

3.9 Stoichiometric Calculations: Amounts of Reactants and Products

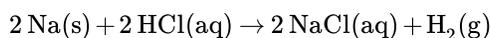
Stoichiometry is a section of chemistry that involves using relationships between reactants and/or products in a chemical reaction to determine desired quantitative data. In Greek, *stoikhein* means element and *metron* means measure, so stoichiometry literally translated means the measure of elements. In order to use stoichiometry to run calculations about chemical reactions, it is important to first understand the relationships that exist between products and reactants and why they exist, which require understanding how to balance reactions.

Balancing

In chemistry, chemical reactions are frequently written as an equation, using chemical symbols. The reactants are displayed on the left side of the equation and the products are shown on the right, with the separation of either a single or double arrow that signifies the direction of the reaction. The significance of single and double arrow is important when discussing solubility constants, but we will not go into detail about it in this module. To balance an equation, it is necessary that there are the same number of atoms on the left side of the equation as the right. One can do this by raising the coefficients.

Reactants to Products

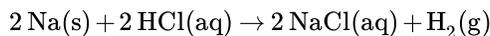
A chemical equation is like a recipe for a reaction so it displays all the ingredients or terms of a chemical reaction. It includes the elements, molecules, or ions in the reactants and in the products as well as their states, and the proportion for how much of each particle reacts or is formed relative to one another, through the stoichiometric coefficient. The following equation demonstrates the typical format of a chemical equation:



In the above equation, the elements present in the reaction are represented by their chemical symbols. Based on the [Law of Conservation of Mass](#), which states that matter is neither created nor destroyed in a chemical reaction, every chemical reaction has the same elements in its reactants and products, though the elements they are paired up with often change in a reaction. In this reaction, sodium (*Na*), hydrogen (*H*), and chloride (*Cl*) are the elements present in both reactants, so based on the law of conservation of mass, they are also present on the product side of the equations. Displaying each element is important when using the chemical equation to convert between elements.

Stoichiometric Coefficients

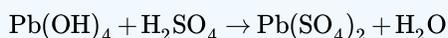
In a balanced reaction, both sides of the equation have the same number of elements. The stoichiometric coefficient is the number written in front of atoms, ion and molecules in a chemical reaction to balance the number of each element on both the reactant and product sides of the equation. Though the stoichiometric coefficients can be fractions, whole numbers are frequently used and often preferred. This stoichiometric coefficients are useful since they establish the mole ratio between reactants and products. In the balanced equation:



we can determine that 2 moles of *HCl* will react with 2 moles of *Na(s)* to form 2 moles of *NaCl(aq)* and 1 mole of *H_{2(g)}*. If we know how many moles of *Na* reacted, we can use the ratio of 2 moles of *NaCl* to 2 moles of *Na* to determine how many moles of *NaCl* were produced or we can use the ratio of 1 mole of *H₂* to 2 moles of *Na* to convert to *NaCl*. This is known as the coefficient factor. The balanced equation makes it possible to convert information about the change in one reactant or product to quantitative data about another reactant or product. Understanding this is essential to solving stoichiometric problems.

✓ Example 1

Lead (IV) hydroxide and sulfuric acid react as shown below. Balance the reaction.



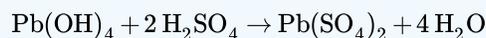
Solution

Start by counting the number of atoms of each element.

UNBALANCED

Element	Reactant (# of atoms)	Product (# of atoms)
Pb	1	1
O	8	9
H	6	2
S	1	2

The reaction is not balanced; the reaction has 16 reactant atoms and only 14 product atoms and does not obey the conservation of mass principle. Stoichiometric coefficients must be added to make the equation balanced. In this example, there are only one sulfur atom present on the reactant side, so a coefficient of 2 should be added in front of H_2SO_4 to have an equal number of sulfur on both sides of the equation. Since there are 12 oxygen on the reactant side and only 9 on the product side, a 4 coefficient should be added in front of H_2O where there is a deficiency of oxygen. Count the number of elements now present on either side of the equation. Since the numbers are the same, the equation is now balanced.



BALANCED

Element	Reactant (# of atoms)	Product (# of atoms)
Pb	1	1
O	12	12
H	8	8
S	2	2

Balancing reactions involves finding least common multiples between numbers of elements present on both sides of the equation. In general, when applying coefficients, add coefficients to the molecules or unpaired elements last.

