H}_2\text{SO}_4$$

Finding the moles of Na₂SO₄ from the moles of H₂SO₄ (if NaOH is limiting):

$$\text{Moles of Na}_2\text{SO}_4 \text{ from H}_2\text{SO}_4 = (0.0735 \text{ mol H}_2\text{SO}_4) \left(\frac{1 \text{ mol Na}_2\text{SO}_4}{1 \text{ mol H}_2\text{SO}_4} \right) = 0.0735 \text{ mol Na}_2\text{SO}_4$$

$$\text{Moles of NaOH} = (0.500 \text{ L}) \left(\frac{0.196 \text{ mol NaOH}}{1 \text{ L}} \right) = 0.0980 \text{ mol NaOH}$$

Finding the moles of Na₂SO₄ from the moles of NaOH (if H₂SO₄ is limiting):

$$\text{Moles of Na}_2\text{SO}_4 \text{ from NaOH} = (0.0980 \text{ mol NaOH}) \left(\frac{1 \text{ mol Na}_2\text{SO}_4}{2 \text{ mol NaOH}} \right) = 0.0490 \text{ mol Na}_2\text{SO}_4$$

NaOH is the limiting reactant and will be used in the remainder of the calculations.

$$\text{Moles of H}_2\text{SO}_4 \text{ that react with NaOH} = (0.0980 \text{ mol NaOH}) \left(\frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}} \right) = 0.0490 \text{ mol H}_2\text{SO}_4$$

$$\begin{aligned} \text{Moles of H}_2\text{SO}_4 \text{ remaining} &= \text{initial moles} - \text{moles reacting with NaOH} \\ &= 0.0735 \text{ mol} - 0.0490 \text{ mol} = \mathbf{0.0245 \text{ mol H}_2\text{SO}_4} \end{aligned}$$

- 3.78 Plan: The first part of the problem is a simple dilution problem ($M_1V_1 = M_2V_2$). The volume in units of gallons can be used. In part b), convert mass of HCl to moles and use the molarity to find the volume that contains that number of moles.

Solution:

$$\text{a) } M_1 = 11.7 \text{ M} \quad V_1 = ? \quad M_2 = 3.5 \text{ M} \quad V_2 = 3.0 \text{ gal}$$

$$V_1 = \frac{M_2 \times V_2}{M_1} = \frac{(3.5 \text{ M})(3.0 \text{ gal})}{11.7 \text{ M}} = 0.897436 \text{ gal}$$

Instructions: Be sure to wear goggles to protect your eyes! Pour approximately 2.0 gal of water into the container. Add slowly and with mixing 0.90 gal of 11.7 M HCl into the water. Dilute to 3.0 gal with water.

b) Converting from mass of HCl to moles of HCl:

$$\text{Moles of HCl} = (9.66 \text{ g HCl}) \left(\frac{1 \text{ mol HCl}}{36.46 \text{ g HCl}} \right) = 0.264948 \text{ mol HCl}$$

Converting from moles of HCl to volume:

$$\begin{aligned} \text{Volume (mL) of solution} &= (0.264948 \text{ mol HCl}) \left(\frac{1 \text{ L}}{11.7 \text{ mol HCl}} \right) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right) \\ &= 22.64513 = \mathbf{22.6 \text{ mL muriatic acid solution}} \end{aligned}$$

- 3.79 Plan: The moles of narceine and the moles of water are required. We can assume any mass of narceine hydrate (we will use 100 g), and use this mass to determine the moles of hydrate. The moles of water in the hydrate is obtained by taking 10.8% of the 100 g mass of hydrate and converting the mass to moles of water. Divide the moles of water by the moles of hydrate to find the value of x.

Solution:

Assuming a 100 g sample of narceine hydrate:

$$\begin{aligned} \text{Moles of narceine hydrate} &= (100 \text{ g narceine hydrate}) \left(\frac{1 \text{ mol narceine hydrate}}{499.52 \text{ g narceine hydrate}} \right) \\ &= 0.20019 \text{ mol narceine hydrate} \end{aligned}$$

$$\text{Mass (g) of H}_2\text{O} = (100 \text{ g narceine hydrate}) \left(\frac{10.8\% \text{ H}_2\text{O}}{100\% \text{ narceine hydrate}} \right) = 10.8 \text{ g H}_2\text{O}$$

$$\text{Moles of H}_2\text{O} = (10.8 \text{ g H}_2\text{O}) \left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) = 0.59933 \text{ mol H}_2\text{O}$$

$$x = \frac{\text{moles of H}_2\text{O}}{\text{moles of hydrate}} = \frac{0.59933 \text{ mol}}{0.20019 \text{ mol}} = \mathbf{3}$$

Thus, there are three water molecules per mole of hydrate. The formula for narceine hydrate is **narceine•3H₂O**.

- 3.80 Plan: Determine the formula and the molar mass of each compound. The formula gives the relative numbers of moles of each element present. Multiply the number of moles of each element by its molar mass to find the total

mass of element in 1 mole of compound. Mass percent = $\frac{\text{total mass of element}}{\text{molar mass of compound}}(100)$. List the compounds

from the highest %H to the lowest.

Solution:

Name	Chemical formula	Molar mass (g/mol)	Mass percent H = $\frac{\text{moles of H} \times \text{molar mass}}{\text{molar mass of compound}}(100)$
Ethane	C ₂ H ₆	30.07	$\frac{6 \text{ mol}(1.008 \text{ g/mol})}{30.07 \text{ g}}(100) = 20.11\% \text{ H}$
Propane	C ₃ H ₈	44.09	$\frac{8 \text{ mol}(1.008 \text{ g/mol})}{44.09 \text{ g}}(100) = 18.29\% \text{ H}$
Benzene	C ₆ H ₆	78.11	$\frac{6 \text{ mol}(1.008 \text{ g/mol})}{78.11 \text{ g}}(100) = 7.743\% \text{ H}$

Ethanol	C ₂ H ₅ OH	46.07	$\frac{6 \text{ mol}(1.008 \text{ g/mol})}{46.07 \text{ g}}(100) = 13.13\% \text{ H}$
Cetyl palmitate	C ₃₂ H ₆₄ O ₂	480.83	$\frac{64 \text{ mol}(1.008 \text{ g/mol})}{480.83 \text{ g}}(100) = 13.42\% \text{ H}$

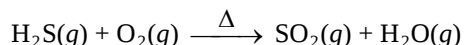
The hydrogen percentage decreases in the following order:

Ethane > Propane > Cetyl palmitate > Ethanol > Benzene

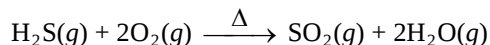
- 3.81 Plan: The names must first be converted to chemical formulas. Balancing is a trial-and-error procedure. Balance one element at a time, placing coefficients where needed to have the same number of atoms of a particular element on each side of the equation. The smallest whole-number coefficients should be used. Remember that oxygen, chlorine, and hydrogen are diatomic.

Solution:

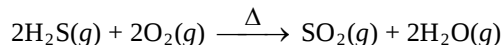
a) All of the substances are gases.



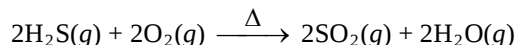
There are 2 O atoms in O₂ on the left and 3 O atoms in SO₂ and H₂O on the right; place a coefficient of 2 in front of H₂O on the right and a coefficient of 2 in front of O₂ on the left for a total of 4 oxygen atoms on each side:



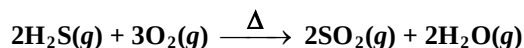
Now the 4 H atoms in 2H₂O on the right require a coefficient of 2 in front of H₂S on the left:



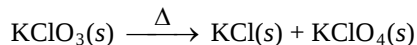
The 2 S atoms in 2H₂S on the left require a coefficient of 2 in front of SO₂ on the right:



Now the O atoms are no longer balanced; the 6 O atoms on the right (4 in 2SO₂ and 2 in 2H₂O) require a coefficient of 3 in front of O₂ on the left:



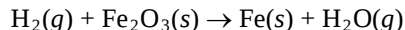
b) All of the substances are solid (crystalline).



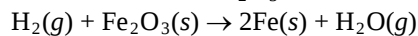
There are 3 O atoms in KClO₃ on the left and 4 O atoms in KClO₄ on the right. Place a coefficient of 4 in front of KClO₃ and a coefficient of 3 in front of KClO₄ for a total of 12 O atoms on each side. The K and Cl atoms are balanced with 4 K atoms and 4 Cl atoms on each side:



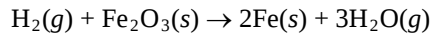
c) Hydrogen and water vapor are gases; iron and iron(III) oxide are solids.



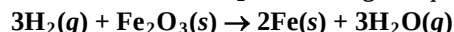
The 2 Fe atoms in Fe₂O₃ on the left require a coefficient of 2 in front of Fe on the right:



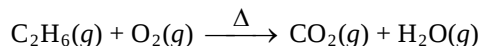
The 3 O atoms in Fe₂O₃ on the left require a coefficient of 3 in front of H₂O on the right:



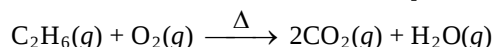
The 6 H atoms in 3H₂O on the right require a coefficient of 3 in front of H₂ on the left:



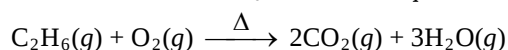
d) All of the substances are gases; combustion required oxygen as a reactant.



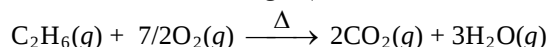
The 2 C atoms in C₂H₆ on the left require a coefficient of 2 in front of CO₂ on the right:



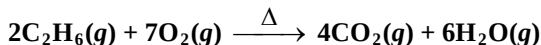
The 6 H atoms in C₂H₆ on the left require a coefficient of 3 in front of H₂O on the right:



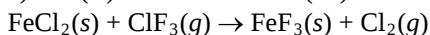
The 7 O atoms on the right (4 in 2CO₂ and 3 in 3H₂O) require a coefficient of 7/2 in front of O₂ on the left:



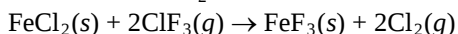
Double all coefficients to get whole number coefficients:



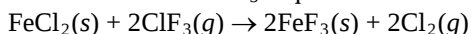
e) Iron(II) chloride and iron(III) fluoride are solids and the other substances are gases.



