CHAPTER 3
MASS RELATIONSHIPS IN CHEMICAL REACTIONS

This chapter reviews the mole concept, balancing chemical equations, and stoichiometry. The topics covered in this chapter are:

- Atomic mass and average atomic mass
- Avogadro's number, mole, and molar mass
- Percent composition calculations
- Empirical and molecular formula determinations
- Chemical equations, amount of reactant and product calculations
- Limiting reagents and reaction yield calculations

> **Take Note:** *It is absolutely essential that you master the mole concept to do well on the quantitative aspects of AP Chemistry!!*
> *When solving quantitative problems on the Free Response section of the AP exam, supporting work must be shown to receive credit. Using dimensional analysis is a very powerful technique in solving problems.*
> *Be sure to report your answer to the correct number of significant figures (see Chapter 1 in this review book).*

Atomic mass and average atomic mass

Atomic mass is the *mass of an atom in atomic mass units (amu)*. **One amu** is defined as *1/12 of one C-12 atom*. The C-12 isotope has a mass of exactly 12.000 amu. The C-12 isotope provides the relative scale for the masses of the other elements.

Average atomic mass is the *value reported on the periodic table, which takes into account the various isotopes of an element and their respective frequencies.* To calculate the average atomic mass of an element, add up the different masses of the isotopes (using amu) multiplied by each isotope's abundance (percent occurrence in nature divided by 100).

Average atomic mass of element x = (isotope$_1$ in amu × abundance$_1$) + (isotope$_2$ × abundance$_2$) + (isotope$_n$ × abundance$_n$)

Example 1. Determining average atomic mass.

The natural abundance of the C-12 isotope in nature is 98.90% and the C-13 isotope is 1.100%. What is the average atomic mass reported on the periodic table for the element carbon? (The atomic masses of C-12 and C-13 are 12.00000 amu and 13.00335 amu, respectively.)

Solution to Example 1.

(12.00 amu × 0.9890) + (13.00335 amu × 0.0110) = *12.01 = average atomic mass of carbon*

Molecular mass refers to the *mass of a molecule expressed in atomic mass units*. The molecular mass is the sum of the atomic masses of the atoms in the molecular formula.

Example 2. Determining molecular mass.

Determine the molecular mass of barium hydroxide, $Ba(OH)_2$.

Solution to Example 2.

1 Ba atom + 2 O atoms + 2 H atoms
137.33 amu + 2(16.00 amu) + 2(1.0079 amu) = *171.34 amu*

Avogadro's number, mole, and molar mass

Avogadro's number refers to *6.022×10^{23} particles*. The *quantity of an element that contains Avogadro's number of particles* is called a **mole**. The **molar mass** of an element or compound is the *mass of one mole of its atoms or molecules expressed in grams*.

Some examples of mole, atom, ion, molecule, and mass equivalencies are:

- 1 mole of C atoms = 6.022×10^{23} atoms of C = 12.01 g of C
- 1 mole of O_2 molecules = 6.022×10^{23} molecules of O_2 = 32.00 g O_2
- 1 mole of of O_2 *molecules* = 2 moles of oxygen *atom*s = 12.044×10^{23} atoms of O
- 1 mole O^{2-} ions = 6.022×10^{23} O^{2-} ions = 16.00 g O^{2-} ions

Percent composition calculations

The **percent composition** of an element in a compound =

$$\frac{\boldsymbol{n \times \text{molar mass of element}}}{\textbf{molar mass of compound}} \times \mathbf{100\%} \quad \text{where}$$

n is the number of moles of the element in the compound

Example 3. Determining the percent composition.

Determine the percent composition of chloride in barium chloride, $BaCl_2$.

$$\frac{\text{Mass of } 2\ Cl^-}{\text{Mass of } BaCl_2} \times 100 = \frac{2(35.45 \text{ g } Cl^-)}{2(35.45 \text{ g } Cl^-) + 137.33 \text{ g Ba}} \times 100\% = \mathit{34.0\ \%\ Cl^-}$$

Empirical and molecular formula determinations

The **empirical formula** is the *most reduced form of the molecular formula*. For instance, sugar, which has the molecular formula of $C_6H_{12}O_6$, has an empirical formula of CH_2O. Empirical formulas can be determined from percent composition data. If the molar mass is also known, the molecular formula can be determined as well (see Example 4).

Example 4. Determining empirical and molecular formulas.

An unknown compound is known to contain 30.43% N and 60.56% O and has a molar mass of 92.00 g. What are the empirical formula and the molecular formula for the compound?

