

CHAPTER 3

MASS RELATIONSHIPS IN CHEMICAL REACTIONS

3.5 $(34.968 \text{ amu})(0.7553) + (36.956 \text{ amu})(0.2447) = 35.45 \text{ amu}$

3.6 **Strategy:** Each isotope contributes to the average atomic mass based on its relative abundance. Multiplying the mass of an isotope by its fractional abundance (not percent) will give the contribution to the average atomic mass of that particular isotope.

It would seem that there are two unknowns in this problem, the fractional abundance of ${}^6\text{Li}$ and the fractional abundance of ${}^7\text{Li}$. However, these two quantities are not independent of each other; they are related by the fact that they must sum to 1. Start by letting x be the fractional abundance of ${}^6\text{Li}$. Since the sum of the two abundance's must be 1, we can write

$$\text{Abundance } {}^7\text{Li} = (1 - x)$$

Solution:

$$\begin{aligned} \text{Average atomic mass of Li} &= 6.941 \text{ amu} = x(6.0151 \text{ amu}) + (1 - x)(7.0160 \text{ amu}) \\ 6.941 &= -1.0009x + 7.0160 \\ 1.0009x &= 0.075 \\ x &= 0.075 \end{aligned}$$

$x = 0.075$ corresponds to a natural abundance of ${}^6\text{Li}$ of **7.5 percent**. The natural abundance of ${}^7\text{Li}$ is $(1 - x) = 0.925$ or **92.5 percent**.

3.7 The unit factor required is $\left(\frac{6.022 \times 10^{23} \text{ amu}}{1 \text{ g}} \right)$

$$? \text{ g} = 13.2 \text{ amu} \times \frac{1 \text{ g}}{6.022 \times 10^{23} \text{ amu}} = 2.19 \times 10^{-23} \text{ g}$$

3.8 The unit factor required is $\left(\frac{6.022 \times 10^{23} \text{ amu}}{1 \text{ g}} \right)$

$$? \text{ amu} = 8.4 \text{ g} \times \frac{6.022 \times 10^{23} \text{ amu}}{1 \text{ g}} = 5.1 \times 10^{24} \text{ amu}$$

3.11 In one year:

$$(6.5 \times 10^9 \text{ people}) \times \frac{365 \text{ days}}{1 \text{ yr}} \times \frac{24 \text{ hr}}{1 \text{ day}} \times \frac{3600 \text{ s}}{1 \text{ hr}} \times \frac{2 \text{ particles}}{1 \text{ person}} = 4.1 \times 10^{17} \text{ particles/yr}$$

$$\text{Total time} = \frac{6.022 \times 10^{23} \text{ particles}}{4.1 \times 10^{17} \text{ particles/yr}} = 1.5 \times 10^6 \text{ yr}$$

3.12 The thickness of the book in miles would be:

$$\frac{0.0036 \text{ in}}{1 \text{ page}} \times \frac{1 \text{ ft}}{12 \text{ in}} \times \frac{1 \text{ mi}}{5280 \text{ ft}} \times (6.022 \times 10^{23} \text{ pages}) = 3.42 \times 10^{16} \text{ mi}$$

The distance, in miles, traveled by light in one year is:

$$1.00 \text{ yr} \times \frac{365 \text{ day}}{1 \text{ yr}} \times \frac{24 \text{ h}}{1 \text{ day}} \times \frac{3600 \text{ s}}{1 \text{ h}} \times \frac{3.00 \times 10^8 \text{ m}}{1 \text{ s}} \times \frac{1 \text{ mi}}{1609 \text{ m}} = 5.88 \times 10^{12} \text{ mi}$$

The thickness of the book in light-years is:

$$(3.42 \times 10^{16} \text{ mi}) \times \frac{1 \text{ light-yr}}{5.88 \times 10^{12} \text{ mi}} = \mathbf{5.8 \times 10^3 \text{ light-yr}}$$

It will take light 5.8×10^3 years to travel from the first page to the last one!

3.13 $5.10 \text{ mol S} \times \frac{6.022 \times 10^{23} \text{ S atoms}}{1 \text{ mol S}} = \mathbf{3.07 \times 10^{24} \text{ S atoms}}$

3.14 $(6.00 \times 10^9 \text{ Co atoms}) \times \frac{1 \text{ mol Co}}{6.022 \times 10^{23} \text{ Co atoms}} = \mathbf{9.96 \times 10^{-15} \text{ mol Co}}$

3.15 $77.4 \text{ g of Ca} \times \frac{1 \text{ mol Ca}}{40.08 \text{ g Ca}} = \mathbf{1.93 \text{ mol Ca}}$

3.16 Strategy: We are given moles of gold and asked to solve for grams of gold. What conversion factor do we need to convert between moles and grams? Arrange the appropriate conversion factor so moles cancel, and the unit grams is obtained for the answer.

Solution: The conversion factor needed to convert between moles and grams is the molar mass. In the periodic table (see inside front cover of the text), we see that the molar mass of Au is 197.0 g. This can be expressed as

$$1 \text{ mol Au} = 197.0 \text{ g Au}$$

From this equality, we can write two conversion factors.

$$\frac{1 \text{ mol Au}}{197.0 \text{ g Au}} \quad \text{and} \quad \frac{197.0 \text{ g Au}}{1 \text{ mol Au}}$$

The conversion factor on the right is the correct one. Moles will cancel, leaving the unit grams for the answer.

We write

$$? \text{ g Au} = 15.3 \text{ mol Au} \times \frac{197.0 \text{ g Au}}{1 \text{ mol Au}} = \mathbf{3.01 \times 10^3 \text{ g Au}}$$

Check: Does a mass of 3010 g for 15.3 moles of Au seem reasonable? What is the mass of 1 mole of Au?

$$3.17 \quad (a) \quad \frac{200.6 \text{ g Hg}}{1 \text{ mol Hg}} \times \frac{1 \text{ mol Hg}}{6.022 \times 10^{23} \text{ Hg atoms}} = 3.331 \times 10^{-22} \text{ g/Hg atom}$$

$$(b) \quad \frac{20.18 \text{ g Ne}}{1 \text{ mol Ne}} \times \frac{1 \text{ mol Ne}}{6.022 \times 10^{23} \text{ Ne atoms}} = 3.351 \times 10^{-23} \text{ g/Ne atom}$$

3.18 (a)

Strategy: We can look up the molar mass of arsenic (As) on the periodic table (74.92 g/mol). We want to find the mass of a single atom of arsenic (unit of g/atom). Therefore, we need to convert from the unit mole in the denominator to the unit atom in the denominator. What conversion factor is needed to convert between moles and atoms? Arrange the appropriate conversion factor so mole in the denominator cancels, and the unit atom is obtained in the denominator.

Solution: The conversion factor needed is Avogadro's number. We have

$$1 \text{ mol} = 6.022 \times 10^{23} \text{ particles (atoms)}$$

From this equality, we can write two conversion factors.

$$\frac{1 \text{ mol As}}{6.022 \times 10^{23} \text{ As atoms}} \quad \text{and} \quad \frac{6.022 \times 10^{23} \text{ As atoms}}{1 \text{ mol As}}$$

The conversion factor on the left is the correct one. Moles will cancel, leaving the unit atoms in the denominator of the answer.

We write

$$? \text{ g/As atom} = \frac{74.92 \text{ g As}}{1 \text{ mol As}} \times \frac{1 \text{ mol As}}{6.022 \times 10^{23} \text{ As atoms}} = 1.244 \times 10^{-22} \text{ g/As atom}$$

(b) Follow same method as part (a).

$$? \text{ g/Ni atom} = \frac{58.69 \text{ g Ni}}{1 \text{ mol Ni}} \times \frac{1 \text{ mol Ni}}{6.022 \times 10^{23} \text{ Ni atoms}} = 9.746 \times 10^{-23} \text{ g/Ni atom}$$

Check: Should the mass of a single atom of As or Ni be a very small mass?

$$3.19 \quad 1.00 \times 10^{12} \text{ Pb atoms} \times \frac{1 \text{ mol Pb}}{6.022 \times 10^{23} \text{ Pb atoms}} \times \frac{207.2 \text{ g Pb}}{1 \text{ mol Pb}} = 3.44 \times 10^{-10} \text{ g Pb}$$

3.20 **Strategy:** The question asks for atoms of Cu. We cannot convert directly from grams to atoms of copper. What unit do we need to convert grams of Cu to in order to convert to atoms? What does Avogadro's number represent?

Solution: To calculate the number of Cu atoms, we first must convert grams of Cu to moles of Cu. We use the molar mass of copper as a conversion factor. Once moles of Cu are obtained, we can use Avogadro's number to convert from moles of copper to atoms of copper.

$$1 \text{ mol Cu} = 63.55 \text{ g Cu}$$

The conversion factor needed is

$$\frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}}$$

Avogadro's number is the key to the second conversion. We have

$$1 \text{ mol} = 6.022 \times 10^{23} \text{ particles (atoms)}$$

From this equality, we can write two conversion factors.

$$\frac{1 \text{ mol Cu}}{6.022 \times 10^{23} \text{ Cu atoms}} \quad \text{and} \quad \frac{6.022 \times 10^{23} \text{ Cu atoms}}{1 \text{ mol Cu}}$$

The conversion factor on the right is the one we need because it has number of Cu atoms in the numerator, which is the unit we want for the answer.

Let's complete the two conversions in one step.

grams of Cu \rightarrow moles of Cu \rightarrow number of Cu atoms

$$\text{? atoms of Cu} = 3.14 \text{ g Cu} \times \frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}} \times \frac{6.022 \times 10^{23} \text{ Cu atoms}}{1 \text{ mol Cu}} = \mathbf{2.98 \times 10^{22} \text{ Cu atoms}}$$

Check: Should 3.14 g of Cu contain fewer than Avogadro's number of atoms? What mass of Cu would contain Avogadro's number of atoms?

3.21 For hydrogen: $1.10 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} \times \frac{6.022 \times 10^{23} \text{ H atoms}}{1 \text{ mol H}} = \mathbf{6.57 \times 10^{23} \text{ H atoms}}$

For chromium: $14.7 \text{ g Cr} \times \frac{1 \text{ mol Cr}}{52.00 \text{ g Cr}} \times \frac{6.022 \times 10^{23} \text{ Cr atoms}}{1 \text{ mol Cr}} = \mathbf{1.70 \times 10^{23} \text{ Cr atoms}}$

There are more hydrogen atoms than chromium atoms.

3.22 $2 \text{ Pb atoms} \times \frac{1 \text{ mol Pb}}{6.022 \times 10^{23} \text{ Pb atoms}} \times \frac{207.2 \text{ g Pb}}{1 \text{ mol Pb}} = 6.881 \times 10^{-22} \text{ g Pb}$

$$(5.1 \times 10^{-23} \text{ mol He}) \times \frac{4.003 \text{ g He}}{1 \text{ mol He}} = 2.0 \times 10^{-22} \text{ g He}$$

2 atoms of lead have a greater mass than 5.1×10^{-23} mol of helium.

3.23 Using the appropriate atomic masses,

- (a) CH_4 $12.01 \text{ amu} + 4(1.008 \text{ amu}) = \mathbf{16.04 \text{ amu}}$
- (b) NO_2 $14.01 \text{ amu} + 2(16.00 \text{ amu}) = \mathbf{46.01 \text{ amu}}$
- (c) SO_3 $32.07 \text{ amu} + 3(16.00 \text{ amu}) = \mathbf{80.07 \text{ amu}}$
- (d) C_6H_6 $6(12.01 \text{ amu}) + 6(1.008 \text{ amu}) = \mathbf{78.11 \text{ amu}}$
- (e) NaI $22.99 \text{ amu} + 126.9 \text{ amu} = \mathbf{149.9 \text{ amu}}$
- (f) K_2SO_4 $2(39.10 \text{ amu}) + 32.07 \text{ amu} + 4(16.00 \text{ amu}) = \mathbf{174.27 \text{ amu}}$
- (g) $\text{Ca}_3(\text{PO}_4)_2$ $3(40.08 \text{ amu}) + 2(30.97 \text{ amu}) + 8(16.00 \text{ amu}) = \mathbf{310.18 \text{ amu}}$

3.24 Strategy: How do molar masses of different elements combine to give the molar mass of a compound?

Solution: To calculate the molar mass of a compound, we need to sum all the molar masses of the elements in the molecule. For each element, we multiply its molar mass by the number of moles of that element in one mole of the compound. We find molar masses for the elements in the periodic table (inside front cover of the text).

(a) **molar mass Li_2CO_3** = $2(6.941 \text{ g}) + 12.01 \text{ g} + 3(16.00 \text{ g}) = \mathbf{73.89 \text{ g}}$

(b) **molar mass CS_2** = $12.01 \text{ g} + 2(32.07 \text{ g}) = \mathbf{76.15 \text{ g}}$

(c) **molar mass CHCl_3** = $12.01 \text{ g} + 1.008 \text{ g} + 3(35.45 \text{ g}) = \mathbf{119.37 \text{ g}}$

(d) **molar mass $\text{C}_6\text{H}_8\text{O}_6$** = $6(12.01 \text{ g}) + 8(1.008 \text{ g}) + 6(16.00 \text{ g}) = \mathbf{176.12 \text{ g}}$

(e) **molar mass KNO_3** = $39.10 \text{ g} + 14.01 \text{ g} + 3(16.00 \text{ g}) = \mathbf{101.11 \text{ g}}$

(f) **molar mass Mg_3N_2** = $3(24.31 \text{ g}) + 2(14.01 \text{ g}) = \mathbf{100.95 \text{ g}}$

3.25 To find the molar mass (g/mol), we simply divide the mass (in g) by the number of moles.

$$\frac{152 \text{ g}}{0.372 \text{ mol}} = \mathbf{409 \text{ g/mol}}$$

3.26 Strategy: We are given grams of ethane and asked to solve for molecules of ethane. We cannot convert directly from grams ethane to molecules of ethane. What unit do we need to obtain first before we can convert to molecules? How should Avogadro's number be used here?

Solution: To calculate number of ethane molecules, we first must convert grams of ethane to moles of ethane. We use the molar mass of ethane as a conversion factor. Once moles of ethane are obtained, we can use Avogadro's number to convert from moles of ethane to molecules of ethane.

$$\text{molar mass of } \text{C}_2\text{H}_6 = 2(12.01 \text{ g}) + 6(1.008 \text{ g}) = 30.068 \text{ g}$$

The conversion factor needed is

$$\frac{1 \text{ mol } \text{C}_2\text{H}_6}{30.068 \text{ g } \text{C}_2\text{H}_6}$$

Avogadro's number is the key to the second conversion. We have

$$1 \text{ mol} = 6.022 \times 10^{23} \text{ particles (molecules)}$$

From this equality, we can write the conversion factor:

$$\frac{6.022 \times 10^{23} \text{ ethane molecules}}{1 \text{ mol ethane}}$$

Let's complete the two conversions in one step.

grams of ethane \rightarrow moles of ethane \rightarrow number of ethane molecules

$$\begin{aligned} \text{? molecules of } \text{C}_2\text{H}_6 &= 0.334 \text{ g } \text{C}_2\text{H}_6 \times \frac{1 \text{ mol } \text{C}_2\text{H}_6}{30.068 \text{ g } \text{C}_2\text{H}_6} \times \frac{6.022 \times 10^{23} \text{ } \text{C}_2\text{H}_6 \text{ molecules}}{1 \text{ mol } \text{C}_2\text{H}_6} \\ &= \mathbf{6.69 \times 10^{21} \text{ } \text{C}_2\text{H}_6 \text{ molecules}} \end{aligned}$$

Check: Should 0.334 g of ethane contain fewer than Avogadro's number of molecules? What mass of ethane would contain Avogadro's number of molecules?

$$\begin{aligned}
 3.27 \quad & 1.50 \text{ g glucose} \times \frac{1 \text{ mol glucose}}{180.2 \text{ g glucose}} \times \frac{6.022 \times 10^{23} \text{ molecules glucose}}{1 \text{ mol glucose}} \times \frac{6 \text{ C atoms}}{1 \text{ molecule glucose}} \\
 & = \mathbf{3.01 \times 10^{22} \text{ C atoms}}
 \end{aligned}$$

The ratio of O atoms to C atoms in glucose is 1:1. Therefore, there are the same number of O atoms in glucose as C atoms, so the number of O atoms = $\mathbf{3.01 \times 10^{22} \text{ O atoms}}$.

The ratio of H atoms to C atoms in glucose is 2:1. Therefore, there are twice as many H atoms in glucose as C atoms, so the number of H atoms = $2(3.01 \times 10^{22} \text{ atoms}) = \mathbf{6.02 \times 10^{22} \text{ H atoms}}$.

