CHAPTER 3
MASS RELATIONSHIPS IN CHEMICAL REACTIONS

PROBLEM-SOLVING STRATEGIES

TYPES OF PROBLEMS

Problem Type 1: Calculating Average Atomic Mass.

Problem Type 2: Calculations Involving Molar Mass of an Element and Avogadro's Number.
- **(a)** Converting between moles of atoms and mass of atoms.
- **(b)** Calculating the mass of a single atom.
- **(c)** Converting mass in grams to number of atoms.

Problem Type 3: Calculations Involving Molecular Mass.
- **(a)** Calculating molecular mass.
- **(b)** Calculating the number of moles in a given amount of a compound.
- **(c)** Calculating the number of atoms in a given amount of a compound.

Problem Type 4: Calculations Involving Percent Composition
- **(a)** Calculating percent composition of a compound.
- **(b)** Determining empirical formula from percent composition.
- **(c)** Calculating mass from percent composition.

Problem Type 5: Experimental Determination of Empirical Formulas.

Problem Type 6: Determining the Molecular Formula of a Compound.

Problem Type 7: Calculating the Amounts of Reactants and Products.

Problem Type 8: Limiting Reagent Calculations.

Problem Type 9: Calculating the Percent Yield of a Reaction.

PROBLEM TYPE 1: CALCULATING AVERAGE ATOMIC MASS

The atomic mass you look up on a periodic table is an average atomic mass. The reason for this is that most naturally occurring elements have more than one isotope. The average atomic mass can be calculated as follows:

Step 1: Convert the percentage of each isotope to fractions. For example, an isotope that is 69.09 percent abundant becomes 69.09/100 = 0.6909.

Step 2: Multiply the mass of each isotope by its abundance and add them together.

average atomic mass = (fraction of isotope A)(mass of isotope A) + (fraction of isotope B) (mass of isotope B) + . . . + (fraction of isotope Z)(mass of isotope Z).

EXAMPLE 3.1

The element lithium has two isotopes that occur in nature: ${}^{6}_{3}Li$ with 7.5 percent abundance and ${}^{7}_{3}Li$ with 92.5 percent abundance. The atomic mass of ${}^{6}_{3}Li$ is 6.01513 amu and that of ${}^{7}_{3}Li$ is 7.01601 amu. Calculate the average atomic mass of lithium.

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.

Solution: Convert the percentage of each isotope to fractions.

$$ {}^{6}_{3}\mathbf{Li}: \ 7.5/100 = 0.075 $$

$$ {}^{7}_{3}\mathbf{Li}: \ 92.5/100 = 0.925 $$

Multiply the mass of each isotope by its abundance and add them together.

$$ \textbf{average atomic mass} = (0.075)(6.01513) + (0.925)(7.01601) = \mathbf{6.94\ amu} $$

PRACTICE EXERCISE

1. The element boron (B) consists of two stable isotopes with atomic masses of 10.0129 amu and 11.0093 amu. The average atomic mass of B is 10.81 amu. Which isotope is more abundant?

Text Problem: 3.6

PROBLEM TYPE 2: CALCULATIONS INVOLVING MOLAR MASS OF AN ELEMENT AND AVOGADRO'S NUMBER

A. Converting between moles of atoms and mass of atoms

In order to convert from one unit to another, you need to be proficient at the dimensional analysis method. See Section 1.9 of your text and Problem Type 5, Chapter 1. Unit conversions can seem daunting, but if you keep track of the units, making sure that the appropriate units cancel, your effort will be rewarded.

Step 1: Map out a strategy to proceed from initial units to final units based on available conversion factors.

Step 2: Use the following method to ensure that you obtain the desired unit.

$$ \text{Given}\cancel{\text{ unit}} \times \left(\frac{\text{desired unit}}{\text{given}\cancel{\text{ unit}}} \right) = \text{desired unit} $$

To convert between moles and mass, you need to use the molar mass of the element as a conversion factor.

$$ \cancel{\text{mol}} \times \frac{\text{g}}{\cancel{\text{mol}}} = \text{g} $$

Also, going in the opposite direction

$$ \cancel{\text{g}} \times \frac{\text{mol}}{\cancel{\text{g}}} = \text{mol} $$

Tip: Whether you are converting from g → mol or from mol → g, you will need to use the molar mass as the conversion factor. The molar mass of an element can be found directly on the periodic table.

EXAMPLE 3.2

How many grams are there in 0.130 mole of Cu?

Strategy: We are given moles of copper and asked to solve for grams of copper. 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 covert 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 Cu is 63.55 g. This can be expressed as

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

From this equality, we can write two conversion factors.

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

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

We write

$$? \textbf{ g Cu} = 0.130 \cancel{\text{mol}} \text{ Cu} \times \frac{63.55 \text{ g Cu}}{1 \cancel{\text{mol}} \text{ Cu}} = \mathbf{8.26 \text{ g Cu}}$$

Check: Does a mass of 8.26 g for 0.130 mole of Cu seem reasonable? What is the mass of 1 mole of Cu?

PRACTICE EXERCISE

2. How many moles of Cu are in 125 g of Cu?

Text Problem: 3.16

B. Calculating the mass of a single atom

To calculate the mass of a single atom, you can use Avogadro's number. The conversion factor is

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

EXAMPLE 3.3

Copper is a minor component of pennies minted since 1981, and it is also used in electrical cables. Calculate the mass (in grams) of a single Cu atom.

Strategy: We can look up the molar mass of copper (Cu) on the periodic table (63.55 g/mol). We want to find the mass of a single atom of copper (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.

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

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

We write

$$? \text{ g/Cu atom} = \frac{63.55 \text{ g Cu}}{1 \cancel{\text{mol}} \text{ Cu}} \times \frac{1 \cancel{\text{mol}} \text{ Cu}}{6.022 \times 10^{23} \text{ Cu atoms}} = \mathbf{1.055 \times 10^{-22} \text{ g/Cu atom}}$$

Check: Should the mass of a single atom of Cu be a very small mass?

PRACTICE EXERCISE

3. Titanium (Ti) is a transition metal with a very high strength-to-weight ratio. For this reason, titanium is used in the construction of aircraft. What is the mass (in grams) of one Ti atom?

Text Problem: 3.18

C. Converting mass in grams to number of atoms

To complete the following conversion, you need to use both molar mass and Avogadro's number as conversion factors.

EXAMPLE 3.4

Zinc is the main component of pennies minted after 1981. How many zinc atoms are present in 20.0 g of Zn?

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

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

$$1 \text{ mol Zn} = 65.39 \text{ g Zn}$$

The conversion factor needed is

$$\frac{1 \text{ mol Zn}}{65.39 \text{ g Zn}}$$

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.

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

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

Let's complete the two conversions in one step.

$$\text{grams of Zn} \rightarrow \text{moles of Zn} \rightarrow \text{number of Zn atoms}$$

$$\textbf{? atoms of Zn} = 20.0\,\cancel{g}\,Zn \times \frac{1\,\cancel{mol}\,Zn}{65.39\,\cancel{g}\,Zn} \times \frac{6.022 \times 10^{23}\text{ Zn atoms}}{1\,\cancel{mol}\,Zn} = \mathbf{1.84 \times 10^{23}\text{ Zn atoms}}$$

Check: Should 20.0 g of Zn contain fewer than Avogadro's number of atoms? What mass of Zn would contain Avogadro's number of atoms?

PRACTICE EXERCISE

4. What is the mass (in grams) of 9.09×10^{23} atoms of Zn?

Text Problem: 3.20

PROBLEM TYPE 3: CALCULATIONS INVOLVING MOLECULAR MASS

A. Calculating molecular mass

The molecular mass is simply the sum of the atomic masses (in amu) of all the atoms in the molecule.

EXAMPLE 3.5

Calculate the molecular mass of carbon tetrachloride (CCl_4).

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

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

$$\text{molecular mass } CCl_4 = (\text{mass of C}) + 4(\text{mass of Cl})$$

$$\textbf{molecular mass } \mathbf{CCl_4} = (12.01\text{ amu}) + 4(35.45\text{ amu}) = \mathbf{153.8\text{ amu}}$$

PRACTICE EXERCISE

5. Bananas owe their characteristic smell and flavor to the ester, isopentyl acetate [$CH_3COOCH_2CH_2CH(CH_3)_2$]. Calculate the molecular mass of isopentyl acetate.

Text Problem: 3.24

B. Calculating the number of moles in a given amount of a compound

To complete this conversion, the only conversion factor needed is the molar mass in units of g/mol. Remember, the molar mass of a compound (in grams) is numerically equal to its molecular mass (in atomic mass units). For example, the molar mass of CCl_4 is 153.8 g/mol, compared to its molecular mass of 153.8 amu.

EXAMPLE 3.6

How many moles of ethane (C_2H_6) are present in 50.3 g of ethane?