General Strategy	**Solution to Example 4**
Assume a 100 g sample.	30.43 g of N 60.56 g O
Convert gram quantities to moles by dividing by the molar mass of the atom.	$\frac{30.43 \text{ g N}}{14.01 \text{ g N/mol N}} = 2.17 \text{ mol N}$ $\frac{60.56 \text{ g O}}{16.00 \text{ g O/ mol O}} = 4.35 \text{ mol}$
Divide by the smallest number of moles to obtain whole number ratios. The whole number ratios are the subscripts in the empirical formula.	$\frac{2.17 \text{ mol N}}{2.17 \text{ mol N}} = 1 \text{ mol N}$ $\frac{4.35 \text{ mol O}}{2.17 \text{ mol N}} = 2 \text{ mol}$ *Empirical formula* = $\mathit{NO_2}$
To determine the molecular formula divide the molar mass by the empirical mass. Multiply the subscripts of the empirical formula by this factor.	Empirical mass of NO_2 = 14.01 g N + 2(16.00) g O = 46.01 g NO_2 $\frac{\text{Molar mass}}{\text{Empirical mass}} = \frac{92.00 \text{ g}}{46.01 \text{ g}} = 2$ Molecular formula is twice the empirical formula: *Molecular formula* = $\mathit{N_2O_4}$

Chemical equations, amount of reactant and product calculations

Chemical equations are a shorthand method for *representing chemical reactions*. *Starting materials* are called **reactants** and are on the left side of the equation. The arrow indicates a reaction has taken place in the forward direction. A double-ended arrow `*substances formed in a chemical reaction*, are on the right side of the equation. The numbers in front of the reactant and product molecules or atoms are called **coefficients**. They refer to the *molar relationships between substances* in a chemical reaction. The coefficients are used to balance the *atoms* in an equation. Atoms and mass are conserved in a reaction, molecules may not be. See Table 3.1.

TABLE 3.1 Interpretation of a Chemical Equation

$2H_2$	$+ O_2$	$\longrightarrow 2H_2O$
Two molecules	+ one molecule	$\longrightarrow$ two molecules
2 moles	+ 1 mole	$\longrightarrow$ 2 moles
$\underbrace{2(2.02\text{ g}) = 4.04\text{ g} + 32.00\text{ g}}_{36.04\text{ g reactants}}$		$\longrightarrow \underbrace{2(18.02\text{ g}) = 36.04\text{ g}}_{36.04\text{ g product}}$

When balancing a chemical equation by inspection, it is usually best to balance the elements of hydrogen and oxygen last. Subscripts in molecules *cannot* be changed to balance a chemical equation. Coefficients are used to balance the equation.

Example 5. Balancing a chemical equation.

Consider the reaction: $C_5H_{12}\,(l) + 8O_2\,(g) \rightarrow 5CO_2\,(g) + 6H_2O\,(g)$
State what quantities are conserved and which are not conserved.

Solution to Example 5.
There are 5 atoms of C, 12 atoms of H, and 16 atoms of O on each side of the equation. The atoms are conserved as the reaction proceeds forward. The mass of reactants is equal to the mass of products. Mass is conserved. There are 9 reactant molecules and 11 product molecules. The number of molecules is not conserved.

The balanced chemical equation can be used to calculate the amount of reactants and products used or produced in a reaction. Examples 6, 7, and 8 represent some common types of stoichiometry problems. Example 6 describes a general strategy that can be applied when solving stoichiometry problems.

Example 6. Stoichiometry problem determining amount of product produced.

Pentane is burned in an excess of oxygen to produce carbon dioxide and water. If 144 g of pentane are burned, how many grams of carbon dioxide are produced?

General Strategy	**Solution to Example 6**
Write and balance the equation.	$C_5H_{12}\ (l) + 8O_2\ (g) \rightarrow 5CO_2\ (g) + 6H_2O\ (g)$
Place the data from the problem underneath the balanced equation. Identify the quantity to be calculated (the *desired* molecule, CO_2 in this problem).	$C_5H_{12}\ (l) + 8O_2\ (g) \rightarrow 5CO_2\ (g) + 6H_2O\ (g)$ 144 g excess <u>?g</u>
Convert the *given* quantity (C_5H_{12} in this problem) to moles using dimensional analysis.	$144\,g\,C_5H_{12} \times \frac{\mathbf{1\,mol\,C_5H_{12}}}{\mathbf{72\,g\,C_5H_{12}}} = \text{moles of } \textit{given} \text{ molecule } C_5H_{12}$
Use the mole ratio from the balanced equation (i.e., the coefficients) to convert the *given* molecule (C_5H_{12}) to the *desired* molecule (CO_2).	$144\,g\,C_5H_{12} \times \frac{1\,mol\,C_5H_{12}}{72\,g\,C_5H_{12}} \times \frac{\mathbf{5\,mol\,CO_2}}{\mathbf{1\,mol\,C_5H_{12}}} =$ moles of *desired* molecule, CO_2
Convert the *desired* molecule to the appropriate units (grams of CO_2).	$144\,g\,C_5H_{12} \times \frac{1\,mol\,C_5H_{12}}{72\,g\,C_5H_{12}} \times \frac{5\,mol\,CO_2}{1\,mol\,C_5H_{12}} \times \frac{\mathbf{44\,g\,CO_2}}{\mathbf{1\,mol\,CO_2}}$ $= 110.\,g\,CO_2$ *produced*

Example 7. Stoichiometry problem determining the amount of reactant consumed.