3.28 Strategy: We are asked to solve for the number of N, C, O, and H atoms in 1.68×10^4 g of urea. We cannot convert directly from grams urea to atoms. What unit do we need to obtain first before we can convert to atoms? How should Avogadro's number be used here? How many atoms of N, C, O, or H are in 1 molecule of urea?

Solution: Let's first calculate the number of N atoms in 1.68×10^4 g of urea. First, we must convert grams of urea to number of molecules of urea. This calculation is similar to Problem 3.26. The molecular formula of urea shows there are two N atoms in one urea molecule, which will allow us to convert to atoms of N. We need to perform three conversions:

grams of urea \rightarrow moles of urea \rightarrow molecules of urea \rightarrow atoms of N

The conversion factors needed for each step are: 1) the molar mass of urea, 2) Avogadro's number, and 3) the number of N atoms in 1 molecule of urea.

We complete the three conversions in one calculation.

$$\begin{aligned}
 \text{? atoms of N} &= (1.68 \times 10^4 \text{ g urea}) \times \frac{1 \text{ mol urea}}{60.062 \text{ g urea}} \times \frac{6.022 \times 10^{23} \text{ urea molecules}}{1 \text{ mol urea}} \times \frac{2 \text{ N atoms}}{1 \text{ molecule urea}} \\
 &= \mathbf{3.37 \times 10^{26} \text{ N atoms}}
 \end{aligned}$$

The above method utilizes the ratio of molecules (urea) to atoms (nitrogen). We can also solve the problem by reading the formula as the ratio of moles of urea to moles of nitrogen by using the following conversions:

grams of urea \rightarrow moles of urea \rightarrow moles of N \rightarrow atoms of N

Try it.

Check: Does the answer seem reasonable? We have 1.68×10^4 g urea. How many atoms of N would 60.06 g of urea contain?

We could calculate the number of atoms of the remaining elements in the same manner, or we can use the atom ratios from the molecular formula. The carbon atom to nitrogen atom ratio in a urea molecule is 1:2, the oxygen atom to nitrogen atom ratio is 1:2, and the hydrogen atom to nitrogen atom ratio is 4:2.

$$\text{? atoms of C} = (3.37 \times 10^{26} \text{ N atoms}) \times \frac{1 \text{ C atom}}{2 \text{ N atoms}} = \mathbf{1.69 \times 10^{26} \text{ C atoms}}$$

$$? \text{ atoms of O} = (3.37 \times 10^{26} \text{ N atoms}) \times \frac{1 \text{ O atom}}{2 \text{ N atoms}} = \mathbf{1.69 \times 10^{26} \text{ O atoms}}$$

$$? \text{ atoms of H} = (3.37 \times 10^{26} \text{ N atoms}) \times \frac{4 \text{ H atoms}}{2 \text{ N atoms}} = \mathbf{6.74 \times 10^{26} \text{ H atoms}}$$

3.29 The molar mass of $\text{C}_{19}\text{H}_{38}\text{O}$ is 282.5 g.

$$1.0 \times 10^{-12} \text{ g} \times \frac{1 \text{ mol}}{282.49 \text{ g}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = \mathbf{2.1 \times 10^9 \text{ molecules}}$$

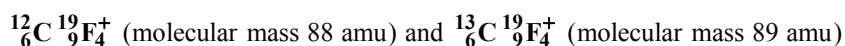
Notice that even though 1.0×10^{-12} g is an extremely small mass, it still is comprised of over a billion pheromone molecules!

3.30 Mass of water = $2.56 \text{ mL} \times \frac{1.00 \text{ g}}{1.00 \text{ mL}} = 2.56 \text{ g}$

Molar mass of H_2O = $(16.00 \text{ g}) + 2(1.008 \text{ g}) = 18.016 \text{ g/mol}$

$$\begin{aligned} ? \text{ H}_2\text{O molecules} &= 2.56 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.016 \text{ g H}_2\text{O}} \times \frac{6.022 \times 10^{23} \text{ molecules H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \\ &= \mathbf{8.56 \times 10^{22} \text{ molecules}} \end{aligned}$$

3.33 Since there are only two isotopes of carbon, there are only two possibilities for CF_4^+ .



There would be two peaks in the mass spectrum.

3.34 Since there are two hydrogen isotopes, they can be paired in three ways: ${}^1\text{H}-{}^1\text{H}$, ${}^1\text{H}-{}^2\text{H}$, and ${}^2\text{H}-{}^2\text{H}$. There will then be three choices for each sulfur isotope. We can make a table showing all the possibilities (masses in amu):

	${}^{32}\text{S}$	${}^{33}\text{S}$	${}^{34}\text{S}$	${}^{36}\text{S}$
${}^1\text{H}_2$	34	35	36	38
${}^1\text{H}{}^2\text{H}$	35	36	37	39
${}^2\text{H}_2$	36	37	38	40

There will be **seven peaks** of the following mass numbers: 34, 35, 36, 37, 38, 39, and 40.

Very accurate (and expensive!) mass spectrometers can detect the mass difference between two ${}^1\text{H}$ and one ${}^2\text{H}$. How many peaks would be detected in such a “high resolution” mass spectrum?

3.39 Molar mass of SnO_2 = $(118.7 \text{ g}) + 2(16.00 \text{ g}) = 150.7 \text{ g}$

$$\% \text{Sn} = \frac{118.7 \text{ g/mol}}{150.7 \text{ g/mol}} \times 100\% = \mathbf{78.77\%}$$

$$\% \text{O} = \frac{(2)(16.00 \text{ g/mol})}{150.7 \text{ g/mol}} \times 100\% = \mathbf{21.23\%}$$

3.40 Strategy: Recall the procedure for calculating a percentage. Assume that we have 1 mole of CHCl_3 . The percent by mass of each element (C, H, and Cl) is given by the mass of that element in 1 mole of CHCl_3 divided by the molar mass of CHCl_3 , then multiplied by 100 to convert from a fractional number to a percentage.

Solution: The molar mass of $\text{CHCl}_3 = 12.01 \text{ g/mol} + 1.008 \text{ g/mol} + 3(35.45 \text{ g/mol}) = 119.4 \text{ g/mol}$. The percent by mass of each of the elements in CHCl_3 is calculated as follows:

$$\%C = \frac{12.01 \text{ g/mol}}{119.4 \text{ g/mol}} \times 100\% = \mathbf{10.06\%}$$

$$\%H = \frac{1.008 \text{ g/mol}}{119.4 \text{ g/mol}} \times 100\% = \mathbf{0.8442\%}$$

$$\%Cl = \frac{3(35.45) \text{ g/mol}}{119.4 \text{ g/mol}} \times 100\% = \mathbf{89.07\%}$$

Check: Do the percentages add to 100%? The sum of the percentages is $(10.06\% + 0.8442\% + 89.07\%) = 99.97\%$. The small discrepancy from 100% is due to the way we rounded off.

3.41 The molar mass of cinnamic alcohol is 134.17 g/mol.

$$\text{(a) } \%C = \frac{(9)(12.01 \text{ g/mol})}{134.17 \text{ g/mol}} \times 100\% = \mathbf{80.56\%}$$

$$\%H = \frac{(10)(1.008 \text{ g/mol})}{134.17 \text{ g/mol}} \times 100\% = \mathbf{7.51\%}$$

$$\%O = \frac{16.00 \text{ g/mol}}{134.17 \text{ g/mol}} \times 100\% = \mathbf{11.93\%}$$

$$\begin{aligned} \text{(b) } 0.469 \text{ g C}_9\text{H}_{10}\text{O} &\times \frac{1 \text{ mol C}_9\text{H}_{10}\text{O}}{134.17 \text{ g C}_9\text{H}_{10}\text{O}} \times \frac{6.022 \times 10^{23} \text{ molecules C}_9\text{H}_{10}\text{O}}{1 \text{ mol C}_9\text{H}_{10}\text{O}} \\ &= \mathbf{2.11 \times 10^{21} \text{ molecules C}_9\text{H}_{10}\text{O}} \end{aligned}$$

3.42	Compound	Molar mass (g)	N% by mass
(a)	$(\text{NH}_2)_2\text{CO}$	60.06	$\frac{2(14.01 \text{ g})}{60.06 \text{ g}} \times 100\% = 46.65\%$
(b)	NH_4NO_3	80.05	$\frac{2(14.01 \text{ g})}{80.05 \text{ g}} \times 100\% = 35.00\%$
(c)	$\text{HNC}(\text{NH}_2)_2$	59.08	$\frac{3(14.01 \text{ g})}{59.08 \text{ g}} \times 100\% = 71.14\%$
(d)	NH_3	17.03	$\frac{14.01 \text{ g}}{17.03 \text{ g}} \times 100\% = 82.27\%$

Ammonia, NH_3 , is the richest source of nitrogen on a mass percentage basis.

3.43 Assume you have exactly 100 g of substance.

$$n_{\text{C}} = 44.4 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 3.697 \text{ mol C}$$

$$n_{\text{H}} = 6.21 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 6.161 \text{ mol H}$$

$$n_{\text{S}} = 39.5 \text{ g S} \times \frac{1 \text{ mol S}}{32.07 \text{ g S}} = 1.232 \text{ mol S}$$

$$n_{\text{O}} = 9.86 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.6163 \text{ mol O}$$

Thus, we arrive at the formula $\text{C}_{3.697}\text{H}_{6.161}\text{S}_{1.232}\text{O}_{0.6163}$. Dividing by the smallest number of moles (0.6163 mole) gives the empirical formula, **$\text{C}_6\text{H}_{10}\text{S}_2\text{O}$** .

To determine the molecular formula, divide the molar mass by the empirical mass.

$$\frac{\text{molar mass}}{\text{empirical molar mass}} = \frac{162 \text{ g}}{162.28 \text{ g}} \approx 1$$

Hence, the molecular formula and the empirical formula are the same, **$\text{C}_6\text{H}_{10}\text{S}_2\text{O}$** .

3.44 METHOD 1:

Step 1: Assume you have exactly 100 g of substance. 100 g is a convenient amount, because all the percentages sum to 100%. The percentage of oxygen is found by difference:

$$100\% - (19.8\% + 2.50\% + 11.6\%) = 66.1\%$$

In 100 g of PAN there will be 19.8 g C, 2.50 g H, 11.6 g N, and 66.1 g O.

Step 2: Calculate the number of moles of each element in the compound. Remember, an *empirical formula* tells us which elements are present and the simplest whole-number ratio of their atoms. This ratio is also a mole ratio. Use the molar masses of these elements as conversion factors to convert to moles.

$$n_{\text{C}} = 19.8 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 1.649 \text{ mol C}$$

$$n_{\text{H}} = 2.50 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 2.480 \text{ mol H}$$

$$n_{\text{N}} = 11.6 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 0.8280 \text{ mol N}$$

$$n_{\text{O}} = 66.1 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 4.131 \text{ mol O}$$

Step 3: Try to convert to whole numbers by dividing all the subscripts by the smallest subscript. The formula is $\text{C}_{1.649}\text{H}_{2.480}\text{N}_{0.8280}\text{O}_{4.131}$. Dividing the subscripts by 0.8280 gives the empirical formula, **$\text{C}_2\text{H}_3\text{NO}_5$** .

To determine the molecular formula, remember that the molar mass/empirical mass will be an integer greater than or equal to one.

$$\frac{\text{molar mass}}{\text{empirical molar mass}} \geq 1 \text{ (integer values)}$$

In this case,

$$\frac{\text{molar mass}}{\text{empirical molar mass}} = \frac{120 \text{ g}}{121.05 \text{ g}} \approx 1$$

Hence, the molecular formula and the empirical formula are the same, **C₂H₃NO₅**.

METHOD 2:

Step 1: Multiply the mass % (converted to a decimal) of each element by the molar mass to convert to grams of each element. Then, use the molar mass to convert to moles of each element.

$$n_{\text{C}} = (0.198) \times (120 \text{ g}) \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 1.98 \text{ mol C} \approx \mathbf{2 \text{ mol C}}$$

$$n_{\text{H}} = (0.0250) \times (120 \text{ g}) \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 2.98 \text{ mol H} \approx \mathbf{3 \text{ mol H}}$$

$$n_{\text{N}} = (0.116) \times (120 \text{ g}) \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 0.994 \text{ mol N} \approx \mathbf{1 \text{ mol N}}$$

$$n_{\text{O}} = (0.661) \times (120 \text{ g}) \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 4.96 \text{ mol O} \approx \mathbf{5 \text{ mol O}}$$

Step 2: Since we used the molar mass to calculate the moles of each element present in the compound, this method directly gives the molecular formula. The formula is **C₂H₃NO₅**.

Step 3: Try to reduce the molecular formula to a simpler whole number ratio to determine the empirical formula. The formula is already in its simplest whole number ratio. The molecular and empirical formulas are the same. The empirical formula is **C₂H₃NO₅**.

$$\mathbf{3.45} \quad 24.6 \text{ g Fe}_2\text{O}_3 \times \frac{1 \text{ mol Fe}_2\text{O}_3}{159.7 \text{ g Fe}_2\text{O}_3} \times \frac{2 \text{ mol Fe}}{1 \text{ mol Fe}_2\text{O}_3} = \mathbf{0.308 \text{ mol Fe}}$$

3.46 Using unit factors we convert:

$$\text{g of Hg} \rightarrow \text{mol Hg} \rightarrow \text{mol S} \rightarrow \text{g S}$$

$$\mathbf{? \text{ g S}} = 246 \text{ g Hg} \times \frac{1 \text{ mol Hg}}{200.6 \text{ g Hg}} \times \frac{1 \text{ mol S}}{1 \text{ mol Hg}} \times \frac{32.07 \text{ g S}}{1 \text{ mol S}} = \mathbf{39.3 \text{ g S}}$$

3.47 The balanced equation is: $2\text{Al}(s) + 3\text{I}_2(s) \longrightarrow 2\text{AlI}_3(s)$

Using unit factors, we convert: g of Al \rightarrow mol of Al \rightarrow mol of I₂ \rightarrow g of I₂

$$20.4 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \times \frac{3 \text{ mol I}_2}{2 \text{ mol Al}} \times \frac{253.8 \text{ g I}_2}{1 \text{ mol I}_2} = \mathbf{288 \text{ g I}_2}$$

3.48 Strategy: Tin(II) fluoride is composed of Sn and F. The mass due to F is based on its percentage by mass in the compound. How do we calculate mass percent of an element?

Solution: First, we must find the mass % of fluorine in SnF₂. Then, we convert this percentage to a fraction and multiply by the mass of the compound (24.6 g), to find the mass of fluorine in 24.6 g of SnF₂.

The percent by mass of fluorine in tin(II) fluoride, is calculated as follows:

$$\begin{aligned}\text{mass \% F} &= \frac{\text{mass of F in 1 mol SnF}_2}{\text{molar mass of SnF}_2} \times 100\% \\ &= \frac{2(19.00 \text{ g})}{156.7 \text{ g}} \times 100\% = 24.25\% \text{ F}\end{aligned}$$

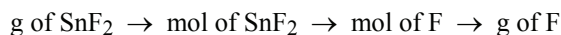
Converting this percentage to a fraction, we obtain $24.25/100 = 0.2425$.

Next, multiply the fraction by the total mass of the compound.

$$\text{? g F in 24.6 g SnF}_2 = (0.2425)(24.6 \text{ g}) = \mathbf{5.97 \text{ g F}}$$

Check: As a ball-park estimate, note that the mass percent of F is roughly 25 percent, so that a quarter of the mass should be F. One quarter of approximately 24 g is 6 g, which is close to the answer.

Note: This problem could have been worked in a manner similar to Problem 3.46. You could complete the following conversions:



3.49 In each case, assume 100 g of compound.

$$\text{(a)} \quad 2.1 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 2.08 \text{ mol H}$$

$$65.3 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 4.081 \text{ mol O}$$

$$32.6 \text{ g S} \times \frac{1 \text{ mol S}}{32.07 \text{ g S}} = 1.017 \text{ mol S}$$

This gives the formula $\text{H}_{2.08}\text{S}_{1.017}\text{O}_{4.081}$. Dividing by 1.017 gives the empirical formula, $\mathbf{H}_2\text{SO}_4$.

$$\text{(b)} \quad 20.2 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} = 0.7487 \text{ mol Al}$$

$$79.8 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} = 2.251 \text{ mol Cl}$$

This gives the formula, $\text{Al}_{0.7487}\text{Cl}_{2.251}$. Dividing by 0.7487 gives the empirical formula, \mathbf{AlCl}_3 .