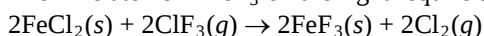
There are 3 Cl atoms on the left (2 in FeCl_2 and 1 in ClF_3) and 2 Cl atoms in Cl_2 on the right. Place a coefficient of 2 in front of Cl_2 and a coefficient of 2 in front of ClF_3 on the left for a total of 4 Cl atoms on each side:



The 6 F atoms in 2ClF_3 require a coefficient of 2 in front of FeF_3 on the right:



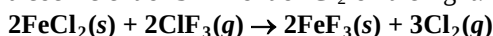
The 2 Fe atoms in FeF_3 on the right require a coefficient of 2 in front of FeCl_2 on the left:



Now the Cl atoms are not balanced with 6 on the left (4 in 2FeCl_2 and 2 in 2ClF_3) and 4 in 2Cl_2 on the right;

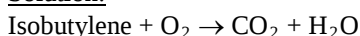
place

a coefficient of 3 in front of Cl_2 on the right:



- 3.82 Plan: In combustion analysis, finding the moles of carbon and hydrogen is relatively simple because all of the carbon present in the sample is found in the carbon of CO_2 , and all of the hydrogen present in the sample is found in the hydrogen of H_2O . Convert the mass of CO_2 to moles and use the ratio between CO_2 and C to find the moles and mass of C present. Do the same to find the moles and mass of H from H_2O . Divide the moles of C and H by the smaller value to convert to whole numbers to get the empirical formula.

Solution:

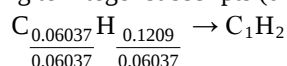


$$\text{Moles of C} = (2.657 \text{ g CO}_2) \left(\frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \right) \left(\frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \right) = 0.06037 \text{ mol C}$$

$$\text{Moles of H} = (1.089 \text{ g H}_2\text{O}) \left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) \left(\frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \right) = 0.1209 \text{ mol H}$$

$$\text{Preliminary formula} = \text{C}_{0.06037}\text{H}_{0.1209}$$

Converting to integer subscripts (dividing all by the smallest subscript):

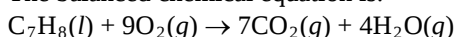


This gives an empirical formula of **CH_2** .

- 3.83 Plan: Write a balanced equation. Use the density of toluene to convert the given volume of toluene to mass and divide by the molar mass of toluene to convert mass to moles. Use the mole ratio between toluene and oxygen to find the moles and then mass of oxygen required for the reaction. The mole ratio between toluene and the gaseous products are used to find the moles of product produced. The moles of water are multiplied by Avogadro's number to find the number of water molecules.

Solution:

The balanced chemical equation is:



$$\text{a) Moles of C}_7\text{H}_8 = (20.0 \text{ mL C}_7\text{H}_8) \left(\frac{0.867 \text{ g C}_7\text{H}_8}{1 \text{ mL C}_7\text{H}_8} \right) \left(\frac{1 \text{ mol C}_7\text{H}_8}{92.13 \text{ g C}_7\text{H}_8} \right) = 0.1882123 \text{ mol C}_7\text{H}_8$$

$$\text{Mass (g) oxygen} = (0.1882123 \text{ mol C}_7\text{H}_8) \left(\frac{9 \text{ mol O}_2}{1 \text{ mol C}_7\text{H}_8} \right) \left(\frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} \right) = 54.20514 = \mathbf{54.2 \text{ g O}_2}$$

$$\text{b) Total moles of gas} = (0.1882123 \text{ mol C}_7\text{H}_8) \left(\frac{11 \text{ mol product gas}}{1 \text{ mol C}_7\text{H}_8} \right) = 2.07034 = \mathbf{2.07 \text{ mol of gas}}$$

The 11 mol of gas is an exact, not measured, number, so it does not affect the significant figures.

$$\text{c) Moles of H}_2\text{O} = (0.1882123 \text{ mol C}_7\text{H}_8) \left(\frac{4 \text{ mol H}_2\text{O}}{1 \text{ mol C}_7\text{H}_8} \right) = 0.7528492 \text{ mol H}_2\text{O}$$

$$\begin{aligned}\text{Molecules of H}_2\text{O} &= (0.7528492 \text{ mol H}_2\text{O}) \left(\frac{6.022 \times 10^{23} \text{ H}_2\text{O molecules}}{1 \text{ mol H}_2\text{O}} \right) \\ &= 4.53366 \times 10^{23} = \mathbf{4.53 \times 10^{23} \text{ molecules H}_2\text{O}}\end{aligned}$$

- 3.84 **Plan:** If 100.0 g of dinitrogen tetroxide reacts with 100.0 g of hydrazine (N_2H_4), what is the theoretical yield of nitrogen if no side reaction takes place? First, we need to identify the limiting reactant. To determine which reactant is limiting, calculate the amount of nitrogen formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the theoretical yield of nitrogen. Then determine the amount of limiting reactant required to produce 10.0 grams of NO. Reduce the amount of limiting reactant by the amount used to produce NO. The reduced amount of limiting reactant is then used to calculate an “actual yield.” The “actual” and theoretical yields will give the maximum percent yield.

Solution:

The balanced reaction is $2\text{N}_2\text{H}_4(l) + \text{N}_2\text{O}_4(l) \rightarrow 3\text{N}_2(g) + 4\text{H}_2\text{O}(g)$

Determining the limiting reactant:

Finding the moles of N_2 from the amount of N_2O_4 (if N_2H_4 is limiting):

$$\text{Moles of N}_2 \text{ from N}_2\text{O}_4 = (100.0 \text{ g N}_2\text{O}_4) \left(\frac{1 \text{ mol N}_2\text{O}_4}{92.02 \text{ g N}_2\text{O}_4} \right) \left(\frac{3 \text{ mol N}_2}{1 \text{ mol N}_2\text{O}_4} \right) = 3.26016 \text{ mol N}_2$$

Finding the moles of N_2 from the amount of N_2H_4 (if N_2O_4 is limiting):

$$\text{N}_2 \text{ from N}_2\text{H}_4 = (100.0 \text{ g N}_2\text{H}_4) \left(\frac{1 \text{ mol N}_2\text{H}_4}{32.05 \text{ g N}_2\text{H}_4} \right) \left(\frac{3 \text{ mol N}_2}{2 \text{ mol N}_2\text{H}_4} \right) = 4.68019 \text{ mol N}_2$$

N_2O_4 is the limiting reactant.

$$\text{Theoretical yield of N}_2 = (100.0 \text{ g N}_2\text{O}_4) \left(\frac{1 \text{ mol N}_2\text{O}_4}{92.02 \text{ g N}_2\text{O}_4} \right) \left(\frac{3 \text{ mol N}_2}{1 \text{ mol N}_2\text{O}_4} \right) \left(\frac{28.02 \text{ g N}_2}{1 \text{ mol N}_2} \right) = 91.3497 \text{ g N}_2$$

How much of the limiting reactant is used to produce 10.0 g NO?

$\text{N}_2\text{H}_4(l) + 2\text{N}_2\text{O}_4(l) \rightarrow 6\text{NO}(g) + 2\text{H}_2\text{O}(g)$

$$\begin{aligned}\text{Mass (g) of N}_2\text{O}_4 \text{ used} &= (10.0 \text{ g NO}) \left(\frac{1 \text{ mol NO}}{30.01 \text{ g NO}} \right) \left(\frac{2 \text{ mol N}_2\text{O}_4}{6 \text{ mol NO}} \right) \left(\frac{92.02 \text{ g N}_2\text{O}_4}{1 \text{ mol N}_2\text{O}_4} \right) \\ &= 10.221 \text{ g N}_2\text{O}_4\end{aligned}$$

$$\begin{aligned}\text{Amount of N}_2\text{O}_4 \text{ available to produce N}_2 &= 100.0 \text{ g N}_2\text{O}_4 - \text{mass of N}_2\text{O}_4 \text{ required to produce 10.0 g NO} \\ &= 100.0 \text{ g} - 10.221 \text{ g} = 89.779 \text{ g N}_2\text{O}_4\end{aligned}$$

Determine the “actual yield” of N_2 from 89.779 g N_2O_4 :

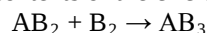
$$\begin{aligned}\text{“Actual yield” of N}_2 &= (89.779 \text{ g N}_2\text{O}_4) \left(\frac{1 \text{ mol N}_2\text{O}_4}{92.02 \text{ g N}_2\text{O}_4} \right) \left(\frac{3 \text{ mol N}_2}{1 \text{ mol N}_2\text{O}_4} \right) \left(\frac{28.02 \text{ g N}_2}{1 \text{ mol N}_2} \right) \\ &= 82.01285 \text{ g N}_2\end{aligned}$$

$$\text{Theoretical yield} = \left(\frac{\text{actual yield}}{\text{theoretical yield}} \right) (100) = \left(\frac{82.01285}{91.3497} \right) (100) = 89.7790 = \mathbf{89.8\%}$$

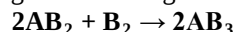
- 3.85 **Plan:** Identify the product molecules and write the balanced equation. To determine the limiting reactant for part b), examine the product circle to see which reactant remains in excess and which reactant was totally consumed. For part c), use the mole ratios in the balanced equation to determine the number of moles of product formed by each reactant, assuming the other reactant is in excess. The reactant that produces fewer moles of product is the limiting reactant. Use the mole ratio between the two reactants to determine the moles of excess reactant required to react with the limiting reactant. The difference between the initial moles of excess reactant and the moles required for reaction is the moles of excess reactant that remain.

Solution:

a) The contents of the circles give:



Balancing the reaction gives:



b) Two B_2 molecules remain after reaction so B_2 is in excess. All of the AB_2 molecules have reacted so AB_2 is the limiting reactant.

c) Finding the moles of AB_3 from the moles of AB_2 (if B_2 is limiting):

$$\text{Moles of AB}_3 \text{ from AB}_2 = (5.0 \text{ mol AB}_2) \left(\frac{2 \text{ mol AB}_3}{2 \text{ mol AB}_2} \right) = 5.0 \text{ mol AB}_3$$

Finding the moles of AB_3 from the moles of B_2 (if AB_2 is limiting):

$$\text{Moles of AB}_3 \text{ from B}_2 = (3.0 \text{ mol B}_2) \left(\frac{2 \text{ mol AB}_3}{1 \text{ mol B}_2} \right) = 6.0 \text{ mol AB}_3$$

AB_2 is the limiting reagent and **5.0 mol of AB_3** is formed.

$$\text{d) Moles of B}_2 \text{ that react with 5.0 mol AB}_2 = (5.0 \text{ mol AB}_2) \left(\frac{1 \text{ mol B}_2}{2 \text{ mol AB}_2} \right) = 2.5 \text{ mol B}_2$$

The unreacted B_2 is $3.0 \text{ mol} - 2.5 \text{ mol} = \mathbf{0.5 \text{ mol B}_2}$.

