STOICHIOMETRY
via ChemLog

3 Mg (s) + N₂ (g) → Mg₃N₂ (s)

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In the following reaction, one mole of reactant A goes to 2 moles of product B.

\[ A \rightarrow 2B \]

This can also be shown using “blocks” from the ChemLog.

\[ A \rightarrow 2B \]

From the ChemLog, how many moles of B are formed for each mole of A that reacts? To answer this question, fill in the appropriate numbers in this sentence: _____ mole(s) of A react(s) to form _____ mole(s) of B.

From the ChemLog, how many moles of A are needed to decompose for each mole of B produced? To answer this question, fill in the appropriate number in this sentence: For every _____ mole(s) of B that is formed, _____ mole(s) of A decomposed.

Another way to say this is that the ratio of moles of A to B is 1 to 2.

\[ \frac{\text{mol A}}{\text{mol B}} \]

The ratio of moles of B to A is 2 to 1.

\[ \frac{\text{mol B}}{\text{mol A}} \]

We can use this ratio to answer the following question: when starting with one mole of A, how many moles of B will you obtain? Let’s set this up mathematically.

You might be wondering why we chose this ratio (with moles of B on top) rather than the other ratio. The trick is to remember to put the units you want (in our case, we want to get to moles of B) on top. In this example, moles of A cancels, and we’re left with moles of B.

We can answer this question simply by using numbers also. Notice that we get the same answer.

\[ \frac{1 \text{ mol A}}{\text{mol A}} \times \frac{2 \text{ mol B}}{1 \text{ mol A}} = 2 \text{ mol B} \]

Now try this one, when starting with four moles of A, how many moles of B will you obtain?

\[ \text{mol A} \times \frac{2 \text{ mol B}}{\text{mol A}} = \]

Fill in the appropriate numbers for this reaction.

\[ \frac{\text{mol A}}{\text{mol A}} \times \frac{\text{mol B}}{\text{mol A}} = \]

Now try this one, when starting with 3 moles of A, how many moles of B will you obtain?

\[ \text{mol A} \times \frac{2 \text{ mol B}}{\text{mol A}} = \]

When starting with one half a mole of A, how many moles of B will you obtain?

\[ \frac{\text{mol A}}{2} \times \frac{2 \text{ mol B}}{\text{mol A}} = \]
Let’s continue working with the following reaction.

\[ A \rightarrow 2B \]

Here’s a slightly different, but similar question: when you get 2 moles of B from this reaction, how many moles of A did you start with? We’ll go about answering it in the same manner.

\[ \text{mol B} \times \frac{\text{mol A}}{\text{mol B}} = \text{mol A} \]

Notice that we’re using a different ratio here. Our answer needs to be in units of “mol A”, so we use the appropriate ratio with moles of A on top.

Fill in the appropriate numbers for this reaction.

\[ \underline{\text{mol B}} \times \underline{\text{mol A}} = \underline{\text{mol A}} \]

When you get 6 moles of B from this reaction, how many moles of A did you start with?

\[ \text{mol B} \times \frac{\text{mol A}}{\text{mol B}} = \text{mol A} \]

Fill in the appropriate numbers for this reaction.

\[ \underline{\text{mol B}} \times \underline{\text{mol A}} = \underline{\text{mol A}} \]

When you get 3 moles of B from this reaction, how many moles of A did you start with?

How many moles of B are produced in this reaction when you start with 3 moles of A?
Here are some more practice questions.

**Chemical Reactions**

To help with answering the following questions, all the possible ratios are given.

**How many moles of A will react to give 6 moles of B?** Five moles of B?

**How much of the reactant is left after the reaction?**
When 7 moles of A react, how many moles of product will be obtained? Three moles of B?

