Balancing Chemical Reactions and Basic Stoichiometry Concepts

**Balancing Chemical Reactions**

A chemical process is represented by a reaction that shows the reactants and products of the process with an arrow separating them. An example is shown below for the reaction that forms water, H2O, from its elements, hydrogen, H2 and oxygen, O2.

 H2 + O2  → H2O

Note that the subscripts show the numbers of each element in the molecule or compound. The subscripts shown for the reactants and products must not be changed during the balancing process. Coefficients are the numbers either shown (or implied as one) in front of each reactant and product. The coefficients of the reaction are the numbers that must be adjusted to balance the reaction.

For a reaction to be balanced, it must show mass balance (conservation of mass), and charge balance (conservation of charge). For mass balance, the reaction must show the same number of each type of atoms on both sides of the arrow. For charge balance, the reaction must show the same net charge on both sides of the arrow. In the reaction shown above, two hydrogen atoms are represented on both the reactant side and the product side. However, only one oxygen atom is represented on the right side, the product side, where the reactant side shows two. The reaction can be balanced by adjusting the coefficients as shown below.

 2 H2 + O2 → 2 H2O

In this balanced reaction, the number of hydrogen atoms in the reactant side and on the product side is four; the number of oxygen atoms on the reactant side and on the product side is two. Charge signs are shown in the upper right hand side of each reactant and product. If the chemical species is neutral no charge is usually shown. The charge is implied as zero charge.

The reactions described below are shown as unbalanced reactions and then as balanced reactions.

The reaction that forms ammonia from its elements:

N2 + H2  → NH3 (unbalanced)

 N2 + 3 H2  → 2 NH3 (balanced)

The combustion of methane, CH4 to form carbon dioxide, CO2 and water, H2O

 CH4 + O2 → CO2 + H2O (unbalanced)

 CH4 + 2 O2 → CO2 + 2 H2O (balanced)

The reaction of concentrated nitric acid on copper metal to form nitrogen dioxide and water:

 Cu(s) + HNO3 (aq) → Cu2+(aq)  + NO3 2-(aq)   + NO2 (g)  + H2O (l) (unbalanced)

 Cu(s) + 4 HNO3 (aq) → Cu2+(aq)  + 2 NO3 2-(aq)   + 2 NO2 (g)  + 2 H2O (l) (balanced)

Note that in the last reaction, states of the substances are show as subscripts in parenthesis. These are solid (s), aqueous (aq), gas (g) and liquid (l). In many cases, showing the states of the reactants and products is important and will be noted. In other cases, these can be omitted without compromising the information of the reaction.

**The Concept of the Chemical Mole and Avogadro’s Number**

The coefficients of a chemical reaction describe the reaction in terms of molecules (or atoms if some of the reactants or products are atomic) or in terms of collective amounts. Because the atomic masses of the Periodic Table show the relative masses of the elements, the chemical mole is a very useful unit to illustrate the coefficients of a reaction. A mole of objects is equivalent to 6.02 x 1023 objects, 602,000,000,000,000,000,000,000 objects. The chemical mole is tied to the Periodic Table in terms of masses of the atoms, in units of grams. A three-way relationship can be used to illustrate this for atoms and for molecules.

 For any atom from the Periodic Table or For any molecule formed from atoms

 **1 mole of C atoms 1 mole of H2O molecules**

$\adots \adots $$\ddots \ddots $ **or** $ \adots \adots $$\ddots \ddots $

 **6.02 x 1023 atoms of C = 12.01 g of C 6.02 x 1023 molecules of H2O = 18.00 g H2O**

The size of Avogadro’s number is beyond comprehension; but, it can be appreciated if one thinks of the number in terms of everyday type of situations.

1 mole of tiny grains of sand: enough sand to fill the state of California with a layer of sand as high

as a ten-story building

1 mole of marshmallows: enough marshmallows to fill the entire US, Alaska and Hawaii included with a layer of marshmallows 600 miles deep

1 mole of pennies: enough money for each person on Earth to spend $1,000,000 each hour, night and day throughout his/her lifetime.

1 mole of counts: it would take a person who could count to 10 in one second, 1.9 million, billion years (1.9 x 1015 years) to count to Avogadro’s number!!

But!! Avogadro’s number of C atoms fits in the palm of a person’s hand and weighs 12.01 g.

This gives an insight of the very small size of atoms in relation to everyday type of things.

**Mole/Mass Calculations and Three-Way Relationships**

Dimensional Analysis or the Factor-Label method can be used with three-way relationships to perform mole/mass calculations.

Example: Calculate the number of grams of lead, Pb, in 5.0 moles of Pb.