- 3.86 Plan: Since 85% of ions in seawater are from NaCl, take 85% of the mass percent of dissolved ions (4.0%) to find the mass % of NaCl in part a). To find the mass % of Na^+ and Cl^- individually in part b), use the ratio of the mass of the two ions to the mass of NaCl. To find the molarity in part c), use the mass of NaCl in 100 g of seawater; convert mass of NaCl to moles and mass of seawater to volume in liters, using the density. Molarity = moles of NaCl/L of seawater.

Solution:

$$\text{a) } (4.0\% \text{ ions}) \left(\frac{85\% \text{ NaCl}}{100\% \text{ ions}} \right) = \mathbf{3.4\% \text{ NaCl}}$$

$$\text{b) \% Na}^+ \text{ ions} = (3.4\% \text{ NaCl}) \left(\frac{22.99 \text{ g Na}^+}{58.44 \text{ g NaCl}} \right) = 1.3375 = \mathbf{1.3\% \text{ Na}^+ \text{ ions}}$$

$$\% \text{ Cl}^- \text{ ions} = (3.4\% \text{ NaCl}) \left(\frac{35.45 \text{ g Cl}^-}{58.44 \text{ g NaCl}} \right) = 2.062 = \mathbf{2.1\% \text{ Cl}^- \text{ ions}}$$

c) Since the mass % of NaCl is 3.4%, there are 3.4 g of NaCl in 100 g of seawater.

$$\text{Moles of NaCl} = (3.4 \text{ g NaCl}) \left(\frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} \right) = 0.0581793 \text{ mol NaCl}$$

$$\text{Volume (L) of 100 g of seawater} = (100 \text{ g seawater}) \left(\frac{1 \text{ mL}}{1.025 \text{ g seawater}} \right) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) = 0.097561 \text{ L}$$

$$M \text{ NaCl} = \frac{\text{moles NaCl}}{\text{L seawater}} = \frac{0.0581793 \text{ mol}}{0.097561 \text{ L}} = 0.596338 = \mathbf{0.60 \text{ M NaCl}}$$

- 3.87 a) **False**, a mole of one substance has the same number of units as a mole of any other substance.
b) **True**
c) **False**, a limiting-reactant problem is present when the quantity of available material is given for more than one reactant.
d) **True**

- 3.88 Plan: Count the total number of spheres in each box. The number in box A divided by the volume change in each part will give the number we are looking for and allow us to match boxes.

Solution:

The number in each box is: A = 12, B = 6, C = 4, and D = 3.

- a) When the volume is tripled, there should be $12/3 = 4$ spheres in a box. This is box **C**.
b) When the volume is doubled, there should be $12/2 = 6$ spheres in a box. This is box **B**.
c) When the volume is quadrupled, there should be $12/4 = 3$ spheres in a box. This is box **D**.

- 3.89 **Plan:** To convert mass to moles, divide the mass by the molar mass of the substance. To convert moles to mass, divide by the molar mass. To obtain number of particles, multiply moles by Avogadro's number. Divide a number of particles by Avogadro's number to obtain moles.

Solution:

a) Since 1 mole of any substance contains Avogadro's number of entities, equal amounts of moles of various substances contain equal numbers of entities. The number of entities (O_3 molecules) in 0.4 mol of O_3 is **equal** to the number of entities (O atoms) in 0.4 mol of O atoms.

b) O_3 has a molar mass of $3(16.0 \text{ g/mol O}) = 48.0 \text{ g/mol}$; O has a molar mass of $1(16.0 \text{ g/mol O}) = 16.0 \text{ g/mol}$. Since O_3 has a larger molar mass than O, **0.4 mol of O_3** has a greater mass than 0.4 mol of O.

$$\text{c) Moles of } \text{N}_2\text{O}_4 = (4.0 \text{ g } \text{N}_2\text{O}_4) \left(\frac{1 \text{ mol } \text{N}_2\text{O}_4}{92.02 \text{ g } \text{N}_2\text{O}_4} \right) = 0.043 \text{ mol } \text{N}_2\text{O}_4$$

$$\text{Moles of } \text{SO}_2 = (3.3 \text{ g } \text{SO}_2) \left(\frac{1 \text{ mol } \text{SO}_2}{64.07 \text{ g } \text{SO}_2} \right) = 0.052 \text{ mol } \text{SO}_2$$

SO_2 is the larger quantity in terms of moles.

$$\text{d) Mass (g) of } \text{C}_2\text{H}_4 = (0.6 \text{ mol } \text{C}_2\text{H}_4) \left(\frac{28.05 \text{ g } \text{C}_2\text{H}_4}{1 \text{ mol } \text{C}_2\text{H}_4} \right) = 17 \text{ g } \text{C}_2\text{H}_4$$

$$\text{Mass (g) of } \text{F}_2 = (0.6 \text{ mol } \text{F}_2) \left(\frac{38.00 \text{ g } \text{F}_2}{1 \text{ mol } \text{F}_2} \right) = 23 \text{ g } \text{F}_2$$

F_2 is the greater quantity in terms of mass.

Note that if each of these values is properly rounded to one significant figure, the answers are identical.

$$\text{e) Total moles of ions in } 2.3 \text{ mol } \text{NaClO}_3 = (2.3 \text{ mol } \text{NaClO}_3) \left(\frac{2 \text{ mol ions}}{1 \text{ mol } \text{NaClO}_3} \right) = 4.6 \text{ mol ions}$$

$$\text{Total moles of ions in } 2.2 \text{ mol } \text{MgCl}_2 = (2.2 \text{ mol } \text{MgCl}_2) \left(\frac{3 \text{ mol ions}}{1 \text{ mol } \text{MgCl}_2} \right) = 6.6 \text{ mol ions}$$

MgCl_2 is the greater quantity in terms of total moles of ions.

f) The compound with the lower molar mass will have more molecules in a given mass. H_2O (18.02 g/mol) has a lower molar mass than H_2O_2 (34.02 g/mol). **1.0 g H_2O** has more molecules than 1.0 g H_2O_2 .

$$\text{g) Moles of } \text{NaBr} = (0.500 \text{ L } \text{NaBr}) \left(\frac{0.500 \text{ mol}}{1 \text{ L}} \right) = 0.250 \text{ mol } \text{NaBr}$$

$$\text{Moles of } \text{Na}^+ = (0.250 \text{ mol } \text{NaBr}) \left(\frac{1 \text{ mol } \text{Na}^+}{1 \text{ mol } \text{NaBr}} \right) = 0.250 \text{ mol } \text{Na}^+$$

$$\text{Moles of } \text{NaCl} = (0.0146 \text{ kg } \text{NaCl}) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left(\frac{1 \text{ mol } \text{NaCl}}{58.44 \text{ g } \text{NaCl}} \right) = 0.250 \text{ mol } \text{NaCl}$$

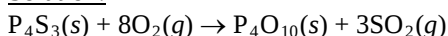
$$\text{Moles of } \text{Na}^+ = (0.250 \text{ mol } \text{NaCl}) \left(\frac{2 \text{ mol ions}}{1 \text{ mol } \text{NaCl}} \right) = 0.250 \text{ mol } \text{Na}^+$$

The two quantities are **equal**.

h) The heavier atoms, ^{238}U , will give a greater total mass since there is an equal number of particles of both.

- 3.90 **Plan:** Write a balanced equation. The coefficients in the balanced equation give the number of molecules or moles of each reactant and product. Moles are converted to amount in grams by multiplying by the molar masses.

Solution:



a) 1 molecule of P_4S_3 reacts with 8 molecules of O_2 to produce 1 molecule of P_4O_{10} and 3 molecules of SO_2 .

b) 1 mol of P_4S_3 reacts with 8 mol of O_2 to produce 1 mol of P_4O_{10} and 3 mol of SO_2 .

c) 220.09 g of P_4S_3 react with $8(32.00 \text{ g/mol O}) = 256.00 \text{ g}$ of O_2 to produce 283.88 g of P_4O_{10} and $3(64.07 \text{ g/mol } \text{SO}_2) = 192.21 \text{ g}$ of SO_2 .

- 3.91 Plan: Write a balanced equation. Use the actual yield (105 kg) and the percent yield (98.8%) to find the theoretical yield of hydrogen. Use the mole ratio between hydrogen and water in the balanced equation to obtain the amount of hydrogen required to produce that theoretical yield of water.

Solution:

The balanced equation is $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g})$

$$\% \text{ yield} = \left(\frac{\text{actual yield}}{\text{theoretical yield}} \right) \times 100\%$$

$$\text{Theoretical yield (g) of H}_2\text{O} = \frac{\text{actual yield}}{\% \text{ yield}}(100) = \frac{105 \text{ kg}}{98.8\%}(100) = 106.2753 \text{ kg H}_2\text{O}$$

$$\begin{aligned} \text{Mass (g) of H}_2 &= (106.2753 \text{ kg H}_2\text{O}) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) \left(\frac{2 \text{ mol H}_2}{2 \text{ mol H}_2\text{O}} \right) \left(\frac{2.016 \text{ g H}_2}{1 \text{ mol H}_2} \right) \\ &= 1.18896 \times 10^4 = \mathbf{1.19 \times 10^4 \text{ g H}_2} \end{aligned}$$

- 3.92 Plan: This problem may be done as two dilution problems with the two final molarities added, or, as done here, it may be done by calculating, then adding the moles and dividing by the total volume.

Solution:

$$M \text{ KBr} = \frac{\text{total moles KBr}}{\text{total volume}} = \frac{\text{moles KBr from solution 1} + \text{moles KBr from solution 2}}{\text{volume solution 1} + \text{volume solution 2}}$$

$$M \text{ KBr} = \frac{\left(\frac{0.053 \text{ mol KBr}}{1 \text{ L}} \right)(0.200 \text{ L}) + \left(\frac{0.078 \text{ mol KBr}}{1 \text{ L}} \right)(0.550 \text{ L})}{0.200 \text{ L} + 0.550 \text{ L}} = 0.071333 = \mathbf{0.071 \text{ M KBr}}$$

- 3.93 Plan: Divide the given mass of a substance by its molar mass to obtain moles; multiply the given moles of a substance by its molar mass to obtain mass in grams. Number of particles is obtained by multiplying an amount in moles by Avogadro's number. Density is used to convert mass to volume.