Strategy: First, calculate the molar mass of ethane. Then, arrange the molar mass as a conversion factor to convert from grams of ethane to moles of ethane.

Solution:

$$\text{molar mass of } C_2H_6 = 2(12.01\text{ g}) + 6(1.008\text{ g}) = 30.07\text{ g}$$

Hence, the conversion factor is

$$\frac{1 \text{ mol } C_2H_6}{30.07 \text{ g } C_2H_6}$$

Using this conversion factor, convert from grams to moles.

$$\textbf{? mol of } \mathbf{C_2H_6} = 50.3\cancel{\text{g}}\,C_2H_6 \times \frac{1 \text{ mol } C_2H_6}{30.07\cancel{\text{g}}\,C_2H_6} = \mathbf{1.67\ mol}$$

PRACTICE EXERCISE

6. What is the mass (in grams) of 0.436 moles of ethane (C_2H_6)?

Text Problem: 3.26

C. Calculating the number of atoms in a given amount of a compound

Again, this is a unit conversion problem. This calculation is more difficult than the conversions above, because you must convert from *grams of compound* to *moles of compound* to *moles of a particular atom* to *number of atoms*. Sound tough? Let's try an example.

EXAMPLE 3.7

How many carbon atoms are present in 50.3 g of ethane (C_2H_6)?

Strategy: We started this problem in Example 3.6 when we calculated the moles of ethane in 50.3 g ethane. To continue, we need two additional conversion factors. One should represent the mole ratio between moles of C atoms and moles of ethane molecules. The other conversion factor needed is Avogadro's number.

Solution: The two conversion factors needed are:

$$\frac{2 \text{ mol C}}{1 \text{ mol } C_2H_6} \qquad \frac{6.022 \times 10^{23} \text{ C atoms}}{1 \text{ mol C}}$$

You should come up with the following strategy.

grams of C_2H_6 → moles of C_2H_6 → moles of C → atoms of C

$$\textbf{? C atoms} = 50.3\cancel{\text{g}}\,C_2H_6 \times \frac{1\,\cancel{\text{mol}}\,C_2H_6}{30.07\cancel{\text{g}}\,C_2H_6} \times \frac{2\,\cancel{\text{mol}}\,\text{C}}{1\,\cancel{\text{mol}}\,C_2H_6} \times \frac{6.022 \times 10^{23} \text{ C atoms}}{1\,\cancel{\text{mol}}\,\text{C}}$$
$$= \mathbf{2.01 \times 10^{24}\ C\ atoms}$$

Check: Does the answer seem reasonable? We have 50.3 g ethane. How many atoms of C would 30.07 g of ethane contain?

PRACTICE EXERCISE

7. Glucose, the sugar used by the cells of our bodies for energy, has the molecular formula, $C_6H_{12}O_6$. How many atoms of *carbon* are present in a 3.50 g sample of glucose?

Text Problem: 3.28

PROBLEM TYPE 4: CALCULATIONS INVOLVING PERCENT COMPOSITION

A. Calculating percent composition of a compound

The *percent composition by mass* is the percent by mass of each element the compound contains. Percent composition is obtained by dividing the mass of each element in 1 mole of the compound by the molar mass of the compound, then multiplying by 100 percent.

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

EXAMPLE 3.8

Calculate the percent composition by mass of all the elements in sodium bicarbonate, $NaHCO_3$.

Strategy: First, calculate the molar mass of sodium bicarbonate. Then, calculate the percent by mass of each element.

Solution:

$$\text{molar mass sodium bicarbonate} = 22.99\text{ g} + 1.008\text{ g} + 12.01\text{ g} + 3(16.00\text{ g}) = 84.01\text{ g}$$

$$\%\mathbf{Na} = \frac{22.99\text{ g}}{84.01\text{ g}} \times 100\% = \mathbf{27.37\%}$$

$$\%\mathbf{H} = \frac{1.008\text{ g}}{84.01\text{ g}} \times 100\% = \mathbf{1.200\%}$$

$$\%\mathbf{C} = \frac{12.01\text{ g}}{84.01\text{ g}} \times 100\% = \mathbf{14.30\%}$$

$$\%\mathbf{O} = \frac{3(16.00\text{ g})}{84.01\text{ g}} \times 100\% = \mathbf{57.14\%}$$

Tip: You can check your work by making sure that the mass percents of all the elements added together equals 100%. Checking above, 27.37% + 1.200% + 14.30% + 57.14% = 100.01% ≈ 100%.

PRACTICE EXERCISE

8. Cinnamic alcohol is used mainly in perfumes, particularly for soaps and cosmetics. Its molecular formula is $C_9H_{10}O$. Calculate the percent composition by mass of *hydrogen* in cinnamic alcohol.

Text Problems: 3.40, 3.42

B. Determining empirical formula from percent composition

The procedure used above to calculate the percent composition of a compound can be reversed. Given the percent composition by mass of a compound, you can determine the empirical formula of the compound.

EXAMPLE 3.9

Dieldrin, like DDT, is an insecticide that contains only C, H, Cl, and O. It is composed of 37.84 percent C, 2.12 percent H, 55.84 percent Cl, and 4.20 percent O. Determine its empirical formula.

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 37.84 g of C, 2.12 g of H, 55.84 g Cl, and 4.20 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

$$n_C = 37.84\,\cancel{g}\,C \times \frac{1 \text{ mol C}}{12.01\,\cancel{g}\,C} = \mathbf{3.151 \text{ mol C}}$$

$$n_H = 2.12\,\cancel{g}\,H \times \frac{1 \text{ mol H}}{1.008\,\cancel{g}\,H} = \mathbf{2.10 \text{ mol H}}$$

$$n_{Cl} = 55.84\,\cancel{g}\,Cl \times \frac{1 \text{ mol Cl}}{35.45\,\cancel{g}\,Cl} = \mathbf{1.575 \text{ mol Cl}}$$

$$n_O = 4.20\,\cancel{g}\,O \times \frac{1 \text{ mol O}}{16.00\,\cancel{g}\,O} = \mathbf{0.263 \text{ mol O}}$$

Thus, we arrive at the formula $C_{3.151}H_{2.10}Cl_{1.575}O_{0.263}$, which gives the identity and the 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.

$$\mathbf{C:}\ \frac{3.151}{0.263} = 12.0 \qquad \mathbf{H:}\ \frac{2.10}{0.263} = 7.98 \approx 8 \qquad \mathbf{Cl:}\ \frac{1.575}{0.263} = 5.99 \approx 6 \qquad \mathbf{O:}\ \frac{0.263}{0.263} = 1$$

This gives us the empirical for dieldrin, $\mathbf{C_{12}H_8Cl_6O}$.

Check: Are the subscripts in $C_{12}H_8Cl_6O$ reduced to the smallest whole numbers?

> **Tip:** It's not always this easy. Dividing by the smallest subscript often does not give all whole numbers. If this is the case, you must multiply all the subscripts by some *integer* to come up with whole number subscripts. Try the practice exercise below.

PRACTICE EXERCISE

9. The substance responsible for the green color on the yolk of a boiled egg is composed of 53.58 percent Fe and 46.42 percent S. Determine its empirical formula.

Text Problems: 3.44, **3.50**, 3.54

C. Calculating mass from percent composition

Step 1: Convert the mass percentage to a fraction. For example, if the mass percent of an element in a compound were 54.73 percent, you would convert this to 54.73/100 = 0.5473.

Step 2: Multiply the fraction by the total mass of the compound. This gives the mass of the particular element in the compound.

EXAMPLE 3.10

Calculate the mass of carbon in exactly 10 g of glucose ($C_6H_{12}O_6$).

Strategy: Glucose is composed of C, H, and O. The mass due to C 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 carbon in $C_6H_{12}O_6$. Then, we convert this percentage to a fraction and multiply by the mass of the compound (10 g), to find the mass of carbon in 10 g of $C_6H_{12}O_6$.

The percent by mass of carbon in glucose, is calculated as follows:

$$\text{mass \% C} = \frac{\text{mass of C in 1 mol of glucose}}{\text{molar mass of glucose}} \times 100\%$$

$$\textbf{mass \% C} = \frac{6(12.01\text{ g})}{180.16\text{ g}} \times 100\% = \textbf{40.00\% C}$$

Converting this percentage to a fraction, we obtain 40.00/100 = **0.4000**

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

$$\textbf{? g C in 10 g glucose} = (0.4000)(10\text{ g}) = \textbf{4.000 g C}$$

Check: Note that the mass percent of C is 40 percent. 40% of 10 g is 4 g.