3.50 (a)

Strategy: In a chemical formula, the subscripts represent the ratio of the number of moles of each element that combine to form the compound. Therefore, we need to convert from mass percent to moles in order to determine the empirical formula. If we assume an exactly 100 g sample of the compound, do we know the mass of each element in the compound? How do we then convert from grams to moles?

Solution: If we have 100 g of the compound, then each percentage can be converted directly to grams. In this sample, there will be 40.1 g of C, 6.6 g of H, and 53.3 g of O. Because the subscripts in the formula represent a mole ratio, we need to convert the grams of each element to moles. The conversion factor needed is the molar mass of each element. Let n represent the number of moles of each element so that

$$n_{\text{C}} = 40.1 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 3.339 \text{ mol C}$$

$$n_{\text{H}} = 6.6 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 6.55 \text{ mol H}$$

$$n_{\text{O}} = 53.3 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 3.331 \text{ mol O}$$

Thus, we arrive at the formula $\text{C}_{3.339}\text{H}_{6.55}\text{O}_{3.331}$, which gives the identity and the mole ratios of atoms present. However, chemical formulas are written with whole numbers. Try to convert to whole numbers by dividing all the subscripts by the smallest subscript (3.331).

$$\text{C} : \frac{3.339}{3.331} \approx 1 \quad \text{H} : \frac{6.55}{3.331} \approx 2 \quad \text{O} : \frac{3.331}{3.331} = 1$$

This gives the empirical formula, **CH₂O**.

Check: Are the subscripts in CH₂O reduced to the smallest whole numbers?

(b) Following the same procedure as part (a), we find:

$$n_{\text{C}} = 18.4 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 1.532 \text{ mol C}$$

$$n_{\text{N}} = 21.5 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 1.535 \text{ mol N}$$

$$n_{\text{K}} = 60.1 \text{ g K} \times \frac{1 \text{ mol K}}{39.10 \text{ g K}} = 1.537 \text{ mol K}$$

Dividing by the smallest number of moles (1.532 mol) gives the empirical formula, **KCN**.

3.51 The molar mass of CaSiO₃ is 116.17 g/mol.

$$\% \text{Ca} = \frac{40.08 \text{ g}}{116.17 \text{ g}} = \mathbf{34.50\%}$$

$$\% \text{Si} = \frac{28.09 \text{ g}}{116.17 \text{ g}} = \mathbf{24.18\%}$$

$$\% \text{O} = \frac{(3)(16.00 \text{ g})}{116.17 \text{ g}} = \mathbf{41.32\%}$$

Check to see that the percentages sum to 100%. (34.50% + 24.18% + 41.32%) = 100.00%

3.52 The empirical molar mass of CH is approximately 13.018 g. Let's compare this to the molar mass to determine the molecular formula.

Recall that the molar mass divided by the empirical mass will be an integer greater than or equal to one.

$$\frac{\text{molar mass}}{\text{empirical molar mass}} \geq 1 \text{ (integer values)}$$

In this case,

$$\frac{\text{molar mass}}{\text{empirical molar mass}} = \frac{78 \text{ g}}{13.018 \text{ g}} \approx 6$$

Thus, there are six CH units in each molecule of the compound, so the molecular formula is $(\text{CH})_6$, or C_6H_6 .

3.53 Find the molar mass corresponding to each formula.

$$\text{For } \text{C}_4\text{H}_5\text{N}_2\text{O}: \quad 4(12.01 \text{ g}) + 5(1.008 \text{ g}) + 2(14.01 \text{ g}) + (16.00 \text{ g}) = 97.10 \text{ g}$$

$$\text{For } \text{C}_8\text{H}_{10}\text{N}_4\text{O}_2: \quad 8(12.01 \text{ g}) + 10(1.008 \text{ g}) + 4(14.01 \text{ g}) + 2(16.00 \text{ g}) = 194.20 \text{ g}$$

The molecular formula is $\text{C}_8\text{H}_{10}\text{N}_4\text{O}_2$.

3.54 **METHOD 1:**

Step 1: Assume you have exactly 100 g of substance. 100 g is a convenient amount, because all the percentages sum to 100%. In 100 g of MSG there will be 35.51 g C, 4.77 g H, 37.85 g O, 8.29 g N, and 13.60 g Na.

Step 2: Calculate the number of moles of each element in the compound. Remember, an *empirical formula* tells us which elements are present and the simplest whole-number ratio of their atoms. This ratio is also a mole ratio. Let n_{C} , n_{H} , n_{O} , n_{N} , and n_{Na} be the number of moles of elements present. Use the molar masses of these elements as conversion factors to convert to moles.

$$n_{\text{C}} = 35.51 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 2.9567 \text{ mol C}$$

$$n_{\text{H}} = 4.77 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 4.732 \text{ mol H}$$

$$n_{\text{O}} = 37.85 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.3656 \text{ mol O}$$

$$n_{\text{N}} = 8.29 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 0.5917 \text{ mol N}$$

$$n_{\text{Na}} = 13.60 \text{ g Na} \times \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} = 0.59156 \text{ mol Na}$$

Thus, we arrive at the formula $\text{C}_{2.9567}\text{H}_{4.732}\text{O}_{2.3656}\text{N}_{0.5917}\text{Na}_{0.59156}$, which gives the identity and the ratios of atoms present. However, chemical formulas are written with whole numbers.

Step 3: Try to convert to whole numbers by dividing all the subscripts by the smallest subscript.

$$\begin{array}{lll} \text{C: } \frac{2.9567}{0.59156} = 4.9981 \approx 5 & \text{H: } \frac{4.732}{0.59156} = 7.999 \approx 8 & \text{O: } \frac{2.3656}{0.59156} = 3.9989 \approx 4 \\ \text{N: } \frac{0.5917}{0.59156} = 1.000 & \text{Na: } \frac{0.59156}{0.59156} = 1 & \end{array}$$

This gives us the empirical formula for MSG, $\text{C}_5\text{H}_8\text{O}_4\text{NNa}$.

To determine the molecular formula, remember that the molar mass/empirical mass will be an integer greater than or equal to one.

$$\frac{\text{molar mass}}{\text{empirical molar mass}} \geq 1 \text{ (integer values)}$$

In this case,

$$\frac{\text{molar mass}}{\text{empirical molar mass}} = \frac{169 \text{ g}}{169.11 \text{ g}} \approx 1$$

Hence, the molecular formula and the empirical formula are the same, **C₅H₈O₄NNa**. It should come as no surprise that the empirical and molecular formulas are the same since MSG stands for *monosodium*glutamate.

METHOD 2:

Step 1: Multiply the mass % (converted to a decimal) of each element by the molar mass to convert to grams of each element. Then, use the molar mass to convert to moles of each element.

$$n_{\text{C}} = (0.3551) \times (169 \text{ g}) \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 5.00 \text{ mol C}$$

$$n_{\text{H}} = (0.0477) \times (169 \text{ g}) \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 8.00 \text{ mol H}$$

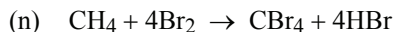
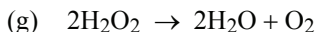
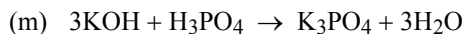
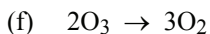
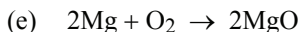
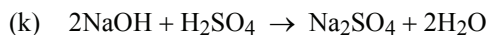
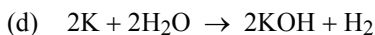
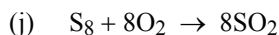
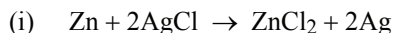
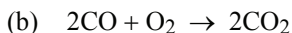
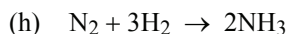
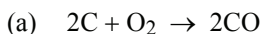
$$n_{\text{O}} = (0.3785) \times (169 \text{ g}) \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 4.00 \text{ mol O}$$

$$n_{\text{N}} = (0.0829) \times (169 \text{ g}) \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 1.00 \text{ mol N}$$

$$n_{\text{Na}} = (0.1360) \times (169 \text{ g}) \times \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} = 1.00 \text{ mol Na}$$

Step 2: Since we used the molar mass to calculate the moles of each element present in the compound, this method directly gives the molecular formula. The formula is **C₅H₈O₄NNa**.

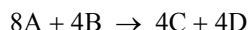
3.59 The balanced equations are as follows:



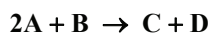
3.60 The balanced equations are as follows:

- | | |
|--|---|
| (a) $2\text{N}_2\text{O}_5 \rightarrow 2\text{N}_2\text{O}_4 + \text{O}_2$ | (h) $2\text{Al} + 3\text{H}_2\text{SO}_4 \rightarrow \text{Al}_2(\text{SO}_4)_3 + 3\text{H}_2$ |
| (b) $2\text{KNO}_3 \rightarrow 2\text{KNO}_2 + \text{O}_2$ | (i) $\text{CO}_2 + 2\text{KOH} \rightarrow \text{K}_2\text{CO}_3 + \text{H}_2\text{O}$ |
| (c) $\text{NH}_4\text{NO}_3 \rightarrow \text{N}_2\text{O} + 2\text{H}_2\text{O}$ | (j) $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$ |
| (d) $\text{NH}_4\text{NO}_2 \rightarrow \text{N}_2 + 2\text{H}_2\text{O}$ | (k) $\text{Be}_2\text{C} + 4\text{H}_2\text{O} \rightarrow 2\text{Be}(\text{OH})_2 + \text{CH}_4$ |
| (e) $2\text{NaHCO}_3 \rightarrow \text{Na}_2\text{CO}_3 + \text{H}_2\text{O} + \text{CO}_2$ | (l) $3\text{Cu} + 8\text{HNO}_3 \rightarrow 3\text{Cu}(\text{NO}_3)_2 + 2\text{NO} + 4\text{H}_2\text{O}$ |
| (f) $\text{P}_4\text{O}_{10} + 6\text{H}_2\text{O} \rightarrow 4\text{H}_3\text{PO}_4$ | (m) $\text{S} + 6\text{HNO}_3 \rightarrow \text{H}_2\text{SO}_4 + 6\text{NO}_2 + 2\text{H}_2\text{O}$ |
| (g) $2\text{HCl} + \text{CaCO}_3 \rightarrow \text{CaCl}_2 + \text{H}_2\text{O} + \text{CO}_2$ | (n) $2\text{NH}_3 + 3\text{CuO} \rightarrow 3\text{Cu} + \text{N}_2 + 3\text{H}_2\text{O}$ |

3.63 On the reactants side there are 8 A atoms and 4 B atoms. On the products side, there are 4 C atoms and 4 D atoms. Writing an equation,

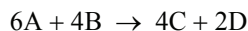


Chemical equations are typically written with the smallest set of whole number coefficients. Dividing the equation by four gives,

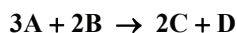


The correct answer is choice (c).

3.64 On the reactants side there are 6 A atoms and 4 B atoms. On the products side, there are 4 C atoms and 2 D atoms. Writing an equation,



Chemical equations are typically written with the smallest set of whole number coefficients. Dividing the equation by two gives,



The correct answer is choice (d).

3.65 The mole ratio from the balanced equation is 2 moles CO_2 : 2 moles CO .

$$3.60 \text{ mol CO} \times \frac{2 \text{ mol CO}_2}{2 \text{ mol CO}} = 3.60 \text{ mol CO}_2$$

3.66 $\text{Si}(s) + 2\text{Cl}_2(g) \longrightarrow \text{SiCl}_4(l)$

Strategy: Looking at the balanced equation, how do we compare the amounts of Cl_2 and SiCl_4 ? We can compare them based on the mole ratio from the balanced equation.

Solution: Because the balanced equation is given in the problem, the mole ratio between Cl_2 and SiCl_4 is known: 2 moles $\text{Cl}_2 \approx 1$ mole SiCl_4 . From this relationship, we have two conversion factors.

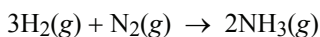
$$\frac{2 \text{ mol Cl}_2}{1 \text{ mol SiCl}_4} \quad \text{and} \quad \frac{1 \text{ mol SiCl}_4}{2 \text{ mol Cl}_2}$$

Which conversion factor is needed to convert from moles of SiCl_4 to moles of Cl_2 ? The conversion factor on the left is the correct one. Moles of SiCl_4 will cancel, leaving units of "mol Cl_2 " for the answer. We calculate moles of Cl_2 reacted as follows:

$$\text{? mol Cl}_2 \text{ reacted} = 0.507 \text{ mol SiCl}_4 \times \frac{2 \text{ mol Cl}_2}{1 \text{ mol SiCl}_4} = \mathbf{1.01 \text{ mol Cl}_2}$$

Check: Does the answer seem reasonable? Should the moles of Cl_2 reacted be *double* the moles of SiCl_4 produced?

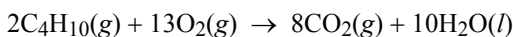
- 3.67** Starting with the amount of ammonia produced (6.0 moles), we can use the mole ratio from the balanced equation to calculate the moles of H_2 and N_2 that reacted to produce 6.0 moles of NH_3 .



$$\text{? mol H}_2 = 6.0 \text{ mol NH}_3 \times \frac{3 \text{ mol H}_2}{2 \text{ mol NH}_3} = \mathbf{9.0 \text{ mol H}_2}$$

$$\text{? mol N}_2 = 6.0 \text{ mol NH}_3 \times \frac{1 \text{ mol N}_2}{2 \text{ mol NH}_3} = \mathbf{3.0 \text{ mol N}_2}$$

- 3.68** Starting with the 5.0 moles of C_4H_{10} , we can use the mole ratio from the balanced equation to calculate the moles of CO_2 formed.



$$\text{? mol CO}_2 = 5.0 \text{ mol C}_4\text{H}_{10} \times \frac{8 \text{ mol CO}_2}{2 \text{ mol C}_4\text{H}_{10}} = 20 \text{ mol CO}_2 = \mathbf{2.0 \times 10^1 \text{ mol CO}_2}$$

- 3.69** It is convenient to use the unit ton-mol in this problem. We normally use a g-mol. 1 g-mol SO_2 has a mass of 64.07 g. In a similar manner, 1 ton-mol of SO_2 has a mass of 64.07 tons. We need to complete the following conversions: tons $\text{SO}_2 \rightarrow$ ton-mol $\text{SO}_2 \rightarrow$ ton-mol S \rightarrow ton S.

$$(2.6 \times 10^7 \text{ tons SO}_2) \times \frac{1 \text{ ton-mol SO}_2}{64.07 \text{ ton SO}_2} \times \frac{1 \text{ ton-mol S}}{1 \text{ ton-mol SO}_2} \times \frac{32.07 \text{ ton S}}{1 \text{ ton-mol S}} = \mathbf{1.3 \times 10^7 \text{ tons S}}$$

- 3.70 (a)** $2\text{NaHCO}_3 \longrightarrow \text{Na}_2\text{CO}_3 + \text{H}_2\text{O} + \text{CO}_2$

- (b)** Molar mass $\text{NaHCO}_3 = 22.99 \text{ g} + 1.008 \text{ g} + 12.01 \text{ g} + 3(16.00 \text{ g}) = 84.008 \text{ g}$
 Molar mass $\text{CO}_2 = 12.01 \text{ g} + 2(16.00 \text{ g}) = 44.01 \text{ g}$

The balanced equation shows one mole of CO_2 formed from two moles of NaHCO_3 .

$$\begin{aligned} \text{mass NaHCO}_3 &= 20.5 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{2 \text{ mol NaHCO}_3}{1 \text{ mol CO}_2} \times \frac{84.008 \text{ g NaHCO}_3}{1 \text{ mol NaHCO}_3} \\ &= \mathbf{78.3 \text{ g NaHCO}_3} \end{aligned}$$

3.71 The balanced equation shows a mole ratio of 1 mole HCN : 1 mole KCN.

$$0.140 \text{ g KCN} \times \frac{1 \text{ mol KCN}}{65.12 \text{ g KCN}} \times \frac{1 \text{ mol HCN}}{1 \text{ mol KCN}} \times \frac{27.03 \text{ g HCN}}{1 \text{ mol HCN}} = \mathbf{0.0581 \text{ g HCN}}$$

3.72 $\text{C}_6\text{H}_{12}\text{O}_6 \longrightarrow 2\text{C}_2\text{H}_5\text{OH} + 2\text{CO}_2$
 glucose ethanol

Strategy: We compare glucose and ethanol based on the *mole ratio* in the balanced equation. Before we can determine moles of ethanol produced, we need to convert to moles of glucose. What conversion factor is needed to convert from grams of glucose to moles of glucose? Once moles of ethanol are obtained, another conversion factor is needed to convert from moles of ethanol to grams of ethanol.