Solution:

$$\text{a) Moles of NH}_4\text{Br} = (0.588 \text{ g NH}_4\text{Br}) \left(\frac{1 \text{ mol NH}_4\text{Br}}{97.94 \text{ g NH}_4\text{Br}} \right) = 0.0060037 = \mathbf{0.00600 \text{ mol NH}_4\text{Br}}$$

$$\text{b) Moles of KNO}_3 = (88.5 \text{ g KNO}_3) \left(\frac{1 \text{ mol KNO}_3}{101.11 \text{ g KNO}_3} \right) = 0.875284 \text{ mol KNO}_3$$

$$\begin{aligned} \text{Number of K}^+ \text{ ions} &= (0.875284 \text{ mol KNO}_3) \left(\frac{1 \text{ mol K}^+}{1 \text{ mol KNO}_3} \right) \left(\frac{6.022 \times 10^{23} \text{ K}^+ \text{ ions}}{1 \text{ mol K}^+} \right) \\ &= 5.27096 \times 10^{23} = \mathbf{5.27 \times 10^{23} \text{ K}^+ \text{ ions}} \end{aligned}$$

$$\text{c) Mass (g) of C}_3\text{H}_8\text{O}_3 = (5.85 \text{ mol C}_3\text{H}_8\text{O}_3) \left(\frac{92.09 \text{ g C}_3\text{H}_8\text{O}_3}{1 \text{ mol C}_3\text{H}_8\text{O}_3} \right) = 538.7265 = \mathbf{539 \text{ g C}_3\text{H}_8\text{O}_3}$$

$$\text{d) Mass (g) of CHCl}_3 = (2.85 \text{ mol CHCl}_3) \left(\frac{119.37 \text{ g CHCl}_3}{1 \text{ mol CHCl}_3} \right) = 340.2045 \text{ g CHCl}_3$$

$$\text{Volume (mL) of CHCl}_3 = (340.2045 \text{ g CHCl}_3) \left(\frac{\text{mL}}{1.48 \text{ g CHCl}_3} \right) = 229.868 = \mathbf{230. \text{ mL CHCl}_3}$$

$$\text{e) Moles of Na}^+ = (2.11 \text{ mol Na}_2\text{CO}_3) \left(\frac{2 \text{ mol Na}^+}{1 \text{ mol Na}_2\text{CO}_3} \right) = 4.22 \text{ mol Na}^+$$

$$\text{Number of Na}^+ = (4.22 \text{ mol Na}^+) \left(\frac{6.022 \times 10^{23} \text{ Na}^+ \text{ ions}}{1 \text{ mol Na}^+} \right) = 2.54128 \times 10^{24} = \mathbf{2.54 \times 10^{24} \text{ Na}^+ \text{ ions}}$$

$$\text{f) Moles of Cd atoms} = (25.0\mu\text{g Cd}) \left(\frac{10^{-6}\text{ g}}{1\mu\text{g}} \right) \left(\frac{1\text{ mol Cd}}{112.4\text{ g Cd}} \right) = 2.224199 \times 10^{-7}\text{ mol Cd atoms}$$

$$\begin{aligned}\text{Number of Cd atoms} &= (2.224199 \times 10^{-7}\text{ mol Cd}) \left(\frac{6.022 \times 10^{23}\text{ Cd atoms}}{1\text{ mol Cd}} \right) \\ &= 1.3394126 \times 10^{17} = \mathbf{1.34 \times 10^{17}\text{ Cd atoms}}\end{aligned}$$

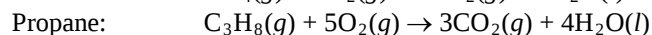
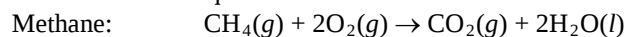
$$\begin{aligned}\text{g) Number of F atoms} &= (0.0015\text{ mol F}_2) \left(\frac{2\text{ mol F}}{1\text{ mol F}_2} \right) \left(\frac{6.022 \times 10^{23}\text{ F atoms}}{1\text{ mol F}} \right) \\ &= 1.8066 \times 10^{21} = \mathbf{1.8 \times 10^{21}\text{ F atoms}}\end{aligned}$$

3.94 Neither A nor B has any XY_3 molecules. Both C and D have XY_3 molecules. D shows both XY_3 and XY molecules. Only C has a single XY_3 product, thus the answer is C.

3.95 Plan: Deal with the methane and propane separately, and combine the results. Balanced equations are needed for each hydrocarbon. The total mass and the percentages will give the mass of each hydrocarbon. The mass of each hydrocarbon is changed to moles, and through the balanced chemical equation the amount of CO_2 produced by each gas may be found. Summing the amounts of CO_2 gives the total from the mixture. For part b), let x and $252 - x$ represent the masses of CH_4 and C_3H_8 , respectively.

Solution:

a) The balanced chemical equations are:



Mass (g) of CO_2 from each:

$$\text{Methane: } (200.\text{g Mixture}) \left(\frac{25.0\%}{100\%} \right) \left(\frac{1\text{ mol CH}_4}{16.04\text{ g CH}_4} \right) \left(\frac{1\text{ mol CO}_2}{1\text{ mol CH}_4} \right) \left(\frac{44.01\text{ g CO}_2}{1\text{ mol CO}_2} \right) = 137.188\text{ g CO}_2$$

$$\text{Propane: } (200.\text{g Mixture}) \left(\frac{75.0\%}{100\%} \right) \left(\frac{1\text{ mol C}_3\text{H}_8}{44.09\text{ g C}_3\text{H}_8} \right) \left(\frac{3\text{ mol CO}_2}{1\text{ mol C}_3\text{H}_8} \right) \left(\frac{44.01\text{ g CO}_2}{1\text{ mol CO}_2} \right) = 449.183\text{ g CO}_2$$

$$\text{Total CO}_2 = 137.188\text{ g} + 449.183\text{ g} = 586.371 = \mathbf{586\text{ g CO}_2}$$

b) Since the mass of CH_4 + the mass of C_3H_8 = 252 g, let x = mass of CH_4 in the mixture and $252 - x$ = mass of C_3H_8 in the mixture. Use mole ratios to calculate the amount of CO_2 formed from x amount of CH_4 and the amount of CO_2 formed from $252 - x$ amount of C_3H_8 .

The total mass of CO_2 produced = 748 g.

$$\text{The total moles of CO}_2 \text{ produced} = (748\text{ g CO}_2) \left(\frac{1\text{ mol CO}_2}{44.01\text{ g CO}_2} \right) = 16.996\text{ mol CO}_2$$

$$16.996\text{ mol CO}_2 =$$

$$(x\text{ g CH}_4) \left(\frac{1\text{ mol CH}_4}{16.04\text{ g CH}_4} \right) \left(\frac{1\text{ mol CO}_2}{1\text{ mol CH}_4} \right) + (252 - x\text{ g C}_3\text{H}_8) \left(\frac{1\text{ mol C}_3\text{H}_8}{44.09\text{ g C}_3\text{H}_8} \right) \left(\frac{3\text{ mol CO}_2}{1\text{ mol C}_3\text{H}_8} \right)$$

$$16.996\text{ mol CO}_2 = \frac{x}{16.04}\text{ mol CO}_2 + \frac{3(252 - x)}{44.09}\text{ mol CO}_2$$

$$16.996\text{ mol CO}_2 = \frac{x}{16.04}\text{ mol CO}_2 + \frac{756 - 3x}{44.09}\text{ mol CO}_2$$

$$16.996\text{ mol CO}_2 = 0.06234x\text{ mol CO}_2 + (17.147 - 0.06804x\text{ mol CO}_2)$$

$$16.996 = 17.147 - 0.0057x$$

$$x = 26.49\text{ g CH}_4 \quad 252 - x = 252\text{ g} - 26.49\text{ g} = 225.51\text{ g C}_3\text{H}_8$$

$$\text{Mass \% CH}_4 = \frac{\text{mass of CH}_4}{\text{mass of mixture}}(100) = \frac{26.49\text{ g CH}_4}{252\text{ g mixture}}(100) = \mathbf{10.5\% \text{ CH}_4}$$

$$\text{Mass \% C}_3\text{H}_8 = \frac{\text{mass of C}_3\text{H}_8}{\text{mass of mixture}}(100) = \frac{225.51 \text{ g C}_3\text{H}_8}{252 \text{ g mixture}}(100) = \mathbf{89.5\% \text{ C}_3\text{H}_8}$$

- 3.96 Plan: If we assume a 100-gram sample of fertilizer, then the 30:10:10 percentages become the masses, in grams, of N, P₂O₅, and K₂O. These masses may be changed to moles of substance, and then to moles of each element. To get the desired x:y:1.0 ratio, divide the moles of each element by the moles of potassium.

Solution:

A 100-gram sample of 30:10:10 fertilizer contains 30 g N, 10 g P₂O₅, and 10 g K₂O.

$$\text{Moles of N} = (30 \text{ g N}) \left(\frac{1 \text{ mol N}}{14.01 \text{ g N}} \right) = 2.1413 \text{ mol N}$$

$$\text{Moles of P} = (10 \text{ g P}_2\text{O}_5) \left(\frac{1 \text{ mol P}_2\text{O}_5}{141.94 \text{ g P}_2\text{O}_5} \right) \left(\frac{2 \text{ mol P}}{1 \text{ mol P}_2\text{O}_5} \right) = 0.14090 \text{ mol P}$$

$$\text{Moles of K} = (10 \text{ g K}_2\text{O}) \left(\frac{1 \text{ mol K}_2\text{O}}{94.20 \text{ g K}_2\text{O}} \right) \left(\frac{2 \text{ mol K}}{1 \text{ mol K}_2\text{O}} \right) = 0.21231 \text{ mol K}$$

This gives a N:P:K ratio of 2.1413:0.14090:0.21231

The ratio must be divided by the moles of K and rounded.

$$\frac{2.1413 \text{ mol N}}{0.21231} = 10.086 \quad \frac{0.14090 \text{ mol P}}{0.21231} = 0.66365 \quad \frac{0.21231 \text{ mol K}}{0.21231} = 1$$

10.086:0.66365:1.000 or **10:0.66:1.0**

- 3.97 Plan: If we assume a 100-gram sample of fertilizer, then the 10:10:10 percentages become the masses, in grams, of N, P₂O₅, and K₂O. These masses may be changed to moles of substance, and then to moles of each element. Use the mole ratio between N and ammonium sulfate, P and ammonium hydrogen phosphate, and K and potassium chloride to find the mass of each compound required to provide the needed amount of the respective element. Divide the mass of each compound by the total mass of sample, 100 g, and multiply by 100 for mass %.