The Haber process is a reaction in which ammonia is produced by combining elemental hydrogen and nitrogen gas. How many grams of nitrogen are required to completely react with 6.0 grams of hydrogen?

General Strategy	Solution to Example 7
Write and balance the equation.	$3H_2\,(g) + N_2\,(g) \Leftrightarrow 2NH_3\,(g)$
Place the data from the problem underneath the balanced equation. Identify the quantity to be calculated (the *desired* molecule, N_2 in this problem).	$3H_2\,(g) + N_2\,(g) \Leftrightarrow 2NH_3\,(g)$ 6.0 g ?g
Convert the *given* quantity (H_2 in this problem) to moles using dimensional analysis.	$6.0\,g\,H_2 \times \frac{\mathbf{1\,mol\,H_2}}{\mathbf{2.0\,g\,H_2}} = \text{moles of } \textit{given} \text{ molecule, } H_2$
Use the mole ratio from the balanced equation (i.e., the coefficients) to convert the *given* molecule (H_2) to the *desired* molecule (N_2).	$6.0\,g\,H_2 \times \frac{1\,mol\,H_2}{2.0\,g\,H_2} \times \frac{\mathbf{1\,mol\,N_2}}{\mathbf{3\,mol\,H_2}}$ moles of *desired* molecule, N_2
Convert *desired* molecule to the appropriate units (grams of N_2).	$6.0\,g\,H_2 \times \frac{1\,mol\,H_2}{2.0\,g\,H_2} \times \frac{1\,mol\,N_2}{3\,mol\,H_2} \times \frac{\mathbf{28.0\,g\,N_2}}{\mathbf{1\,mol\,N_2}}$ $= 28\,g\,N_2$ *required*

> **Take Note:** *When doing stoichiometry problems on the Free Response section of the AP exam, you must show supporting work describing the solution process to receive credit for the problem.*
> *It is critical that equations be written and balanced correctly. It is essential that you be familiar with polyatomic ions, correct oxidation numbers for all species, common organic molecules, and the diatomic elements. You must also be proficient at balancing equations.*

Limiting reagents and reaction yield calculations

Reactants in a chemical reaction may not be present in stoichiometric amounts. One of the reactants, known as the **limiting reactant,** is *consumed completely*; the other reactant, known as the **excess reactant,** is only *partially consumed* in the reaction. The *limiting reactant determines how much product is formed.*

For example, if 1 mol of iodine is reacted with 0.50 mol of hydrogen according to the following reaction, only 1.0 mol of hydrogen iodide is produced.

$I_2\,(g)$	+	$H_2\,(g)$	$\Leftrightarrow$	$2HI\,(g)$
1.0 mol		0.5 mol		
excess reactant		limiting reactant		

H_2 is the limiting reactant in this reaction. When all the H_2 is consumed, 0.50 mol of I_2 are left over. H_2 determines the amount of HI produced.

$$0.50\,\text{mol}\,\text{H}_2 \times \frac{2\,\text{mol}\,\text{HI}}{1\,\text{mol}\,\text{H}_2} = 1.0\,\text{mol}\,\text{HI}\,\text{produced}$$

When more than one reactant quantity is given in a problem, it is likely that one of the reactants will be consumed completely (the limiting reactant) while the other reactant is not (the excess reactant). To determine which reactant is limiting, determine the amount of mole product produced by each reactant. The reactant that produces the least amount of product is the limiting reactant.

Example 8. Stoichiometry problem involving limiting reactants.

When 56 g of silicon are combined with 35 g of chlorine gas in a reaction vessel:

a. How many moles of $SiCl_4$ are formed?
b. What is the limiting reactant?
c. How many moles of the excess reactant are left?

General Strategy	**Solution to Example 8**
Write and balance the equation.	Si (*s*) + $2Cl_2$ (*g*) → $SiCl_4$ (*l*)
Place the data from the problem underneath the balanced equation. Identify the quantity to be calculated (the *desired* molecule, $SiCl_4$ in this problem).	Si (*s*) + 2 Cl_2 (*g*) → $SiCl_4$ (*l*) 56 g 35 g ?g
Convert the *given* quantities to moles using dimensional analysis.	$56\,\text{g}\,\text{Si} \times \frac{1\,\text{mol}\,\text{Si}}{28.1\,\text{g}\,\text{Si}} = 2.0\,\text{mol}\,\text{Si}$ $35\,\text{g}\,\text{Cl}_2 \times \frac{1\,\text{mol}\,\text{Cl}_2}{71.0\,\text{g}\,\text{Cl}_2} = 0.50\,\text{mol}\,\text{Cl}_2$
Use the mole ratio from the balanced equation (i.e., the coefficients) to convert the *given* molecules to the *desired* product molecule and compare. *The reactant molecule that produces the least product molecule is the limiting reactant.*	$2.0\,\text{mol}\,\text{Si} \times \frac{1\,\text{mol}\,\text{SiCl}_4}{1\,\text{mol}\,\text{Si}} = 2.0$ mol $SiCl_4$ are formed *if* all Si reacted. $0.50\,mol\,Cl_2 \times \frac{1\,\text{mol}\,\text{SiCl}_4}{2\,\text{mol}\,\text{Cl}_2} = 0.25$ mol $SiCl_4$ are formed *if* all of Cl_2 reacted. Cl_2 produces less product, so *Cl_2 is the limiting reactant.*