A balanced equation ultimately has to satisfy two conditions.

1. The numbers of each element on the left and right side of the equation must be equal.
2. The charge on both sides of the equation must be equal. It is especially important to pay attention to charge when [balancing redox reactions](#).

Stoichiometry and Balanced Equations

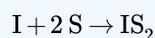
In stoichiometry, balanced equations make it possible to compare different elements through the **stoichiometric factor** discussed earlier. This is the mole ratio between two factors in a chemical reaction found through the ratio of stoichiometric coefficients. Here is a real world example to show how stoichiometric factors are useful.

✓ Example 2

There are 12 party invitations and 20 stamps. Each party invitation needs 2 stamps to be sent. How many party invitations can be sent?

Solution

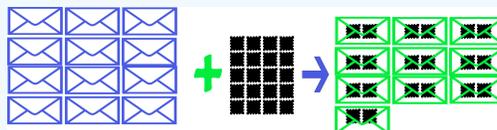
The equation for this can be written as



where

- I represents invitations,
- S represents stamps, and
- IS_2 represents the sent party invitations consisting of one invitation and two stamps.

Based on this, we have the ratio of 2 stamps for 1 sent invite, based on the balanced equation.



Invitations Stamps Party Invitations Sent

In this example are all the reactants (stamps and invitations) used up? No, and this is normally the case with chemical reactions. There is often excess of one of the reactants. The limiting reagent, the one that runs out first, prevents the reaction from continuing and determines the maximum amount of product that can be formed.

✓ Example 3

What is the **limiting reagent** in this example?

Solution

Stamps, because there was only enough to send out invitations, whereas there were enough invitations for 12 complete party invitations. Aside from just looking at the problem, the problem can be solved using stoichiometric factors.

$$12 \text{ I} \times (1\text{IS}_2/1\text{I}) = 12 \text{ IS}_2 \text{ possible}$$

$$20 \text{ S} \times (1\text{IS}_2/2\text{S}) = 10 \text{ IS}_2 \text{ possible}$$

When there is no limiting reagent because the ratio of all the reactants caused them to run out at the same time, it is known as *stoichiometric proportions*.

Types of Reactions

There are 6 basic types of reactions.

- **Combustion:** Combustion is the formation of CO_2 and H_2O from the reaction of a chemical and O_2
- **Combination (synthesis):** Combination is the addition of 2 or more simple reactants to form a complex product.
- **Decomposition:** Decomposition is when complex reactants are broken down into simpler products.
- **Single Displacement:** Single displacement is when an element from one reactant switches with an element of the other to form two new reactants.
- **Double Displacement:** Double displacement is when two elements from one reactants switched with two elements of the other to form two new reactants.
- **Acid-Base:** Acid-base reactions are when two reactants form salts and water.

Molar Mass

Before applying stoichiometric factors to chemical equations, you need to understand molar mass. Molar mass is a useful chemical ratio between mass and moles. The atomic mass of each individual element as listed in the periodic table established this relationship for atoms or ions. For compounds or molecules, you have to take the sum of the atomic mass times the number of each atom in order to determine the molar mass

✓ Example 4

What is the molar mass of H_2O ?

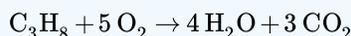
Solution

$$\text{Molar mass} = 2 \times (1.00794 \text{ g/mol}) + 1 \times (15.9994 \text{ g/mol}) = 18.01528 \text{ g/mol}$$

Using molar mass and coefficient factors, it is possible to convert mass of reactants to mass of products or vice versa.

✓ Example 5: Combustion of Propane

Propane (C_3H_8) burns in this reaction:



If 200 g of propane is burned, how many g of H_2O is produced?

Solution

Steps to getting this answer: Since you cannot calculate from grams of reactant to grams of products you must convert from grams of C_3H_8 to moles of C_3H_8 then from moles of C_3H_8 to moles of H_2O . Then convert from moles of H_2O to grams of H_2O .

- Step 1: 200 g C_3H_8 is equal to 4.54 mol C_3H_8 .
- Step 2: Since there is a ratio of 4:1 H_2O to C_3H_8 , for every 4.54 mol C_3H_8 there are 18.18 mol H_2O .
- Step 3: Convert 18.18 mol H_2O to g H_2O . 18.18 mol H_2O is equal to 327.27 g H_2O .

