

## Chapter 3

# Calculations with Chemical Formulas and Equations

### Concept Check 3.1

You have 1.5 moles of tricycles.

- How many moles of seats do you have?
- How many moles of tires do you have?
- How could you use parts a. and b. as an analogy to teach a friend about the number of moles of  $\text{OH}^-$  ions in 1.5 moles of  $\text{Mg}(\text{OH})_2$ ?

### Solution

- Each tricycle has one seat, so you have a total of 1.5 mol of seats.
- Each tricycle has three tires, so you have  $1.5 \text{ mol} \times 3 = 4.5 \text{ mol}$  of tires.
- Each  $\text{Mg}(\text{OH})_2$  has two  $\text{OH}^-$  ions, so there are  $1.5 \text{ mol} \times 2 = 3.0 \text{ mol}$   $\text{OH}^-$  ions.

### Concept Check 3.2

You perform combustion analysis on a compound that contains only C and H.

- Considering the fact that the combustion products  $\text{CO}_2$  and  $\text{H}_2\text{O}$  are colorless, how can you tell if some of the product got trapped in the  $\text{CuO}$  pellets (see Figure 3.6)?
- Would your calculated results of mass percentage of C and H be affected if some of the combustion products got trapped in the  $\text{CuO}$  pellets? If your answer is yes, how might your results differ from the expected value for the compound?

### Solution

- When conducting this type of experiment, you are assuming that all of the carbon and hydrogen show up in the  $\text{CO}_2$  and  $\text{H}_2\text{O}$ , respectively. In this experiment, where all of the carbon and hydrogen do not show up, when you analyze the  $\text{CO}_2$  for carbon and  $\text{H}_2\text{O}$  for hydrogen, you will find that the weights in the products are less than those in the carbon and hydrogen you started with.
- Since you collected less carbon and hydrogen than were present in the original sample, the calculated mass percentage will be less than the expected (real) value. For example, say you have a 10.0-g sample that contains 7.5 g of carbon. You run the experiment on the 10.0-g sample and collect only 5.0 g of carbon. The calculated percent carbon based on your experimental results would be 50% instead of the correct amount of 75%.

### Concept Check 3.3

A friend has some questions about empirical formulas and molecular formulas. You can assume that he is good at performing the calculations.

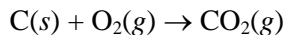
- For a problem that asked him to determine the empirical formula, he found the answer  $\text{C}_2\text{H}_8\text{O}_2$ . Is this a possible answer to the problem? If not, what guidance would you offer your friend?
- For another problem he found the answer  $\text{C}_{1.5}\text{H}_4$  as the empirical formula. Is this answer correct? Once again, if it isn't correct, what could you do to help your friend?
- Since you have been a big help, your friend asks one more question. He completed a problem of the same type as Example 3.11. His answers indicate that the compound has an empirical formula of  $\text{C}_3\text{H}_8\text{O}$  and the molecular formula  $\text{C}_3\text{H}_8\text{O}$ . Is this result possible?

### Solution

- $\text{C}_2\text{H}_8\text{O}_2$  is not an empirical formula, since each of the subscripts can be divided by two to obtain a possible empirical formula of  $\text{CH}_4\text{O}$ . (The empirical formula is not the smallest integer ratio of subscripts.)
- $\text{C}_{1.5}\text{H}_4$  is not a correct empirical formula, because one of the subscripts is not an integer. Multiply each of the subscripts by two to obtain the possible empirical formula  $\text{C}_3\text{H}_8$ . (Since the subscript of carbon is the decimal number 1.5, the empirical formula is not the smallest integer ratio of subscripts.)
- Yes, the empirical formula and the molecular formula can be the same, as is the case in this problem where the formula is written with the smallest integer subscripts.

