

### 3. Chemical Stoichiometry - other applications for ratios of *mass of*

Chemical compounds have a constant composition, so it follows that mass ratios of components in compounds must also be constant. *Mass ratios* are really conversion factors used in a variety of settings for compounds, reactants, and products. Consider the following examples:

- A. i. How many grams of carbon are in 57.83 grams of  $C_8H_{16}O_2$  ?  
 ii. How many grams of  $C_8H_{16}O_2$  can be prepared from 14.275 g hydrogen?
- B. How many grams of sulfur will combine with 4.829 g of sodium, to form  $NaHSO_4$  ?
- C. How many grams of mercury(II) oxide are required to produce 2.594 g of oxygen?

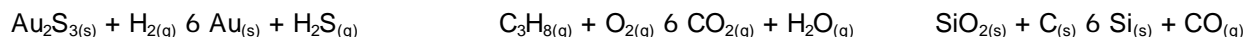
### 4. Chemical Stoichiometry - applied to chemical reactions as represented by *chemical equations*

Masses of all starting reactants in every chemical reaction, must equal masses of all final products. This is an expression of the Law of Conservation of Matter - matter cannot be created or destroyed by ordinary chemical means. It is for this reason that *all chemical equations must be balanced*, i.e., there must be the same number and type of starting atoms on the reactant side (the left side) as there are on the product side. Equations are balanced by applying integer multipliers which increment the entire formula they precede. The multipliers are called coefficients. Coefficients are applied to achieve the same numbers and types of atoms on both sides of the equation. Coefficients increment the entire chemical formula they precede, including any ionic charges. (Coefficients are really ...

Reading Chemistry (revisited...) Numbers or Mass? Microscopic or macroscopic?

A balanced chemical equation is interpreted in several ways depending on context and viewpoint. From a macroscopic viewpoint, coefficients in the equation represents the *number of* moles of reactants/products, and their masses as well. From a microscopic viewpoint, coefficients represent the *number of* atoms, or molecules, or formula units - and their *masses* as well. It depends on context. Recall, in a macroscopic world, subscripts in chemical formulas represent MOLES, and coefficients in balanced equations also represent MOLES.

A. *Balancing of equations* - Balance (and read) the following chemical equations:



Write balanced equations for the reaction of sodium,  $Na(s)$ , with the following nonmetals to form ionic solids:  
 nitrogen, oxygen forming oxide ions ( $O^{2-}$ ), sulfur, bromine, iodine.

Reaction for decomposition of dinitrogen oxide gas to its elements.

Reaction of solid calcium carbide ( $CaC_2$ ) with water to form calcium hydroxide and acetylene ( $C_2H_2$ ).

B. *mass/mole calculations involving balanced equations*. Write a balanced equation for:

reaction of magnesium oxide with iron to form iron(III) oxide and magnesium. HOW MANY...

- i. mole MgO will react with 8.46 mole iron?  
 ii. mole  $Fe_2O_3$  will form when 0.735 grams of Mg are produced?  
 iii. mole  $Fe_2O_3$  will form from  $5.75 \times 10^{20}$  Fe atoms/?
- i. grams MgO are required to form 4.37 grams of  $Fe_2O_3$ ?  
 ii. grams MgO will form 5.382 moles of  $Fe_2O_3$ ?  
 iii. grams  $Fe_2O_3$  will be formed along with  $2.58 \times 10^{25}$  formula units of MgO?
- i. atoms Mg will be formed from 0.836 moles of Fe?  
 ii. formula units MgO will react with  $1.57 \times 10^{12}$  atoms of Fe?  
 iii. formula units MgO are necessary to form 5.33 grams of  $Fe_2O_3$ ?

- C. *limiting reagents* - It is impractical (if not impossible) to measure exact quantities of reactants required for complete conversion to products, so that no reactant materials remain afterwards. In reality, some reactant(s) must be present in an excess amount before reaction, and they will remain so after the reaction is completed.

For reactions that go to completion (i.e., reactions having very large equilibrium constants) let us presume that at least one reactant will be completely used up during the reaction so none of it will remain after reaction. Let's call this reactant the limiting reagent. The amount of product formed as a result of a chemical reaction will depend only on the one limiting reagent, and is independent of all other reactant(s) present in excess. Let's take some examples:

Write a balanced equation for the combustion of propane ( $C_3H_8$ ).