Determine how many moles of excess reactant are left.	$0.50\,\text{mol}\,Cl_2 \times \dfrac{1\,\text{mol}\,Si}{2\,\text{mol}\,Cl_2} = 0.25$ mol of Si are consumed in the reaction. 2.0 mol Si available – 0.25 mol Si reacted = 1.75 mol of Si are left over = *1.8 mol of Si* (correct number of significant figures).
Answers to Example 8.	*a. 0.25 mol $SiCl_4$ are produced* *b. Cl_2 is the limiting reactant* *c. 1.8 mol Si remain unreacted*

Percent yield (sometimes called **reaction yield**) refers to the *actual yield of a reaction*. In Example 8 above the theoretical yield was calculated by *assuming* that all the reactant had been converted to product. In real life, all the reactant may not react and less than theoretical yield is produced. The formula to calculate percent yield is:

$$\textbf{\% yield} = \frac{\textbf{Actual yield}}{\textbf{Theortical yield}} \times \textbf{100}$$

In Example 8, the theoretical yield is 0.25 mol of $SiCl_4$. If in the lab only 0.20 mol of $SiCl_4$ were produced, then the percent yield would be:

$$\%\,yield = \frac{0.20}{0.25} \times 100\% = 80.\%$$

SAMPLE MULTIPLE CHOICE QUESTIONS

1. How many atoms of helium are there in a balloon that contains 40. 0 g of helium gas?

   A. 3.01×10^{23}
   B. 6.02×10^{23}
   C. 12.04×10^{23}
   D. 6.02×10^{24}
   E. 12.04×10^{24}

2. What is the mass percent of nitrogen in ammonium carbonate, $(NH_4)_2CO_3$?

A. 29%
B. 14.5%
C. 42%
D. 50%
E. 58%

3. Octane fuel is burned in air. When the equation is balanced with the lowest whole-number coefficients, the coefficient for the water molecule is:

A. 2
B. 9
C. 12
D. 16
E. 18

4. Iron rusts readily in air according to the reaction:

$$4Fe\ (s) + 3O_2\ (g) \rightarrow 2Fe_2O_3\ (s)$$

When 112 g of iron rust, how much iron(III) oxide is formed?

A. 160 g
B. 320 g
C. 676 g
D. 722 g
E. 1280 g

5. When ammonia gas reacts with hydrogen chloride gas, a white solid, ammonium chloride, forms.

$$NH_3\ (g) \quad + \quad HCl\ (g) \quad \rightarrow \quad NH_4Cl\ (s)$$

If 6.02×10^{23} molecules of ammonia react with 12.04×10^{23} molecules of hydrogen chloride, how many molecules are in the reaction vessel when the reaction is complete?

A. 1.00×10^{23}
B. 3.01×10^{23}
C. 6.02×10^{23}
D. 12.04×10^{23}
E. None of these

6. Which of the following nitrogen oxide compounds is the empirical formula for a compound that is analyzed as 47% nitrogen and 53% oxygen?

   A. NO
   B. N_2O_3
   C. N_2O_4
   D. NO_2
   E. N_2O_2

7. A 100.0 g sample of impure calcium carbonate was heated. It decomposed to form carbon dioxide gas and calcium oxide. After heating, the solid residue weighed 78 grams. What was the percent of calcium carbonate by mass in the original sample?

   A. 10.%
   B. 15%
   C. 25%
   D. 50.%
   E. 75%

8. ___KOH + ___H_3PO_4 $\rightarrow$ ___K_3PO_4 +___ H_2O

   When 1 mol of KOH neutralizes H_3PO_4 according to this equation, how many moles of water are formed?