PRACTICE EXERCISE

10. Calculate the mass of hydrogen in exactly 10 grams of glucose ($C_6H_{12}O_6$).

Text Problems: 3.46, **3.48**

PROBLEM TYPE 5: EXPERIMENTAL DETERMINATION OF EMPIRICAL FORMULAS

See Section 3.6 of your text for a description of the experimental setup. To solve this type of problem, you must recognize that all of the carbon in the sample is converted to CO_2 and all the hydrogen in the sample is converted to H_2O. Then, you can calculate the mass of C in CO_2 and the mass of H in H_2O. Finally, you can calculate the mass of oxygen by difference, if necessary.

EXAMPLE 3.11

When a 0.761-g sample of a compound containing only carbon and hydrogen is burned in an apparatus with CO_2 and H_2O absorbers, 2.23 g CO_2 and 1.37 g H_2O are collected. Determine the empirical formula of the compound.

Strategy: Calculate the moles of C in 2.23 g CO_2, and the moles of H in 1.37 g H_2O. In this problem, we do not need to convert to grams of C and H, because there are no other elements in the compound. To calculate the moles of each component, you need the molar masses and the correct mole ratio.

You should come up with the following strategy.

$$\text{g } CO_2 \rightarrow \text{mol } CO_2 \rightarrow \text{mol C}$$

Next, determine the smallest whole number ratio in which the elements combine.

Solution:

$$\text{? mol C} = 2.23\,\cancel{\text{g}}\,CO_2 \times \frac{1\,\cancel{\text{mol}}\,CO_2}{44.01\,\cancel{\text{g}}\,CO_2} \times \frac{1\,\cancel{\text{mol}}\,C}{1\,\cancel{\text{mol}}\,CO_2} = 0.507\text{ mol C}$$

Similarly,

$$\text{? mol H} = 1.37\,\cancel{\text{g}}\,H_2O \times \frac{1\,\cancel{\text{mol}}\,H_2O}{18.02\,\cancel{\text{g}}\,H_2O} \times \frac{2\,\cancel{\text{mol}}\,H}{1\,\cancel{\text{mol}}\,H_2O} = 0.152\text{ mol H}$$

Thus, we arrive at the formula $C_{0.0507}H_{0.152}$, which gives the identity and the 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.

$$\mathbf{C}: \frac{0.0507}{0.0507} = 1.00 \qquad \mathbf{H}: \frac{0.152}{0.0507} = 3.00$$

This gives the empirical formula, $\mathbf{CH_3}$.

PRACTICE EXERCISE

11. Diethyl ether, commonly known as "ether", was used as an anesthetic for many years. Diethyl ether contains C, H, and O. When a 1.45 g sample of ether is burned in an apparatus such as that shown in Figure 3.6 of the text, 2.77 g of CO_2 and 1.70 g of H_2O are collected. Determine the empirical formula of diethyl ether.

Text Problem: 3.136

PROBLEM TYPE 6: DETERMINING THE MOLECULAR FORMULA OF A COMPOUND

To determine the molecular formula of a compound, we must know both the *approximate* molar mass and the empirical formula of the compound. The molecular formula will either be equal to the empirical formula or be some integral multiple of it. Thus, the molar mass divided by the empirical mass will be an integer greater than or equal to one.

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

EXAMPLE 3.12

A mass spectrum obtained on the compound in Example 3.11, shows its molecular mass to be about 31 g/mol. What is its molecular formula?

Strategy: First, determine the empirical formula. Then compare the molar mass to the empirical molar mass to determine the molecular formula.

Solution: The empirical formula was determined in the previous example to be $\mathbf{CH_3}$.

Next, calculate the empirical molar mass.

$$\text{empirical molar mass} = 12.01\text{ g} + 3(1.008\text{ g}) = 15.03\text{ g/mol}$$

Determine the number of (CH_3) units present in the molecular formula. This number is found by taking the ratio

$$\frac{\text{molar mass}}{\text{empirical molar mass}} = \frac{31\text{ g}}{15.03\text{ g}} = 2.1 \approx 2$$

Thus, there are two CH_3 units in each molecule of the compound, so the molecular formula is $(CH_3)_2$, or $\mathbf{C_2H_6}$.

PRACTICE EXERCISE

12. In Example 3.9, the empirical formula of dieldrin was determined to be $C_{12}H_8Cl_6O$. If the molar mass of dieldrin is 381 ± 10 g/mol, what is the molecular formula of dieldrin?

Text Problems: 3.52, 3.54

PROBLEM TYPE 7: CALCULATING THE AMOUNTS OF REACTANTS AND PRODUCTS

These types of problems are dimensional analysis problems. You must always remember to start this type of problem with a balanced chemical equation. The typical approach is given below. See Section 3.8 of your text for a step-by-step method.

Amount of reactant (expressed in grams, kg, etc.)	→	Moles of reactant	→	Moles of product	→	Amount of product (expressed in grams, kg, etc.)

Tip: Always try to be flexible when solving problems. Most problems of this type will follow an approach similar to the one above, but you may have to modify it sometimes.

EXAMPLE 3.13

Sulfur dioxide can be removed from stack gases by reaction with quicklime (CaO):

$$SO_2(g) + CaO(s) \longrightarrow CaSO_3(s)$$

If 975 kg of SO_2 are to be removed from stack gases by the above reaction, how many kilograms of CaO are required?

Strategy: We compare SO_2 and CaO based on the *mole ratio* in the balanced equation. Before we can determine moles of CaO required, we need to convert to moles of SO_2. What conversion factor is needed to convert from grams of SO_2 to moles of SO_2? Once moles of CaO are obtained, another conversion factor is needed to convert from moles of CaO to grams of CaO.

Solution: The molar mass of SO_2 will allow us to convert from grams of SO_2 to moles of SO_2. The molar mass of SO_2 = 32.07 g + 2(16.00 g) = 64.07 g. The balanced equation is given, so the mole ratio between SO_2 and CaO is known, that is, 1 mole $SO_2 \simeq$ 1 mole CaO. Finally, the molar mass of CaO will convert moles of CaO to grams of CaO. This sequence of conversions is summarized as follows:

$$\text{kg SO}_2 \rightarrow \text{g SO}_2 \rightarrow \text{moles SO}_2 \rightarrow \text{moles CaO} \rightarrow \text{g CaO} \rightarrow \text{kg CaO}$$

$$\textbf{? kg CaO} = 975\ \cancel{\text{kg}}\ \text{SO}_2 \times \frac{1000\ \cancel{\text{g}}\ \text{SO}_2}{1\ \cancel{\text{kg}}\ \text{SO}_2} \times \frac{1\ \cancel{\text{mol}}\ \text{SO}_2}{64.07\ \cancel{\text{g}}\ \text{SO}_2} \times \frac{1\ \cancel{\text{mol}}\ \text{CaO}}{1\ \cancel{\text{mol}}\ \text{SO}_2} \times \frac{56.08\ \cancel{\text{g}}\ \text{CaO}}{1\ \cancel{\text{mol}}\ \text{CaO}} \times \frac{1\ \text{kg CaO}}{1000\ \cancel{\text{g}}\ \text{CaO}}$$

$$= \textbf{853 kg CaO}$$

Tip: Notice that the approach followed was a slight modification of the flow diagram given above. We went from mass of one reactant, to moles of that reactant, to moles of a second reactant, and finally to mass of second reactant.

PRACTICE EXERCISE

13. Carbon dioxide in the air of a spacecraft can be removed by its reaction with a lithium hydroxide solution.

$$CO_2(g) + 2LiOH(aq) \longrightarrow Li_2CO_3(aq) + H_2O(l)$$

On average, a person will exhale about 1 kg of CO_2/day. How many kilograms of LiOH are required to react with 1.0 kg of CO_2?

Text Problems: 3.66, 3.68, 3.70, **3.72**, 3.74, 3.76, 3.78

PROBLEM TYPE 8: LIMITING REAGENT CALCULATIONS

When a chemist carries out a reaction, the reactants are usually not present in exact **stoichiometric amounts**, that is, in the proportions indicated by the balanced equation. The reactant used up first in a reaction is called the **limiting reagent**. When this reactant is used up, no more product can be formed.

Typically, the only difference between this type of problem and Problem Type 7, Calculating the Amounts of Reactants and Products, is that you must first determine which reactant is the limiting reagent.

EXAMPLE 3.14

Phosphine (PH_3) burns in oxygen (O_2) to produce phosphorus pentoxide and water.

$$2PH_3(g) + 4O_2(g) \longrightarrow P_2O_5(s) + 3H_2O(l)$$

How many grams of P_2O_5 will be produced when 17.0 g of phosphine are reacted with 16.0 g of O_2?

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, and then compare the moles of product, P_2O_5, formed by the given amounts of PH_3 and O_2 to determine which reactant is the limiting reagent.