Solution:

Assume a 100 g sample. 10:10:10 indicates 10 g N, 10 g P₂O₅ and 10 g K₂O.

$$\text{Moles of N} = (10 \text{ g N}) \left(\frac{1 \text{ mol N}}{14.01 \text{ g N}} \right) = 0.713776 \text{ mol N}$$

$$\text{Moles of P} = (10 \text{ g P}_2\text{O}_5) \left(\frac{1 \text{ mol P}_2\text{O}_5}{141.94 \text{ g P}_2\text{O}_5} \right) \left(\frac{2 \text{ mol P}}{1 \text{ mol P}_2\text{O}_5} \right) = 0.14090 \text{ mol P}$$

$$\text{Moles of K} = (10 \text{ g K}_2\text{O}) \left(\frac{1 \text{ mol K}_2\text{O}}{94.20 \text{ g K}_2\text{O}} \right) \left(\frac{2 \text{ mol K}}{1 \text{ mol K}_2\text{O}} \right) = 0.21231 \text{ mol K}$$

To obtain 0.713776 mol N from (NH₄)₂SO₄:

$$(0.713776 \text{ mol N}) \left(\frac{1 \text{ mol (NH}_4)_2\text{SO}_4}{2 \text{ mol N}} \right) \left(\frac{132.15 \text{ g (NH}_4)_2\text{SO}_4}{1 \text{ mol (NH}_4)_2\text{SO}_4} \right) = 47.1627 \text{ g (NH}_4)_2\text{SO}_4$$

$$\text{Mass \% (NH}_4)_2\text{SO}_4 = \frac{\text{mass of (NH}_4)_2\text{SO}_4}{\text{mass of mixture}}(100) = \frac{47.1627 \text{ g (NH}_4)_2\text{SO}_4}{100 \text{ g mixture}}(100)$$

= 47.1627% = **47.2% (NH₄)₂SO₄**

To obtain 0.14090 mol P from (NH₄)₂HPO₄:

$$(0.14090 \text{ mol P}) \left(\frac{1 \text{ mol (NH}_4)_2\text{HPO}_4}{1 \text{ mol P}} \right) \left(\frac{132.06 \text{ g (NH}_4)_2\text{HPO}_4}{1 \text{ mol (NH}_4)_2\text{HPO}_4} \right) = 18.6073 \text{ g (NH}_4)_2\text{HPO}_4$$

$$\text{Mass \% (NH}_4)_2\text{HPO}_4 = \frac{\text{mass of (NH}_4)_2\text{HPO}_4}{\text{mass of mixture}}(100) = \frac{18.6073 \text{ g (NH}_4)_2\text{HPO}_4}{100 \text{ g mixture}}(100)$$

= 18.6073% = **18.6% (NH₄)₂HPO₄**

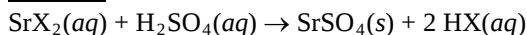
To obtain 0.21231 mol K from KCl:

$$(0.21231 \text{ mol K}) \left(\frac{1 \text{ mol KCl}}{1 \text{ mol K}} \right) \left(\frac{74.55 \text{ g KCl}}{1 \text{ mol KCl}} \right) = 15.8277 \text{ g KCl}$$

$$\text{Mass \% KCl} = \frac{\text{mass of KCl}}{\text{mass of mixture}} (100) = \frac{15.8277 \text{ g KCl}}{100 \text{ g mixture}} (100) = 15.8277\% = \mathbf{15.8\% \text{ KCl}}$$

- 3.98 Plan: Write a balanced equation. Convert the mass of strontium sulfate produced to moles and use the mole ratio in the balanced equation to find the moles of strontium halide required to produce that amount of product. Divide the given mass of strontium halide by the moles of strontium halide to obtain its molar mass. Subtracting the molar mass of strontium from the molar mass of compound gives the molar mass of the halogen in the formula. The molar mass of the halogen is used to identify the halogen.

Solution:



$$0.652 \text{ g} \qquad \qquad \qquad 0.755 \text{ g}$$

$$\text{Moles SrX}_2 = (0.755 \text{ g SrSO}_4) \left(\frac{1 \text{ mol SrSO}_4}{183.69 \text{ g SrSO}_4} \right) \left(\frac{1 \text{ mol SrX}_2}{1 \text{ mol SrSO}_4} \right) = 0.004110186 \text{ mol SrX}_2$$

The 0.652 g sample of $\text{SrX}_2 = 0.004110186 \text{ mol}$

$$\text{SrX}_2 = \frac{0.652 \text{ g}}{0.004110186 \text{ mol}} = 158.630 \text{ g/mol} = \text{molar mass}$$

Molar mass of $\text{X}_2 = 158.630 \text{ g} - \text{molar mass of Sr}$

Molar mass of $\text{X}_2 = 158.630 \text{ g} - 87.62 \text{ g} = 71.01 \text{ g} = \text{X}_2$

Molar mass of $\text{X} = 71.01 \text{ g}/2 = 35.505 = 35.5 \text{ g/mol} = \text{Cl}$ The original halide formula is **SrCl_2** .

- 3.99 Plan: Assume 100 grams of mixture. This means the mass of each compound, in grams, is the same as its percentage. Find the mass of C from CO and from CO_2 and add these masses together. For mass %, divide the total mass of C by the mass of the mixture (100 g) and multiply by 100.

Solution:

100 g of mixture = 35 g CO and 65 g CO_2 .

$$\text{Mass (g) of C from CO} = (35.0 \text{ g CO}) \left(\frac{1 \text{ mol CO}}{28.01 \text{ g CO}} \right) \left(\frac{1 \text{ mol C}}{1 \text{ mol CO}} \right) \left(\frac{12.01 \text{ g C}}{1 \text{ mol C}} \right) = 15.007 \text{ g C}$$

$$\text{Mass (g) of C from CO}_2 = (65.0 \text{ g CO}_2) \left(\frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \right) \left(\frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \right) \left(\frac{12.01 \text{ g C}}{1 \text{ mol C}} \right) = 17.738 \text{ g C}$$

Total mass (g) of C = $15.007 \text{ g} + 17.738 \text{ g} = 32.745 \text{ g C}$

$$\text{Mass \% C} = \frac{\text{mass of C}}{\text{mass of mixture}} (100) = \frac{32.745 \text{ g C}}{100 \text{ g mixture}} (100) = 32.745 = \mathbf{32.7\% \text{ C}}$$

- 3.100 Plan: Write a balanced equation for the reaction. Count the molecules of each reactant to obtain the moles of each reactant present. Use the mole ratios in the equation to calculate the amount of product formed. Only 87.0% of the calculated amount of product actually forms, so the actual yield is 87.0% of the theoretical yield.

Solution:

The balanced equation is $\text{SiH}_4 + \text{N}_2\text{F}_4 \rightarrow \text{SiF}_4 + \text{N}_2 + 2\text{H}_2$.

$$\text{Moles of SiH}_4 = (3 \text{ SiH}_4 \text{ molecules}) \left(\frac{1.25 \times 10^{-2} \text{ mol}}{1 \text{ molecule}} \right) = 0.0375 \text{ mol SiH}_4$$

$$\text{Moles of N}_2\text{F}_4 = (3 \text{ N}_2\text{F}_4 \text{ molecules}) \left(\frac{1.25 \times 10^{-2} \text{ mol}}{1 \text{ molecule}} \right) = 0.0375 \text{ mol N}_2\text{F}_4$$

Since there is an equal amount of each reactant and the ratio between each reactant and SiF_4 is 1:1, neither reactant is in excess and either may be used to calculate the amount of SiF_4 produced.

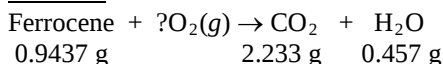
$$\text{Mass (g) of SiF}_4 = (0.0375 \text{ mol SiH}_4) \left(\frac{1 \text{ mol SiF}_4}{1 \text{ mol SiH}_4} \right) \left(\frac{104.09 \text{ g SiF}_4}{1 \text{ mol SiF}_4} \right) = 3.903375 \text{ g SiF}_4$$

$$\% \text{ yield} = \left(\frac{\text{actual Yield}}{\text{theoretical Yield}} \right) \times 100\%$$

$$\text{Actual yield (g) of SiF}_4 = \frac{\% \text{ yield}}{100\%} (\text{theoretical yield}) = \frac{87\%}{100\%} (3.903375 \text{ g SiF}_4) = 3.3959 = \mathbf{3.4 \text{ g SiF}_4}$$

- 3.101 Plan: In combustion analysis, finding the moles of carbon and hydrogen is relatively simple because all of the carbon present in the sample is found in the carbon of CO₂, and all of the hydrogen present in the sample is found in the hydrogen of H₂O. Convert the mass of CO₂ to moles and use the ratio between CO₂ and C to find the moles and mass of C present. Do the same to find the moles and mass of H from H₂O. Subtracting the masses of C and H from the mass of the sample gives the mass of Fe. Convert the mass of Fe to moles of Fe. Take the moles of C, H, and Fe and divide by the smallest value to convert to whole numbers to get the empirical formula.

Solution:



$$\text{Moles of C} = (2.233 \text{ g CO}_2) \left(\frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \right) \left(\frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \right) = 0.050738 \text{ mol C}$$

$$\text{Mass (g) of C} = (0.050738 \text{ mol C}) \left(\frac{12.01 \text{ g C}}{1 \text{ mol C}} \right) = 0.60936 \text{ g C}$$

$$\text{Moles of H} = (0.457 \text{ g H}_2\text{O}) \left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) \left(\frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \right) = 0.050721 \text{ mol H}$$

$$\text{Mass (g) of H} = (0.050721 \text{ mol H}) \left(\frac{1.008 \text{ g H}}{1 \text{ mol H}} \right) = 0.051127 \text{ g H}$$

$$\begin{aligned} \text{Mass (g) of Fe} &= \text{Sample mass} - (\text{mass of C} + \text{mass of H}) \\ &= 0.9437 \text{ g} - (0.60936 \text{ g C} + 0.052217 \text{ g H}) = 0.283213 \text{ g Fe} \end{aligned}$$

$$\text{Moles of Fe} = (0.283213 \text{ g Fe}) \left(\frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} \right) = 0.005071 \text{ mol Fe}$$

$$\text{Preliminary formula} = \text{C}_{0.050738} \text{H}_{0.050721} \text{Fe}_{0.005071}$$

Converting to integer subscripts (dividing all by the smallest subscript):

$$\frac{\text{C}_{0.050738}}{0.005071} \frac{\text{H}_{0.050721}}{0.005071} \frac{\text{Fe}_{0.005071}}{0.005071} \rightarrow \text{C}_{10} \text{H}_{10} \text{Fe}_1$$

Empirical formula = C₁₀H₁₀Fe

- 3.102 Plan: Determine the molecular formula from the figure. Once the molecular formula is known, use the periodic table to determine the molar mass. Convert the volume of lemon juice in part b) from qt to mL and use the density to convert from mL to mass in g. Take 6.82% of that mass to find the mass of citric acid and use the molar mass to convert to moles.