   A. 1
   B. 2
   C. 3
   D. 4
   E. 5

9. A solution contains 0.10 mol of $Pb(NO_3)_2$ and 0.050 mol of BaI_2. How many moles of PbI_2 will precipitate?

   A. 0.050
   B. 0.10
   C. 0.15
   D. 0.20
   E. None of these

10. What mass of oxygen gas is produced when 0.10 mol of water is electrolyzed?

   A. 0.32 g
   B. 3.2 g
   C. 1.6 g
   D. 16 g
   E. 32 g

Comprehension Questions

1) The main group element gallium is one of the very few metallic elements that can exist in the liquid state at room temperature, that is, providing that it's a warm summer day. Gallium's melting point is 29.8 ºC or about 86 ºF. Compounds of gallium have unique electrical properties and have therefore found use in products such as light emitting diodes (LEDs). This element has two stable isotopes, Ga = 69 (atomic mass = 68.926 amu) and Ga = 71 (atomic mass = 70.925), and an average atomic mass of 69.723 amu. Calculate the percent abundance of each isotope.

2) Twelve-gauge copper wire, like the kind commonly used in residential electrical systems, is roughly cylindrical and has a diameter of approximately 0.1040 in. Copper's density is 8.92 g/cm^3 and copper atoms have an approximate atomic radius of 135 pm.

a) Calculate the number of atoms it would take to span the thickness, that is, diameter, of one of these wires. Express this value as a number of atoms and a number of moles of atoms.

b) Calculate the mass, in grams, of a 100-ft piece of copper wire.

c) How many moles of copper atoms would be found in a piece of this wire that is exactly 100 ft long? How many atoms?

3) Oxidation of carbon-containing compounds can take place not only through reaction with molecular oxygen, as in common combustion, but also by reaction with a variety of other oxidizing agents. Sugar-containing candies, which we will represent with the formula $C_{12}H_{22}O_{11}$, react violently at elevated temperatures with the strong oxidant potassium chlorate, $KClO_3$, according to the following reaction, which closely resembles combustion:

$$C_{12}H_{22}O_{11} + KClO_3 \rightarrow CO_2 + H_2O + KCl$$

a) Provide coefficients to balance the equation.

b) What quantity of carbon dioxide could be produced from the reaction of 1.50 mol of sugar with an excess of potassium chlorate? Express your answer in grams and moles.

c) What minimum amount of $KClO_3$, in grams, would need to be reacted to produce 3.25 g of water?

d) If a reaction is set up in which 16.1 g of $KClO_3$ is combined with 3.42 g of candy, what quantity of H_2O will be formed? Which reactant is in excess and by how much? Give both answers in grams.

4) The arthritis drug Celebrex is a *selective* inhibitor of the enzyme that causes inflammation in humans and consequently has very few, if any, of the side effects associated with traditional nonsteroidal anti-inflammatory drugs, NSAIDs. (These compounds also inhibit enzymes responsible for noninflammatory processes.) It has therefore found widespread use in patients suffering from many inflammatory disorders. Celebrex's molecular formula is $C_{17}H_{14}N_3SO_2F_3$.

a) Calculate the molecular mass and percent composition of Celebrex.

b) This anti-inflammatory agent is synthesized from the condensation of 4-sulphonamidophenyl hydrazine, $C_6H_9N_3SO_2$, and the Claisen condensation product of 4-methyl acetophenone and ethyl trifluoroacetate, $C_{11}H_9O_2F_3$ according to the following reaction:

$$C_6H_9N_3SO_2 + C_{11}H_9O_2F_3 \rightarrow C_{17}H_{14}N_3SO_2F_3 + 2H_2O$$

Suppose that a chemist sets up a reaction to prepare Celebrex by combining 20.0 g of each of the above reactants. How much of the anti-inflammatory compound could be synthesized from this reaction? Express your answer in grams and moles.

c) Which reactant is in excess? Which is limiting?

d) By what amount, expressed in grams and moles, is the excess reactant in excess?

e) Suppose the chemist isolates 20.8 g of the purified drug from this reaction. What is the percent yield for this process?

ANSWERS TO SAMPLE MULTIPLE CHOICE QUESTIONS

1. D

Dimensional analysis yields the following solution:

$$40.0\,\text{g He} \times \frac{1\,\text{mol He}}{4.00\,\text{g He}} \times \frac{6.022 \times 10^{23}}{1\,\text{mol He}} = \mathit{6.02 \times 10^{24}\ atoms\ He}$$

2. A

Percent nitrogen is the mass of nitrogen divided by molar mass times 100:

$$\frac{2\text{N}}{(\text{NH}_4)_2\text{CO}_3} \times 100\% = \frac{2(14)}{2(14)+8(1)+12+3(16)} \times 100\% = \mathit{29\%\ N}$$

3. E

The first step is to write the equation correctly. This is a combustion reaction involving the burning of a hydrocarbon in which CO_2 and H_2O are produced:

$C_8H_{18} + O_2 \rightarrow CO_2 + H_2O$

Using coefficients, balance the equation by inspection. Begin by balancing the carbons (step 1) , then hydrogen (step 2), and then oxygen (step 3), and then multiply by 2 to obtain whole-number coefficients (step 4).