Solution: The molar mass of glucose will allow us to convert from grams of glucose to moles of glucose. The molar mass of glucose = $6(12.01 \text{ g}) + 12(1.008 \text{ g}) + 6(16.00 \text{ g}) = 180.16 \text{ g}$. The balanced equation is given, so the mole ratio between glucose and ethanol is known; that is 1 mole glucose \simeq 2 moles ethanol. Finally, the molar mass of ethanol will convert moles of ethanol to grams of ethanol. This sequence of three conversions is summarized as follows:

grams of glucose \rightarrow moles of glucose \rightarrow moles of ethanol \rightarrow grams of ethanol

$$\begin{aligned} ? \text{ g C}_2\text{H}_5\text{OH} &= 500.4 \text{ g C}_6\text{H}_{12}\text{O}_6 \times \frac{1 \text{ mol C}_6\text{H}_{12}\text{O}_6}{180.16 \text{ g C}_6\text{H}_{12}\text{O}_6} \times \frac{2 \text{ mol C}_2\text{H}_5\text{OH}}{1 \text{ mol C}_6\text{H}_{12}\text{O}_6} \times \frac{46.068 \text{ g C}_2\text{H}_5\text{OH}}{1 \text{ mol C}_2\text{H}_5\text{OH}} \\ &= \mathbf{255.9 \text{ g C}_2\text{H}_5\text{OH}} \end{aligned}$$

Check: Does the answer seem reasonable? Should the mass of ethanol produced be approximately half the mass of glucose reacted? Twice as many moles of ethanol are produced compared to the moles of glucose reacted, but the molar mass of ethanol is about one-fourth that of glucose.

The liters of ethanol can be calculated from the density and the mass of ethanol.

$$\text{volume} = \frac{\text{mass}}{\text{density}}$$

$$\mathbf{\text{Volume of ethanol obtained}} = \frac{255.9 \text{ g}}{0.789 \text{ g/mL}} = 324 \text{ mL} = \mathbf{0.324 \text{ L}}$$

3.73 The mass of water lost is just the difference between the initial and final masses.

$$\text{Mass H}_2\text{O lost} = 15.01 \text{ g} - 9.60 \text{ g} = 5.41 \text{ g}$$

$$\text{moles of H}_2\text{O} = 5.41 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.016 \text{ g H}_2\text{O}} = \mathbf{0.300 \text{ mol H}_2\text{O}}$$

3.74 The balanced equation shows that eight moles of KCN are needed to combine with four moles of Au.

$$? \text{ mol KCN} = 29.0 \text{ g Au} \times \frac{1 \text{ mol Au}}{197.0 \text{ g Au}} \times \frac{8 \text{ mol KCN}}{4 \text{ mol Au}} = \mathbf{0.294 \text{ mol KCN}}$$

3.75 The balanced equation is: $\text{CaCO}_3(s) \longrightarrow \text{CaO}(s) + \text{CO}_2(g)$

$$1.0 \text{ kg CaCO}_3 \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol CaCO}_3}{100.09 \text{ g CaCO}_3} \times \frac{1 \text{ mol CaO}}{1 \text{ mol CaCO}_3} \times \frac{56.08 \text{ g CaO}}{1 \text{ mol CaO}} = 5.6 \times 10^2 \text{ g CaO}$$

3.76 (a) $\text{NH}_4\text{NO}_3(s) \longrightarrow \text{N}_2\text{O}(g) + 2\text{H}_2\text{O}(g)$

(b) Starting with moles of NH_4NO_3 , we can use the mole ratio from the balanced equation to find moles of N_2O . Once we have moles of N_2O , we can use the molar mass of N_2O to convert to grams of N_2O . Combining the two conversions into one calculation, we have:

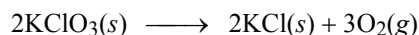
$$\text{mol NH}_4\text{NO}_3 \rightarrow \text{mol N}_2\text{O} \rightarrow \text{g N}_2\text{O}$$

$$? \text{ g N}_2\text{O} = 0.46 \text{ mol NH}_4\text{NO}_3 \times \frac{1 \text{ mol N}_2\text{O}}{1 \text{ mol NH}_4\text{NO}_3} \times \frac{44.02 \text{ g N}_2\text{O}}{1 \text{ mol N}_2\text{O}} = 2.0 \times 10^1 \text{ g N}_2\text{O}$$

3.77 The quantity of ammonia needed is:

$$1.00 \times 10^8 \text{ g (NH}_4)_2\text{SO}_4 \times \frac{1 \text{ mol (NH}_4)_2\text{SO}_4}{132.15 \text{ g (NH}_4)_2\text{SO}_4} \times \frac{2 \text{ mol NH}_3}{1 \text{ mol (NH}_4)_2\text{SO}_4} \times \frac{17.034 \text{ g NH}_3}{1 \text{ mol NH}_3} \times \frac{1 \text{ kg}}{1000 \text{ g}} \\ = 2.58 \times 10^4 \text{ kg NH}_3$$

3.78 The balanced equation for the decomposition is :



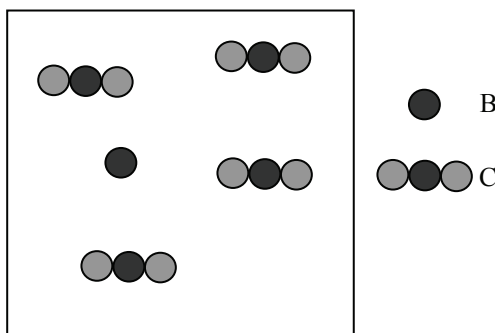
$$? \text{ g O}_2 = 46.0 \text{ g KClO}_3 \times \frac{1 \text{ mol KClO}_3}{122.55 \text{ g KClO}_3} \times \frac{3 \text{ mol O}_2}{2 \text{ mol KClO}_3} \times \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} = 18.0 \text{ g O}_2$$

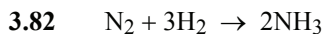
3.81 $2\text{A} + \text{B} \rightarrow \text{C}$

(a) The number of B atoms shown in the diagram is 5. The balanced equation shows 2 moles A \simeq 1 mole B. Therefore, we need 10 atoms of A to react completely with 5 atoms of B. There are only 8 atoms of A present in the diagram. There are not enough atoms of A to react completely with B.

A is the limiting reagent.

(b) There are 8 atoms of A. Since the mole ratio between A and B is 2:1, 4 atoms of B will react with 8 atoms of A, leaving 1 atom of B in excess. The mole ratio between A and C is also 2:1. When 8 atoms of A react, 4 molecules of C will be produced.

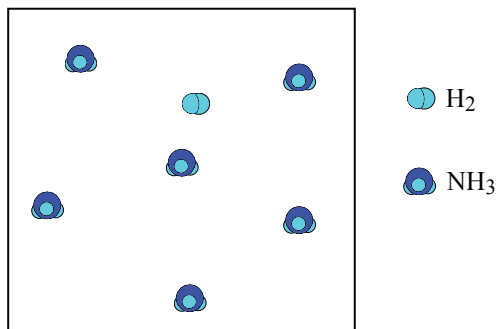




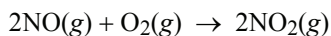
- (a) The number of N_2 molecules shown in the diagram is 3. The balanced equation shows 3 moles $\text{H}_2 \simeq 1$ mole N_2 . Therefore, we need 9 molecules of H_2 to react completely with 3 molecules of N_2 . There are 10 molecules of H_2 present in the diagram. H_2 is in excess.

N_2 is the limiting reagent.

- (b) 9 molecules of H_2 will react with 3 molecules of N_2 , leaving 1 molecule of H_2 in excess. The mole ratio between N_2 and NH_3 is 1:2. When 3 molecules of N_2 react, 6 molecules of NH_3 will be produced.



- 3.83 This is a limiting reagent problem. Let's calculate the moles of NO_2 produced assuming complete reaction for each reactant.



$$0.886 \text{ mol NO} \times \frac{2 \text{ mol NO}_2}{2 \text{ mol NO}} = 0.886 \text{ mol NO}_2$$

$$0.503 \text{ mol O}_2 \times \frac{2 \text{ mol NO}_2}{1 \text{ mol O}_2} = 1.01 \text{ mol NO}_2$$

NO is the **limiting reagent**; it limits the amount of product produced. The amount of product produced is **0.886 mole NO_2** .

- 3.84 **Strategy:** Note that this reaction gives the amounts of both reactants, so it is likely to be a limiting reagent problem. The reactant that produces fewer moles of product is the limiting reagent because it limits the amount of product that can be produced. How do we convert from the amount of reactant to amount of product? Perform this calculation for each reactant, then compare the moles of product, NO_2 , formed by the given amounts of O_3 and NO to determine which reactant is the limiting reagent.

Solution: We carry out two separate calculations. First, starting with 0.740 g O_3 , we calculate the number of moles of NO_2 that could be produced if all the O_3 reacted. We complete the following conversions.

$$\text{grams of O}_3 \rightarrow \text{moles of O}_3 \rightarrow \text{moles of NO}_2$$

Combining these two conversions into one calculation, we write

$$? \text{ mol NO}_2 = 0.740 \text{ g O}_3 \times \frac{1 \text{ mol O}_3}{48.00 \text{ g O}_3} \times \frac{1 \text{ mol NO}_2}{1 \text{ mol O}_3} = 0.01542 \text{ mol NO}_2$$

Second, starting with 0.670 g of NO, we complete similar conversions.

grams of NO → moles of NO → moles of NO₂

Combining these two conversions into one calculation, we write

$$? \text{ mol NO}_2 = 0.670 \text{ g NO} \times \frac{1 \text{ mol NO}}{30.01 \text{ g NO}} \times \frac{1 \text{ mol NO}_2}{1 \text{ mol NO}} = 0.02233 \text{ mol NO}_2$$

The initial amount of O₃ limits the amount of product that can be formed; therefore, it is the **limiting reagent**.

The problem asks for grams of NO₂ produced. We already know the moles of NO₂ produced, 0.01542 mole. Use the molar mass of NO₂ as a conversion factor to convert to grams (Molar mass NO₂ = 46.01 g).

$$? \text{ g NO}_2 = 0.01542 \text{ mol NO}_2 \times \frac{46.01 \text{ g NO}_2}{1 \text{ mol NO}_2} = \mathbf{0.709 \text{ g NO}_2}$$

Check: Does your answer seem reasonable? 0.01542 mole of product is formed. What is the mass of 1 mole of NO₂?

Strategy: Working backwards, we can determine the amount of NO that reacted to produce 0.01542 mole of NO₂. The amount of NO left over is the difference between the initial amount and the amount reacted.

Solution: Starting with 0.01542 mole of NO₂, we can determine the moles of NO that reacted using the mole ratio from the balanced equation. We can calculate the initial moles of NO starting with 0.670 g and using molar mass of NO as a conversion factor.

$$\text{mol NO reacted} = 0.01542 \text{ mol NO}_2 \times \frac{1 \text{ mol NO}}{1 \text{ mol NO}_2} = 0.01542 \text{ mol NO}$$

$$\text{mol NO initial} = 0.670 \text{ g NO} \times \frac{1 \text{ mol NO}}{30.01 \text{ g NO}} = 0.02233 \text{ mol NO}$$

$$\text{mol NO remaining} = \text{mol NO initial} - \text{mol NO reacted.}$$

$$\mathbf{\text{mol NO remaining} = 0.02233 \text{ mol NO} - 0.01542 \text{ mol NO} = \mathbf{0.0069 \text{ mol NO}}$$

3.85 (a) The balanced equation is: $\text{C}_3\text{H}_8(\text{g}) + 5\text{O}_2(\text{g}) \longrightarrow 3\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\text{l})$

(b) The balanced equation shows a mole ratio of 3 moles CO₂ : 1 mole C₃H₈. The mass of CO₂ produced is:

$$3.65 \text{ mol C}_3\text{H}_8 \times \frac{3 \text{ mol CO}_2}{1 \text{ mol C}_3\text{H}_8} \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = \mathbf{482 \text{ g CO}_2}$$

3.86 This is a limiting reagent problem. Let's calculate the moles of Cl₂ produced assuming complete reaction for each reactant.

$$0.86 \text{ mol MnO}_2 \times \frac{1 \text{ mol Cl}_2}{1 \text{ mol MnO}_2} = 0.86 \text{ mol Cl}_2$$

$$48.2 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.458 \text{ g HCl}} \times \frac{1 \text{ mol Cl}_2}{4 \text{ mol HCl}} = 0.3305 \text{ mol Cl}_2$$

HCl is the limiting reagent; it limits the amount of product produced. It will be used up first. The amount of product produced is 0.3305 mole Cl_2 . Let's convert this to grams.

$$? \text{ g Cl}_2 = 0.3305 \text{ mol Cl}_2 \times \frac{70.90 \text{ g Cl}_2}{1 \text{ mol Cl}_2} = \mathbf{23.4 \text{ g Cl}_2}$$

3.89 The balanced equation is given: $\text{CaF}_2 + \text{H}_2\text{SO}_4 \longrightarrow \text{CaSO}_4 + 2\text{HF}$

The balanced equation shows a mole ratio of 2 moles HF : 1 mole CaF_2 . The theoretical yield of HF is:

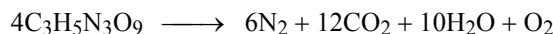
$$(6.00 \times 10^3 \text{ g CaF}_2) \times \frac{1 \text{ mol CaF}_2}{78.08 \text{ g CaF}_2} \times \frac{2 \text{ mol HF}}{1 \text{ mol CaF}_2} \times \frac{20.008 \text{ g HF}}{1 \text{ mol HF}} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 3.075 \text{ kg HF}$$

The actual yield is given in the problem (2.86 kg HF).

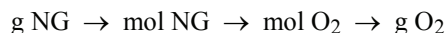
$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

$$\% \text{ yield} = \frac{2.86 \text{ kg}}{3.075 \text{ kg}} \times 100\% = \mathbf{93.0\%}$$

3.90 (a) Start with a balanced chemical equation. It's given in the problem. We use NG as an abbreviation for nitroglycerin. The molar mass of NG = 227.1 g/mol.



Map out the following strategy to solve this problem.



Calculate the grams of O_2 using the strategy above.

$$? \text{ g O}_2 = 2.00 \times 10^2 \text{ g NG} \times \frac{1 \text{ mol NG}}{227.1 \text{ g NG}} \times \frac{1 \text{ mol O}_2}{4 \text{ mol NG}} \times \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} = \mathbf{7.05 \text{ g O}_2}$$

(b) The theoretical yield was calculated in part (a), and the actual yield is given in the problem (6.55 g). The percent yield is:

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

$$\% \text{ yield} = \frac{6.55 \text{ g O}_2}{7.05 \text{ g O}_2} \times 100\% = \mathbf{92.9\%}$$

3.91 The balanced equation shows a mole ratio of 1 mole TiO_2 : 1 mole FeTiO_3 . The molar mass of FeTiO_3 is 151.73 g/mol, and the molar mass of TiO_2 is 79.88 g/mol. The theoretical yield of TiO_2 is:

$$8.00 \times 10^6 \text{ g FeTiO}_3 \times \frac{1 \text{ mol FeTiO}_3}{151.73 \text{ g FeTiO}_3} \times \frac{1 \text{ mol TiO}_2}{1 \text{ mol FeTiO}_3} \times \frac{79.88 \text{ g TiO}_2}{1 \text{ mol TiO}_2} \times \frac{1 \text{ kg}}{1000 \text{ g}} \\ = \mathbf{4.21 \times 10^3 \text{ kg TiO}_2}$$

The actual yield is given in the problem (3.67×10^3 kg TiO_2).

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{3.67 \times 10^3 \text{ kg}}{4.21 \times 10^3 \text{ kg}} \times 100\% = \mathbf{87.2\%}$$

- 3.92** The actual yield of ethylene is 481 g. Let's calculate the yield of ethylene if the reaction is 100 percent efficient. We can calculate this from the definition of percent yield. We can then calculate the mass of hexane that must be reacted.