Solution:

a) The formula of citric acid obtained by counting the number of carbon atoms, oxygen atoms, and hydrogen atoms is **C₆H₈O₇**.

$$\text{Molar mass} = (6 \times 12.01 \text{ g/mol C}) + (8 \times 1.008 \text{ g/mol H}) + (7 \times 16.00 \text{ g/mol O}) = \mathbf{192.12 \text{ g/mol}}$$

b) Converting volume of lemon juice in qt to mL:

$$\text{Volume (mL) of lemon juice} = (1.50 \text{ qt}) \left(\frac{1 \text{ L}}{1.057 \text{ qt}} \right) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right) = 1419.111 \text{ mL}$$

Converting volume to mass in grams:

$$\text{Mass (g) of lemon juice} = (1419.111 \text{ mL}) \left(\frac{1.09 \text{ g}}{\text{mL}} \right) = 1546.831 \text{ g lemon juice}$$

$$\text{Mass (g) of C}_6\text{H}_8\text{O}_7 = (1546.831 \text{ g lemon juice}) \left(\frac{6.82\% \text{ C}_6\text{H}_8\text{O}_7}{100\% \text{ lemon juice}} \right) = 105.494 \text{ g C}_6\text{H}_8\text{O}_7$$

$$\text{Moles of C}_6\text{H}_8\text{O}_7 = (105.494 \text{ g C}_6\text{H}_8\text{O}_7) \left(\frac{1 \text{ mol C}_6\text{H}_8\text{O}_7}{192.12 \text{ g C}_6\text{H}_8\text{O}_7} \right) = 0.549104 = \mathbf{0.549 \text{ mol C}_6\text{H}_8\text{O}_7}$$

- 3.103 Plan: For parts a) and b), convert the masses to moles. Take the moles and divide by the smallest value to convert to whole numbers to get the empirical formula. For part c), write the two balanced equations and use two equations as shown.

Solution:

$$\text{a) Moles of Pt} = (0.327 \text{ g Pt}) \left(\frac{1 \text{ mol Pt}}{195.1 \text{ g Pt}} \right) = 0.001676 \text{ mol Pt}$$

$$\text{Mass (g) of F} = \text{mass of product} - \text{mass of Pt} = 0.519 \text{ g} - 0.327 \text{ g} = 0.192 \text{ g F}$$

$$\text{Moles of F} = (0.192 \text{ g F}) \left(\frac{1 \text{ mol F}}{19.00 \text{ g F}} \right) = 0.010105 \text{ mol F}$$

$$\text{Preliminary formula} = \text{Pt}_{0.001676}\text{F}_{0.010105}$$

Converting to integer subscripts (dividing all by the smallest subscript):

$$\frac{\text{Pt}_{0.001676}\text{F}_{0.010105}}{0.001676 \quad 0.001676} \rightarrow \text{Pt}_1\text{F}_6$$

$$\text{Empirical formula} = \mathbf{PtF_6}$$

$$\text{b) Moles of PtF}_6 = (0.265 \text{ g PtF}_6) \left(\frac{1 \text{ mol PtF}_6}{309.1 \text{ g PtF}_6} \right) = 0.0008576 \text{ mol PtF}_6$$

$$\text{Mass of Xe} = \text{mass of product} - \text{mass of PtF}_6 = 0.378 \text{ g} - 0.265 \text{ g} = 0.113 \text{ g Xe}$$

$$\text{Moles of Xe} = (0.113 \text{ g Xe}) \left(\frac{1 \text{ mol Xe}}{131.3 \text{ g Xe}} \right) = 0.0008606 \text{ mol Xe}$$

$$\text{Preliminary formula} = \text{Xe}_{0.0008606}(\text{PtF}_6)_{0.0008576}$$

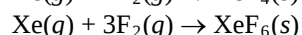
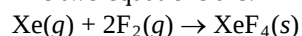
Converting to integer subscripts (dividing all by the smallest subscript):

$$\frac{\text{Xe}_{0.0008606}(\text{PtF}_6)_{0.0008576}}{0.0008576 \quad 0.0008576} \rightarrow \text{Xe}_1(\text{PtF}_6)_1$$

$$\text{Empirical formula} = \mathbf{XePtF_6}$$

- c) This problem can be solved as a system of two equations and two unknowns.

The two equations are:



$$\text{Moles of Xe consumed} = 1.85 \times 10^{-4} \text{ mol present} - 9.00 \times 10^{-6} \text{ mol excess} = 1.76 \times 10^{-4} \text{ mol Xe}$$

$$\text{Then } x + y = 1.76 \times 10^{-4} \text{ mol Xe consumed}$$

$$2x + 3y = 5.00 \times 10^{-4} \text{ mol F}_2 \text{ consumed}$$

Solve for x using the first equation and substitute the value of x into the second equation:

$$x = 1.76 \times 10^{-4} - y$$

$$2(1.76 \times 10^{-4} - y) + 3y = 5.00 \times 10^{-4}$$

$$(3.52 \times 10^{-4}) - 2y + 3y = 5.00 \times 10^{-4}$$

$$y = (5.00 \times 10^{-4}) - (3.52 \times 10^{-4}) = 1.48 \times 10^{-4} \text{ mol XeF}_6$$

$$x = (1.76 \times 10^{-4}) - (1.48 \times 10^{-4}) = 2.8 \times 10^{-5} \text{ mol XeF}_4$$

Converting moles of each product to grams using the molar masses:

$$\text{Mass (g) of XeF}_4 = (2.8 \times 10^{-5} \text{ mol XeF}_4) \left(\frac{207.3 \text{ g XeF}_4}{1 \text{ mol XeF}_4} \right) = 5.8044 \times 10^{-3} \text{ g XeF}_4$$

$$\text{Mass (g) of XeF}_6 = (1.48 \times 10^{-4} \text{ mol XeF}_6) \left(\frac{245.3 \text{ g XeF}_6}{1 \text{ mol XeF}_6} \right) = 3.63044 \times 10^{-2} \text{ g XeF}_6$$

Calculate the percent of each compound using the total weight of the products:

$$(5.8044 \times 10^{-3} + 3.63044 \times 10^{-2}) \text{ g} = 0.0421088 \text{ g}$$

$$\text{Mass \% XeF}_4 = \frac{\text{mass of XeF}_4}{\text{total mass}}(100) = \frac{5.8044 \times 10^{-3} \text{ g XeF}_4}{0.0421088 \text{ g}}(100) = 13.784 = \mathbf{14\% \text{ XeF}_4}$$

$$\text{Mass \% XeF}_6 = \frac{\text{mass of XeF}_6}{\text{total mass}}(100) = \frac{3.63044 \times 10^{-2} \text{ g XeF}_6}{0.0421088 \text{ g}}(100) = 86.2157 = \mathbf{86.2\% \text{ XeF}_6}$$

- 3.104 Plan: Use the mass percent to find the mass of heme in the sample; use the molar mass to convert the mass of heme to moles. Then find the mass of Fe in the sample by using the mole ratio between heme and iron. The mass of hemin is found by using the mole ratio between heme and hemoglobin.

Solution:

$$\text{a) Mass (g) of heme} = (0.65 \text{ g hemoglobin}) \left(\frac{6.0\% \text{ heme}}{100\% \text{ hemoglobin}} \right) = \mathbf{0.039 \text{ g heme}}$$

$$\text{b) Moles of heme} = (0.039 \text{ g heme}) \left(\frac{1 \text{ mol heme}}{616.49 \text{ g heme}} \right) = 6.32614 \times 10^{-5} = \mathbf{6.3 \times 10^{-5} \text{ mol heme}}$$

$$\begin{aligned} \text{c) Mass (g) of Fe} &= (6.32614 \times 10^{-5} \text{ mol heme}) \left(\frac{1 \text{ mol Fe}}{1 \text{ mol heme}} \right) \left(\frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} \right) \\ &= 3.5331 \times 10^{-3} = \mathbf{3.5 \times 10^{-3} \text{ g Fe}} \end{aligned}$$

$$\begin{aligned} \text{d) Mass (g) of hemin} &= (6.32614 \times 10^{-5} \text{ mol heme}) \left(\frac{1 \text{ mol hemin}}{1 \text{ mol heme}} \right) \left(\frac{651.94 \text{ g hemin}}{1 \text{ mol hemin}} \right) \\ &= 4.1243 \times 10^{-2} = \mathbf{4.1 \times 10^{-2} \text{ g hemin}} \end{aligned}$$

- 3.105 Plan: Find the Mn:O ratio in the two oxides. Write two equations to solve simultaneously; one equation shows that the sum of the ratio of Mn in the two oxides will equal the ratio of Mn in the sample and the other equation shows that the total amount of oxide in the sample is the sum of the amounts of the two oxides. The two equations will give the mole ratio of the two oxides. Convert moles of each oxide to mass to obtain the mass ratio of the two oxides from which the mass % of each can be calculated. Use that mass % of each to find the mass of each in the sample. For part b), the moles of Mn^{3+} come from the Mn_2O_3 and the moles of Mn^{2+} come from the MnO .

Solution:

Mn:O ratio:

In sample:	1.00:1.42	or	0.704
In braunite:	2.00:3.00	or	0.667
In manganosite:	1.00:1.00	or	1.00

- a) The total amount of ore is equal to the amount of braunite (B) + the amount of manganosite (M).

$$B + M = 1.00$$

$$M = 1.00 - B$$

The amount of Mn is dependent on the sample's composition.

$$M(1.00) + B(0.667) = 0.704$$

$$(1.00 - B)(1.00) + B(0.667) = 0.704$$

$$1.00 - 1.00B + 0.667B = 0.704$$

$$0.296 = 0.333B$$

$$B = 0.888889 \text{ mol braunite}$$

$$M = 1.00 - B = 1.00 - 0.888889 = 0.111111 \text{ mol manganosite}$$

$$\text{Mass (g) of braunite} = (0.888889 \text{ mol}) \left(\frac{157.88 \text{ g}}{1 \text{ mol}} \right) = 140.338 \text{ g braunite}$$

$$\text{Mass (g) of manganosite} = (0.111111 \text{ mol}) \left(\frac{70.94 \text{ g}}{1 \text{ mol}} \right) = 7.88221 \text{ g manganosite}$$

There are 140.338 g of braunite for every 7.88221 g of manganosite. Finding mass % of each:

$$\text{Mass \% braunite} = \frac{\text{mass of braunite}}{\text{mass of braunite} + \text{manganosite}}(100) = \frac{140.338 \text{ g}}{140.338 + 7.88221 \text{ g}}(100) = 94.6821\%$$

$$\begin{aligned}\text{Mass \% manganosite} &= \frac{\text{mass of manganosite}}{\text{mass of braunite} + \text{manganosite}}(100) = \frac{7.88221 \text{ g}}{140.338 + 7.88221 \text{ g}}(100) \\ &= 5.3179\%\end{aligned}$$

In the 542.3 g sample:

$$\text{Mass (g) of braunite} = (542.3 \text{ g sample}) \left(\frac{94.6821 \text{ braunite}}{100\% \text{ sample}} \right) = 513.461 = \mathbf{513 \text{ g braunite}}$$

$$\text{Mass (g) of manganosite} = (542.3 \text{ g sample}) \left(\frac{5.3179\% \text{ manganosite}}{100\% \text{ sample}} \right) = 28.839 = \mathbf{28.8 \text{ g manganosite}}$$

b) Each mole of braunite, Mn_2O_3 , contains 2 moles of Mn^{3+} while each mole of manganosite, MnO , contains 1 mole of Mn^{2+} .