Variation in Stoichiometric Equations

Almost every quantitative relationship can be converted into a ratio that can be useful in data analysis.

Density

Density (ρ) is calculated as mass/volume. This ratio can be useful in determining the volume of a solution, given the mass or useful in finding the mass given the volume. In the latter case, the inverse relationship would be used.

$$\text{Volume} \times (\text{Mass/Volume}) = \text{Mass}$$

$$\text{Mass} \times (\text{Volume/Mass}) = \text{Volume}$$

Percent Mass

Percents establish a relationship as well. A percent mass states how many grams of a mixture are of a certain element or molecule. The percent X% states that of every 100 grams of a mixture, X grams are of the stated element or compound. This is useful in determining mass of a desired substance in a molecule.

✓ Example 6

A substance is 5% carbon by mass. If the total mass of the substance is 10 grams, what is the mass of carbon in the sample? How many moles of carbon are there?

Solution

$$10 \text{ g sample} \times (5 \text{ g carbon}/100 \text{ g sample}) = 0.5 \text{ g carbon}$$

$$0.5 \text{ g carbon} \times (1 \text{ mol carbon}/12.011 \text{ g carbon}) = 0.0416 \text{ mol carbon}$$

Molarity

Molarity (moles/L) establishes a relationship between moles and liters. Given volume and molarity, it is possible to calculate mole or use moles and molarity to calculate volume. This is useful in chemical equations and dilutions.

✓ Example 7

How much 5 M stock solution is needed to prepare 100 mL of 2 M solution?

Solution

100 mL of dilute solution (1 L/1000 mL)(2 mol/1L solution)(1 L stock solution/5 mol solution)(1000 ml stock solution/1L stock solution) = 40 mL stock solution.

These ratios of molarity, density, and mass percent are useful in complex examples ahead.

Determining Empirical Formulas

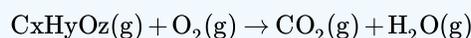
An empirical formula can be determined through chemical stoichiometry by determining which elements are present in the molecule and in what ratio. The ratio of elements is determined by comparing the number of moles of each element present.

✓ Example 8: Combustion of Organic Molecules

1.000 gram of an organic molecule burns completely in the presence of excess oxygen. It yields 0.0333 mol of CO₂ and 0.599 g of H₂O. What is the empirical formula of the organic molecule?

Solution

This is a combustion reaction. The problem requires that you know that organic molecules consist of some combination of carbon, hydrogen, and oxygen elements. With that in mind, write the chemical equation out, replacing unknown numbers with variables. Do not worry about coefficients here.



Since all the moles of C and H in CO₂ and H₂O, respectively have to have come from the 1 gram sample of unknown, start by calculating how many moles of each element were present in the unknown sample.

$$0.0333 \text{ mol CO}_2 (1 \text{ mol C} / 1 \text{ mol CO}_2) = 0.0333 \text{ mol C in unknown}$$

$$0.599 \text{ g H}_2\text{O} (1 \text{ mol H}_2\text{O} / 18.01528 \text{ g H}_2\text{O})(2 \text{ mol H} / 1 \text{ mol H}_2\text{O}) = 0.0665 \text{ mol H in unknown}$$

Calculate the final moles of oxygen by taking the sum of the moles of oxygen in CO₂ and H₂O. This will give you the number of moles from both the unknown organic molecule and the O₂ so you must subtract the moles of oxygen transferred from the O₂.

Moles of oxygen in CO₂:

$$0.0333 \text{ mol CO}_2 (2 \text{ mol O} / 1 \text{ mol CO}_2) = 0.0666 \text{ mol O}$$

Moles of oxygen in H₂O:

$$0.599 \text{ g H}_2\text{O} (1 \text{ mol H}_2\text{O} / 18.01528 \text{ g H}_2\text{O})(1 \text{ mol O} / 1 \text{ mol H}_2\text{O}) = 0.0332 \text{ mol O}$$

Using the Law of Conservation, we know that the mass before a reaction must equal the mass after a reaction. With this we can use the difference of the final mass of products and initial mass of the unknown organic molecule to determine the mass of the O₂ reactant.