- i. What is the limiting reagent when 1.48 moles of propane are reacted with 4.87 moles of oxygen? How many moles of excess reagent remain after reaction?
  - ii. What is the limiting reagent when 10.27 g of propane is reacted with 56.31 g oxygen? How many grams of water will be formed? How many grams of excess reagent remain after reaction?
  - iii. Suppose equal weights of zinc and iodine are reacted to form zinc iodide,  $ZnI_2$  ... What fraction by weight of the element present in excess remains unreacted?
- D. *reaction efficiencies - per cent yield* Write a balanced equation for reaction of:
- i. calcium cyanamide ( $CaCN_2$ ) with water, to form calcium carbonate and ammonia gas.  
How many grams of ammonia can be formed from 5.3279 g of calcium cyanamide, if the yield for this reaction is 70 %?
  - ii. boron trifluoride gas with water, to form hydrogen fluoride and solid boric acid ( $H_3BO_3$ ).  
How many grams of boron trifluoride are necessary to obtain 24.000 g HF, if the yield for this reaction is 83 %?

## Answers

### 3. Chemical Stoichiometry - other applications for ratios of mass of

- A. i.  $C_8H_{16}O_2$  has MW =  $8(12)+16(1)+2(16) = 144$  g/mole,  
of which  $8(12) = 96$  g is Carbon.

$$? \text{ g C} = 57.83 \text{ g compd} \left[ \frac{96 \text{ g C}}{144 \text{ g compd}} \right] = 38.55 \text{ g carbon}$$

- ii. Of the 144 parts by weight of compound, hydrogen accounts for 16 parts.

$$? \text{ g compd} = 14.275 \text{ g H} \left[ \frac{144 \text{ g compd}}{16 \text{ g H}} \right] = 128.5 \text{ g compd}$$

- B.  $NaHSO_4$  has a FW =  $23+1+32+4(16) = 120$  g/mole,  
of which sodium accounts for 23 parts and sulfur for 32 parts.

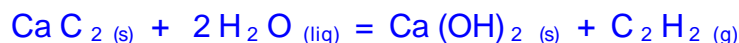
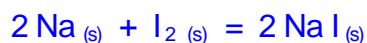
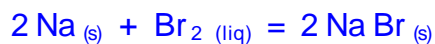
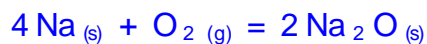
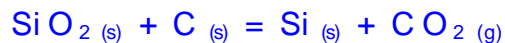
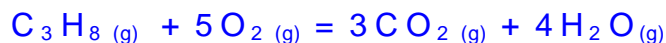
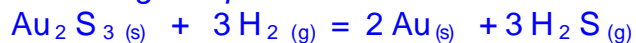
$$? \text{ g S} = 4.829 \text{ g Na} \left[ \frac{32 \text{ g S in NaHSO}_4}{23 \text{ g Na in NaHSO}_4} \right] = 6.719 \text{ g sulfur}$$

- C.  $HgO$  has a FW =  $200.6+16 = 216.6$  g/mole, of which oxygen accounts for 16 parts.

$$? \text{ g HgO} = 2.594 \text{ g O} \left[ \frac{216.6 \text{ g HgO}}{16 \text{ g O}} \right] = 35.12 \text{ g HgO}$$

### 4. Chemical Stoichiometry - applied to chemical reactions as represented by chemical equations

#### A. Balancing of equations



B. mass/mole calculations for the reaction:  $3 \text{ Mg O}_{(s)} + 2 \text{ Fe}_{(s)} = \text{Fe}_2 \text{ O}_3_{(s)} + 3 \text{ Mg}_{(s)}$

$$\text{i. } ? \text{ mole MgO} = 8.46 \text{ mole Fe} \left[ \frac{3 \text{ mole MgO}}{2 \text{ mole Fe}} \right] = 12.7 \text{ mole MgO}$$

$$\text{ii. } ? \text{ mole Fe}_2 \text{ O}_3 = 0.735 \text{ g Mg} \left[ \frac{1 \text{ mole Fe}_2 \text{ O}_3}{3(24) \text{ g Mg}} \right] = 1.02 \text{ E} - 2 \text{ mole Fe}_2 \text{ O}_3$$

$$\text{iii. } ? \text{ mole Fe}_2 \text{ O}_3 = 5.75 \text{ E} + 20 \text{ Fe atoms} \left[ \frac{1 \text{ mole Fe}_2 \text{ O}_3}{2(6.02 \text{ E} + 23) \text{ iron atoms}} \right]$$

$$= 4.78 \text{ E} - 4 \text{ moles Fe}_2 \text{ O}_3$$

$$\text{iv. } ? \text{ g MgO} = 4.37 \text{ g Fe}_2 \text{ O}_3 \left[ \frac{3(24+16) \text{ g MgO}}{(2(55.85) + 3(16)) \text{ g Fe}_2 \text{ O}_3} \right] = 3.28 \text{ g MgO}$$