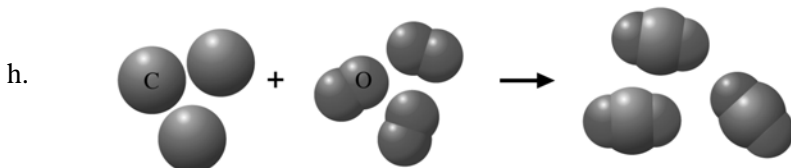
### Concept Check 3.4

The main reaction of a charcoal grill is



Which of the statements below are incorrect? Why?

- 1 atom of carbon reacts with 1 molecule of oxygen to produce 1 molecule of  $\text{CO}_2$ .
- 1 g of C reacts with 1 g of  $\text{O}_2$  to produce 2 g of  $\text{CO}_2$ .
- 1 g of C reacts with 0.5 g of  $\text{O}_2$  to produce 1 g of  $\text{CO}_2$ .
- 12 g of C reacts with 32 g of  $\text{O}_2$  to produce 44 g of  $\text{CO}_2$ .
- 1 mol of C reacts with 1 mol of  $\text{O}_2$  to produce 1 mol of  $\text{CO}_2$ .
- 1 mol of C reacts with 0.5 mol of  $\text{O}_2$  to produce 1 mol of  $\text{CO}_2$ .



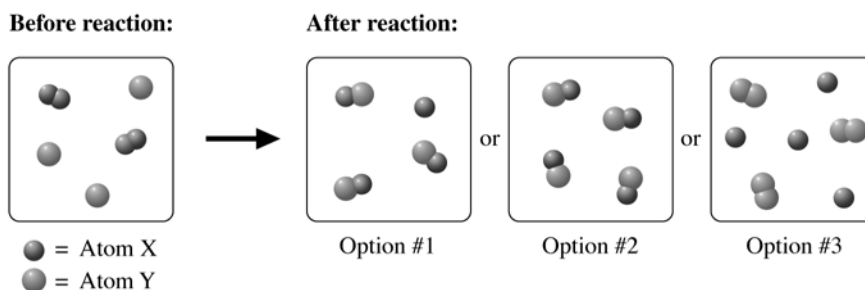
### Solution

- Correct. The coefficients in balanced equations can represent amounts in atoms and molecules.
- Incorrect. The coefficients in a balanced chemical equation do not represent amounts in grams. One gram of carbon and one gram of oxygen represent different molar amounts.
- Incorrect. The coefficients in a balanced chemical equation do not represent amounts in grams.
- Correct. You might initially think that this is an incorrect representation; however, 12 g of C, 32 g of  $\text{O}_2$ , and 44 g of  $\text{CO}_2$  each represent one mole of the substance, so the relationship of the chemical reaction is obeyed.
- Correct. The coefficients in balanced equations can represent amounts in moles.
- Incorrect. The amount of  $\text{O}_2$  present is not enough to react completely with one mole of carbon. Only one-half of the carbon would react, and one-half mole of  $\text{CO}_2$  would form.
- Incorrect. In this representation, oxygen is being shown as individual atoms of O, not as molecules of  $\text{O}_2$ , so the drawing is not correctly depicting the chemical reaction.
- Correct. The molecular models correctly depict a balanced chemical reaction since the same number of atoms of each element appears on both sides of the equation.

### Concept Check 3.5

You perform the hypothetical reaction of an element,  $X_2(g)$ , with another element,  $Y(g)$ , to produce  $XY(g)$ .

- Write the balanced chemical equation for the reaction.
- If  $X_2$  and  $Y$  were mixed in the quantities shown in the container below and allowed to react, which of the following options is the correct representation of the contents of the container after the reaction has occurred?
- Using the information presented in part b., identify the limiting reactant.

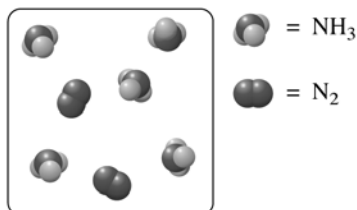


### Solution

- $X_2(g) + 2Y(g) \rightarrow 2XY(g)$
- Since the product consists of a combination of X and Y in a 1:1 ratio, it must consist of two atoms hooked together. If you count the total number of X atoms (split apart the  $X_2$  molecules) and Y atoms present prior to the reaction there are four X atoms and three Y atoms. From these starting quantities, you are limited to three XY molecules and left with an unreacted Y. Option #1 represents this situation and is therefore the correct answer.
- Since  $X_2(g)$  was completely used up during the course of the reaction, it is the limiting reactant.