Solution: We carry out two separate calculations. First, starting with 17.0 g PH_3, we calculate the number of moles of P_2O_5 that could be produced if all the PH_3 reacted. We complete the following conversions.

$$\text{grams of } PH_3 \rightarrow \text{moles of } PH_3 \rightarrow \text{moles of } P_2O_5$$

Combining these two conversions into one calculation, we write

$$? \text{ mol } P_2O_5 = 17.0 \cancel{g} PH_3 \times \frac{1 \cancel{mol} PH_3}{33.99 \cancel{g} PH_3} \times \frac{1 \text{ mol } P_2O_5}{2 \cancel{mol} PH_3} = 0.250 \text{ mol } P_2O_5$$

Second, starting with 16.0 g of O_2, we complete similar conversions.

$$\text{grams of } O_2 \rightarrow \text{moles of } O_2 \rightarrow \text{moles of } P_2O_5$$

Combining these two conversions into one calculation, we write

$$? \text{ mol } P_2O_5 = 16.0 \cancel{g} O_2 \times \frac{1 \cancel{mol} O_2}{32.0 \cancel{g} O_2} \times \frac{1 \text{ mol } P_2O_5}{4 \cancel{mol} O_2} = 0.125 \text{ mol } P_2O_5$$

The initial amount of O_2 limits the amount of product that can be formed; therefore, it is the limiting reagent.

The problem asks for grams of P_2O_5 produced. We already know the moles of P_2O_5 produced, 0.125 mole. Use the molar mass of P_2O_5 as a conversion factor to convert to grams.

$$\mathbf{? \text{ g } P_2O_5} = 0.125 \cancel{mol} P_2O_5 \times \frac{141.94 \text{ g } P_2O_5}{1 \cancel{mol} P_2O_5} = \mathbf{17.7 \text{ g } P_2O_5}$$

Check: Does your answer seem reasonable? 0.125 mole of product is formed. What is the mass of 1 mole of P_2O_5?

PRACTICE EXERCISE

14. Iron can be produced by reacting iron ore with carbon. The iron produced can then be used to make steel. The reaction is

$$2Fe_2O_3(s) + 3C(s) \xrightarrow{\text{heat}} 4Fe(l) + 3CO_2(g)$$

(a) How many grams of Fe can be produced from a mixture of 200.0 g of Fe_2O_3 and 300.0 g C?
(b) How many grams of excess reagent will remain after the reaction ceases?

Text Problems: 3.84, 3.86

PROBLEM TYPE 9: CALCULATING THE PERCENT YIELD OF A REACTION

The **theoretical yield** is the amount of product that would result if all the limiting reagent reacted. This is the maximum obtainable yield predicted by the balanced equation. However, the amount of product obtained is almost always less than the theoretical yield. The **actual yield** is the quantity of product that actually results from a reaction.

To determine the efficiency of a reaction, chemists often calculate the **percent yield**, which describes the proportion of the actual yield to the theoretical yield. The percent yield is calculated as follows:

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

EXAMPLE 3.15

In Example 3.14, the theoretical yield of P_2O_5 was determined to be 17.7 g. If only 12.6 g of P_2O_5 are actually obtained, what is the percent yield of the reaction?

Solution:

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

$$\textbf{\% yield} = \frac{12.6\text{ g}}{17.7\text{ g}} \times 100\% = \mathbf{71.2\%}$$

PRACTICE EXERCISE

15. Refer back to Practice Exercise 14 to answer this question. If the actual yield of Fe is 110 g, what is the percent yield of Fe?

Text Problems: 3.90, 3.92, 3.94

ANSWERS TO PRACTICE EXERCISES

1. ^{11}B

2. 1.97 moles Cu

3. 7.951×10^{-23} g/Ti atom

4. 98.7 g Zn

5. 130.18 amu

6. 13.1 g ethane

7. 7.02×10^{22} C atoms

8. 7.513 percent H by mass

9. Fe_2S_3

10. 0.67 g H

11. C_2H_6O

12. $C_{12}H_8Cl_6O$

13. 1.1 kg LiOH

14. **(a)** 139.9 g Fe
(b) 277 g C

15. 78.6 percent yield

SOLUTIONS TO SELECTED TEXT PROBLEMS

3.6 This is a variation of Problem Type 1, Calculating Average Atomic Mass.

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}Li and the fractional abundance of ^{7}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}Li. Since the sum of the two abundance's must be 1, we can write

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

Solution:

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

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

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

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

3.12 The thickness of the book in miles would be:

$$\frac{0.0036 \cancel{\text{in}}}{1 \cancel{\text{page}}} \times \frac{1 \cancel{\text{ft}}}{12 \cancel{\text{in}}} \times \frac{1 \text{ mi}}{5280 \cancel{\text{ft}}} \times (6.022 \times 10^{23} \cancel{\text{pages}}) = 3.4 \times 10^{16} \text{ mi}$$

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

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

The thickness of the book in light-years is:

$$(3.4 \times 10^{16} \cancel{\text{mi}}) \times \frac{1 \text{ light-yr}}{5.88 \times 10^{12} \cancel{\text{mi}}} = \mathbf{5.8 \times 10^3\ light\text{-}yr}$$

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

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

3.16 Converting between moles of atoms and mass of atoms, Problem Type 2A.

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 covert 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

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

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

3.18 Calculating the mass of a single atom, Problem Type 2B.

(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

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

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

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

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

3.20 Converting mass in grams to number of atoms, Problem Type 2a.

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 → moles of Cu → number of Cu atoms

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

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?

3.22

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

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

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

3.24 Calculating molar mass, modification of Problem Type 3A.

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 g) + 12.01 g + 3(16.00 g) = **73.89 g**

(b) **molar mass CS_2** = 12.01 g + 2(32.07 g) = **76.15 g**

(c) **molar mass $CHCl_3$** = 12.01 g + 1.008 g + 3(35.45 g) = **119.37 g**

(d) **molar mass $C_6H_8O_6$** = 6(12.01 g) + 8(1.008 g) + 6(16.00 g) = **176.12 g**

(e) **molar mass KNO_3** = 39.10 g + 14.01 g + 3(16.00 g) = **101.11 g**

(f) **molar mass Mg_3N_2** = 3(24.31 g) + 2(14.01 g) = **100.95 g**

3.26 Calculating the number of molecules in a given amount of compound, similar to Problem Type 3B.

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.

$$\text{molar mass of } C_2H_6 = 2(12.01\text{ g}) + 6(1.008\text{ g}) = 30.068\text{ g}$$

The conversion factor needed is

$$\frac{1\text{ mol } C_2H_6}{30.068\text{ g } C_2H_6}$$

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 → moles of ethane → number of ethane molecules

$$\textbf{? molecules of } \mathbf{C_2H_6} = 0.334\,\cancel{\text{g}}\,C_2H_6 \times \frac{1\,\cancel{\text{mol}}\,C_2H_6}{30.07\,\cancel{\text{g}}\,C_2H_6} \times \frac{6.022 \times 10^{23}\,C_2H_6\text{ molecules}}{1\,\cancel{\text{mol}}\,C_2H_6}$$

$$= \mathbf{6.69 \times 10^{21}\ C_2H_6\ molecules}$$

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?

3.28 Calculating the number of atoms in a given amount of a compound, Problem Type 3C.

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 → moles of urea → molecules of urea → 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.

$$\textbf{? atoms of N} = (1.68 \times 10^4\ \cancel{g}\ \text{urea}) \times \frac{1\ \cancel{mol}\ \text{urea}}{60.06\ \cancel{g}\ \text{urea}} \times \frac{6.022 \times 10^{23}\ \text{urea}\ \cancel{\text{molecules}}}{1\ \cancel{mol}\ \text{urea}} \times \frac{2\ \text{N atoms}}{1\ \cancel{\text{molecule}}\ \text{urea}}$$

$$= \mathbf{3.37 \times 10^{26}\ N\ atoms}$$

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 → moles of urea → moles of N → 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 ration is 4:2.

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

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

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

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

Molar mass of H_2O = (16.00 g) + 2(1.008 g) = 18.02 g/mol

$$\textbf{? H}_2\textbf{O molecules} = 2.56\ \cancel{g}\ H_2O \times \frac{1\ \cancel{mol}\ H_2O}{18.02\ \cancel{g}\ H_2O} \times \frac{6.022 \times 10^{23}\ \text{molecules}\ H_2O}{1\ \cancel{mol}\ H_2O}$$

$$= \mathbf{8.56 \times 10^{22}\ molecules}$$

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

	^{32}S	^{33}S	^{34}S	^{36}S
$^{1}H_2$	34	35	36	38
$^{1}H^{2}H$	35	36	37	39
$^{2}H_2$	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}H and one ^{2}H. How many peaks would be detected in such a "high resolution" mass spectrum?

3.40 Calculating percent composition of a compound, Problem Type 4A.

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 $CHCl_3$ = 12.01 g/mol + 1.008 g/mol + 3(35.45 g/mol) = 119.4 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.