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

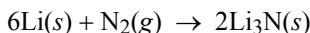
$$42.5\% \text{ yield} = \frac{481 \text{ g C}_2\text{H}_4}{\text{theoretical yield}} \times 100\%$$

$$\text{theoretical yield C}_2\text{H}_4 = 1.132 \times 10^3 \text{ g C}_2\text{H}_4$$

The mass of hexane that must be reacted is:

$$(1.132 \times 10^3 \text{ g C}_2\text{H}_4) \times \frac{1 \text{ mol C}_2\text{H}_4}{28.052 \text{ g C}_2\text{H}_4} \times \frac{1 \text{ mol C}_6\text{H}_{14}}{1 \text{ mol C}_2\text{H}_4} \times \frac{86.172 \text{ g C}_6\text{H}_{14}}{1 \text{ mol C}_6\text{H}_{14}} = \mathbf{3.48 \times 10^3 \text{ g C}_6\text{H}_{14}}$$

- 3.93** This is a limiting reagent problem. Let's calculate the moles of Li_3N produced assuming complete reaction for each reactant.



$$12.3 \text{ g Li} \times \frac{1 \text{ mol Li}}{6.941 \text{ g Li}} \times \frac{2 \text{ mol Li}_3\text{N}}{6 \text{ mol Li}} = 0.5907 \text{ mol Li}_3\text{N}$$

$$33.6 \text{ g N}_2 \times \frac{1 \text{ mol N}_2}{28.02 \text{ g N}_2} \times \frac{2 \text{ mol Li}_3\text{N}}{1 \text{ mol N}_2} = 2.398 \text{ mol Li}_3\text{N}$$

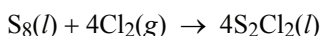
Li is the limiting reagent; it limits the amount of product produced. The amount of product produced is 0.5907 mole Li_3N . Let's convert this to grams.

$$? \text{ g Li}_3\text{N} = 0.5907 \text{ mol Li}_3\text{N} \times \frac{34.833 \text{ g Li}_3\text{N}}{1 \text{ mol Li}_3\text{N}} = \mathbf{20.6 \text{ g Li}_3\text{N}}$$

This is the theoretical yield of Li_3N . The actual yield is given in the problem (5.89 g). The percent yield is:

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{5.89 \text{ g}}{20.6 \text{ g}} \times 100\% = \mathbf{28.6\%}$$

- 3.94** This is a limiting reagent problem. Let's calculate the moles of S_2Cl_2 produced assuming complete reaction for each reactant.



$$4.06 \text{ g S}_8 \times \frac{1 \text{ mol S}_8}{256.56 \text{ g S}_8} \times \frac{4 \text{ mol S}_2\text{Cl}_2}{1 \text{ mol S}_8} = 0.0633 \text{ mol S}_2\text{Cl}_2$$

$$6.24 \text{ g Cl}_2 \times \frac{1 \text{ mol Cl}_2}{70.90 \text{ g Cl}_2} \times \frac{4 \text{ mol S}_2\text{Cl}_2}{4 \text{ mol Cl}_2} = 0.0880 \text{ mol S}_2\text{Cl}_2$$

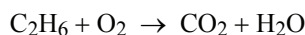
S₈ is the limiting reagent; it limits the amount of product produced. The amount of product produced is 0.0633 mole S₂Cl₂. Let's convert this to grams.

$$? \text{ g S}_2\text{Cl}_2 = 0.0633 \text{ mol S}_2\text{Cl}_2 \times \frac{135.04 \text{ g S}_2\text{Cl}_2}{1 \text{ mol S}_2\text{Cl}_2} = \mathbf{8.55 \text{ g S}_2\text{Cl}_2}$$

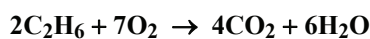
This is the theoretical yield of S₂Cl₂. The actual yield is given in the problem (6.55 g). The percent yield is:

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{6.55 \text{ g}}{8.55 \text{ g}} \times 100\% = \mathbf{76.6\%}$$

- 3.95** All the carbon from the hydrocarbon reactant ends up in CO₂, and all the hydrogen from the hydrocarbon reactant ends up in water. In the diagram, we find 4 CO₂ molecules and 6 H₂O molecules. This gives a ratio between carbon and hydrogen of 4:12. We write the formula C₄H₁₂, which reduces to the empirical formula CH₃. The empirical molar mass equals approximately 15 g, which is half the molar mass of the hydrocarbon. Thus, the molecular formula is double the empirical formula or C₂H₆. Since this is a combustion reaction, the other reactant is O₂. We write:



Balancing the equation,



- 3.96** $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g})$

We start with 8 molecules of H₂ and 3 molecules of O₂. The balanced equation shows 2 moles H₂ ≅ 1 mole O₂. If 3 molecules of O₂ react, 6 molecules of H₂ will react, leaving 2 molecules of H₂ in excess. The balanced equation also shows 1 mole O₂ ≅ 2 moles H₂O. If 3 molecules of O₂ react, 6 molecules of H₂O will be produced.

After complete reaction, there will be **2 molecules of H₂** and **6 molecules of H₂O**. The correct diagram is choice **(b)**.

- 3.97** First, let's convert to moles of HNO₃ produced.

$$1.00 \text{ ton HNO}_3 \times \frac{2000 \text{ lb}}{1 \text{ ton}} \times \frac{453.6 \text{ g}}{1 \text{ lb}} \times \frac{1 \text{ mol HNO}_3}{63.018 \text{ g HNO}_3} = 1.44 \times 10^4 \text{ mol HNO}_3$$

Now, we will work in the reverse direction to calculate the amount of reactant needed to produce 1.44×10^3 mol of HNO₃. Realize that since the problem says to assume an 80% yield for each step, the amount of reactant needed in each step will be **larger** by a factor of $\frac{100\%}{80\%}$, compared to a standard stoichiometry calculation where a 100% yield is assumed.

Referring to the balanced equation in the *last step*, we calculate the moles of NO₂.

$$(1.44 \times 10^4 \text{ mol HNO}_3) \times \frac{2 \text{ mol NO}_2}{1 \text{ mol HNO}_3} \times \frac{100\%}{80\%} = 3.60 \times 10^4 \text{ mol NO}_2$$

Now, let's calculate the amount of NO needed to produce 3.60×10^4 mol NO₂. Following the same procedure as above, and referring to the balanced equation in the *middle step*, we calculate the moles of NO.

$$(3.60 \times 10^4 \text{ mol NO}_2) \times \frac{1 \text{ mol NO}}{1 \text{ mol NO}_2} \times \frac{100\%}{80\%} = 4.50 \times 10^4 \text{ mol NO}$$

Now, let's calculate the amount of NH₃ needed to produce 4.5×10^4 mol NO. Referring to the balanced equation in the *first step*, the moles of NH₃ is:

$$(4.50 \times 10^4 \text{ mol NO}) \times \frac{4 \text{ mol NH}_3}{4 \text{ mol NO}} \times \frac{100\%}{80\%} = 5.625 \times 10^4 \text{ mol NH}_3$$

Finally, converting to grams of NH₃:

$$5.625 \times 10^4 \text{ mol NH}_3 \times \frac{17.034 \text{ g NH}_3}{1 \text{ mol NH}_3} = \mathbf{9.58 \times 10^5 \text{ g NH}_3}$$

- 3.98** We assume that all the Cl in the compound ends up as HCl and all the O ends up as H₂O. Therefore, we need to find the number of moles of Cl in HCl and the number of moles of O in H₂O.

$$\text{mol Cl} = 0.233 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.458 \text{ g HCl}} \times \frac{1 \text{ mol Cl}}{1 \text{ mol HCl}} = 0.006391 \text{ mol Cl}$$

$$\text{mol O} = 0.403 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.016 \text{ g H}_2\text{O}} \times \frac{1 \text{ mol O}}{1 \text{ mol H}_2\text{O}} = 0.02237 \text{ mol O}$$

Dividing by the smallest number of moles (0.006391 mole) gives the formula, ClO_{3.5}. Multiplying both subscripts by two gives the empirical formula, **Cl₂O₇**.

- 3.99** The number of moles of Y in 84.10 g of Y is:

$$27.22 \text{ g X} \times \frac{1 \text{ mol X}}{33.42 \text{ g X}} \times \frac{1 \text{ mol Y}}{1 \text{ mol X}} = 0.81448 \text{ mol Y}$$

The molar mass of Y is:

$$\text{molar mass Y} = \frac{84.10 \text{ g Y}}{0.81448 \text{ mol Y}} = 103.3 \text{ g/mol}$$

The atomic mass of Y is **103.3 amu**.

- 3.100** The symbol "O" refers to moles of oxygen atoms, not oxygen molecule (O₂). Look at the molecular formulas given in parts (a) and (b). What do they tell you about the relative amounts of carbon and oxygen?

(a) $0.212 \text{ mol C} \times \frac{1 \text{ mol O}}{1 \text{ mol C}} = \mathbf{0.212 \text{ mol O}}$

(b) $0.212 \text{ mol C} \times \frac{2 \text{ mol O}}{1 \text{ mol C}} = \mathbf{0.424 \text{ mol O}}$

3.101 The observations mean either that the amount of the more abundant isotope was increasing or the amount of the less abundant isotope was decreasing. One possible explanation is that the less abundant isotope was undergoing radioactive decay, and thus its mass would decrease with time.

3.102 This is a calculation involving percent composition. Remember,

$$\text{percent by mass of each element} = \frac{\text{mass of element in 1 mol of compound}}{\text{molar mass of compound}} \times 100\%$$

The molar masses are: Al, 26.98 g/mol; $\text{Al}_2(\text{SO}_4)_3$, 342.17 g/mol; H_2O , 18.016 g/mol. Thus, using x as the number of H_2O molecules,

$$\text{mass \% Al} = \left(\frac{2(\text{molar mass of Al})}{\text{molar mass of Al}_2(\text{SO}_4)_3 + x(\text{molar mass of H}_2\text{O})} \right) \times 100\%$$

$$8.20\% = \left(\frac{2(26.98 \text{ g})}{342.17 \text{ g} + x(18.016 \text{ g})} \right) \times 100\%$$

$$x = 17.53$$

Rounding off to a whole number of water molecules, $x = 18$. Therefore, the formula is $\text{Al}_2(\text{SO}_4)_3 \cdot 18 \text{ H}_2\text{O}$.

3.103 Molar mass of $\text{C}_4\text{H}_8\text{Cl}_2\text{S} = 4(12.01 \text{ g}) + 8(1.008 \text{ g}) + 2(35.45 \text{ g}) + 32.07 \text{ g} = 159.07 \text{ g}$

$$\% \text{C} = \frac{4(12.01 \text{ g/mol})}{159.07 \text{ g/mol}} \times 100\% = 30.20\%$$

$$\% \text{H} = \frac{8(1.008 \text{ g/mol})}{159.07 \text{ g/mol}} \times 100\% = 5.069\%$$

$$\% \text{Cl} = \frac{2(35.45 \text{ g/mol})}{159.07 \text{ g/mol}} \times 100\% = 44.57\%$$

$$\% \text{S} = \frac{32.07 \text{ g/mol}}{159.07 \text{ g/mol}} \times 100\% = 20.16\%$$

3.104 The number of carbon atoms in a 24-carat diamond is:

$$24 \text{ carat} \times \frac{200 \text{ mg C}}{1 \text{ carat}} \times \frac{0.001 \text{ g C}}{1 \text{ mg C}} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} \times \frac{6.022 \times 10^{23} \text{ atoms C}}{1 \text{ mol C}} = 2.4 \times 10^{23} \text{ atoms C}$$

3.105 The amount of Fe that reacted is: $\frac{1}{8} \times 664 \text{ g} = 83.0 \text{ g}$ reacted

The amount of Fe remaining is: $664 \text{ g} - 83.0 \text{ g} = 581 \text{ g remaining}$

Thus, 83.0 g of Fe reacts to form the compound Fe_2O_3 , which has two moles of Fe atoms per 1 mole of compound. The mass of Fe_2O_3 produced is:

$$83.0 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} \times \frac{1 \text{ mol Fe}_2\text{O}_3}{2 \text{ mol Fe}} \times \frac{159.7 \text{ g Fe}_2\text{O}_3}{1 \text{ mol Fe}_2\text{O}_3} = 119 \text{ g Fe}_2\text{O}_3$$

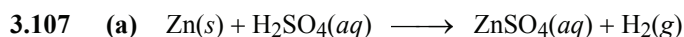
The final mass of the iron bar and rust is: $581 \text{ g Fe} + 119 \text{ g Fe}_2\text{O}_3 = 700 \text{ g}$

- 3.106** The mass of oxygen in MO is $39.46 \text{ g} - 31.70 \text{ g} = 7.76 \text{ g O}$. Therefore, for every 31.70 g of M, there is 7.76 g of O in the compound MO. The molecular formula shows a mole ratio of 1 mole M : 1 mole O. First, calculate moles of M that react with 7.76 g O .

$$\text{mol M} = 7.76 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} \times \frac{1 \text{ mol M}}{1 \text{ mol O}} = 0.485 \text{ mol M}$$

$$\text{molar mass M} = \frac{31.70 \text{ g M}}{0.485 \text{ mol M}} = 65.4 \text{ g/mol}$$

Thus, the atomic mass of M is **65.4 amu**. The metal is most likely **Zn**.



- (b)** We assume that a pure sample would produce the theoretical yield of H_2 . The balanced equation shows a mole ratio of 1 mole H_2 : 1 mole Zn. The theoretical yield of H_2 is:

$$3.86 \text{ g Zn} \times \frac{1 \text{ mol Zn}}{65.39 \text{ g Zn}} \times \frac{1 \text{ mol H}_2}{1 \text{ mol Zn}} \times \frac{2.016 \text{ g H}_2}{1 \text{ mol H}_2} = 0.119 \text{ g H}_2$$

$$\text{percent purity} = \frac{0.0764 \text{ g H}_2}{0.119 \text{ g H}_2} \times 100\% = \mathbf{64.2\%}$$

- (c)** We assume that the impurities are inert and do not react with the sulfuric acid to produce hydrogen.

- 3.108** The wording of the problem suggests that the actual yield is less than the theoretical yield. The percent yield will be equal to the percent purity of the iron(III) oxide. We find the theoretical yield :

$$\begin{aligned} (2.62 \times 10^3 \text{ kg Fe}_2\text{O}_3) \times \frac{1000 \text{ g Fe}_2\text{O}_3}{1 \text{ kg Fe}_2\text{O}_3} \times \frac{1 \text{ mol Fe}_2\text{O}_3}{159.7 \text{ g Fe}_2\text{O}_3} \times \frac{2 \text{ mol Fe}}{1 \text{ mol Fe}_2\text{O}_3} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} \times \frac{1 \text{ kg Fe}}{1000 \text{ g Fe}} \\ = 1.833 \times 10^3 \text{ kg Fe} \end{aligned}$$

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

$$\text{percent yield} = \frac{1.64 \times 10^3 \text{ kg Fe}}{1.833 \times 10^3 \text{ kg Fe}} \times 100\% = \mathbf{89.5\% = \text{purity of Fe}_2\text{O}_3}$$

- 3.109** The balanced equation is: $\text{C}_6\text{H}_{12}\text{O}_6 + 6\text{O}_2 \longrightarrow 6\text{CO}_2 + 6\text{H}_2\text{O}$

$$\begin{aligned} \frac{5.0 \times 10^2 \text{ g glucose}}{1 \text{ day}} \times \frac{1 \text{ mol glucose}}{180.16 \text{ g glucose}} \times \frac{6 \text{ mol CO}_2}{1 \text{ mol glucose}} \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} \times \frac{365 \text{ days}}{1 \text{ yr}} \times (6.5 \times 10^9 \text{ people}) \\ = \mathbf{1.7 \times 10^{15} \text{ g CO}_2/\text{yr}} \end{aligned}$$

- 3.110** The carbohydrate contains 40 percent carbon; therefore, the remaining 60 percent is hydrogen and oxygen. The problem states that the hydrogen to oxygen ratio is 2:1. We can write this 2:1 ratio as H₂O.

Assume 100 g of compound.

$$40.0 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 3.331 \text{ mol C}$$

$$60.0 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.016 \text{ g H}_2\text{O}} = 3.330 \text{ mol H}_2\text{O}$$

Dividing by 3.330 gives **CH₂O** for the empirical formula.

To find the molecular formula, divide the molar mass by the empirical mass.

$$\frac{\text{molar mass}}{\text{empirical mass}} = \frac{178 \text{ g}}{30.026 \text{ g}} \approx 6$$

Thus, there are six CH₂O units in each molecule of the compound, so the molecular formula is (CH₂O)₆, or **C₆H₁₂O₆**.

- 3.111** The molar mass of chlorophyll is 893.48 g/mol. Finding the mass of a 0.0011-mol sample:

$$0.0011 \text{ mol chlorophyll} \times \frac{893.48 \text{ g chlorophyll}}{1 \text{ mol chlorophyll}} = 0.98 \text{ g chlorophyll}$$

The chlorophyll sample has the greater mass.