$ \frac{5.0 moles of Pb}{1}$ **x** $\frac{207.2 g Pb}{1 mole Pb} =$ **1,036 g of Pb**

Example: Calculate the number of moles of CO2 in 98 g of CO2.

$$\frac{98.0 g of CO2}{1} x \frac{ 1 mole CO2}{44.0 g CO2}=2.23 moles of CO2$$

Example: How many molecules of sucrose (C12H22O11) are in 8.42 g of sugar?

$\frac{8.42 g of sucrose}{1} x \frac{1 mole sucrose}{342 g of sucrose} x \frac{6.02 x 10^{23 }molecules}{1 mole molecules} =1.53 x 10^{22} $**molecules**

**Chemical Stoichiometry; Limiting Reagent and Theoretical Yield**

Chemical Stoichiometry is called chemical arithmetic or ratios of chemical change. This type of calculation involves determining the amount of product that can be formed when two or more substances react together, or when a compound decomposes to form different substances. Chemical stoichiometry requires that a balanced chemical reaction be used. Mole/mass unit conversions are used to allow the calculation of the final answer. Dimensional analysis is a good method that can be used to work stoichiometry type of problems.

Definitions:

Limiting Reactant or Limiting Reagent – the reactant or reagent that will be exhausted first in a chemical reaction

Theoretical Yield – the amount of a given product, in moles or in mass amounts, that can be formed in a reaction assuming that all of the limiting reagent is consumed.

Percentage Yield – the actual amount of a given product obtained from a reaction, on a percentage basis, compared to the theoretical yield

Useful Equations in Working with Stoichiometry Problems

$Moles=\frac{Mass}{Formula Weight}$ **Mass = Moles x Formula Weight**

 **Percentage Yield =** $\frac{Actual Yield}{Theoretical Yield}$ **x 100**

**Example:**

**How many moles of urea could be formed from 3.5 moles of ammonia?**

**2NH3(*g*)+ CO2(*g*) → 1 (NH2)2CO(a*q*)+ *H2O(l*)**

**Example:**

**A chemist needs 58.75 grams of urea, how many grams of ammonia are needed to produce this amount?**

***Strategy:* grams → moles → mole ratio → grams**

 **58.75 g (NH2)2CO x** $\frac{(1 mole (NH2)2CO )}{58.06 g (NH2)2CO }x \frac{2 moles NH3}{1 mole (NH2)2CO }x \frac{17.04 g NH3}{1 mole NH3 }$ = **34.49 g NH3**

**Use of Molecular Models to Illustrate Chemical Stoichiometry**

Use a set of molecular models to illustrate the maximum number of product molecules that can be produced from the following situations.

Reaction of molecular hydrogen and molecular oxygen to form water:

 2 H2 + O2 → 2 H2O (balanced)

Start with molecular models consisting of 2 molecules of H2 and one molecule of O2. Disconnect the molecules to H atoms and O atoms. Reconnect the atoms to make H2O molecules. Make note of how many molecules of H2O can be made from the original two molecules of hydrogen and one molecule of oxygen? Has the stoichiometry of the reaction been obeyed? Reactions as the one just illustrated, in which the exact ratio of reactant molecules as called for by the balanced chemical reaction is used, are called stoichiometric reactions.

Now start instead with models consisting of 4 molecules of H2 and 3 molecules of O2. Disconnect and then connect the hydrogen atoms and oxygen atoms to form as many molecules of H2O as possible. Reconnect any molecules of hydrogen or oxygen left over.

How many H2O molecules were formed?

Which reactant is the limiting reactant (the one exhausted first)?

Which reactant is the reactant in excess?

Conclusion: A product can be made only until the limiting reaction is consumed. In problems where amounts of more than one reactant are provided, It is important to identify the limiting reactant in a stoichiometry problem and use this reactant to determine the theoretical yield of product.

Now consider the reaction of N2 and H2 to form ammonia, NH3.

 N2 + 3 H2  → 2 NH3 (balanced)

To completely react 6 molecules of H2, how many molecules of N2 are required? How many molecules of NH3 can be formed?

If in another reaction, 2 molecules of N2 are reacted with seven molecules of H2, how many molecules of NH3 will be formed? Which reactant is limiting?

Inspect the balanced reaction and predict how many molecules of NH3 can form if 10 molecules of N2 and 30 molecules of H2 react. (You will not have enough models to illustrate this one.)

Finally, consider the reaction of methane, CH4 and oxygen, O2 to form carbon dioxide, CO2 and water, H2O.

 CH4 + 2 O2 → CO2 + 2 H2O

What is the theoretical yield of water molecules that can result from 3 molecules of methane and four molecules of O2?