$$0.333 \text{ mol CO}_2 (44.0098 \text{ g CO}_2 / 1 \text{ mol CO}_2) = 1.466 \text{ g CO}_2$$

$$1.466 \text{ g CO}_2 + 0.599 \text{ g H}_2\text{O} - 1.000 \text{ g unknown organic} = 1.065 \text{ g O}_2$$

Moles of oxygen in O₂

$$1.065 \text{ g O}_2 (1 \text{ mol O}_2 / 31.9988 \text{ g O}_2)(2 \text{ mol O} / 1 \text{ mol O}_2) = 0.0666 \text{ mol O}$$

Moles of oxygen in unknown

$$(0.0666 \text{ mol O} + 0.0332 \text{ mol O}) - 0.0666 \text{ mol O} = 0.0332 \text{ mol O}$$

Construct a mole ratio for C, H, and O in the unknown and divide by the smallest number.

$$(1/0.0332)(0.0333 \text{ mol C} : 0.0665 \text{ mol H} : 0.0332 \text{ mol O}) \Rightarrow 1 \text{ mol C} : 2 \text{ mol H} : 1 \text{ mol O}$$

From this ratio, the empirical formula is calculated to be CH₂O.

Determining Molecular Formulas

To determine a molecular formula, first determine the empirical formula for the compound as shown in the section above and then determine the molecular mass experimentally. Next, divide the molecular mass by the molar mass of the empirical formula (calculated by finding the sum the total atomic masses of all the elements in the empirical formula). Multiply the subscripts of the molecular formula by this answer to get the molecular formula.

✓ Example 9

In the example above, it was determined that the unknown molecule had an empirical formula of CH_2O .

1. Find the molar mass of the empirical formula CH_2O .

$$12.011\text{g C} + (1.008\text{ g H}) * (2\text{ H}) + 15.999\text{g O} = 30.026\text{ g/mol CH}_2\text{O}$$

2. Determine the molecular mass experimentally. For our compound, it is 120.056 g/mol.

3. Divide the experimentally determined molecular mass by the mass of the empirical formula.

$$(120.056\text{ g/mol}) / (30.026\text{ g/mol}) = 3.9984$$

4. Since 3.9984 is very close to four, it is possible to safely round up and assume that there was a slight error in the experimentally determined molecular mass. If the answer is not close to a whole number, there was either an error in the calculation of the empirical formula or a large error in the determination of the molecular mass.

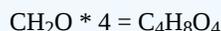
5. Multiply the ratio from step 4 by the subscripts of the empirical formula to get the molecular formula.

$$\text{CH}_2\text{O} * 4 = ?$$

$$\text{C: } 1 * 4 = 4$$

$$\text{H: } 2 * 4 = 8$$

$$\text{O } 1 * 4 = 4$$



6. Check your result by calculating the molar mass of the molecular formula and comparing it to the experimentally determined mass.

$$\text{molar mass of C}_4\text{H}_8\text{O}_4 = 120.104\text{ g/mol}$$

$$\text{experimentally determined mass} = 120.056\text{ g/mol}$$

$$\% \text{ error} = | \text{theoretical} - \text{experimental} | / \text{theoretical} * 100\%$$

$$\% \text{ error} = | 120.104\text{ g/mol} - 120.056\text{ g/mol} | / 120.104\text{ g/mol} * 100\%$$

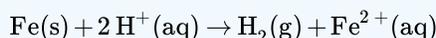
$$\% \text{ error} = 0.040\%$$

✓ Example 10: Complex Stoichiometry Problem

An amateur welder melts down two metals to make an alloy that is 45% copper by mass and 55% iron(II) by mass. The alloy's density is 3.15 g/L. One liter of alloy completely fills a mold of volume 1000 cm^3 . He accidentally breaks off a 1.203 cm^3 piece of the homogenous mixture and sweeps it outside where it reacts with acid rain over years. Assuming the acid reacts with all the iron(II) and not with the copper, how many grams of $\text{H}_2(\text{g})$ are released into the atmosphere because of the amateur's carelessness? (Note that the situation is fiction.)

Solution

Step 1: Write a balanced equation after determining the products and reactants. In this situation, since we assume copper does not react, the reactants are only $\text{H}^+(\text{aq})$ and $\text{Fe}(\text{s})$. The given product is $\text{H}_2(\text{g})$ and based on knowledge of redox reactions, the other product must be $\text{Fe}^{2+}(\text{aq})$.