### Conceptual Problem 3.13

You react nitrogen and hydrogen in a container to produce ammonia,  $NH_3(g)$ . The following figure depicts the contents of the container after the reaction is complete.



- Write a balanced chemical equation for the reaction.
- What is the limiting reactant?

- c. How many molecules of the limiting reactant would you need to add to the container in order to have a complete reaction (convert all reactants to products)?

### Solution

- a.  $3\text{H}_2(\text{g}) + \text{N}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$
- b. Since there is no  $\text{H}_2$  present in the container, it was entirely consumed during the reaction which makes it the limiting reactant.
- c. According to the chemical reaction, three molecules of  $\text{H}_2$  are required for every one molecule of  $\text{N}_2$ . Since there are two moles of unreacted  $\text{N}_2$ , you would need six additional moles of  $\text{H}_2$  to complete the reaction.

### Conceptual Problem 3.14

Propane,  $\text{C}_3\text{H}_8$ , is the fuel of choice in a gas barbecue. When propane burns, the balanced equation is  $\text{C}_3\text{H}_8 + 5\text{O}_2 \rightarrow 3\text{CO}_2 + 4\text{H}_2\text{O}$ .

- a. What is the limiting reactant when cooking with a gas grill?
- b. If the grill will not light and you know that you have an ample flow of propane to the burner, what is the limiting reactant?
- c. When using a gas grill you can sometimes turn the gas up to the point at which the flame becomes yellow and smoky. In terms of the chemical reaction, what is happening?

### Solution

- a. The limiting reactant when cooking with a gas grill would be the propane. This makes sense since propane is the material that you must purchase in order to cook your food.
- b. Since the chemical reaction only requires propane and oxygen, if the grill will not light with ample propane present, then the limiting reactant must be oxygen.
- c. Once again, here is a case where you have adequate propane, so you can conclude that a yellow flame indicates that not enough oxygen is present to combust all of the propane. If there is not enough  $\text{O}_2$  available for complete combustion, a reasonable assumption is that some of the products will have fewer oxygen atoms than  $\text{CO}_2$ . Therefore, a mixture of products would be obtained in this case, including carbon monoxide ( $\text{CO}$ ) and soot (carbon particles).

### Conceptual Problem 3.15

A critical point to master in becoming proficient at solving problems is evaluating whether or not your answer is reasonable. A friend asks you to look over her homework to see if she has done the calculations correctly. Shown below are descriptions of some of her answers. Without using your calculator or doing calculations on paper, see if you can judge the

answers below as being reasonable or ones that will require her to go back and work the problems again.

- 0.33 mol of an element has a mass of  $1.0 \times 10^{-3}$  g.
- The mass of one molecule of water is  $1.80 \times 10^{-10}$  g.
- There are  $3.01 \times 10^{23}$  atoms of Na in 0.500 mol of Na.
- The molar mass of  $\text{CO}_2$  is 44.0 kg/mol.

### Solution

- This answer is unreasonable because  $1.0 \times 10^{-3}$  g is too small for 0.33 mole of an element. For example, 0.33 mole of hydrogen, the lightest element, would have a mass of  $3.3 \times 10^{-1}$  g.
- This answer is unreasonable because  $1.80 \times 10^{-10}$  g is too large for one water molecule. (The mass of one water molecule is  $2.99 \times 10^{-23}$  g.)
- This answer is reasonable because  $3.01 \times 10^{23}$  is one-half of Avogadro's number.
- This answer is unreasonable because the units for molar mass should be g/mol, so this quantity is 1000 times too large.