In this reaction, which reactant, the CH4 or the O2 is the limiting reactant?

**Example Problems for Continued Work**

**Balancing Chemical Reactions**

1. What is the coefficient of H2O when the following equation is properly balanced with the smallest set of whole numbers?
 \_\_\_ Na + \_\_\_ H2O → \_\_\_ NaOH + \_\_\_ H2
2. 1 B. 2 C. 3 D. 4 E. 5

 2) What is the coefficient of H2O when the following equation is properly balanced with smallest set of whole numbers?
 \_\_\_ Al4C3 + \_\_\_ H2O → \_\_\_ Al(OH)3 + \_\_\_ CH4

A. 3 B. 4 C. 6 D. 12 E. 24

3) When balanced with smallest set of whole numbers, the coefficient of O2 in the following equation is:
 \_\_ C2H4 + \_\_ O2 → \_\_ CO2 + \_\_ H2O

A. 1 B. 2 C. 3 D. 4 E. 6

4) When a chemical equation is balanced, it will have a set of whole number coefficients that cannot be reduced to smaller whole numbers. What is the coefficient for O2 when the following combustion reaction of a hydrocarbon is balanced?

 \_\_\_ C7H14 + \_\_\_ O2 → \_\_\_ CO2 + \_\_\_ H2O

A. 42 B. 21 C. 11 D. 10 E. None of the answers is correct.

5) What is the coefficient of H2SO4 when the following equation is properly balanced with the smallest set of whole numbers?

 \_\_\_ Ca3(PO4)2 + \_\_\_ H2SO4 → \_\_\_ CaSO4 + \_\_\_ H3PO4

1. 3 B. 8 C. 10 D. 11 E. None of the answers is correct.

 6) Balance the equation below using the smallest set of whole numbers. What is the coefficient of H2O?

 \_\_\_ PCl3(l) + \_\_\_ H2O(l) → \_\_\_ H3PO3(aq) + \_\_\_ HCl(aq)

A.1 B. 2 C. 3 D. 5 E. None of the answers is correct.

7) What is the coefficient of O2 when the following equation is properly balanced with the smallest set of whole numbers?

\_\_\_ CH3OH + \_\_\_ O2 → \_\_\_ CO2 + \_\_\_ H2O

1. 1 B. 2 C. 3 D. 7 E. None of the answers is correct.

**Mole/Mass Calculations and Three-Way Relationships**

1) What is the average mass, in grams, of one atom of iron (NA = 6.022 × 1023 mol-1)?

A. 6.02 × 1023 g

B. 1.66 × 10-24 g

C. 9.28 × 10-23 g

D. 55.85 g

E. 55.85 × 10-23 g

2) The mass of 1.21 × 1020 atoms of sulfur (NA = 6.022 × 1023 mol-1) is:

A. 3.88 × 1021 g

B. 2.00 mg

C. 32.06 g

D. 6.44 mg

E. 2.00 × 10-4 g

3) Determine the number of ammonia molecules in 4.85 g of ammonia (NA = 6.022 × 1023 mol-1).

A. 2.92 x 1023

B. 4.73 x 10-25

C. 1.24 x 1023

D. 5.83 x 10-24

E. 1.71 x 1023

4) What is the mass in grams of 0.250 mol of the common antacid calcium carbonate?

A. 4.00 × 102 g

B. 25.0 g

C. 17.0 g

D. 4.00 × 10-2 g

E. 2.50 × 10-3 g

5) Calculate the number of moles in 17.8 g of the antacid magnesium hydroxide, Mg(OH)2.

A. 3.28 mol

B. 2.32 mol

C. 0.431 mol

D. 0.305 mol

E. 0.200 mol

6) Phosphorus pentachloride, a white solid that has a pungent, unpleasant odor, is used as a catalyst for certain organic reactions. Calculate the number of moles in 38.7 g of phosphorus pentachloride.

A. 5.38 mol

B. 3.55 mol

C. 0.583 mol

D. 0.282 mol

E. 0.186 mol

7) Calculate the mass in grams of 8.35 × 1022 molecules of CBr4.

A. 0.0217 g

B. 0.139 g

C. 7.21 g

D. 12.7 g

E. 46.0 g

8) The number of hydrogen atoms in 0.050 mol of C3H8O3 is:

A. 3.0 × 1022 H atoms

B. 1.2 × 1023 H atoms

C. 2.4 × 1023 H atoms

D. 4.8 × 1023 H atoms

E. None of the answers is correct.

9) Calculate the mass of 3.00 moles of CF2Cl2.