3.42

	Compound	Molar mass (g)	N% by mass
(a)	$(NH_2)_2CO$	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)	$HNC(NH_2)_2$	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.

3.44 **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.

$$n_C = 19.8\cancel{g}\,C \times \frac{1\text{ mol C}}{12.01\cancel{g}\,C} = 1.65\text{ mol C}$$

$$n_H = 2.50\cancel{g}\,H \times \frac{1\text{ mol H}}{1.008\cancel{g}\,H} = 2.48\text{ mol H}$$

$$n_N = 11.6\cancel{g}\,N \times \frac{1\text{ mol N}}{14.01\cancel{g}\,N} = 0.828\text{ mol N}$$

$$n_O = 66.1\cancel{g}\,O \times \frac{1\text{ mol O}}{16.00\cancel{g}\,O} = 4.13\text{ mol O}$$

Step 3: Try to convert to whole numbers by dividing all the subscripts by the smallest subscript. The formula is $C_{1.65}H_{2.48}N_{0.828}O_{4.13}$. Dividing the subscripts by 0.828 gives the empirical formula, $\mathbf{C_2H_3NO_5}$.

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, $\mathbf{C_2H_3NO_5}$.

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.

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

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

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

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

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 $\mathbf{C_2H_3NO_5}$.

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 $\mathbf{C_2H_3NO_5}$.

3.46 Using unit factors we convert:

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

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

3.48 Calculating mass from percent composition, Problem Type 4C.

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 SnF_2. 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 SnF_2.

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

$$\text{mass \% F} = \frac{\text{mass of F in 1 mol SnF}_2}{\text{molar mass of SnF}_2} \times 100\%$$

$$= \frac{2(19.00\ \text{g})}{156.7\ \text{g}} \times 100\% = 24.25\%\ \text{F}$$

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

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

$$\textbf{? g F in 24.6 g SnF}_\mathbf{2} = (0.2425)(24.6\ \text{g}) = \mathbf{5.97\ g\ F}$$

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:

$$\text{g of SnF}_2 \rightarrow \text{mol of SnF}_2 \rightarrow \text{mol of F} \rightarrow \text{g of F}$$

3.50 Determining empirical formula from percent composition, Problem Type 4C.

(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

$$n_C = 40.1\cancel{g}\,C \times \frac{1\text{ mol C}}{12.01\cancel{g}\,C} = 3.34\text{ mol C}$$

$$n_H = 6.6\cancel{g}\,H \times \frac{1\text{ mol H}}{1.008\cancel{g}\,H} = 6.55\text{ mol H}$$

$$n_O = 53.3\cancel{g}\,O \times \frac{1\text{ mol O}}{16.00\cancel{g}\,O} = 3.33\text{ mol O}$$

Thus, we arrive at the formula $C_{3.34}H_{6.5}O_{3.33}$, 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).

$$\mathbf{C}: \frac{3.34}{3.33} \approx 1 \qquad \mathbf{H}: \frac{6.55}{3.33} \approx 2 \qquad \mathbf{O}: \frac{3.33}{3.33} = 1$$

This gives the empirical formula, $\mathbf{CH_2O}$.

Check: Are the subscripts in CH_2O reduced to the smallest whole numbers?

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

$$n_C = 18.4\cancel{g}\,C \times \frac{1\text{ mol C}}{12.01\cancel{g}\,C} = 1.53\text{ mol C}$$

$$n_N = 21.5\cancel{g}\,N \times \frac{1\text{ mol N}}{14.01\cancel{g}\,N} = 1.53\text{ mol N}$$

$$n_K = 60.1\cancel{g}\,K \times \frac{1\text{ mol K}}{39.10\cancel{g}\,K} = 1.54\text{ mol K}$$

Dividing by the smallest number of moles (1.53 mol) gives the empirical formula, **KCN**.

3.52 The empirical molar mass of CH is approximately 13.02 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,

$$\frac{\text{molar mass}}{\text{empirical molar mass}} = \frac{78\text{ g}}{13.02\text{ g}} \approx 6$$

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

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.

$$n_C = 35.51\,\cancel{g}\,C \times \frac{1\text{ mol C}}{12.01\,\cancel{g}\,C} = 2.957\text{ mol C}$$

$$n_H = 4.77\,\cancel{g}\,H \times \frac{1\text{ mol H}}{1.008\,\cancel{g}\,H} = 4.73\text{ mol H}$$

$$n_O = 37.85\,\cancel{g}\,O \times \frac{1\text{ mol O}}{16.00\,\cancel{g}\,O} = 2.366\text{ mol O}$$

$$n_N = 8.29\,\cancel{g}\,N \times \frac{1\text{ mol N}}{14.01\,\cancel{g}\,N} = 0.592\text{ mol N}$$

$$n_{Na} = 13.60\,\cancel{g}\,Na \times \frac{1\text{ mol Na}}{22.99\,\cancel{g}\,Na} = 0.5916\text{ mol Na}$$

Thus, we arrive at the formula $C_{2.957}H_{4.73}O_{2.366}N_{0.592}Na_{0.5916}$, 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.

$$\textbf{C}: \frac{2.957}{0.59156} = 4.998 \approx 5 \qquad \textbf{H}: \frac{4.73}{0.59156} = 8.00 \qquad \textbf{O}: \frac{2.366}{0.59156} = 3.999 \approx 4$$

$$\textbf{N}: \frac{0.592}{0.59156} = 1.00 \qquad \textbf{Na}: \frac{0.5916}{0.5916} = 1$$

This gives us the empirical formula for MSG, $C_5H_8O_4NNa$.

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, $\mathbf{C_5H_8O_4NNa}$. 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.

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

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

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

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

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

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 $\mathbf{C_5H_8O_4NNa}$.

3.60 The balanced equations are as follows:

(a) $2N_2O_5 \rightarrow 2N_2O_4 + O_2$

(b) $2KNO_3 \rightarrow 2KNO_2 + O_2$

(c) $NH_4NO_3 \rightarrow N_2O + 2H_2O$

(d) $NH_4NO_2 \rightarrow N_2 + 2H_2O$

(e) $2NaHCO_3 \rightarrow Na_2CO_3 + H_2O + CO_2$

(f) $P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$

(g) $2HCl + CaCO_3 \rightarrow CaCl_2 + H_2O + CO_2$

(h) $2Al + 3H_2SO_4 \rightarrow Al_2(SO_4)_3 + 3H_2$

(i) $CO_2 + 2KOH \rightarrow K_2CO_3 + H_2O$

(j) $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

(k) $Be_2C + 4H_2O \rightarrow 2Be(OH)_2 + CH_4$

(l) $3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$

(m) $S + 6HNO_3 \rightarrow H_2SO_4 + 6NO_2 + 2H_2O$

(n) $2NH_3 + 3CuO \rightarrow 3Cu + N_2 + 3H_2O$

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,

$$6A + 4B \rightarrow 4C + 2D$$

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

$$\mathbf{3A + 2B \rightarrow 2C + D}$$

The correct answer is choice **(d)**.

3.66 Calculating the Amounts of Reactants and Products, Problem Type 7.

$$Si(s) + 2Cl_2(g) \longrightarrow SiCl_4(l)$$

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 $Cl_2 \simeq$ 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:

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

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

3.68 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.

$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(l)$

$$?\ \text{mol CO}_2 = 5.0\ \cancel{\text{mol}}\ \text{C}_4\text{H}_{10} \times \frac{8\ \text{mol CO}_2}{2\ \cancel{\text{mol}}\ \text{C}_4\text{H}_{10}} = 20\ \text{mol CO}_2 = \mathbf{2.0 \times 10^1\ mol\ CO_2}$$

3.70 **(a)** $2NaHCO_3 \longrightarrow Na_2CO_3 + H_2O + CO_2$

(b) Molar mass $NaHCO_3$ = 22.99 g + 1.008 g + 12.01 g + 3(16.00 g) = 84.01 g
Molar mass CO_2 = 12.01 g + 2(16.00 g) = 44.01 g

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

$$\textbf{mass NaHCO}_3 = 20.5\ \cancel{\text{g}}\ \text{CO}_2 \times \frac{1\ \cancel{\text{mol}}\ \text{CO}_2}{44.01\ \cancel{\text{g}}\ \text{CO}_2} \times \frac{2\ \cancel{\text{mol}}\ \text{NaHCO}_3}{1\ \cancel{\text{mol}}\ \text{CO}_2} \times \frac{84.01\ \text{g NaHCO}_3}{1\ \cancel{\text{mol}}\ \text{NaHCO}_3}$$

$$= \textbf{78.3 g NaHCO}_3$$

3.72 Calculating the Amounts of Reactants and Products, Problem Type 7.

$$\underset{\text{glucose}}{C_6H_{12}O_6} \longrightarrow \underset{\text{ethanol}}{2C_2H_5OH} + 2CO_2$$

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 g) + 12(1.008 g) + 6(16.00 g) = 180.16 g. The balanced equation is given, so the mole ratio between glucose and ethanol is known; that is 1 mole glucose ≏ 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 → moles of glucose → moles of ethanol → grams of ethanol

$$\textbf{? g C}_2\textbf{H}_5\textbf{OH} = 500.4\ \cancel{\text{g}}\ \text{C}_6\text{H}_{12}\text{O}_6 \times \frac{1\ \cancel{\text{mol}}\ \text{C}_6\text{H}_{12}\text{O}_6}{180.16\ \cancel{\text{g}}\ \text{C}_6\text{H}_{12}\text{O}_6} \times \frac{2\ \cancel{\text{mol}}\ \text{C}_2\text{H}_5\text{OH}}{1\ \cancel{\text{mol}}\ \text{C}_6\text{H}_{12}\text{O}_6} \times \frac{46.07\ \text{g C}_2\text{H}_5\text{OH}}{1\ \cancel{\text{mol}}\ \text{C}_2\text{H}_5\text{OH}}$$

$$= \textbf{255.9 g C}_2\textbf{H}_5\textbf{OH}$$

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.