(step 1)	$C_8H_{18} + O_2 \rightarrow 8CO_2 + H_2O$
(step 2)	$C_8H_{18} + O_2 \rightarrow 8CO_2 + 9H_2O$
(step 3)	$C_8H_{18} + 25/2O_2 \rightarrow 8CO_2 + 9H_2O$
(step 4)	$2C_8H_{18} + 25\ O_2 \rightarrow 16CO_2 + \mathit{18}H_2O$

4. A

Dimensional analysis yields the following solution:

$$112\,\text{g Fe} \times \frac{1\,\text{mol Fe}}{56\,\text{g Fe}} \times \frac{2\,\text{mol Fe}_2\text{O}_3}{4\,\text{mol Fe}} \times \frac{160\,\text{g Fe}_2\text{O}_3}{1\,\text{mol Fe}_2\text{O}_3} = \mathit{160.\,g\ Fe_2O_3}$$

5. D

This is a limiting reactant type of problem. The general strategy is to convert all quantities to moles and then determine which reactant is totally consumed. The limiting reagent (also called limiting reactant) determines the amount of product produced.

NH_3	+ HCl	$\rightarrow$	NH_4Cl
6.02×10^{23}	12.04×10^{23}		
1 mol	2 mol		
limiting reactant	excess reactant		

Since there is a 1 to 1 ratio between NH_3 and HCl, all of the NH_3 is consumed and 1 mol of HCl is left unreacted. One mole of NH_4Cl is formed.

Net result:

1 mol HCl left unreacted + 1 mol NH_4Cl produced = 2 mol of molecules

$= \mathit{12 \times 10^{23}\ molecules}$

6. A

This is a traditional empirical formula problem. Begin by assuming a 100-g sample, convert all elements to moles by dividing by atomic weight, divide by the smallest number of moles to get whole-number ratios that serve as the subscripts in the empirical formula.

General Strategy	***Example 6* worked out**
Assume a 100-g sample and convert the percentages of elements to grams.	46.7 g N 53.3 g O
Convert gram quantities to moles by dividing by the molar mass of the atom.	$46.7\,\text{g N} \times \frac{1\,\text{mol N}}{14.0\,\text{g}} = 3.33\,\text{mol N}$ $53.3\,\text{g O} \times \frac{1\,\text{mol O}}{16.0\,\text{g O}} = 3.33\,\text{mol O}$
Divide by the smallest number of moles to get whole-number ratios. The whole number ratios are the subscripts in the empirical formula.	$N_{\frac{3.33}{3.33}}\ O_{\frac{3.33}{3.33}} = NO$ *Empirical formula = NO*

7. D

The 100.0 g is made of $CaCO_3$(*s*) and some other material. The weight 'loss' as the reaction proceeds is due to carbon dioxide gas being released.

$$CaCO_3\,(s) \rightarrow CaO\,(s) + CO_2\,(g)$$

By converting the mass of CO_2 to moles of CO_2 it is possible to calculate the moles of $CaCO_3$(*s*) in the original sample, since every mole of CO_2 came from a mole of $CaCO_3$.

100.0 g solid reactant – 78 g solid product residue = 22 g CO_2 product gas

$$22\,\text{g CO}_2 \times \frac{1\,\text{mol CO}_2}{44\,\text{g CO}_2} \times \frac{1\ \text{mol CaCO}_3}{1\,\text{mol CO}_2} \times \frac{100.0\,\text{g CaCO}_3}{1\,\text{mol CaCO}_3} = 50.\,\text{g CaCO}_3$$

$$\%\,\text{CaCO}_3 \text{ in sample} = \frac{\text{part}}{\text{whole}} \times 100\% = \frac{50.\,\text{g CaCO}_3}{100\,\text{g sample}} \times 100 = \mathit{50.\%\,CaCO_3}$$

8. A

When dealing with a stoichiometry problem, the first step is always to balance the equation and then use the coefficients of the balanced equation to relate one molecule to another using dimensional analysis.

$3KOH + H_3PO_4 \rightarrow K_3PO_4 + 3H_2O$
1 mol ? mol

$$1\,\text{mol KOH} \times \frac{3\,\text{mol } H_2O}{3\,\text{mol KOH}} = 1\,\text{mol } H_2O$$

9. A

This problem is based on the mole relationships between molecules and ions.

0.10 mol of $Pb(NO_3)_2$ contains 0.10 mol Pb^{2+} ion since there is a 1 to 1 ratio between $Pb(NO_3)_2$ and Pb^{2+}.