A. 3.00 g

B. 174 g

C. 363 g

D. 1.81 × 1024 g

E. 40.3

**Chemical Stoichiometry; Limiting Reagent and Theoretical Yield**

1) Aluminum will react with bromine to form aluminum bromide (used as an acid catalyst in organic synthesis).
 Al(s) + Br2(l) → Al2Br6(s) [unbalanced]

How many moles of Al are needed to form 2.43 mol of Al2Br6?

A. 7.29 mol

B. 4.86 mol

C. 2.43 mol

D. 1.62 mol

E. 1.22 mol

2) Ammonia will react with fluorine to produce dinitrogen tetrafluoride and hydrogen fluoride (used in production of aluminum, in uranium processing, and in frosting of light bulbs).

2NH3(g) + 5F2(g) → N2F4(g) + 6HF(g)

How many moles of NH3 are needed to react completely with 13.6 mol of F2?

A. 34.0 mol

B. 27.2 mol

C. 6.80 mol

D. 5.44 mol

E. 2.27 mol

3) How many grams of ammonia are formed when 5.36 g of nitrogen reacts with hydrogen?

A. 3.26 g

B. 0.629 g

C. 13.04 g

D. 17.63 g

E. 6.52 g

 4) How many grams of nitrogen are required to react with 2.79 g of hydrogen to produce ammonia?

A. 25.8 g

B. 13.0 g

C. 78.2 g

D. 38.7 g

E. 77.4 g

4) How many grams of nitrogen are required to react with 2.79 g of hydrogen to produce ammonia?

A. 25.8 g

B. 13.0 g

C. 78.2 g

D. 38.7 g

E. 77.4 g

5) How many grams of nitrogen are required to react with hydrogen to produce 13.6 g of ammonia?

A. 11.2 g

B. 0.06 g

C. 22.4 g

D. 16.5 g

E. 44.8 g

6) What mass, in grams, of sodium carbonate is required for complete reaction with 8.35 g of nitric acid to produce sodium nitrate, carbon dioxide, and water?

A. 28.1 g

B. 14.04 g

C. 4.96 g

D. 7.02 g

E. 400.0 g

7) How many grams of lead (II) chloride are produced when 13.87 g lead (II) nitrate combines with hydrochloric acid to produce lead (II) chloride and nitric acid?

A. 5.82 g

B. 1.19 g

C. 0.086 g

D. 11.65 g

E. 16.52 g

8) How many grams of calcium are required to react with 7.75 g of water to produce calcium hydroxide and hydrogen gas?

A. 8.62 g

B. 34.5 g

C. 4.31 g

D. 3.48 g

E. 17.2 g

9) How many grams of oxygen are required to react with calcium to produce 44.8 g calcium oxide?

A. 12.8 g

B. 25.6 g

C. 6.4 g

D. 0.05 g

E. 51.1 g

10) How many grams of sodium fluoride (used in water fluoridation and manufacture of insecticides) are needed to form 485 g of sulfur tetrafluoride?

3SCl2(l) + 4NaF(s) → SF4(g) + S2Cl2(l) + 4NaCl(s)

A. 1940 g

B. 1510 g

C. 754 g

D. 205 g

E. 51.3 g

11) Phosphine, an extremely poisonous and highly reactive gas, will react with oxygen to form tetraphosphorus decaoxide and water.

PH3(g) + O2(g) → P4O10(s) + H2O(g)

Calculate the mass of P4O10(s) formed when 225 g of PH3 reacts with excess oxygen.

A. 1880 g

B. 940. g

C. 900. g

D. 470 g

E. 56.3 g

12) Potassium chlorate (used in fireworks, flares and safety matches) forms oxygen and potassium chloride when heated.

KClO3(s) → KCl(s) + O2(g)

How many grams of oxygen are formed when 26.4 g of potassium chlorate is heated?

A. 223 g

B. 99.1 g

C. 10.3 g

D. 6.86 g

E. 4.60 g

13) Aluminum metal reacts with chlorine gas to form solid aluminum trichloride. What mass of chlorine gas is needed to react completely with 163 g of aluminum?

A. 214 g

B. 245 g

C. 321 g

D. 489 g

E. 643 g

14) Lead (II) sulfide was once used in glazing earthenware. It will also react with hydrogen peroxide to form lead (II) sulfate and water. How many grams of hydrogen peroxide are needed to react completely with 265 g of lead (II) sulfide?