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

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

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

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

3.76 (a) $NH_4NO_3(s) \longrightarrow N_2O(g) + 2H_2O(g)$

(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:

$$\text{mol } NH_4NO_3 \rightarrow \text{mol } N_2O \rightarrow \text{g } N_2O$$

$$\textbf{? g N}_\mathbf{2}\textbf{O} = 0.46\ \cancel{mol}\ NH_4NO_3 \times \frac{1\ \cancel{mol}\ N_2O}{1\ \cancel{mol}\ NH_4NO_3} \times \frac{44.02\ \text{g } N_2O}{1\ \cancel{mol}\ N_2O} = \mathbf{2.0 \times 10^1\ g\ N_2O}$$

3.78 The balanced equation for the decomposition is :

$$2KClO_3(s) \longrightarrow 2KCl(s) + 3O_2(g)$$

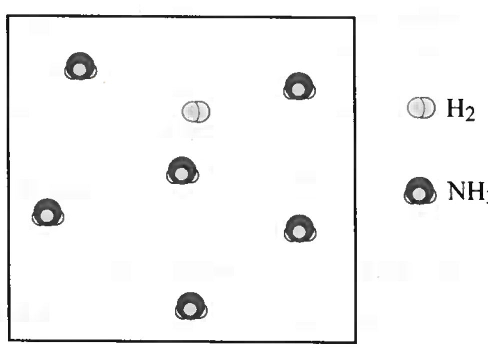
$$\textbf{? g O}_\mathbf{2} = 46.0\ \cancel{g}\ KClO_3 \times \frac{1\ \cancel{mol}\ KClO_3}{122.55\ \cancel{g}\ KClO_3} \times \frac{3\ \cancel{mol}\ O_2}{2\ \cancel{mol}\ KClO_3} \times \frac{32.00\ \text{g } O_2}{1\ \cancel{mol}\ O_2} = \mathbf{18.0\ g\ O_2}$$

3.82 $N_2 + 3H_2 \rightarrow 2NH_3$

(a) The number of N_2 molecules shown in the diagram is 3. The balanced equation shows 3 moles $H_2 \simeq 1$ 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.



3.84 Limiting Reagent Calculation, Problem Type 8.

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

$$? \text{ mol } NO_2 = 0.740 \cancel{\text{g}}\, O_3 \times \frac{1 \cancel{\text{mol}}\, O_3}{48.00 \cancel{\text{g}}\, O_3} \times \frac{1 \text{ mol } NO_2}{1 \cancel{\text{mol}}\, O_3} = 0.0154 \text{ mol } NO_2$$

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

$$\text{grams of NO} \rightarrow \text{moles of NO} \rightarrow \text{moles of } NO_2$$

Combining these two conversions into one calculation, we write

$$? \text{ mol } NO_2 = 0.670 \cancel{\text{g}}\, \text{NO} \times \frac{1 \cancel{\text{mol}}\, \text{NO}}{30.01 \cancel{\text{g}}\, \text{NO}} \times \frac{1 \text{ mol } NO_2}{1 \cancel{\text{mol}}\, \text{NO}} = 0.0223 \text{ mol } NO_2$$

The initial amount of $\mathbf{O_3}$ limits the amount of product that can be formed; therefore, it is the **limiting reagent**.

The problem asks for grams of NO_2 produced. We already know the moles of NO_2 produced, 0.0154 mole. Use the molar mass of NO_2 as a conversion factor to convert to grams (Molar mass NO_2 = 46.01 g).

$$\mathbf{? \text{ g } NO_2} = 0.0154 \cancel{\text{mol}}\, NO_2 \times \frac{46.01 \text{ g } NO_2}{1 \cancel{\text{mol}}\, NO_2} = \mathbf{0.709 \text{ g } NO_2}$$

Check: Does your answer seem reasonable? 0.0154 mole of product is formed. What is the mass of 1 mole of NO_2?

Strategy: Working backwards, we can determine the amount of NO that reacted to produce 0.0154 mole of NO_2. The amount of NO left over is the difference between the initial amount and the amount reacted.

Solution: Starting with 0.0154 mole of NO_2, 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.

$$\text{mol NO reacted} = 0.0154 \text{ mol } NO_2 \times \frac{1 \text{ mol NO}}{1 \text{ mol } NO_2} = 0.0154 \text{ mol NO}$$

$$\text{mol NO initial} = 0.670 \text{ g NO} \times \frac{1 \text{ mol NO}}{30.01 \text{ g NO}} = 0.0223 \text{ mol NO}$$

mol NO remaining = mol NO initial – mol NO reacted.

mol NO remaining = 0.0223 mol NO – 0.0154 mol NO = **0.0069 mol NO**

3.86 This is a limiting reagent problem. Let's calculate the moles of Cl_2 produced assuming complete reaction for each reactant.

$$0.86\ \cancel{mol}\ MnO_2 \times \frac{1\ mol\ Cl_2}{1\ \cancel{mol}\ MnO_2} = 0.86\ mol\ Cl_2$$

$$48.2\ \cancel{g}\ HCl \times \frac{1\ \cancel{mol}\ HCl}{36.458\ \cancel{g}\ HCl} \times \frac{1\ mol\ Cl_2}{4\ \cancel{mol}\ HCl} = 0.330\ mol\ Cl_2$$

HCl is the limiting reagent; it limits the amount of product produced. It will be used up first. The amount of product produced is 0.330 mole Cl_2. Let's convert this to grams.

$$?\ \mathbf{g\ Cl_2} = 0.330\ \cancel{mol}\ Cl_2 \times \frac{70.90\ g\ Cl_2}{1\ \cancel{mol}\ Cl_2} = \mathbf{23.4\ g\ Cl_2}$$

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.

$$4C_3H_5N_3O_9 \longrightarrow 6N_2 + 12CO_2 + 10H_2O + O_2$$

Map out the following strategy to solve this problem.

$$g\ NG \rightarrow mol\ NG \rightarrow mol\ O_2 \rightarrow g\ O_2$$

Calculate the grams of O_2 using the strategy above.

$$?\ \mathbf{g\ O_2} = 2.00 \times 10^2\ \cancel{g}\ NG \times \frac{1\ \cancel{mol}\ NG}{227.1\ \cancel{g}\ NG} \times \frac{1\ \cancel{mol}\ O_2}{4\ \cancel{mol}\ NG} \times \frac{32.00\ g\ O_2}{1\ \cancel{mol}\ O_2} = \mathbf{7.05\ g\ O_2}$$

(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:

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

$$\mathbf{\%\ yield} = \frac{6.55\ g\ O_2}{7.05\ g\ O_2} \times 100\% = \mathbf{92.9\%}$$

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.