0. 050 mol BaI_2 contains 0.10 mol I^- ion since there is a 1 to 2 ratio between BaI_2 and I^- as illustrated in the following step:

$$0.050\,\text{mol } BaI_2 \times \frac{2\,\text{mol } I^-}{1\,\text{mol } BaI_2} = 0.10\,\text{mol } I^-$$

This is a limiting reactant type problem (refer to Example 8). You must determine which reactant ion limits the amount of precipitate, PbI_2, produced. The ion that forms the lesser amount limits the reaction to that amount.

$$0.10\,\text{mol of } Pb^{2+} \times \frac{1\,\text{mol } PbI_2}{1\,\text{mol } Pb^{2+}} = 0.10\,\text{mol } PbI_2$$

$$0.10\,\text{mol } I^- \times \frac{1\,\text{mol } PbI_2}{1\,\text{mol } I^-} = \mathit{0.050\ mol\ PbI_2 = amount\ of\ precipitate\ formed}$$

10. C

Write a balanced equation for the reaction and then use dimensional analysis to obtain the answer.

$2\,H_2O \rightarrow 2H_2 + O_2$
0.10 mol ? g

$$0.10\,\text{mol } H_2O \times \frac{1\,\text{mol } O_2}{2\,\text{mol } H_2O} \times \frac{32.0\,\text{g } O_2}{1\,\text{mol } O_2} = \mathit{1.6\ g\ O_2}$$

Answers to Comprehension Questions

1) The average atomic mass of an element is the weighted average, that is, the sum of the masses of the various isotopes multiplied by their percent abundance. In this case, you are not given the percent abundance of either isotope, that is, two variables and only one equation. Solution of the problem only becomes possible with the recognition that the sum of the percents abundance is 100 and the relative abundance of each isotope can then be represented as X and $100 - X$ (or $1 - X$ if the percents are represented as decimals). Therefore:

$$69.723 = 68.926X + 70.925(1 - X)$$

$$69.723 = 68.926X + 70.925 - 70.925\,X$$

$$-10.202 = -10.999X$$

$X = 0.92754$ or **92.754%;** in this case X represents the percent abundance of the isotope with a mass of 68.926 amu. The percent abundance of the isotope weighing 70.925 amu is 100% – 92.754% or **7.246%**.

2) a) Solution to this part of the problem simply involves conversion to a common unit for length and then division of the wire's diameter by the diameter of a single atom. Therefore:

diameter of wire = 0.1040 in. × (2.54 cm / 1 in.) = 0.2642 cm

0.2642 cm × (1 m / 100 cm) × (10^{12} pm / 1 m) = 2.642×10^{9} pm / 135 pm/atom = **1.96×10^{7} atoms per diameter of wire**

b) The mass of 100 ft of this wire will be determined by multiplying the volume of that quantity of wire by the density of copper. Since the wire is cylindrical, its volume can be calculated using the formula $\pi r^2 l$ (r = radius, l = length of cylinder)

Substituting:

the radius of the wire equals ½ the diameter, so 0.2642 cm × 0.5 = 0.1321 cm

the volume of the wire = $\pi \times (0.1321\text{ cm})^2 \times (100\text{ ft} \times (12\text{ in.} / \text{ft}) \times (2.54\text{ cm} / \text{in})) = 167.1\text{ cm}^3$

the mass of the wire = volume × density = 167.1 cm^3 × 8.92 g/cm^3 = **1491 g**

c) The number of moles of copper found in 100 ft or 1491 g of this wire can be determined by dividing the mass of the wire by the molar mass of copper, 63.546 g / mol. The number of atoms can be determined by multiplying the moles of copper by Avogadro's number, 6.022×10^{23} particles per mole. Therefore:

1491 g Cu / 63.546 g/mol = **23.46 mol of Cu**

23.46 mol of Cu × 6.022×10^{23} atoms/mol = **1.412×10^{25} atoms of Cu**

3) a) $C_{12}H_{22}O_{11}$ + **8**$KClO_3$ → **12**CO_2 + **11**H_2O + **8**KCl

b) The coefficients of a balanced equation represent mole ratios in which the reactants combine and the products are produced. The mass of a product or reactant may be determined by multiplying the moles of that substance by its molar mass. Therefore:

1.50$C_{12}H_{22}O_{11}$ × (12 mol CO_2/1 mol $C_{12}H_{22}O_{11}$) = **18.0 mol CO_2**

18.0 mol CO_2 × 44.01 g CO_2/mol = **792 g of CO_2**

c) Beginning with the quantity of a reagent in grams, you must first convert to moles to use the mole ratios found in the balanced equation. The moles of $KClO_3$ necessary will then have to be re-converted to grams to provide an answer:

3.25 g H_2O × (1 mol H_2O/18.02 g) × (8 mol $KClO_3$/11 mol H_2O) × (122.55 g $KClO_3$/ 1 mol $KClO_3$) = **16.1 g of $KClO_3$**

d) A convenient way to determine the limiting reactant in this situation is to calculate the maximum amount of product, H_2O in this case, that can be formed from each reactant assuming the other to be in excess. The answer with the smaller amount of product will have started with the limiting reactant. According to part c above, 16.1 g of $KClO_3$ will form **3.25 g of H_2O**.