- 3.112** If we assume 100 g of compound, the masses of Cl and X are 67.2 g and 32.8 g, respectively. We can calculate the moles of Cl.

$$67.2 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} = 1.896 \text{ mol Cl}$$

Then, using the mole ratio from the chemical formula (XCl₃), we can calculate the moles of X contained in 32.8 g.

$$1.896 \text{ mol Cl} \times \frac{1 \text{ mol X}}{3 \text{ mol Cl}} = 0.6320 \text{ mol X}$$

0.6320 mole of X has a mass of 32.8 g. Calculating the molar mass of X:

$$\frac{32.8 \text{ g X}}{0.6320 \text{ mol X}} = \mathbf{51.9 \text{ g/mol}}$$

The element is most likely **chromium** (molar mass = 52.00 g/mol).

- 3.113 (a)** The molar mass of hemoglobin is:

$$\begin{aligned} &2952(12.01 \text{ g}) + 4664(1.008 \text{ g}) + 812(14.01 \text{ g}) + 832(16.00 \text{ g}) + 8(32.07 \text{ g}) + 4(55.85 \text{ g}) \\ &= \mathbf{6.532 \times 10^4 \text{ g}} \end{aligned}$$

- (b)** To solve this problem, the following conversions need to be completed:

L → mL → red blood cells → hemoglobin molecules → mol hemoglobin → mass hemoglobin

We will use the following abbreviations: RBC = red blood cells, HG = hemoglobin

$$5.00 \text{ L} \times \frac{1 \text{ mL}}{0.001 \text{ L}} \times \frac{5.0 \times 10^9 \text{ RBC}}{1 \text{ mL}} \times \frac{2.8 \times 10^8 \text{ HG molecules}}{1 \text{ RBC}} \times \frac{1 \text{ mol HG}}{6.022 \times 10^{23} \text{ molecules HG}} \times \frac{6.532 \times 10^4 \text{ g HG}}{1 \text{ mol HG}}$$

$$= 7.6 \times 10^2 \text{ g HG}$$

3.114 A 100 g sample of myoglobin contains 0.34 g of iron (0.34% Fe). The number of moles of Fe is:

$$0.34 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} = 6.09 \times 10^{-3} \text{ mol Fe}$$

Since there is one Fe atom in a molecule of myoglobin, the moles of myoglobin also equal 6.09×10^{-3} mole. The molar mass of myoglobin can be calculated.

$$\text{molar mass myoglobin} = \frac{100 \text{ g myoglobin}}{6.09 \times 10^{-3} \text{ mol myoglobin}} = 1.6 \times 10^4 \text{ g/mol}$$

3.115 (a) $8.38 \text{ g KBr} \times \frac{1 \text{ mol KBr}}{119.0 \text{ g KBr}} \times \frac{6.022 \times 10^{23} \text{ KBr}}{1 \text{ mol KBr}} \times \frac{1 \text{ K}^+ \text{ ion}}{1 \text{ KBr}} = 4.24 \times 10^{22} \text{ K}^+ \text{ ions}$

Since there is one Br^- for every one K^+ , the number of Br^- ions = $4.24 \times 10^{22} \text{ Br}^-$ ions

(b) $5.40 \text{ g Na}_2\text{SO}_4 \times \frac{1 \text{ mol Na}_2\text{SO}_4}{142.05 \text{ g Na}_2\text{SO}_4} \times \frac{6.022 \times 10^{23} \text{ Na}_2\text{SO}_4}{1 \text{ mol Na}_2\text{SO}_4} \times \frac{2 \text{ Na}^+ \text{ ions}}{1 \text{ Na}_2\text{SO}_4} = 4.58 \times 10^{22} \text{ Na}^+ \text{ ions}$

Since there are two Na^+ for every one SO_4^{2-} , the number of SO_4^{2-} ions = $2.29 \times 10^{22} \text{ SO}_4^{2-}$ ions

(c) $7.45 \text{ g Ca}_3(\text{PO}_4)_2 \times \frac{1 \text{ mol Ca}_3(\text{PO}_4)_2}{310.18 \text{ g Ca}_3(\text{PO}_4)_2} \times \frac{6.022 \times 10^{23} \text{ Ca}_3(\text{PO}_4)_2}{1 \text{ mol Ca}_3(\text{PO}_4)_2} \times \frac{3 \text{ Ca}^{2+} \text{ ions}}{1 \text{ Ca}_3(\text{PO}_4)_2}$

$$= 4.34 \times 10^{22} \text{ Ca}^{2+} \text{ ions}$$

Since there are three Ca^{2+} for every two PO_4^{3-} , the number of PO_4^{3-} ions is:

$$4.34 \times 10^{22} \text{ Ca}^{2+} \text{ ions} \times \frac{2 \text{ PO}_4^{3-} \text{ ions}}{3 \text{ Ca}^{2+} \text{ ions}} = 2.89 \times 10^{22} \text{ PO}_4^{3-} \text{ ions}$$

3.116 If we assume 100 g of the mixture, then there are 29.96 g of Na in the mixture (29.96% Na by mass). This amount of Na is equal to the mass of Na in NaBr plus the mass of Na in Na_2SO_4 .

$$29.96 \text{ g Na} = \text{mass of Na in NaBr} + \text{mass of Na in Na}_2\text{SO}_4$$

To calculate the mass of Na in each compound, grams of compound need to be converted to grams of Na using the mass percentage of Na in the compound. If x equals the mass of NaBr, then the mass of Na_2SO_4 is $100 - x$. Recall that we assumed 100 g of the mixture. We set up the following expression and solve for x .

$$29.96 \text{ g Na} = \text{mass of Na in NaBr} + \text{mass of Na in Na}_2\text{SO}_4$$

$$29.96 \text{ g Na} = \left[x \text{ g NaBr} \times \frac{22.99 \text{ g Na}}{102.89 \text{ g NaBr}} \right] + \left[(100 - x) \text{ g Na}_2\text{SO}_4 \times \frac{(2)(22.99 \text{ g Na})}{142.05 \text{ g Na}_2\text{SO}_4} \right]$$

$$29.96 = 0.22344x + 32.369 - 0.32369x$$

$$0.10025x = 2.409$$

$$x = 24.03 \text{ g, which equals the mass of NaBr.}$$

The mass of Na_2SO_4 is $100 - x$ which equals 75.97 g.

Because we assumed 100 g of compound, the mass % of NaBr in the mixture is **24.03%** and the mass % of Na_2SO_4 is **75.97%**.

$$3.117 \quad (\text{a}) \quad 0.400 \text{ g aspirin} \times \frac{1 \text{ mol aspirin}}{180.15 \text{ g aspirin}} \times \frac{1 \text{ mol salicylic acid}}{1 \text{ mol aspirin}} \times \frac{138.12 \text{ g salicylic acid}}{1 \text{ mol salicylic acid}} = \mathbf{0.307 \text{ g salicylic acid}}$$

$$(\text{b}) \quad 0.307 \text{ g salicylic acid} \times \frac{1}{0.749} = \mathbf{0.410 \text{ g salicylic acid}}$$

If you have trouble deciding whether to multiply or divide by 0.749 in the calculation, remember that if only 74.9% of salicylic acid is converted to aspirin, a larger amount of salicylic acid will need to be reacted to yield the same amount of aspirin.

$$(\text{c}) \quad 9.26 \text{ g salicylic acid} \times \frac{1 \text{ mol salicylic acid}}{138.12 \text{ g salicylic acid}} \times \frac{1 \text{ mol aspirin}}{1 \text{ mol salicylic acid}} = 0.06704 \text{ mol aspirin}$$

$$8.54 \text{ g acetic anhydride} \times \frac{1 \text{ mol acetic anhydride}}{102.09 \text{ g acetic anhydride}} \times \frac{1 \text{ mol aspirin}}{1 \text{ mol acetic anhydride}} = 0.08365 \text{ mol aspirin}$$

The limiting reagent is salicylic acid. The theoretical yield of aspirin is:

$$0.06704 \text{ mol aspirin} \times \frac{180.15 \text{ g aspirin}}{1 \text{ mol aspirin}} = \mathbf{12.1 \text{ g aspirin}}$$

The percent yield is:

$$\% \text{ yield} = \frac{10.9 \text{ g}}{12.1 \text{ g}} \times 100\% = \mathbf{90.1\%}$$

3.118 The mass percent of an element in a compound can be calculated as follows:

$$\text{percent by mass of each element} = \frac{\text{mass of element in 1 mol of compound}}{\text{molar mass of compound}} \times 100\%$$

The molar mass of $\text{Ca}_3(\text{PO}_4)_2 = 310.18 \text{ g/mol}$

$$\% \text{ Ca} = \frac{(3)(40.08 \text{ g})}{310.18 \text{ g}} \times 100\% = \mathbf{38.76\% \text{ Ca}}$$

$$\% \text{ P} = \frac{(2)(30.97 \text{ g})}{310.18 \text{ g}} \times 100\% = \mathbf{19.97\% \text{ P}}$$

$$\% \text{ O} = \frac{(8)(16.00 \text{ g})}{310.18 \text{ g}} \times 100\% = \mathbf{41.27\% \text{ O}}$$

- 3.119 (a)** First, calculate the mass of C in CO₂, the mass of H in H₂O, and the mass of N in NH₃. For now, we will carry more than 3 significant figures and then round to the correct number at the end of the problem.

$$? \text{ g C} = 3.94 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 1.075 \text{ g C}$$

$$? \text{ g H} = 1.89 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 0.2114 \text{ g H}$$

$$? \text{ g N} = 0.436 \text{ g NH}_3 \times \frac{1 \text{ mol NH}_3}{17.03 \text{ g NH}_3} \times \frac{1 \text{ mol N}}{1 \text{ mol NH}_3} \times \frac{14.01 \text{ g N}}{1 \text{ mol N}} = 0.3587 \text{ g N}$$

Next, we can calculate the %C, %H, and the %N in each sample, then we can calculate the %O by difference.

$$\% \text{C} = \frac{1.075 \text{ g C}}{2.175 \text{ g sample}} \times 100\% = 49.43\% \text{ C}$$

$$\% \text{H} = \frac{0.2114 \text{ g H}}{2.175 \text{ g sample}} \times 100\% = 9.720\% \text{ H}$$

$$\% \text{N} = \frac{0.3587 \text{ g N}}{1.873 \text{ g sample}} \times 100\% = 19.15\% \text{ N}$$

The % O = 100% - (49.43% + 9.720% + 19.15%) = 21.70% O

Assuming 100 g of compound, calculate the moles of each element.

$$? \text{ mol C} = 49.43 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 4.116 \text{ mol C}$$

$$? \text{ mol H} = 9.720 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 9.643 \text{ mol H}$$

$$? \text{ mol N} = 19.15 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 1.367 \text{ mol N}$$

$$? \text{ mol O} = 21.70 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 1.356 \text{ mol O}$$

Thus, we arrive at the formula C_{4.116}H_{9.643}N_{1.367}O_{1.356}. Dividing by 1.356 gives the empirical formula, **C₃H₇NO**.

- (b)** The empirical molar mass is 73.10 g. Since the approximate molar mass of lysine is 150 g, we have:

$$\frac{150 \text{ g}}{73.10 \text{ g}} \approx 2$$

Therefore, the molecular formula is (C₃H₇NO)₂ or **C₆H₁₄N₂O₂**.

- 3.120 Yes.** The number of hydrogen atoms in one gram of hydrogen molecules is the same as the number in one gram of hydrogen atoms. There is no difference in mass, only in the way that the particles are arranged.

Would the mass of 100 dimes be the same if they were stuck together in pairs instead of separated?

- 3.121** The mass of one fluorine atom is 19.00 amu. The mass of one mole of fluorine atoms is 19.00 g. Multiplying the mass of one atom by Avogadro's number gives the mass of one mole of atoms. We can write:

$$\frac{19.00 \text{ amu}}{1 \text{ F atom}} \times (6.022 \times 10^{23} \text{ F atoms}) = 19.00 \text{ g F}$$

or,

$$6.022 \times 10^{23} \text{ amu} = 1 \text{ g}$$

This is why Avogadro's numbers has sometimes been described as a conversion factor between amu and grams.

- 3.122** Since we assume that water exists as either H₂O or D₂O, the natural abundances are 99.985 percent and 0.015 percent, respectively. If we convert to molecules of water (both H₂O or D₂O), we can calculate the molecules that are D₂O from the natural abundance (0.015%).

The necessary conversions are:

mL water → g water → mol water → molecules water → molecules D₂O

$$\begin{aligned} 400 \text{ mL water} &\times \frac{1 \text{ g water}}{1 \text{ mL water}} \times \frac{1 \text{ mol water}}{18.02 \text{ g water}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol water}} \times \frac{0.015\% \text{ molecules D}_2\text{O}}{100\% \text{ molecules water}} \\ &= \mathbf{2.01 \times 10^{21} \text{ molecules D}_2\text{O}} \end{aligned}$$

- 3.123** There can only be one chlorine per molecule, since two chlorines have a combined mass in excess of 70 amu. Since the ³⁵Cl isotope is more abundant, let's subtract 35 amu from the mass corresponding to the more intense peak.

$$50 \text{ amu} - 35 \text{ amu} = 15 \text{ amu}$$

15 amu equals the mass of one ¹²C and three ¹H. To explain the two peaks, we have:

$$\text{molecular mass } ^{12}\text{C}^1\text{H}_3^{35}\text{Cl} = 12 \text{ amu} + 3(1 \text{ amu}) + 35 \text{ amu} = 50 \text{ amu}$$

$$\text{molecular mass } ^{12}\text{C}^1\text{H}_3^{37}\text{Cl} = 12 \text{ amu} + 3(1 \text{ amu}) + 37 \text{ amu} = 52 \text{ amu}$$

³⁵Cl is three times more abundant than ³⁷Cl; therefore, the 50 amu peak will be three times more intense than the 52 amu peak.

- 3.124** First, we can calculate the moles of oxygen.

$$2.445 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} \times \frac{1 \text{ mol O}}{1 \text{ mol C}} = 0.2036 \text{ mol O}$$

Next, we can calculate the molar mass of oxygen.

$$\text{molar mass O} = \frac{3.257 \text{ g O}}{0.2036 \text{ mol O}} = 16.00 \text{ g/mol}$$

If 1 mole of oxygen atoms has a mass of 16.00 g, then 1 atom of oxygen has an **atomic mass of 16.00 amu**.

- 3.125** The molecular formula for Cl_2O_7 means that there are 2 Cl atoms for every 7 O atoms or 2 moles of Cl atoms for every 7 moles of O atoms. We can write:

$$\text{mole ratio} = \frac{2 \text{ mol Cl}}{7 \text{ mol O}} = \frac{1 \text{ mol Cl}_2}{3.5 \text{ mol O}_2}$$

- 3.126 (a)** The mass of chlorine is **5.0 g**.
- (b)** From the percent by mass of Cl, we can calculate the mass of chlorine in 60.0 g of NaClO_3 .

$$\text{mass \% Cl} = \frac{35.45 \text{ g Cl}}{106.44 \text{ g compound}} \times 100\% = 33.31\% \text{ Cl}$$

$$\text{mass Cl} = 60.0 \text{ g} \times 0.3331 = \mathbf{20.0 \text{ g Cl}}$$

- (c)** 0.10 mol of KCl contains 0.10 mol of Cl.

$$0.10 \text{ mol Cl} \times \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}} = \mathbf{3.5 \text{ g Cl}}$$

- (d)** From the percent by mass of Cl, we can calculate the mass of chlorine in 30.0 g of MgCl_2 .

$$\text{mass \% Cl} = \frac{(2)(35.45 \text{ g Cl})}{95.21 \text{ g compound}} \times 100\% = 74.47\% \text{ Cl}$$

$$\text{mass Cl} = 30.0 \text{ g} \times 0.7447 = \mathbf{22.3 \text{ g Cl}}$$

- (e)** The mass of Cl can be calculated from the molar mass of Cl_2 .

$$0.50 \text{ mol Cl}_2 \times \frac{70.90 \text{ g Cl}}{1 \text{ mol Cl}_2} = \mathbf{35.45 \text{ g Cl}}$$

Thus, **(e) 0.50 mol Cl_2** contains the greatest mass of chlorine.