$$\text{v. } ? \text{ g MgO} = 5.382 \text{ moles Fe}_2 \text{ O}_3 \left[ \frac{3(40) \text{ g MgO}}{1 \text{ mole Fe}_2 \text{ O}_3} \right] = 645.8 \text{ g MgO}$$

$$\text{vi. } ? \text{ g Fe}_2 \text{ O}_3 = 2.58 \text{ E} + 25 \text{ formula units MgO} \left[ \frac{159.7 \text{ g Fe}_2 \text{ O}_3}{3(6.02 \text{ E} + 23) \text{ formula units MgO}} \right]$$

$$= 2.28 \text{ E} + 3 \text{ g Fe}_2 \text{ O}_3$$

$$\text{vii. } ? \text{ atoms Mg} = 0.836 \text{ moles Fe} \left[ \frac{3(6.02 \text{ E} + 23) \text{ atoms Mg}}{2 \text{ moles Fe}} \right] = 7.55 \text{ E} + 23 \text{ atoms Mg}$$

$$\text{viii. } ? \text{ formula units MgO} = 1.57 \text{ E} + 12 \text{ atoms Fe} \left[ \frac{3(6.02 \text{ E} + 23) \text{ formula units MgO}}{2(6.02 \text{ E} + 23) \text{ atoms Fe}} \right]$$

$$= 2.36 \text{ E} + 12 \text{ formula units MgO}$$

$$\text{ix. } ? \text{ formula units MgO} = 5.33 \text{ g Fe}_2 \text{ O}_3 \left[ \frac{3(6.02 \text{ E} + 23) \text{ formula units MgO}}{159.7 \text{ g Fe}_2 \text{ O}_3} \right]$$

$$= 6.03 \text{ E} + 22 \text{ formula units MgO}$$

C. *limiting reagents*



When the amounts of two reactants is known, then one is most likely to be present in an excess

amount, and the other is said to be the limiting reagent. The limiting reagent controls the amount of product that can be formed. In order to identify the limiting reagent, trial calculations are performed to show the amount of a product that each reactant will form. The reactant forming the LEAST amount of product is the limiting reagent, and the LEAST amount of product formed is also the maximum amount that could be formed.

Suggest performing trial calculations to find the amount of  $C O_2$  that 1.48 moles of propane can form, and also to find the amount of  $C O_2$  that 4.87 moles of oxygen can form.

$$? \text{ mole water} = 1.48 \text{ mole } C_3 H_8 \left[ \frac{3 \text{ mole } CO}{1 \text{ mole propane}} \right] = 4.44 \text{ mole } C O_2$$

$$? \text{ mole water} = 4.87 \text{ mole } O_2 \left[ \frac{3 \text{ mole } CO_2}{5 \text{ mole } O_2} \right] = 2.92 \text{ mole } C O_2$$

The LEAST amount of  $C O_2$  formed is from 4.87 mole oxygen, so  $O_2$  is the limiting reagent. Only 2.92 mole  $C O_2$  can be formed b/c there is not enough oxygen to react completely with all of the 1.48 mole of  $C_3 H_8$  present.  $C_3 H_8$  is present in excess.

The amount of  $C_3 H_8$  that reacts can be found finding how much  $C_3 H_8$  will (1) react with 4.87 mole oxygen, or (2) be required to form 2.92 mole  $C O_2$ . Either calculation yields the same result.

$$? \text{ mole } C_3 H_8 \text{ reacts} = 4.87 \text{ mole } O_2 \left[ \frac{1 \text{ mole } C_3 H_8}{5 \text{ mole } O_2} \right] = 0.974 \text{ moles of } C_3 H_8 \text{ react}$$

Number of moles of excess reagent (i.e., propane) remaining after reaction is the difference between starting amount and amount reacted.

starting amt propane	1.48 mole
amount reacted	<u>-0.974</u>
	0.51 mole of propane is amount remaining

- ii. Again, perform trial calculations with both reactants; the one forming the LEAST of a common product is the limiting reagent.

$$? \text{ g water} = 10.27 \text{ g } C_3 H_8 \left[ \frac{4(18) \text{ g } H_2O}{(3(12)+8(1)) \text{ g } C_3 H_8} \right] = 16.81 \text{ g } H_2 O$$

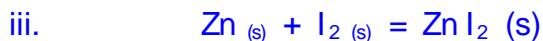
$$? \text{ g water} = 56.31 \text{ g } O_2 \left[ \frac{72 \text{ g } H_2O}{5(32) \text{ g } O_2} \right] = 25.34 \text{ g } H_2 O$$

The LEAST amount of water is formed from 10.27 g  $C_3 H_8$ , so propane is the limiting reagent in this case - and only 16.81 g of water can be produced from this set of starting chemicals.