A. 151 g

B. 123 g

C. 50.3 g

D. 37.7 g

E. 9.41 g

 15) How many grams of Cl2 can be prepared from the reaction of 16.0 g of MnO2 and 30.0 g of HCl according to the following chemical equation?
MnO2 + 4HCl → MnCl2 + Cl2 + 2H2O

A. 0.82 g

B. 5.8 g

C. 13.0 g

D. 14.6 g

E. 58.4 g

16) Hydrochloric acid can be prepared by the following reaction:

2NaCl(s) + H2SO4(aq) → 2HCl(g) + Na2SO4(s)

How many grams of HCl can be prepared from 2.00 mol H2SO4 and 150 g NaCl?

A. 7.30 g

B. 93.5 g

C. 146 g

D. 150 g

E. 196 g

17) Calculate the mass of FeS formed when 9.42 g of Fe reacts with 8.50 g of S.
Fe(s) + S(s) → FeS(s)

A. 17.9 g

B. 87.9 g

C. 26.0 g

D. 14.8 g

E. 1.91 × 10-3 g

18) What is the limiting reactant when 3.41 g of nitrogen react with 2.79 g of hydrogen to produce ammonia and how many grams of ammonia are produced?

A. Hydrogen is the limiting reactant and 0.22 g of ammonia are produced.

B. Nitrogen is the limiting reactant and 2.07 g of ammonia are produced.

C. Nitrogen is the limiting reactant and 4.15 g of ammonia are produced.

D. Hydrogen is the limiting reactant and 23.5 g of ammonia are produced.

E. Hydrogen is the limiting reactant and 15.8 g of ammonia are produced.

19) What is the limiting reactant when 13.9 g of oxygen react with 33.7 g of calcium to produce calcium oxide and how many grams of calcium oxide are produced?

A. Calcium is the limiting reactant and 47.1 g of calcium oxide are produced.

B. Oxygen is the limiting reactant and 24.4 g of calcium oxide are produced.

C. Calcium is the limiting reactant and 47.1 g of calcium dioxide are produced.

D. Oxygen is the limiting reactant and 48.7 g of calcium oxide are produced.

E. Calcium is the limiting reactant and 23.6 g of calcium oxide are produced.

20) What is the limiting reactant when 31.3 g of manganese (II) chloride, 48.3 g of chlorine, and 25.7 g of water react to produce manganese (IV) oxide and how much hydrochloric acid is produced?

A. Manganese (II) chloride is the limiting reactant and 8.40 g of hydrochloric acid is produced.

B. Chlorine is the limiting reactant and 6.21 g of hydrochloric acid is produced.

C. Chlorine is the limiting reactant and 24.8 g of hydrochloric acid is produced.

D. Manganese (II) chloride is the limiting reactant and 36.3 g of hydrochloric acid is produced.

E. Water is the limiting reactant and 52 g of hydrochloric acid is produced.

21) Aluminum reacts with oxygen to produce aluminum oxide which can be used as an adsorbent, desiccant or catalyst for organic reactions. A mixture of 82.49 g of aluminum and 117.65 g of oxygen is allowed to react. Identify the limiting reactant and determine the mass of the excess reactant present in the vessel when the reaction is complete.

A. Oxygen is the limiting reactant; 19.81 g of aluminum remain.

B. Oxygen is the limiting reactant; 35.16 g of aluminum remain.

C. Aluminum is the limiting reactant; 16.70 g of oxygen remain.

D. Aluminum is the limiting reactant; 35.16 g of oxygen remain.

E. Aluminum is the limiting reactant; 44.24 g of oxygen remain.

22) Magnesium (used in the manufacture of light alloys) reacts with iron (III) chloride to form magnesium chloride and iron. A mixture of 41.0 g of magnesium and 175 g of iron (III) chloride is allowed to react. Identify the limiting reactant and determine the mass of the excess reactant present in the vessel when the reaction is complete.

A. Limiting reactant is Mg; 67 g of FeCl3 remain.

B. Limiting reactant is Mg; 134 g of FeCl3 remain.

C. Limiting reactant is Mg; 104 g of FeCl3 remain.

D. Limiting reactant is FeCl3; 2 g of Mg remain.

E. Limiting reactant is FeCl3; 87 g of Mg remain.

23) Potassium chloride is used as a substitute for sodium chloride for individuals with high blood pressure. Identify the limiting reactant and determine the mass of the excess reactant remaining when 7.00 g of chlorine gas reacts with 5.00 g of potassium to form potassium chloride.

A. Potassium is the limiting reactant; 2.47 g of chlorine remain.

B. Potassium is the limiting reactant; 7.23 g of chlorine remain.

C. Chlorine is the limiting reactant; 4.64 g of potassium remain.

D. Chlorine is the limiting reactant; 2.70 g of potassium remain.

E. No limiting reagent: the reactants are present in the correct stoichiometric ratio.