3.42 g $C_{12}H_{22}O_{11}$ × (1 mol $C_{12}H_{22}O_{11}$/342.34 g $C_{12}H_{22}O_{11}$) × (11 mol H_2O/1 mol $C_{12}H_{22}O_{11}$) × (18.02 g H_2O/1 mol H_2O) = **1.98 g H_2O**

Since there's enough $KClO_3$ to make 3.25 g of H_2O and only enough candy to make 1.98 g H_2O, the candy is the limiting reactant. To determine the amount of $KClO_3$ in excess, the amount necessary to react with 3.42 g of candy is calculated and subtracted from the initial amount, 16.1 g.

3.42 g $C_{12}H_{22}O_{11}$ × (1 mol $C_{12}H_{22}O_{11}$/342.34 g $C_{12}H_{22}O_{11}$) × (8 mol $KClO_3$/1 mol $C_{12}H_{22}O_{11}$) × (122.55 g $KClO_3$/1 mol $KClO_3$) = **9.79 g $KClO_3$** needed to react with 3.42 g of candy

16.1 g $KClO_3$ – 9.79 g $KClO_3$ = **6.3 g $KClO_3$ in excess**

4) a) The molar mass of Celebrex, $C_{17}H_{14}N_3SO_2F_3$, is (17 × 12.01) + (14 × 1.008) + (3 × 14.01) + (1 × 32.07) + (2 × 16.00) + (3 × 19.00) = **381.4 g/mol**

Percent composition is determined by dividing the mass of each element in the compound by the total mass of the compound and then multiplying by 100%.

%C = [((17 × 12.01) g/mol) / (381.4 g/mol)] × 100% = **53.53% C**

%H = [((14 × 1.008) g/mol) / (381.4 g/mol)] × 100% = **3.700% H**

%N = [((3 × 14.01) g/mol) / (381.4 g/mol)] × 100% = **11.02% N**

%S = [((1 × 32.07) g/mol) / (381.4 g/mol)] × 100% = **8.41% S**

%O = [((2 × 16.00) g/mol) / (381.4 g/mol)] × 100% = **8.39% O**

%F = [((3 × 19.00) g/mol) / (381.4 g/mol)] × 100% = **14.94% F**

b) As in question 3d, the maximum amount of product can be calculated by performing two separate calculations each starting with a different reactant.

20.0 g $C_6H_9N_3SO_2$ × (1 mol $C_6H_9N_3SO_2$/187.18 g $C_6H_9N_3SO_2$) × (1 mol $C_{17}H_{14}N_3SO_2F_3$/ 1 mol $C_6H_9N_3SO_2$) = **0.107 mol of Celebrex** × (381.4 g $C_{17}H_{14}N_3SO_2F_3$/1 mol $C_{17}H_{14}N_3SO_2F_3$) = **40.8 g Celebrex**

20.0 g $C_{11}H_9O_2F_3$ × (1 mol $C_{11}H_9O_2F_3$/230.18 g $C_{11}H_9O_2F_3$) × (1 mol $C_{17}H_{14}N_3SO_2F_3$/ 1 mol $C_{11}H_9O_2F_3$) = **0.0869 mol of Celebrex** × (381.4 g $C_{17}H_{14}N_3SO_2F_3$/1 mol $C_{17}H_{14}N_3SO_2F_3$) = **33.1 g Celebrex**

Therefore the maximum amount of Celebrex that can be synthesized is **33.1 g or 0.0869 mol.**

c) The limiting reactant is the Claisen condensation product, $C_{11}H_9O_2F_3$, and the phenylhydrazine compound, $C_6H_9N_3SO_2$, is in excess.

d) As in 3d, the amount of excess reactant can be calculated by starting with the limiting reactant and determining the amount of the other reactant needed to completely react with it.

20.0 g $C_{11}H_9O_2F_3$ × (1 mol $C_{11}H_9O_2F_3$/230.18 g $C_{11}H_9O_2F_3$) × (1 mol $C_6H_9N_3SO_2$/ 1 mol $C_{11}H_9O_2F_3$) × (187.18 g $C_6H_9N_3SO_2$/1 mol $C_6H_9N_3SO_2$) = 16.3 g $C_6H_9N_3SO_2$ needed to react with 20.0 g $C_{11}H_9O_2F_3$

20.0 g $C_6H_9N_3SO_2$ – 16.3 g $C_6H_9N_3SO_2$ = **3.7 g $C_6H_9N_3SO_2$ in excess**

Multiplying by the reciprocal of the molar mass of $C_6H_9N_3SO_2$,

3.7 g $C_6H_9N_3SO_2$ × (1 mol $C_6H_9N_3SO_2$/187.18 g $C_6H_9N_3SO_2$) = **0.0198 mol $C_6H_9N_3SO_2$ in excess**

e) Percent yield is equal to 100% times the ratio of actual yield to theoretical yield (from part b), therefore:

% yield = (20.8 g / 33.1 g) × 100% = **62.8%**