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

$$42.5\%\ yield = \frac{481\ g\ C_2H_4}{theoretical\ yield} \times 100\%$$

$$theoretical\ yield\ C_2H_4 = 1.13 \times 10^3\ g\ C_2H_4$$

The mass of hexane that must be reacted is:

$$(1.13 \times 10^3\ \cancel{g}\ C_2H_4) \times \frac{1\ \cancel{mol}\ C_2H_4}{28.05\ \cancel{g}\ C_2H_4} \times \frac{1\ \cancel{mol}\ C_6H_{14}}{1\ \cancel{mol}\ C_2H_4} \times \frac{86.17\ g\ C_6H_{14}}{1\ \cancel{mol}\ C_6H_{14}} = \mathbf{3.47 \times 10^3\ g\ C_6H_{14}}$$

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

$S_8(l) + 4Cl_2(g) \rightarrow 4S_2Cl_2(l)$

$$4.06\ \cancel{g}\ S_8 \times \frac{1\ \cancel{mol}\ S_8}{256.6\ \cancel{g}\ S_8} \times \frac{4\ mol\ S_2Cl_2}{1\ \cancel{mol}\ S_8} = 0.0633\ mol\ S_2Cl_2$$

$$6.24\ \cancel{g}\ Cl_2 \times \frac{1\ \cancel{mol}\ Cl_2}{70.90\ \cancel{g}\ Cl_2} \times \frac{4\ mol\ S_2Cl_2}{4\ \cancel{mol}\ Cl_2} = 0.0880\ mol\ S_2Cl_2$$

S_8 is the limiting reagent; it limits the amount of product produced. The amount of product produced is 0.0633 mole S_2Cl_2. Let's convert this to grams.

$$?\ g\ S_2Cl_2 = 0.0633\ \cancel{mol}\ S_2Cl_2 \times \frac{135.04\ g\ S_2Cl_2}{1\ \cancel{mol}\ S_2Cl_2} = \mathbf{8.55\ g\ S_2Cl_2}$$

This is the theoretical yield of S_2Cl_2. The actual yield is given in the problem (6.55 g). The percent yield is:

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

3.96 $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$

We start with 8 molecules of H_2 and 3 molecules of O_2. The balanced equation shows 2 moles $H_2 \simeq$ 1 mole O_2. If 3 molecules of O_2 react, 6 molecules of H_2 will react, leaving 2 molecules of H_2 in excess. The balanced equation also shows 1 mole $O_2 \simeq$ 2 moles H_2O. If 3 molecules of O_2 react, 6 molecules of H_2O will be produced.

After complete reaction, there will be **2 molecules of H_2** and **6 molecules of H_2O**. The correct diagram is choice **(b)**.

3.98 We assume that all the Cl in the compound ends up as HCl and all the O ends up as H_2O. Therefore, we need to find the number of moles of Cl in HCl and the number of moles of O in H_2O.

$$mol\ Cl = 0.233\ \cancel{g}\ HCl \times \frac{1\ \cancel{mol}\ HCl}{36.46\ \cancel{g}\ HCl} \times \frac{1\ mol\ Cl}{1\ \cancel{mol}\ HCl} = 0.00639\ mol\ Cl$$

$$mol\ O = 0.403\ \cancel{g}\ H_2O \times \frac{1\ \cancel{mol}\ H_2O}{18.02\ \cancel{g}\ H_2O} \times \frac{1\ mol\ O}{1\ \cancel{mol}\ H_2O} = 0.0224\ mol\ O$$

Dividing by the smallest number of moles (0.00639 mole) gives the formula, $ClO_{3.5}$. Multiplying both subscripts by two gives the empirical formula, $\mathbf{Cl_2O_7}$.

3.100 The symbol "O" refers to moles of oxygen atoms, not oxygen molecule (O_2). 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) $0.212\ \cancel{mol}\ C \times \frac{1\ mol\ O}{1\ \cancel{mol}\ C} = \mathbf{0.212\ mol\ O}$

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

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; $Al_2(SO_4)_3$, 342.17 g/mol; H_2O, 18.016 g/mol. Thus, using x as the number of H_2O molecules,

$$\text{mass \% Al} = \left(\frac{2(\text{molar mass of Al})}{\text{molar mass of } Al_2(SO_4)_3 + x(\text{molar mass of } H_2O)}\right) \times 100\%$$

$$8.20\% = \left(\frac{2(26.98\ \text{g})}{342.2\ \text{g} + x(18.02\ \text{g})}\right) \times 100\%$$

$$\boldsymbol{x = 17.53}$$

Rounding off to a whole number of water molecules, $\boldsymbol{x}$ **= 18**. Therefore, the formula is $\mathbf{Al_2(SO_4)_3 \cdot 18\ H_2O}$.

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

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

3.106 The mass of oxygen in MO is 39.46 g – 31.70 g = 7.76 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.

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

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

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

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 :

$$(2.62 \times 10^3\ \cancel{\text{kg}}\ Fe_2O_3) \times \frac{1000\ \cancel{\text{g}}\ Fe_2O_3}{1\ \cancel{\text{kg}}\ Fe_2O_3} \times \frac{1\ \cancel{\text{mol}}\ Fe_2O_3}{159.7\ \cancel{\text{g}}\ Fe_2O_3} \times \frac{2\ \cancel{\text{mol}}\ \text{Fe}}{1\ \cancel{\text{mol}}\ Fe_2O_3} \times \frac{55.85\ \cancel{\text{g}}\ \text{Fe}}{1\ \cancel{\text{mol}}\ \text{Fe}} \times \frac{1\ \text{kg Fe}}{1000\ \cancel{\text{g}}\ \text{Fe}}$$
$$= 1.83 \times 10^3\ \text{kg Fe}$$

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

$$\textbf{percent yield} = \frac{1.64 \times 10^3\ \text{kg Fe}}{1.83 \times 10^3\ \text{kg Fe}} \times 100\% = \mathbf{89.6\%} = \mathbf{purity\ of\ Fe_2O_3}$$

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 H_2O.

Assume 100 g of compound.

$$40.0\ \cancel{g}\ C \times \frac{1\ \text{mol C}}{12.01\ \cancel{g}\ C} = 3.33\ \text{mol C}$$

$$60.0\ \cancel{g}\ H_2O \times \frac{1\ \text{mol}\ H_2O}{18.016\ \cancel{g}\ H_2O} = 3.33\ \text{mol}\ H_2O$$

Dividing by 3.33 gives $\mathbf{CH_2O}$ for the empirical formula.

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

$$\frac{\text{molar mass}}{\text{empirical mass}} = \frac{178\ \text{g}}{30.03\ \text{g}} \approx 6$$

Thus, there are six CH_2O units in each molecule of the compound, so the molecular formula is $(CH_2O)_6$, or $\mathbf{C_6H_{12}O_6}$.

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.

$$67.2\ \cancel{g}\ Cl \times \frac{1\ \text{mol Cl}}{35.45\ \cancel{g}\ Cl} = 1.90\ \text{mol Cl}$$

Then, using the mole ratio from the chemical formula (XCl_3), we can calculate the moles of X contained in 32.8 g.

$$1.90\ \cancel{\text{mol}}\ Cl \times \frac{1\ \text{mol X}}{3\ \cancel{\text{mol}}\ Cl} = 0.633\ \text{mol X}$$

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

$$\frac{32.8\ \text{g X}}{0.633\ \text{mol X}} = \mathbf{51.8\ g/mol}$$

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

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

$$0.34\ \cancel{g}\ Fe \times \frac{1\ \text{mol Fe}}{55.85\ \cancel{g}\ Fe} = 6.1 \times 10^{-3}\ \text{mol Fe}$$

Since there is one Fe atom in a molecule of myoglobin, the moles of myoglobin also equal 6.1×10^{-3} mole. The molar mass of myoglobin can be calculated.

$$\textbf{molar mass myoglobin} = \frac{100\ \text{g myoglobin}}{6.1 \times 10^{-3}\ \text{mol myoglobin}} = \mathbf{1.6 \times 10^4\ g/mol}$$

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.

$$29.96\ \text{g Na} = \text{mass of Na in NaBr} + \text{mass of Na in}\ Na_2SO_4$$

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.

$$29.96 \text{ g Na} = \text{mass of Na in NaBr} + \text{mass of Na in } Na_2SO_4$$

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

$$29.96 = 0.2234x + 32.37 - 0.3237x$$

$$0.1003x = 2.41$$

$x = 24.0$ g, which equals the mass of NaBr.

The mass of Na_2SO_4 is $100 - x$ which equals 76.0 g.

Because we assumed 100 g of compound, the mass % of NaBr in the mixture is **24.0%** and the mass % of Na_2SO_4 is **76.0%.**

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 $Ca_3(PO_4)_2$ = 310.18 g/mol

$$\textbf{\% Ca} = \frac{(3)(40.08 \cancel{\text{g}})}{310.2 \cancel{\text{g}}} \times 100\% = \textbf{38.76\% Ca}$$

$$\textbf{\% P} = \frac{(2)(30.97 \cancel{\text{g}})}{310.2 \cancel{\text{g}}} \times 100\% = \textbf{19.97\% P}$$

$$\textbf{\% O} = \frac{(8)(16.00 \cancel{\text{g}})}{310.2 \cancel{\text{g}}} \times 100\% = \textbf{41.26\% O}$$

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?

3.122 Since we assume that water exists as either H_2O or D_2O, the natural abundances are 99.985 percent and 0.015 percent, respectively. If we convert to molecules of water (both H_2O or D_2O), we can calculate the molecules that are D_2O from the natural abundance (0.015%).