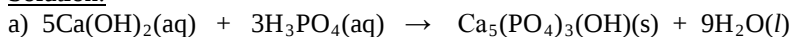
$$\text{Moles of } \text{Mn}^{3+} = 2(0.888889 \text{ mol braunite}) = 1.777778 \text{ mol } \text{Mn}^{3+}$$

$$\text{Moles of } \text{Mn}^{2+} = 1(0.111111 \text{ mol manganosite}) = 0.111111 \text{ mol } \text{Mn}^{2+}$$

$$\text{Mn}^{3+}:\text{Mn}^{2+} = \frac{1.777778 \text{ mol } \text{Mn}^{3+}}{0.111111 \text{ mol } \text{Mn}^{2+}} = 16.000 = \mathbf{16.0}$$

- 3.106 Plan: First, balance the chemical equation. To determine which reactant is limiting, calculate the amount of hydroxyapatite formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the amount of hydroxyapatite formed.

Solution:



b) Find the limiting reagent.

$$\begin{aligned}\text{Moles of } \text{Ca}_5(\text{PO}_4)_3(\text{OH}) \text{ from } \text{Ca(OH)}_2 &= (100. \text{ g } \text{Ca(OH)}_2) \left(\frac{1 \text{ mol } \text{Ca(OH)}_2}{74.10 \text{ g } \text{Ca(OH)}_2} \right) \left(\frac{1 \text{ mol } \text{Ca}_5(\text{PO}_4)_3(\text{OH})}{5 \text{ mol } \text{Ca(OH)}_2} \right) \\ &= 0.2699055 \text{ mol } \text{Ca}_5(\text{PO}_4)_3(\text{OH})\end{aligned}$$

$$\text{Moles of } \text{Ca}_5(\text{PO}_4)_3(\text{OH}) \text{ from } \text{H}_3\text{PO}_4 =$$

$$\begin{aligned}(100. \text{ g } \text{H}_3\text{PO}_4 \text{ solution}) &\left(\frac{85 \text{ g } \text{H}_3\text{PO}_4}{100. \text{ g } \text{H}_3\text{PO}_4 \text{ solution}} \right) \left(\frac{1 \text{ mol } \text{H}_3\text{PO}_4}{97.99 \text{ g } \text{H}_3\text{PO}_4} \right) \left(\frac{1 \text{ mol } \text{Ca}_5(\text{PO}_4)_3(\text{OH})}{3 \text{ mol } \text{H}_3\text{PO}_4} \right) \\ &= 0.2891452 \text{ mol } \text{Ca}_5(\text{PO}_4)_3(\text{OH})\end{aligned}$$

Ca(OH)_2 is the limiting reactant, and will be used to calculate the yield.

$$\begin{aligned}(100. \text{ g } \text{Ca(OH)}_2) &\left(\frac{1 \text{ mol } \text{Ca(OH)}_2}{74.10 \text{ g } \text{Ca(OH)}_2} \right) \left(\frac{1 \text{ mol } \text{Ca}_5(\text{PO}_4)_3(\text{OH})}{5 \text{ mol } \text{Ca(OH)}_2} \right) \left(\frac{502.32 \text{ g } \text{Ca}_5(\text{PO}_4)_3(\text{OH})}{1 \text{ mol } \text{Ca}_5(\text{PO}_4)_3(\text{OH})} \right) \\ &= 135.57893 = \mathbf{140 \text{ g } \text{Ca}_5(\text{PO}_4)_3(\text{OH})}\end{aligned}$$

- 3.107 Plan: To determine which reactant is limiting, calculate the amount of aspirin formed from each reactant, assuming an excess of the other reactant. Use the density of acetic anhydride to determine the amount of this reactant in grams. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the theoretical yield of aspirin. The actual yield divided by the theoretical yield just calculated (with the result multiplied by 100%) gives the percent yield. Use the formula for percent atom economy to determine that quantity.

Solution:

a) Finding the moles of aspirin from the moles of $\text{C}_7\text{H}_6\text{O}_3$ (if $(\text{CH}_3\text{CO})_2\text{O}$ is limiting):

$$\begin{aligned}\text{Moles of aspirin from } \text{C}_7\text{H}_6\text{O}_3 &= (3.077 \text{ g } \text{C}_7\text{H}_6\text{O}_3) \left(\frac{1 \text{ mol } \text{C}_7\text{H}_6\text{O}_3}{138.12 \text{ g } \text{C}_7\text{H}_6\text{O}_3} \right) \left(\frac{1 \text{ mol } \text{C}_9\text{H}_8\text{O}_4}{1 \text{ mol } \text{C}_7\text{H}_6\text{O}_3} \right) \\ &= 0.0222777 \text{ mol } \text{C}_9\text{H}_8\text{O}_4\end{aligned}$$

Finding the moles of aspirin from the moles of $C_4H_6O_3$ (if $C_7H_6O_3$ is limiting):

$$\text{Mass (g) of } (CH_3CO)_2O = (5.50 \text{ mL } (CH_3CO)_2O) \left(\frac{1.080 \text{ g}}{1 \text{ mL}} \right) = 5.94 \text{ g } (CH_3CO)_2O$$

$$\begin{aligned} \text{Moles of aspirin from } (CH_3CO)_2O &= (5.94 \text{ g } (CH_3CO)_2O) \left(\frac{1 \text{ mol } (CH_3CO)_2O}{102.09 \text{ g } (CH_3CO)_2O} \right) \left(\frac{1 \text{ mol } C_9H_8O_4}{1 \text{ mol } (CH_3CO)_2O} \right) \\ &= 0.058183955 \text{ mol } C_9H_8O_4 \end{aligned}$$

The limiting reactant is $C_7H_6O_3$.

b) First, calculate the theoretical yield from the limiting reagent:

$$\begin{aligned} \text{Mass (g) of } C_9H_8O_4 &= (3.077 \text{ g } C_7H_6O_3) \left(\frac{1 \text{ mol } C_7H_6O_3}{138.12 \text{ g } C_7H_6O_3} \right) \left(\frac{1 \text{ mol } C_9H_8O_4}{1 \text{ mol } C_7H_6O_3} \right) \left(\frac{180.15 \text{ g } C_9H_8O_4}{1 \text{ mol } C_9H_8O_4} \right) \\ &= 4.01333 \text{ g } C_9H_8O_4 \end{aligned}$$

$$\text{Percent yield} = \left(\frac{\text{actual yield}}{\text{theoretical yield}} \right) \times 100\% = \left(\frac{3.281 \text{ g}}{4.01333 \text{ g}} \right) \times 100\% = 81.7526 = \mathbf{81.75\% \text{ yield}}$$

- 3.108 Plan: Determine the formula and the molar mass of each compound. The formula gives the relative number of moles of nitrogen present. Multiply the number of moles of nitrogen by its molar mass to find the total mass of nitrogen in 1 mole of compound. Mass percent = $\frac{\text{total mass of element}}{\text{molar mass of compound}}(100)$. For part b), convert mass of ornithine to moles, use the mole ratio between ornithine and urea to find the moles of urea, and then use the ratio between moles of urea and nitrogen to find the moles and mass of nitrogen produced.

Solution:

a) Urea: CH_4N_2O , $\mathcal{M} = 60.06 \text{ g/mol}$

There are 2 moles of N in 1 mole of CH_4N_2O .

$$\text{Mass (g) of N} = (2 \text{ mol N}) \left(\frac{14.01 \text{ g N}}{1 \text{ mol N}} \right) = 28.02 \text{ g N}$$

$$\text{Mass percent} = \frac{\text{total mass N}}{\text{molar mass of compound}}(100) = \frac{28.02 \text{ g N}}{60.06 \text{ g } CH_4N_2O}(100) = 46.6533 = \mathbf{46.65\% \text{ N in urea}}$$

Arginine: $C_6H_{15}N_4O_2$, $\mathcal{M} = 175.22 \text{ g/mol}$

There are 4 moles of N in 1 mole of $C_6H_{15}N_4O_2$.

$$\text{Mass (g) of N} = (4 \text{ mol N}) \left(\frac{14.01 \text{ g N}}{1 \text{ mol N}} \right) = 56.04 \text{ g N}$$

$$\begin{aligned} \text{Mass percent} &= \frac{\text{total mass N}}{\text{molar mass of compound}}(100) = \frac{56.04 \text{ g N}}{175.22 \text{ g } C_6H_{15}N_4O_2}(100) \\ &= 31.98265 = \mathbf{31.98\% \text{ N in arginine}} \end{aligned}$$

Ornithine: $C_5H_{13}N_2O_2$, $\mathcal{M} = 133.17 \text{ g/mol}$

There are 2 moles of N in 1 mole of $C_5H_{13}N_2O_2$.

$$\text{Mass (g) of N} = (2 \text{ mol N}) \left(\frac{14.01 \text{ g N}}{1 \text{ mol N}} \right) = 28.02 \text{ g N}$$

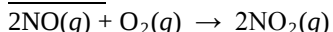
$$\begin{aligned} \text{Mass percent} &= \frac{\text{total mass N}}{\text{molar mass of compound}}(100) = \frac{28.02 \text{ g N}}{133.17 \text{ g } C_5H_{13}N_2O_2}(100) \\ &= 21.04077 = \mathbf{21.04\% \text{ N in ornithine}} \end{aligned}$$

$$\text{b) Moles of urea} = (135.2 \text{ g } C_5H_{13}N_2O_2) \left(\frac{1 \text{ mol } C_5H_{13}N_2O_2}{133.17 \text{ g } C_5H_{13}N_2O_2} \right) \left(\frac{1 \text{ mol } CH_4N_2O}{1 \text{ mol } C_5H_{13}N_2O_2} \right) = 1.015244 \text{ mol urea}$$

$$\text{Mass (g) of nitrogen} = (1.015244 \text{ mol } CH_4N_2O) \left(\frac{2 \text{ mol N}}{1 \text{ mol } CH_4N_2O} \right) \left(\frac{14.01 \text{ g N}}{1 \text{ mol N}} \right) = 28.447 = \mathbf{28.45 \text{ g N}}$$

- 3.109 Plan: Write and balance the chemical reaction. Use the mole ratio to find the amount of product that should be produced and take 66% of that amount to obtain the actual yield.