- 3.127** The mass percent of Cl is given. From the mass of the compound and the number of hydrogen atoms given, we can calculate the mass percent of H. The mass percent of carbon is then obtained by difference. Once the mass percentages of each element are known, the empirical formula can be determined.

$$4.19 \times 10^{23} \text{ H atoms} \times \frac{1 \text{ mol H}}{6.022 \times 10^{23} \text{ H atoms}} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 0.7013 \text{ g H}$$

$$\text{mass \% H} = \frac{0.7013 \text{ g H}}{9.00 \text{ g compound}} \times 100\% = 7.792\% \text{ H}$$

$$\text{mass \% C} = 100\% - (55.0\% + 7.792\%) = 37.21\% \text{ C}$$

To determine the empirical formula, assume 100 g of compound and convert to moles of each element present.

$$\text{mol C} = 37.21 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 3.098 \text{ mol C}$$

$$\text{mol H} = 7.792 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 7.730 \text{ mol H}$$

$$\text{mol Cl} = 55.0 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} = 1.551 \text{ mol Cl}$$

Thus, we arrive at the formula $\text{C}_{3.098}\text{H}_{7.730}\text{Cl}_{1.551}$, which gives the identity and the mole ratios of atoms present. However, chemical formulas are written with whole numbers. Try to convert to whole numbers by dividing each of the subscripts by the smallest subscript (1.551). This gives the empirical formula $\text{C}_2\text{H}_5\text{Cl}$.

- 3.128** Both compounds contain only Pt and Cl. The percent by mass of Pt can be calculated by subtracting the percent Cl from 100 percent.

Compound A: Assume 100 g of compound.

$$26.7 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} = 0.753 \text{ mol Cl}$$

$$73.3 \text{ g Pt} \times \frac{1 \text{ mol Pt}}{195.1 \text{ g Pt}} = 0.376 \text{ mol Pt}$$

Dividing by the smallest number of moles (0.376 mole) gives the empirical formula, PtCl_2 .

Compound B: Assume 100 g of compound.

$$42.1 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} = 1.19 \text{ mol Cl}$$

$$57.9 \text{ g Pt} \times \frac{1 \text{ mol Pt}}{195.1 \text{ g Pt}} = 0.297 \text{ mol Pt}$$

Dividing by the smallest number of moles (0.297 mole) gives the empirical formula, PtCl_4 .

- 3.129** The mass of the metal (X) in the metal oxide is 1.68 g. The mass of oxygen in the metal oxide is $2.40 \text{ g} - 1.68 \text{ g} = 0.72 \text{ g}$ oxygen. Next, find the number of moles of the metal and of the oxygen.

$$\text{moles X} = 1.68 \text{ g} \times \frac{1 \text{ mol X}}{55.9 \text{ g X}} = 0.0301 \text{ mol X}$$

$$\text{moles O} = 0.72 \text{ g} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.045 \text{ mol O}$$

This gives the formula $\text{X}_{0.0301}\text{O}_{0.045}$. Dividing by the smallest number of moles (0.0301 moles) gives the formula $\text{X}_{1.00}\text{O}_{1.5}$. Multiplying by two gives the empirical formula, X_2O_3 .

The balanced equation is: $\text{X}_2\text{O}_3(\text{s}) + 3\text{CO}(\text{g}) \longrightarrow 2\text{X}(\text{s}) + 3\text{CO}_2(\text{g})$

- 3.130** Both compounds contain only Mn and O. When the first compound is heated, oxygen gas is evolved. Let's calculate the empirical formulas for the two compounds, then we can write a balanced equation.

(a) Compound X: Assume 100 g of compound.

$$63.3 \text{ g Mn} \times \frac{1 \text{ mol Mn}}{54.94 \text{ g Mn}} = 1.15 \text{ mol Mn}$$

$$36.7 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.29 \text{ mol O}$$

Dividing by the smallest number of moles (1.15 moles) gives the empirical formula, **MnO₂**.

Compound Y: Assume 100 g of compound.

$$72.0 \text{ g Mn} \times \frac{1 \text{ mol Mn}}{54.94 \text{ g Mn}} = 1.31 \text{ mol Mn}$$

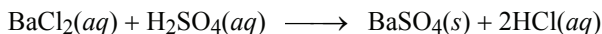
$$28.0 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 1.75 \text{ mol O}$$

Dividing by the smallest number of moles gives MnO_{1.33}. Recall that an empirical formula must have whole number coefficients. Multiplying by a factor of 3 gives the empirical formula **Mn₃O₄**.

(b) The unbalanced equation is: $\text{MnO}_2 \longrightarrow \text{Mn}_3\text{O}_4 + \text{O}_2$

Balancing by inspection gives: **$3\text{MnO}_2 \longrightarrow \text{Mn}_3\text{O}_4 + \text{O}_2$**

- 3.131** The mass of the water is the difference between 1.936 g of the hydrate and the mass of *water-free* (anhydrous) BaCl₂. First, we need to start with a balanced equation for the reaction. Upon treatment with sulfuric acid, BaCl₂ dissolves, losing its waters of hydration.



Next, calculate the mass of anhydrous BaCl₂ based on the amount of BaSO₄ produced.

$$1.864 \text{ g BaSO}_4 \times \frac{1 \text{ mol BaSO}_4}{233.4 \text{ g BaSO}_4} \times \frac{1 \text{ mol BaCl}_2}{1 \text{ mol BaSO}_4} \times \frac{208.2 \text{ g BaCl}_2}{1 \text{ mol BaCl}_2} = 1.663 \text{ g BaCl}_2$$

The mass of water is (1.936 g – 1.663 g) = 0.273 g H₂O. Next, we convert the mass of H₂O and the mass of BaCl₂ to moles to determine the formula of the hydrate.

$$0.273 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 0.0151 \text{ mol H}_2\text{O}$$

$$1.663 \text{ g BaCl}_2 \times \frac{1 \text{ mol BaCl}_2}{208.2 \text{ g BaCl}_2} = 0.00799 \text{ mol BaCl}_2$$

The ratio of the number of moles of H₂O to the number of moles of BaCl₂ is 0.0151/0.00799 = **1.89**. We round this number to **2**, which is the value of *x*. The formula of the hydrate is **BaCl₂ · 2H₂O**.

- 3.132** SO₂ is converted to H₂SO₄ by reaction with water. The mole ratio between SO₂ and H₂SO₄ is 1:1.

This is a unit conversion problem. You should come up with the following strategy to solve the problem.

tons SO₂ → ton-mol SO₂ → ton-mol H₂SO₄ → tons H₂SO₄

$$\begin{aligned} ? \text{ tons H}_2\text{SO}_4 &= (4.0 \times 10^5 \text{ tons SO}_2) \times \frac{1 \text{ ton-mol SO}_2}{64.07 \text{ tons SO}_2} \times \frac{1 \text{ ton-mol H}_2\text{SO}_4}{1 \text{ ton-mol SO}_2} \times \frac{98.09 \text{ tons H}_2\text{SO}_4}{1 \text{ ton-mol H}_2\text{SO}_4} \\ &= \mathbf{6.1 \times 10^5 \text{ tons H}_2\text{SO}_4} \end{aligned}$$

Tip: You probably won't come across a ton-mol that often in chemistry. However, it was convenient to use in this problem. We normally use a g-mol. 1 g-mol SO₂ has a mass of 64.07 g. In a similar manner, 1 ton-mol of SO₂ has a mass of 64.07 tons.

- 3.133** The mass of water lost upon heating the mixture is $(5.020 \text{ g} - 2.988 \text{ g}) = 2.032 \text{ g}$ water. Next, if we let $x = \text{mass of CuSO}_4 \cdot 5\text{H}_2\text{O}$, then the mass of $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ is $(5.020 - x)\text{g}$. We can calculate the amount of water lost by each salt based on the mass % of water in each hydrate. We can write:

$$(\text{mass CuSO}_4 \cdot 5\text{H}_2\text{O})(\% \text{H}_2\text{O}) + (\text{mass MgSO}_4 \cdot 7\text{H}_2\text{O})(\% \text{H}_2\text{O}) = \text{total mass H}_2\text{O} = 2.032 \text{ g H}_2\text{O}$$

Calculate the % H₂O in each hydrate.

$$\% \text{H}_2\text{O (CuSO}_4 \cdot 5\text{H}_2\text{O}) = \frac{(5)(18.02 \text{ g})}{249.7 \text{ g}} \times 100\% = 36.08\% \text{ H}_2\text{O}$$

$$\% \text{H}_2\text{O (MgSO}_4 \cdot 7\text{H}_2\text{O}) = \frac{(7)(18.02 \text{ g})}{246.5 \text{ g}} \times 100\% = 51.17\% \text{ H}_2\text{O}$$

Substituting into the equation above gives:

$$(x)(0.3608) + (5.020 - x)(0.5117) = 2.032 \text{ g}$$

$$0.1509x = 0.5367$$

$$x = 3.557 \text{ g} = \text{mass of CuSO}_4 \cdot 5\text{H}_2\text{O}$$

Finally, the percent by mass of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ in the mixture is:

$$\frac{3.557 \text{ g}}{5.020 \text{ g}} \times 100\% = \mathbf{70.86\%}$$

- 3.134** We assume that the increase in mass results from the element nitrogen. The mass of nitrogen is:

$$0.378 \text{ g} - 0.273 \text{ g} = 0.105 \text{ g N}$$

The empirical formula can now be calculated. Convert to moles of each element.

$$0.273 \text{ g Mg} \times \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}} = 0.0112 \text{ mol Mg}$$

$$0.105 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 0.00749 \text{ mol N}$$

Dividing by the smallest number of moles gives $\text{Mg}_{1.5}\text{N}$. Recall that an empirical formula must have whole number coefficients. Multiplying by a factor of 2 gives the empirical formula $\mathbf{Mg}_3\mathbf{N}_2$. The name of this compound is **magnesium nitride**.

- 3.135** The balanced equations are:



If we let $x = \text{mass of CH}_4$, then the mass of C_2H_6 is $(13.43 - x) \text{ g}$.

Next, we need to calculate the mass of CO₂ and the mass of H₂O produced by both CH₄ and C₂H₆. The sum of the masses of CO₂ and H₂O will add up to 64.84 g.

$$? \text{ g CO}_2 \text{ (from CH}_4\text{)} = x \text{ g CH}_4 \times \frac{1 \text{ mol CH}_4}{16.04 \text{ g CH}_4} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol CH}_4} \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = 2.744x \text{ g CO}_2$$

$$? \text{ g H}_2\text{O (from CH}_4\text{)} = x \text{ g CH}_4 \times \frac{1 \text{ mol CH}_4}{16.04 \text{ g CH}_4} \times \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol CH}_4} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 2.247x \text{ g H}_2\text{O}$$

$$\begin{aligned} ? \text{ g CO}_2 \text{ (from C}_2\text{H}_6\text{)} &= (13.43 - x) \text{ g C}_2\text{H}_6 \times \frac{1 \text{ mol C}_2\text{H}_6}{30.07 \text{ g C}_2\text{H}_6} \times \frac{4 \text{ mol CO}_2}{2 \text{ mol C}_2\text{H}_6} \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} \\ &= 2.927(13.43 - x) \text{ g CO}_2 \end{aligned}$$

$$\begin{aligned} ? \text{ g H}_2\text{O (from C}_2\text{H}_6\text{)} &= (13.43 - x) \text{ g C}_2\text{H}_6 \times \frac{1 \text{ mol C}_2\text{H}_6}{30.07 \text{ g C}_2\text{H}_6} \times \frac{6 \text{ mol H}_2\text{O}}{2 \text{ mol C}_2\text{H}_6} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \\ &= 1.798(13.43 - x) \text{ g H}_2\text{O} \end{aligned}$$

Summing the masses of CO₂ and H₂O:

$$2.744x \text{ g} + 2.247x \text{ g} + 2.927(13.43 - x) \text{ g} + 1.798(13.43 - x) \text{ g} = 64.84 \text{ g}$$

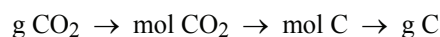
$$0.266x = 1.383$$

$$x = 5.20 \text{ g}$$

The fraction of CH₄ in the mixture is $\frac{5.20 \text{ g}}{13.43 \text{ g}} = \mathbf{0.387}$

- 3.136 Step 1:** Calculate the mass of C in 55.90 g CO₂, and the mass of H in 28.61 g H₂O. This is a dimensional analysis problem. To calculate the mass of each component, you need the molar masses and the correct mole ratio.

You should come up with the following strategy:



$$\text{Step 2: } ? \text{ g C} = 55.90 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 15.25 \text{ g C}$$

Similarly,

$$? \text{ g H} = 28.61 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 3.201 \text{ g H}$$

Since the compound contains C, H, and Pb, we can calculate the mass of Pb by difference.

$$51.36 \text{ g} = \text{mass C} + \text{mass H} + \text{mass Pb}$$

$$51.36 \text{ g} = 15.25 \text{ g} + 3.201 \text{ g} + \text{mass Pb}$$

$$\text{mass Pb} = 32.91 \text{ g Pb}$$

Step 3: Calculate the number of moles of each element present in the sample. Use molar mass as a conversion factor.

$$? \text{ mol C} = 15.25 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 1.270 \text{ mol C}$$

Similarly,

$$? \text{ mol H} = 3.201 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 3.176 \text{ mol H}$$

$$? \text{ mol Pb} = 32.91 \text{ g Pb} \times \frac{1 \text{ mol Pb}}{207.2 \text{ g Pb}} = 0.1588 \text{ mol Pb}$$

Thus, we arrive at the formula $\text{Pb}_{0.1588}\text{C}_{1.270}\text{H}_{3.176}$, which gives the identity and the ratios of atoms present. However, chemical formulas are written with whole numbers.

Step 4: Try to convert to whole numbers by dividing all the subscripts by the smallest subscript.

$$\text{Pb: } \frac{0.1588}{0.1588} = 1.00 \quad \text{C: } \frac{1.270}{0.1588} \approx 8 \quad \text{H: } \frac{3.176}{0.1588} \approx 20$$

This gives the empirical formula, **PbC₈H₂₀**.

3.137 First, calculate the mass of C in CO₂ and the mass of H in H₂O.

$$? \text{ g C} = 30.2 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 8.24 \text{ g C}$$

$$? \text{ g H} = 14.8 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 1.66 \text{ g H}$$

Since the compound contains C, H, and O, we can calculate the mass of O by difference.

$$12.1 \text{ g} = \text{mass C} + \text{mass H} + \text{mass O}$$

$$12.1 \text{ g} = 8.24 \text{ g} + 1.66 \text{ g} + \text{mass O}$$

$$\text{mass O} = 2.2 \text{ g O}$$

Next, calculate the moles of each element.

$$? \text{ mol C} = 8.24 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 0.686 \text{ mol C}$$

$$? \text{ mol H} = 1.66 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 1.65 \text{ mol H}$$

$$? \text{ mol O} = 2.2 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.14 \text{ mol O}$$

Thus, we arrive at the formula $\text{C}_{0.686}\text{H}_{1.65}\text{O}_{0.14}$. Dividing by 0.14 gives the empirical formula, **C₅H₁₂O**.

- 3.138 (a)** The following strategy can be used to convert from the volume of the Mg cube to the number of Mg atoms.