### Conceptual Problem 3.16

An exciting, and often loud, chemical demonstration involves the simple reaction of hydrogen gas and oxygen gas to produce water vapor:  $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g})$ . The reaction is carried out in soap bubbles or balloons filled with the reactant gases. You get the reaction to proceed by igniting the bubbles or balloons. The more  $\text{H}_2\text{O}$  that is formed during the reaction, the bigger the bang. Explain the following observations.

- A bubble containing just  $\text{H}_2$  makes a quiet "fffft" sound when ignited.
- When a bubble containing equal amounts of  $\text{H}_2$  and  $\text{O}_2$  is ignited, a sizable bang results.
- When a bubble containing a ratio of 2 to 1 in the amounts of  $\text{H}_2$  and  $\text{O}_2$  is ignited, the loudest bang results.
- When a bubble containing just  $\text{O}_2$  is ignited, virtually no sound is made.

### Solution

- In order to have a complete reaction, a ratio of two moles of hydrogen to every mole of oxygen is required. In this case, there is not enough oxygen in the air outside of the bubble for the complete reaction of hydrogen.
- In this case you have a ratio of one mole of  $\text{H}_2$  to one mole of  $\text{O}_2$ . According to the balanced chemical reaction, every one mole of  $\text{O}_2$  can react with two moles of  $\text{H}_2$ . In this case, when 0.5 mole of  $\text{O}_2$  has reacted, all of the  $\text{H}_2$  (one mole) will be consumed, leaving behind 0.5 mole of unreacted  $\text{O}_2$ .
- In this case you have a ratio of two moles of  $\text{H}_2$  to one mole of  $\text{O}_2$ , which is the correct stoichiometric amount, so all of the hydrogen and all of the oxygen react completely.

- d. In order for the reaction to occur, both oxygen and hydrogen must be present. Oxygen does not combust, and there is no hydrogen present to burn so no reaction occurs.

### Conceptual Problem 3.17

High cost and limited availability of a reactant often dictate which reactant is limiting in a particular process. Identify the limiting reactant when running the reactions below and give a reason to support your decision.

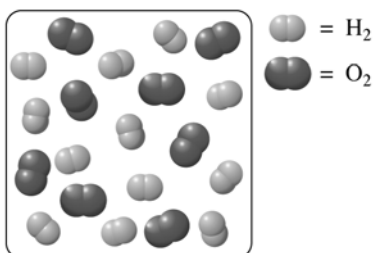
- Burning charcoal on a grill:  $C(s) + O_2(g) \rightarrow CO_2(g)$
- Burning a chunk of Mg in water:  $Mg(s) + 2H_2O(l) \rightarrow Mg(OH)_2(aq) + H_2(g)$
- The Haber process of ammonia production:  $3H_2(g) + N_2(g) \rightarrow 2NH_3(g)$

#### Solution

- The limiting reactant would be the charcoal because the air would supply as much oxygen as needed.
- The limiting reactant would be the magnesium because the beaker would contain much more water than is needed for the reaction (approximately 18 mL of water is one mole).
- The limiting reactant would be the  $H_2$  because the air could supply as much nitrogen as is needed.

### Conceptual Problem 3.18

A few hydrogen and oxygen molecules are introduced into a container in the quantities depicted in the following drawing. The gases are then ignited by a spark causing them to react and form  $H_2O$ .



- What is the maximum number of water molecules that can be formed in the chemical reaction?
- Draw a molecular level representation of the container's contents after the chemical reaction.

#### Solution

- a. Since the balanced chemical equation for the reaction is  $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$ , in order to form the water, you need two molecules of hydrogen for every one molecule of oxygen. Given the quantities of reactants present in the container (12  $\text{H}_2$  molecules and 8  $\text{O}_2$  molecules) and applying the 2:1 ratio, you can produce a maximum of 12 molecules of water.
- b. The drawing of the container after the reaction should contain 12  $\text{H}_2\text{O}$  molecules and two  $\text{O}_2$  molecules.