Solution:



With 6 molecules of NO and 3 molecules of O₂ reacting, 6 molecules of NO₂ can be produced.

If the reaction only has a 66% yield, then (0.66)(6) = 4 molecules of NO₂ will be produced. **Circle A** shows the formation of 4 molecules of NO₂. Circle B also shows the formation of 4 molecules of NO₂ but also has 2 unreacted molecules of NO and 1 unreacted molecule of O₂. Since neither reactant is limiting, there will be no unreacted reactant remaining after the reaction is over.

- 3.110 Plan: First balance the given chemical equation. To determine which reactant is limiting, calculate the amount of ZnS formed from each reactant, assuming an excess of the other reactant. The reactant that produces less product is the limiting reagent. Use the limiting reagent and the mole ratio from the balanced chemical equation to determine the theoretical yield of ZnS. The actual yield divided by the theoretical yield just calculated (with the result multiplied by 100%) gives the percent yield. For part b), determine the mass of Zn that does not produce ZnS; use that amount of zinc and the mole ratio between Zn and ZnO in that reaction to determine the mass of ZnO produced. Find the moles of S₈ in the reactant and the moles of S₈ in the product ZnS. The difference between these two amounts is the moles of S₈ in SO₂.

Solution:

a) The balanced equation is $8\text{Zn}(s) + \text{S}_8(s) \rightarrow 8\text{ZnS}(s)$.

Finding the limiting reagent:

Finding the moles of ZnS from the moles of Zn (if S₈ is limiting):

$$\text{Moles of ZnS from Zn} = (83.2 \text{ g Zn}) \left(\frac{1 \text{ mol Zn}}{65.41 \text{ g Zn}} \right) \left(\frac{8 \text{ mol ZnS}}{8 \text{ mol Zn}} \right) = 1.27198 \text{ mol ZnS}$$

Finding the moles of ZnS from the moles of S₈ (if Zn is limiting):

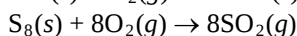
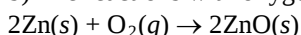
$$\text{Moles of ZnS from S}_8 = (52.4 \text{ g S}_8) \left(\frac{1 \text{ mol S}_8}{256.56 \text{ g S}_8} \right) \left(\frac{8 \text{ mol ZnS}}{1 \text{ mol S}_8} \right) = 1.6339 \text{ mol ZnS}$$

The zinc will produce less zinc sulfide, thus, zinc is the limiting reactant and will first be used to determine the theoretical yield and then the percent yield.

$$\begin{aligned} \text{Theoretical yield (g) of ZnS} &= (83.2 \text{ g Zn}) \left(\frac{1 \text{ mol Zn}}{65.41 \text{ g Zn}} \right) \left(\frac{8 \text{ mol ZnS}}{8 \text{ mol Zn}} \right) \left(\frac{97.48 \text{ g ZnS}}{1 \text{ mol ZnS}} \right) \\ &= 123.9923 \text{ g ZnS (unrounded)} \end{aligned}$$

$$\text{Percent yield} = \left(\frac{\text{actual Yield}}{\text{theoretical Yield}} \right) \times 100\% = \left(\frac{104.4 \text{ g}}{123.9923 \text{ g}} \right) \times 100\% = 84.1988 = \mathbf{84.2\% \text{ yield}}$$

b) The reactions with oxygen are:



The theoretical yield indicates that 84.2% of the zinc produced zinc sulfide so (100 – 84.2)% = 15.8% of the zinc became zinc oxide. This allows the calculation of the amount of zinc oxide formed.

$$\text{Mass (g) of Zn that does not produce ZnS} = (83.2 \text{ g Zn}) \left(\frac{15.8\%}{100\%} \right) = 13.1456 \text{ g ZnS}$$

$$\text{Mass (g) of ZnO} = (13.1456 \text{ g Zn}) \left(\frac{1 \text{ mol Zn}}{65.41 \text{ g Zn}} \right) \left(\frac{2 \text{ mol ZnO}}{2 \text{ mol Zn}} \right) \left(\frac{81.41 \text{ g ZnO}}{1 \text{ mol ZnO}} \right) = 16.3612 = \mathbf{16.4 \text{ g ZnO}}$$

The calculation is slightly different for the sulfur. We need to determine the amount of sulfur not in zinc sulfide. The sulfur not in the zinc sulfide must be in sulfur dioxide. The amount of sulfur not in zinc sulfide will be converted to the mass of sulfur dioxide.

$$\text{Moles of S}_8 \text{ in original S}_8 \text{ reactant} = (52.4 \text{ g S}_8) \left(\frac{1 \text{ mol S}_8}{256.56 \text{ g S}_8} \right) = 0.204241 \text{ mol S}_8$$

$$\text{Moles of S}_8 \text{ in ZnS product} = (104.4 \text{ g ZnS}) \left(\frac{1 \text{ mol ZnS}}{97.48 \text{ g ZnS}} \right) \left(\frac{1 \text{ mol S}_8}{8 \text{ mol ZnS}} \right) = 0.133874 \text{ mol S}_8$$

$$\text{Moles of S}_8 \text{ in SO}_2 = 0.204241 - 0.133874 \text{ mol} = 0.070367 \text{ mol S}_8$$

$$\text{Mass (g) of SO}_2 = (0.070367 \text{ mol S}_8) \left(\frac{8 \text{ mol SO}_2}{1 \text{ mol S}_8} \right) \left(\frac{64.07 \text{ g SO}_2}{1 \text{ mol SO}_2} \right) = 36.0673 = \mathbf{36.1 \text{ g SO}_2}$$

- 3.111 Plan: Use the given values of x to find the molar mass of each compound. . To determine which reactant is limiting, calculate the amount of either product formed from each reactant, assuming an excess of the other reactants. The reactant that produces the smallest amount of product is the limiting reagent. To find the mass of excess reactants, find the mass of each excess reactant that is required to react with the limiting reagent and subtract that mass from the starting mass.

a) x = 0

$$\text{La}_2\text{Sr}_0\text{CuO}_4 = 2(138.9 \text{ g/mol La}) + 0(87.62 \text{ g/mol Sr}) + 1(63.55 \text{ g/mol Cu}) + 4(16.00 \text{ g/mol O}) = \mathbf{405.4 \text{ g/mol}}$$

x = 1

$$\text{La}_1\text{Sr}_1\text{CuO}_4 = 1(138.9 \text{ g/mol La}) + 1(87.62 \text{ g/mol Sr}) + 1(63.55 \text{ g/mol Cu}) + 4(16.00 \text{ g/mol O}) = \mathbf{354.1 \text{ g/mol}}$$

x = 0.163

$$\begin{aligned} \text{La}_{(2-0.163)}\text{Sr}_{0.163}\text{CuO}_4 &= \text{La}_{1.837}\text{Sr}_{0.163}\text{CuO}_4 \\ &= 1.837(138.9 \text{ g/mol La}) + 0.163(87.62 \text{ g/mol Sr}) + 1(63.55 \text{ g/mol Cu}) + 4(16.00 \text{ g/mol O}) \\ &= \mathbf{397.0 \text{ g/mol}} \end{aligned}$$

b) Assuming x grams to be the “equal” mass leads to:

$$\begin{aligned} \text{Moles of product from BaCO}_3 &= (x \text{ g BaCO}_3) \left(\frac{1 \text{ mol BaCO}_3}{197.3 \text{ g BaCO}_3} \right) \left(\frac{2 \text{ mol YBa}_2\text{Cu}_3\text{O}_7}{4 \text{ mol BaCO}_3} \right) \\ &= 0.002534x \text{ mol product} \end{aligned}$$

$$\text{Moles of product from CuO} = (x \text{ g CuO}) \left(\frac{1 \text{ mol CuO}}{79.55 \text{ g CuO}} \right) \left(\frac{2 \text{ mol YBa}_2\text{Cu}_3\text{O}_7}{6 \text{ mol CuO}} \right) = 0.004190x \text{ mol product}$$

$$\text{Moles of product from Y}_2\text{O}_3 = (x \text{ g Y}_2\text{O}_3) \left(\frac{1 \text{ mol Y}_2\text{O}_3}{225.82 \text{ g Y}_2\text{O}_3} \right) \left(\frac{2 \text{ mol YBa}_2\text{Cu}_3\text{O}_7}{1 \text{ mol Y}_2\text{O}_3} \right) = 0.008857x \text{ mol product}$$

BaCO₃ is the limiting reactant.

c) These calculations are based on the limiting reactant.

BaCO₃ remaining = 0% (limiting reagent)

$$\begin{aligned} \text{CuO remaining} &= x \text{ g CuO} - (x \text{ g BaCO}_3) \left(\frac{1 \text{ mol BaCO}_3}{197.3 \text{ g BaCO}_3} \right) \left(\frac{6 \text{ mol CuO}}{4 \text{ mol BaCO}_3} \right) \left(\frac{79.55 \text{ g CuO}}{1 \text{ mol CuO}} \right) \\ &= 0.39521x \text{ g CuO} \end{aligned}$$

$$\text{Percent CuO} = \left(\frac{0.39521 x \text{ g}}{x \text{ g}} \right) \times 100\% = 39.521 = \mathbf{39.52\% \text{ CuO remaining}}$$

$$\begin{aligned} \text{Y}_2\text{O}_3 \text{ remaining} &= x \text{ g Y}_2\text{O}_3 - (x \text{ g BaCO}_3) \left(\frac{1 \text{ mol BaCO}_3}{197.3 \text{ g BaCO}_3} \right) \left(\frac{1 \text{ mol Y}_2\text{O}_3}{4 \text{ mol BaCO}_3} \right) \left(\frac{225.82 \text{ g Y}_2\text{O}_3}{1 \text{ mol Y}_2\text{O}_3} \right) \\ &= 0.713862x \text{ g Y}_2\text{O}_3 \end{aligned}$$

$$\text{Percent Y}_2\text{O}_3 = \left(\frac{0.713862 x \text{ g}}{x \text{ g}} \right) \times 100\% = 71.3862 = \mathbf{71.39\% \text{ Y}_2\text{O}_3 \text{ remaining}}$$