Step 2: Write down all the given information

$$\text{Alloy density} = (3.15\text{g alloy} / 1\text{L alloy})$$

$$x\text{ grams of alloy} = 45\% \text{ copper} = (45\text{g Cu}(\text{s}) / 100\text{g alloy})$$

$$x\text{ grams of alloy} = 55\% \text{ iron(II)} = (55\text{g Fe}(\text{s}) / 100\text{g alloy})$$

$$1\text{ liter alloy} = 1000\text{cm}^3\text{ alloy}$$

alloy sample = 1.203cm³ alloy

Step 3: Answer the question of what is being asked. The question asks how much H₂(g) was produced. You are expected to solve for the amount of product formed.

Step 4: Start with the compound you know the most about and use given ratios to convert it to the desired compound.

Convert the given amount of alloy reactant to solve for the moles of Fe(s) reacted.

$$1.203\text{cm}^3 \text{ alloy} (1\text{liter alloy}/1000\text{cm}^3 \text{ alloy}) (3.15\text{g alloy}/1\text{liter alloy}) (55\text{g Fe(s)}/100\text{g alloy}) (1\text{mol Fe(s)}/55.8\text{g Fe(s)}) = 3.74 \times 10^{-5} \text{ mol Fe(s)}$$

Make sure all the units cancel out to give you moles of Fe(s). The above conversion involves using multiple stoichiometric relationships from density, percent mass, and molar mass.

The balanced equation must now be used to convert moles of Fe(s) to moles of H₂(g). Remember that the balanced equation's coefficients state the stoichiometric factor or mole ratio of reactants and products.

$$3.74 \times 10^{-5} \text{ mol Fe (s)} (1\text{mol H}_2\text{(g)}/1\text{mol Fe(s)}) = 3.74 \times 10^{-5} \text{ mol H}_2\text{(g)}$$

Step 5: Check units

The question asks for how many grams of H₂(g) were released so the moles of H₂(g) must still be converted to grams using the molar mass of H₂(g). Since there are two H in each H₂, its molar mass is twice that of a single H atom.

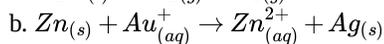
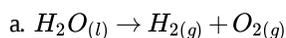
$$\text{molar mass} = 2(1.00794\text{g/mol}) = 2.01588\text{g/mol}$$

$$3.74 \times 10^{-5} \text{ mol H}_2\text{(g)} (2.01588\text{g H}_2\text{(g)}/1\text{mol H}_2\text{(g)}) = \mathbf{7.53 \times 10^{-5} \text{ g H}_2\text{(g) released}$$

Problems

Stoichiometry and balanced equations make it possible to use one piece of information to calculate another. There are countless ways stoichiometry can be used in chemistry and everyday life. Try and see if you can use what you learned to solve the following problems.

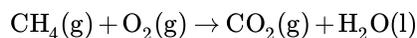
1) Why are the following equations not considered balanced?



2) Hydrochloric acid reacts with a solid chunk of aluminum to produce hydrogen gas and aluminum ions. Write the balanced chemical equation for this reaction.

3) Given a 10.1M stock solution, how many mL must be added to water to produce 200 mL of 5M solution?

4) If 0.502g of methane gas react with 0.27g of oxygen to produce carbon dioxide and water, what is the limiting reagent and how many moles of water are produced? The unbalanced equation is provided below.



5) A 0.777g sample of an organic compound is burned completely. It produces 1.42g CO₂ and 0.388g H₂O. Knowing that all the carbon and hydrogen atoms in CO₂ and H₂O came from the 0.777g sample, what is the empirical formula of the organic compound?

Weblinks for further reference

- 1. Refer to <http://chemistry.about.com/cs/stoich.../aa042903a.htm> as an outside resource on how to balance chemical reactions.
- 2. Refer to <http://www.learnchem.net/tutorials/stoich.shtml> as an outside resource on stoichiometry.

References

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2. J. C. Kotz P.M. Treichel, J. Townsend. Chemistry and Chemical Reactivity. Brooks Cole, February 7, 2008.
3. Petrucci, harwood, Herring, Madura. General Chemistry Principles & Modern Applications. Prentice Hall. New Jersey, 2007.

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