The necessary conversions are:

mL water $\rightarrow$ g water $\rightarrow$ mol water $\rightarrow$ molecules water $\rightarrow$ molecules D_2O

$$400 \cancel{\text{mL}} \text{ water} \times \frac{1 \cancel{\text{g}} \text{ water}}{1 \cancel{\text{mL}} \text{ water}} \times \frac{1 \cancel{\text{mol}} \text{ water}}{18.02 \cancel{\text{g}} \text{ water}} \times \frac{6.022 \times 10^{23} \cancel{\text{molecules}}}{1 \cancel{\text{mol}} \text{ water}} \times \frac{0.015\% \text{ molecules } D_2O}{100\% \cancel{\text{molecules}} \text{ water}}$$

$$= \mathbf{2.01 \times 10^{21}} \textbf{ molecules } \mathbf{D_2O}$$

3.124 First, we can calculate the moles of oxygen.

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

Next, we can calculate the molar mass of oxygen.

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

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**.

3.126 **(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 $NaClO_3$.

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

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

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

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

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

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

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

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

$$0.50\ \cancel{mol}\ Cl_2 \times \frac{70.90\ g\ Cl}{1\ \cancel{mol}\ Cl_2} = \mathbf{35.45\ g\ Cl}$$

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

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.

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

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

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

Compound B: Assume 100 g of compound.

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

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

Dividing by the smallest number of moles (0.297 mole) gives the empirical formula, $\mathbf{PtCl_4}$.

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.

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

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

Dividing by the smallest number of moles (1.15 moles) gives the empirical formula, $\mathbf{MnO_2}$.

Compound Y: Assume 100 g of compound.

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

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

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 $\mathbf{Mn_3O_4}$.

(b) The unbalanced equation is: $MnO_2 \longrightarrow Mn_3O_4 + O_2$

Balancing by inspection gives: $\mathbf{3MnO_2 \longrightarrow Mn_3O_4 + O_2}$

3.132 SO_2 is converted to H_2SO_4 by reaction with water. The mole ratio between SO_2 and H_2SO_4 is 1:1.

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

tons SO_2 → ton-mol SO_2 → ton-mol H_2SO_4 → tons H_2SO_4

$$\textbf{? tons } \mathbf{H_2SO_4} = (4.0 \times 10^5\ \cancel{\text{tons}}\ SO_2) \times \frac{1\ \cancel{\text{ton-mol}}\ SO_2}{64.07\ \cancel{\text{tons}}\ SO_2} \times \frac{1\ \cancel{\text{ton-mol}}\ H_2SO_4}{1\ \cancel{\text{ton-mol}}\ SO_2} \times \frac{98.09 \text{ tons } H_2SO_4}{1\ \cancel{\text{ton-mol}}\ H_2SO_4}$$

$$= \mathbf{6.1 \times 10^5 \textbf{ tons } H_2SO_4}$$

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 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.

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.

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

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

Dividing by the smallest number of moles gives $Mg_{1.5}N$. Recall that an empirical formula must have whole number coefficients. Multiplying by a factor of 2 gives the empirical formula $\mathbf{Mg_3N_2}$. The name of this compound is **magnesium nitride**.

3.136 *Step 1:* Calculate the mass of C in 55.90 g CO_2, and the mass of H in 28.61 g H_2O. 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:

$$\text{g CO}_2 \rightarrow \text{mol CO}_2 \rightarrow \text{mol C} \rightarrow \text{g C}$$

Step 2:

$$?\text{ g C} = 55.90\,\cancel{\text{g}}\,\text{CO}_2 \times \frac{1\,\cancel{\text{mol}}\,\text{CO}_2}{44.01\,\cancel{\text{g}}\,\text{CO}_2} \times \frac{1\,\cancel{\text{mol}}\,\text{C}}{1\,\cancel{\text{mol}}\,\text{CO}_2} \times \frac{12.01\text{ g C}}{1\,\cancel{\text{mol}}\,\text{C}} = 15.25\text{ g C}$$

Similarly,

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

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.

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

Similarly,

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

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

Thus, we arrive at the formula $Pb_{0.1588}C_{1.270}H_{3.176}$, 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.

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

This gives the empirical formula, $\mathbf{PbC_8H_{20}}$.

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

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

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

(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:

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

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.

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

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:

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

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 3.1×10^2 cm^2.

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

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

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

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

3.144 **(a)** The balanced chemical equation is:

$$C_3H_8(g) + 3H_2O(g) \longrightarrow 3CO(g) + 7H_2(g)$$

(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.

$$\text{kg } C_3H_8 \rightarrow \text{mol } C_3H_8 \rightarrow \text{mol } H_2 \rightarrow \text{kg } H_2$$

$$\textbf{? kg } \mathbf{H_2} = (2.84 \times 10^3\ \cancel{\text{kg}}\ C_3H_8) \times \frac{1\ \cancel{\text{kg-mol}}\ C_3H_8}{44.09\ \cancel{\text{kg}}\ C_3H_8} \times \frac{7\ \cancel{\text{kg-mol}}\ H_2}{1\ \cancel{\text{kg-mol}}\ C_3H_8} \times \frac{2.016 \text{ kg } H_2}{1\ \cancel{\text{kg-mol}}\ H_2}$$

$$= \mathbf{9.09 \times 10^2\ kg\ H_2}$$

3.146 **(a)** 16 amu, CH_4 17 amu, NH_3 18 amu, H_2O 64 amu, SO_2

(b) The formula C_3H_8 can also be written as $CH_3CH_2CH_3$. A CH_3 fragment could break off from this molecule giving a peak at 15 amu. No fragment of CO_2 can have a mass of 15 amu. Therefore, the substance responsible for the mass spectrum is most likely C_3H_8.

(c) First, let's calculate the masses of CO_2 and C_3H_8.

$$\text{molecular mass } CO_2 = 12.00000 \text{ amu} + 2(15.99491 \text{ amu}) = 43.98982 \text{ amu}$$

$$\text{molecular mass } C_3H_8 = 3(12.00000 \text{ amu}) + 8(1.00797 \text{ amu}) = 44.06376 \text{ amu}$$

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

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

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

3.148 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.

$$Mg_3N_2(s) + 6H_2O(l) \rightarrow 3Mg(OH)_2(s) + 2NH_3(g)$$

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.

$$2.813\ \cancel{g}\ NH_3 \times \frac{1\ \cancel{mol}\ NH_3}{17.03\ \cancel{g}\ NH_3} \times \frac{1\ \cancel{mol}\ Mg_3N_2}{2\ \cancel{mol}\ NH_3} \times \frac{100.95\ g\ Mg_3N_2}{1\ \cancel{mol}\ Mg_3N_2} = \mathbf{8.337\ g\ Mg_3N_2}$$

The mass of Mg in 8.337 g Mg_3N_2 can be determined from the mass percentage of Mg in Mg_3N_2.

$$\frac{(3)(24.31\ g\ Mg)}{100.95\ \cancel{g}\ Mg_3N_2} \times 8.337\ \cancel{g}\ Mg_3N_2 = 6.023\ g\ Mg$$

The mass of Mg in the product MgO is obtained by difference: 21.496 g Mg − 6.023 g Mg = 15.473 g Mg

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

$$\frac{40.31\ g\ MgO}{24.31\ \cancel{g}\ Mg} \times 15.473\ \cancel{g}\ Mg = \mathbf{25.66\ g\ MgO}$$

3.150 The decomposition of $KClO_3$ produces oxygen gas (O_2) which reacts with Fe to produce Fe_2O_3.

$$4Fe + 3O_2 \rightarrow 2Fe_2O_3$$

When the 15.0 g of Fe is heated in the presence of O_2 gas, any increase in mass is due to oxygen. The mass of oxygen reacted is:

$$17.9\ g - 15.0\ g = 2.9\ g\ O_2$$

From this mass of O_2, we can now calculate the mass of Fe_2O_3 produced and the mass of $KClO_3$ decomposed.

$$2.9\ \cancel{g}\ O_2 \times \frac{1\ \cancel{mol}\ O_2}{32.00\ \cancel{g}\ O_2} \times \frac{2\ \cancel{mol}\ Fe_2O_3}{3\ \cancel{mol}\ O_2} \times \frac{159.7\ g\ Fe_2O_3}{1\ \cancel{mol}\ Fe_2O_3} = \mathbf{9.6\ g\ Fe_2O_3}$$

The balanced equation for the decomposition of $KClO_3$ is: $2KClO_3 \rightarrow 2KCl + 3O_2$. The mass of $KClO_3$ decomposed is:

$$2.9\ \cancel{g}\ O_2 \times \frac{1\ \cancel{mol}\ O_2}{32.00\ \cancel{g}\ O_2} \times \frac{2\ \cancel{mol}\ KClO_3}{3\ \cancel{mol}\ O_2} \times \frac{122.55\ g\ KClO_3}{1\ \cancel{mol}\ KClO_3} = \mathbf{7.4\ g\ KClO_3}$$