$\text{cm}^3 \rightarrow \text{grams} \rightarrow \text{moles} \rightarrow \text{atoms}$

$$1.0 \text{ cm}^3 \times \frac{1.74 \text{ g Mg}}{1 \text{ cm}^3} \times \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}} \times \frac{6.022 \times 10^{23} \text{ Mg atoms}}{1 \text{ mol Mg}} = \mathbf{4.3 \times 10^{22} \text{ Mg atoms}}$$

- (b)** Since 74 percent of the available space is taken up by Mg atoms, 4.3×10^{22} atoms occupy the following volume:

$$0.74 \times 1.0 \text{ cm}^3 = 0.74 \text{ cm}^3$$

We are trying to calculate the radius of a single Mg atom, so we need the volume occupied by a single Mg atom.

$$\text{volume Mg atom} = \frac{0.74 \text{ cm}^3}{4.3 \times 10^{22} \text{ Mg atoms}} = 1.7 \times 10^{-23} \text{ cm}^3/\text{Mg atom}$$

The volume of a sphere is $\frac{4}{3}\pi r^3$. Solving for the radius:

$$V = 1.7 \times 10^{-23} \text{ cm}^3 = \frac{4}{3}\pi r^3$$

$$r^3 = 4.1 \times 10^{-24} \text{ cm}^3$$

$$r = 1.6 \times 10^{-8} \text{ cm}$$

Converting to picometers:

$$\text{radius Mg atom} = (1.6 \times 10^{-8} \text{ cm}) \times \frac{0.01 \text{ m}}{1 \text{ cm}} \times \frac{1 \text{ pm}}{1 \times 10^{-12} \text{ m}} = \mathbf{1.6 \times 10^2 \text{ pm}}$$

- 3.139** The balanced equations for the combustion of sulfur and the reaction of SO_2 with CaO are:



First, find the amount of sulfur present in the daily coal consumption.

$$(6.60 \times 10^6 \text{ kg coal}) \times \frac{1.6\% \text{ S}}{100\%} = 1.06 \times 10^5 \text{ kg S} = 1.06 \times 10^8 \text{ g S}$$

The daily amount of CaO needed is:

$$(1.06 \times 10^8 \text{ g S}) \times \frac{1 \text{ mol S}}{32.07 \text{ g S}} \times \frac{1 \text{ mol SO}_2}{1 \text{ mol S}} \times \frac{1 \text{ mol CaO}}{1 \text{ mol SO}_2} \times \frac{56.08 \text{ g CaO}}{1 \text{ mol CaO}} \times \frac{1 \text{ kg}}{1000 \text{ g}} = \mathbf{1.85 \times 10^5 \text{ kg CaO}}$$

- 3.140** The molar mass of air can be calculated by multiplying the mass of each component by its abundance and adding them together. Recall that nitrogen gas and oxygen gas are diatomic.

$$\text{molar mass air} = (0.7808)(28.02 \text{ g/mol}) + (0.2095)(32.00 \text{ g/mol}) + (0.0097)(39.95 \text{ g/mol}) = \mathbf{28.97 \text{ g/mol}}$$

- 3.141 (a)** Assuming the die pack with no empty space between die, the volume of one mole of die is:

$$(1.5 \text{ cm})^3 \times (6.022 \times 10^{23}) = 2.0 \times 10^{24} \text{ cm}^3$$

- (b)** $6371 \text{ km} = 6.371 \times 10^8 \text{ cm}$

Volume = area \times height (h)

$$h = \frac{\text{volume}}{\text{area}} = \frac{2.0 \times 10^{24} \text{ cm}^3}{4\pi(6.371 \times 10^8 \text{ cm})^2} = 3.9 \times 10^5 \text{ cm} = 3.9 \times 10^3 \text{ m}$$

- 3.142** The surface area of the water can be calculated assuming that the dish is circular.

$$\text{surface area of water} = \pi r^2 = \pi(10 \text{ cm})^2 = 3.1 \times 10^2 \text{ cm}^2$$

The cross-sectional area of one stearic acid molecule in cm^2 is:

$$0.21 \text{ nm}^2 \times \left(\frac{1 \times 10^{-9} \text{ m}}{1 \text{ nm}} \right)^2 \times \left(\frac{1 \text{ cm}}{0.01 \text{ m}} \right)^2 = 2.1 \times 10^{-15} \text{ cm}^2/\text{molecule}$$

Assuming that there is no empty space between molecules, we can calculate the number of stearic acid molecules that will fit in an area of $3.1 \times 10^2 \text{ cm}^2$.

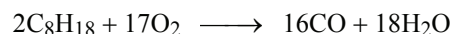
$$(3.1 \times 10^2 \text{ cm}^2) \times \frac{1 \text{ molecule}}{2.1 \times 10^{-15} \text{ cm}^2} = 1.5 \times 10^{17} \text{ molecules}$$

Next, we can calculate the moles of stearic acid in the $1.4 \times 10^{-4} \text{ g}$ sample. Then, we can calculate Avogadro's number (the number of molecules per mole).

$$1.4 \times 10^{-4} \text{ g stearic acid} \times \frac{1 \text{ mol stearic acid}}{284.5 \text{ g stearic acid}} = 4.9 \times 10^{-7} \text{ mol stearic acid}$$

$$\text{Avogadro's number } (N_A) = \frac{1.5 \times 10^{17} \text{ molecules}}{4.9 \times 10^{-7} \text{ mol}} = 3.1 \times 10^{23} \text{ molecules/mol}$$

- 3.143** The balanced equations for the combustion of octane are:



The quantity of octane burned is 2650 g (1 gallon with a density of 2.650 kg/gallon). Let x be the mass of octane converted to CO_2 ; therefore, $(2650 - x)$ g is the mass of octane converted to CO.

The amounts of CO_2 and H_2O produced by x g of octane are:

$$x \text{ g C}_8\text{H}_{18} \times \frac{1 \text{ mol C}_8\text{H}_{18}}{114.2 \text{ g C}_8\text{H}_{18}} \times \frac{16 \text{ mol CO}_2}{2 \text{ mol C}_8\text{H}_{18}} \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = 3.083x \text{ g CO}_2$$

$$x \text{ g C}_8\text{H}_{18} \times \frac{1 \text{ mol C}_8\text{H}_{18}}{114.2 \text{ g C}_8\text{H}_{18}} \times \frac{18 \text{ mol H}_2\text{O}}{2 \text{ mol C}_8\text{H}_{18}} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 1.420x \text{ g H}_2\text{O}$$

The amounts of CO and H₂O produced by (2650 - x) g of octane are:

$$(2650 - x) \text{ g C}_8\text{H}_{18} \times \frac{1 \text{ mol C}_8\text{H}_{18}}{114.2 \text{ g C}_8\text{H}_{18}} \times \frac{16 \text{ mol CO}}{2 \text{ mol C}_8\text{H}_{18}} \times \frac{28.01 \text{ g CO}}{1 \text{ mol CO}} = (5200 - 1.962x) \text{ g CO}$$

$$(2650 - x) \text{ g C}_8\text{H}_{18} \times \frac{1 \text{ mol C}_8\text{H}_{18}}{114.2 \text{ g C}_8\text{H}_{18}} \times \frac{18 \text{ mol H}_2\text{O}}{2 \text{ mol C}_8\text{H}_{18}} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = (3763 - 1.420x) \text{ g H}_2\text{O}$$

The total mass of CO₂ + CO + H₂O produced is 11530 g. We can write:

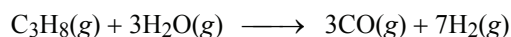
$$11530 \text{ g} = 3.083x + 1.420x + 5200 - 1.962x + 3763 - 1.420x$$

$$x = 2290 \text{ g}$$

Since x is the amount of octane converted to CO₂, we can now calculate the efficiency of the process.

$$\text{efficiency} = \frac{\text{g octane converted}}{\text{g octane total}} \times 100\% = \frac{2290 \text{ g}}{2650 \text{ g}} \times 100\% = \mathbf{86.49\%}$$

3.144 (a) The balanced chemical equation is:



(b) You should come up with the following strategy to solve this problem. In this problem, we use kg-mol to save a couple of steps.



$$\begin{aligned} ? \text{ kg H}_2 &= (2.84 \times 10^3 \text{ kg C}_3\text{H}_8) \times \frac{1 \text{ kg-mol C}_3\text{H}_8}{44.09 \text{ kg C}_3\text{H}_8} \times \frac{7 \text{ kg-mol H}_2}{1 \text{ kg-mol C}_3\text{H}_8} \times \frac{2.016 \text{ kg H}_2}{1 \text{ kg-mol H}_2} \\ &= \mathbf{9.09 \times 10^2 \text{ kg H}_2} \end{aligned}$$

3.145 For the first step of the synthesis, the yield is 90% or 0.9. For the second step, the yield will be 90% of 0.9 or (0.9 × 0.9) = 0.81. For the third step, the yield will be 90% of 0.81 or (0.9 × 0.9 × 0.9) = 0.73. We see that the yield will be:

$$\text{Yield} = (0.9)^n$$

where n = number of steps in the reaction. For 30 steps,

$$\text{Yield} = (0.9)^{30} = 0.04 = \mathbf{4\%}$$

3.146 (a) 16 amu, CH₄ 17 amu, NH₃ 18 amu, H₂O 64 amu, SO₂

(b) The formula C₃H₈ can also be written as CH₃CH₂CH₃. A CH₃ fragment could break off from this molecule giving a peak at 15 amu. No fragment of CO₂ can have a mass of 15 amu. Therefore, the substance responsible for the mass spectrum is most likely C₃H₈.

(c) First, let's calculate the masses of CO₂ and C₃H₈.

$$\text{molecular mass CO}_2 = 12.00000 \text{ amu} + 2(15.99491 \text{ amu}) = 43.98982 \text{ amu}$$

$$\text{molecular mass C}_3\text{H}_8 = 3(12.00000 \text{ amu}) + 8(1.00797 \text{ amu}) = 44.06376 \text{ amu}$$

These masses differ by only 0.07394 amu. The measurements must be precise to ± 0.030 amu.

$$43.98982 + 0.030 \text{ amu} = 44.02 \text{ amu}$$

$$44.06376 - 0.030 \text{ amu} = 44.03 \text{ amu}$$

3.147 (a) We need to compare the mass % of K in both KCl and K_2SO_4 .

$$\%K \text{ in KCl} = \frac{39.10 \text{ g}}{74.55 \text{ g}} \times 100\% = 52.45\% \text{ K}$$

$$\%K \text{ in K}_2\text{SO}_4 = \frac{2(39.10 \text{ g})}{174.27 \text{ g}} \times 100\% = 44.87\% \text{ K}$$

The price is dependent on the %K.

$$\frac{\text{Price of K}_2\text{SO}_4}{\text{Price of KCl}} = \frac{\%K \text{ in K}_2\text{SO}_4}{\%K \text{ in KCl}}$$

$$\text{Price of K}_2\text{SO}_4 = \text{Price of KCl} \times \frac{\%K \text{ in K}_2\text{SO}_4}{\%K \text{ in KCl}}$$

$$\text{Price of K}_2\text{SO}_4 = \frac{\$0.55}{\text{kg}} \times \frac{44.87\%}{52.45\%} = \mathbf{\$0.47/\text{kg}}$$

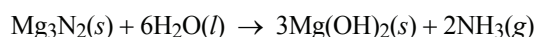
(b) First, calculate the number of moles of K in 1.00 kg of KCl.

$$(1.00 \times 10^3 \text{ g KCl}) \times \frac{1 \text{ mol KCl}}{74.55 \text{ g KCl}} \times \frac{1 \text{ mol K}}{1 \text{ mol KCl}} = 13.4 \text{ mol K}$$

Next, calculate the amount of K_2O needed to supply 13.4 mol K.

$$13.4 \text{ mol K} \times \frac{1 \text{ mol K}_2\text{O}}{2 \text{ mol K}} \times \frac{94.20 \text{ g K}_2\text{O}}{1 \text{ mol K}_2\text{O}} \times \frac{1 \text{ kg}}{1000 \text{ g}} = \mathbf{0.631 \text{ kg K}_2\text{O}}$$

3.148 When magnesium burns in air, magnesium oxide (MgO) and magnesium nitride (Mg_3N_2) are produced. Magnesium nitride reacts with water to produce ammonia gas.



From the amount of ammonia produced, we can calculate the mass of Mg_3N_2 produced. The mass of Mg in that amount of Mg_3N_2 can be determined, and then the mass of Mg in MgO can be determined by difference. Finally, the mass of MgO can be calculated.

$$2.813 \text{ g NH}_3 \times \frac{1 \text{ mol NH}_3}{17.034 \text{ g NH}_3} \times \frac{1 \text{ mol Mg}_3\text{N}_2}{2 \text{ mol NH}_3} \times \frac{100.95 \text{ g Mg}_3\text{N}_2}{1 \text{ mol Mg}_3\text{N}_2} = \mathbf{8.335 \text{ g Mg}_3\text{N}_2}$$

The mass of Mg in 8.335 g Mg_3N_2 can be determined from the mass percentage of Mg in Mg_3N_2 .

$$\frac{(3)(24.31 \text{ g Mg})}{100.95 \text{ g Mg}_3\text{N}_2} \times 8.335 \text{ g Mg}_3\text{N}_2 = 6.022 \text{ g Mg}$$

The mass of Mg in the product MgO is obtained by difference: $21.496 \text{ g Mg} - 6.022 \text{ g Mg} = 15.474 \text{ g Mg}$

The mass of MgO produced can now be determined from this mass of Mg and the mass percentage of Mg in MgO.

$$\frac{40.31 \text{ g MgO}}{24.31 \text{ g Mg}} \times 15.474 \text{ g Mg} = \mathbf{25.66 \text{ g MgO}}$$

- 3.149** Possible formulas for the metal bromide could be MBr, MBr₂, MBr₃, etc. Assuming 100 g of compound, the moles of Br in the compound can be determined. From the mass and moles of the metal for each possible formula, we can calculate a molar mass for the metal. The molar mass that matches a metal on the periodic table would indicate the correct formula.

Assuming 100 g of compound, we have 53.79 g Br and 46.21 g of the metal (M). The moles of Br in the compound are:

$$53.79 \text{ g Br} \times \frac{1 \text{ mol Br}}{79.90 \text{ g Br}} = 0.67322 \text{ mol Br}$$

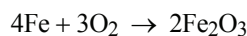
If the formula is MBr, the moles of M are also 0.67322 mole. If the formula is MBr₂, the moles of M are 0.67322/2 = 0.33661 mole, and so on. For each formula (MBr, MBr₂, and MBr₃), we calculate the molar mass of the metal.

$$\text{MBr: } \frac{46.21 \text{ g M}}{0.67322 \text{ mol M}} = 68.64 \text{ g/mol (no such metal)}$$

$$\text{MBr}_2: \frac{46.21 \text{ g M}}{0.33661 \text{ mol M}} = 137.3 \text{ g/mol (The metal is Ba. The formula is } \mathbf{BaBr_2})$$

$$\text{MBr}_3: \frac{46.21 \text{ g M}}{0.22441 \text{ mol M}} = 205.9 \text{ g/mol (no such metal)}$$

- 3.150** The decomposition of KClO₃ produces oxygen gas (O₂) which reacts with Fe to produce Fe₂O₃.



When the 15.0 g of Fe is heated in the presence of O₂ gas, any increase in mass is due to oxygen. The mass of oxygen reacted is:

$$17.9 \text{ g} - 15.0 \text{ g} = 2.9 \text{ g O}_2$$

From this mass of O₂, we can now calculate the mass of Fe₂O₃ produced and the mass of KClO₃ decomposed.

$$2.9 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \times \frac{2 \text{ mol Fe}_2\text{O}_3}{3 \text{ mol O}_2} \times \frac{159.7 \text{ g Fe}_2\text{O}_3}{1 \text{ mol Fe}_2\text{O}_3} = \mathbf{9.6 \text{ g Fe}_2\text{O}_3}$$

The balanced equation for the decomposition of KClO₃ is: 2KClO₃ → 2KCl + 3O₂. The mass of KClO₃ decomposed is:

$$2.9 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \times \frac{2 \text{ mol KClO}_3}{3 \text{ mol O}_2} \times \frac{122.55 \text{ g KClO}_3}{1 \text{ mol KClO}_3} = \mathbf{7.4 \text{ g KClO}_3}$$

3.151 Assume 100 g of sample. Then,

$$\text{mol Na} = 32.08 \text{ g Na} \times \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} = 1.395 \text{ mol Na}$$

$$\text{mol O} = 36.01 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.251 \text{ mol O}$$

$$\text{mol Cl} = 19.51 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} = 0.5504 \text{ mol Cl}$$

Since Cl is only contained in NaCl, the moles of Cl equals the moles of Na contained in NaCl.

$$\text{mol Na (in NaCl)} = 0.5504 \text{ mol}$$

The number of moles of Na in the remaining two compounds is: $1.395 \text{ mol} - 0.5504 \text{ mol} = 0.8446 \text{ mol Na}$.

To solve for moles of the remaining two compounds, let

$$x = \text{moles of Na}_2\text{SO}_4$$

$$y = \text{moles of NaNO}_3$$

Then, from the mole ratio of Na and O in each compound, we can write

$$2x + y = \text{mol Na} = 0.8446 \text{ mol}$$

$$4x + 3y = \text{mol O} = 2.251 \text{ mol}$$

Solving two equations with two unknowns gives

$$x = 0.1414 = \text{mol Na}_2\text{SO}_4 \quad \text{and} \quad y = 0.5618 = \text{mol NaNO}_3$$

Finally, we convert to mass of each compound to calculate the mass percent of each compound in the sample. Remember, the sample size is 100 g.

$$\text{mass \% NaCl} = 0.5504 \text{ mol NaCl} \times \frac{58.44 \text{ g NaCl}}{1 \text{ mol NaCl}} \times \frac{1}{100 \text{ g sample}} \times 100\% = \mathbf{32.17\% NaCl}$$

$$\text{mass \% Na}_2\text{SO}_4 = 0.1414 \text{ mol Na}_2\text{SO}_4 \times \frac{142.1 \text{ g Na}_2\text{SO}_4}{1 \text{ mol Na}_2\text{SO}_4} \times \frac{1}{100 \text{ g sample}} \times 100\% = \mathbf{20.09\% Na}_2\text{SO}_4$$

$$\text{mass \% NaNO}_3 = 0.5618 \text{ mol NaNO}_3 \times \frac{85.00 \text{ g NaNO}_3}{1 \text{ mol NaNO}_3} \times \frac{1}{100 \text{ g sample}} \times 100\% = \mathbf{47.75\% NaNO}_3$$