Oxygen is the reagent present in an excess, and the amount remaining after reaction can be determined by calculating the amount actually used in this reaction.

$$? \text{ g O}_2 = 16.81 \text{ g H}_2\text{O} \left[ \frac{160 \text{ g O}}{4(18) \text{ g H}_2\text{O}} \right] 37.36 \text{ g O}_2 \text{ reacts}$$

$$\begin{array}{r} \text{starting amount O}_2 \quad 56.31 \text{ g} \\ \text{amount O}_2 \text{ reacted} \quad \underline{-37.36 \text{ g}} \\ \hline 18.95 \text{ g O}_2 \text{ remain after reaction is complete} \end{array}$$



The equation can be read as follows:

ONE mole Zn reacts with ONE mole I<sub>2</sub> to form ONE mole ZnI<sub>2</sub>, or as...

$$65.4 \text{ g Zn reacts with } 254 \text{ g I}_2 \text{ to form } 319.4 \text{ g ZnI}_2$$

Although any mass may be used solving this problem, it may be simplifying to use one of the standard masses. Suppose a mass of 65.4 g is used for both zinc and iodine.

From reading the equation it becomes obvious that 65.4 g of Zn requires 254 g of I<sub>2</sub> for complete reaction. However, the conditions of the problem state that only 65.4 g of both elements is present (i.e., only equal masses of both are present). Consequently, zinc is present in an excess amount and iodine is the limiting reagent. To find the ratio of zinc that remains unreacted, it is necessary to find the mass of zinc that has reacted...

$$? \text{ g Zn} = 65.4 \text{ g I}_2 \left[ \frac{65.4 \text{ g Zn}}{254 \text{ g I}_2} \right] = 16.84 \text{ g Zn has reacted.}$$

The fraction of the element present in excess that remains unreacted is given by...

$$\left[ \frac{\text{unreacted g Zn}}{\text{initial g Zn}} \right] = (65.4 - 16.84) / 65.4 = 0.743$$

D. *reaction efficiencies - per cent yield*

Most chemical reactions are not perfectly efficient, i.e., they do not produce the maximum amount of product possible. The amount of product actually formed is expressed as a per cent yield. When per cent yields are less than 100%, then the amount of product formed is LESS than that calculated for a given amount of reactant, and reactant required is MORE than that calculated for a given amount of product.



Perform a trial calculation based on 100 % yield before considering the actual yield

$$? \text{ g NH}_3 = 5.3279 \text{ g CaCN}_2 \left[ \frac{2(17) \text{ g NH}_3}{(40+12+2(14)) \text{ g CaCN}_2} \right] = 2.264 \text{ g NH}_3$$

So, if the reaction was 100% efficient, then 2.264 g ammonia would be formed.  $\text{NH}_3$  is a *product* so LESS than that calculated will be formed b/c the reaction has only 70 % yield. How much product ammonia will be actually formed?

$$? \text{ g NH}_3 \text{ (for 70% yield)} = 2.264 \text{ g NH}_3 \left[ \frac{70 \%}{100 \%} \right] = 1.585 \text{ g NH}_3 \text{ actually formed.}$$



(Note that the 24.000 g HF is to be associated with 83 % yield, b/c it is the amount of product produced from a reaction that is only 83 % efficient.)

Perform a trial calculation based on 83 % yield of HF (24.000 g) before considering the actual yield.

$$? \text{ g BF}_3 = 24.000 \text{ g HF} \left[ \frac{67.8 \text{ g BF}_3}{3(20) \text{ g HF}} \right] = 27.12 \text{ g BF}_3$$

So, if all  $\text{BF}_3$  were converted to HF (i.e., if the reaction was 100 % efficient from the reactant viewpoint) then only 27.12 g  $\text{BF}_3$  would be required to form 24.000 g HF. However,  $\text{BF}_3$  is a reactant in a reaction that has only 83% yield, so MORE than that calculated will be required. How much reactant  $\text{BF}_3$  will be actually needed?

$$? \text{ g BF}_3 \text{ (for 83% yield)} = 27.12 \text{ g BF}_3 \left[ \frac{100 \%}{83 \%} \right] = 32.67 \text{ g BF}_3 \text{ actually required.}$$