**Challenge:** How many moles of C will form from 10 moles of A and 1 mole of B?
MASS PERCENT

Determining the mass percent composition of a compound refers to the proportion of one element expressed as a percentage of the total mass of the compound. Knowing the mass percent composition of a compound can help determine environmental effects from that compound. For example, carbon dioxide ($\text{CO}_2$) from burning fossil fuels may contribute to global warming. Methane ($\text{CH}_4$) and butane ($\text{C}_4\text{H}_{10}$) are both fossil fuels that, when burned, produce $\text{CO}_2$. Which one will produce less $\text{CO}_2$? Well, it's the one that contains the least amount of carbon as a percentage of the total compound. Let's use mass percent calculations to determine this.

**Mass Percent of Carbon ($\text{C}$)**

$$\text{Mass Percent of Carbon (C)} = \frac{\text{Mass of Carbon (C)}}{\text{Total Mass of Compound}} \times 100\%$$

First, let us determine the mass percent composition of carbon in butane.

$$\text{Mass Percent of Carbon (C)} = \frac{4 \times 12.011 \text{ g C}}{(4 \times 12.011 \text{ g C}) + (10 \times 1.008 \text{ g H})} \times 100\%$$

$$= \frac{48.044 \text{ g C}}{48.044 \text{ g C} + 10.080 \text{ g H}} \times 100\%$$

$$= \frac{48.044 \text{ g C}}{58.124 \text{ g C}_4\text{H}_{10}} \times 100\%$$

$$= 82.66\%$$

Now, let's consider methane.

Use the same technique to calculate the mass percent of carbon in methane.

Which compound, methane or butane, contains a higher mass percent of carbon? Which one will produce more $\text{CO}_2$ when burned?
**LIMITING REACTANTS**

When carrying out a chemical reaction, we may use the exact amount of each reactant needed. Or, we may use an excess of some reactants and a limited amount of others. We may do this if one reactant is very expensive and others are inexpensive so that we can use all of the expensive compound. It can be more cost effective, even if we are wasting money on the excess reactants. The reactant that governs the maximum yield of a product is the limiting reactant.

\[ \text{CaC}_2(s) + 2 \text{H}_2\text{O}(l) \rightarrow \text{Ca(OH)}_2(aq) + \text{C}_2\text{H}_2(g) \]

**Which is the limiting reactant when 100 g of water reacts with 100 g of calcium carbide?**

First, determine the moles of each reactant that we start with.

Moles of CaC\(_2\)(s) = \( \frac{100 \text{ g CaC}_2(s)}{64.10 \text{ g CaC}_2(s)} \) = 1.56 mol CaC\(_2\)(s)

Moles of H\(_2\text{O}\)(l) = \( \frac{100 \text{ g H}_2\text{O}(l)}{18.02 \text{ g H}_2\text{O}(l)} \) = 5.55 mol H\(_2\text{O}\)(l)

**How many moles of H\(_2\text{O}\) react with 1 mole CaC\(_2\) in this reaction?**

\[ \text{CaC}_2(s) \times \frac{1 \text{ mol CaC}_2(s)}{2 \text{ mol H}_2\text{O}(l)} = 2.78 \text{ mol CaC}_2(s) \]

**Next, what is the amount of CaC\(_2\) that is needed to react with 100 g of H\(_2\text{O}\)?** Convert moles to grams.

5.55 mol H\(_2\text{O}\)(l) x \( \frac{1 \text{ mol CaC}_2(s)}{2 \text{ mol H}_2\text{O}(l)} \) = 2.78 mol CaC\(_2\)(s)

**Compare this answer with what we actually start with.**

2.78 mol CaC\(_2\) (s) > 1.56 mol CaC\(_2\)(s)

**What is the amount of H\(_2\text{O}\) that is needed to react with 100 g of CaC\(_2\)?** Convert moles to grams.

1.56 mol CaC\(_2\)(s) x \( \frac{2 \text{ mol H}_2\text{O}(l)}{1 \text{ mol CaC}_2(s)} \) = 3.12 mol H\(_2\text{O}\)(l)

**Compare this answer with what we actually start with.**

3.12 mol H\(_2\text{O}\)(l) < 5.55 mol H\(_2\text{O}\)(l)

Because 3.12 mol H\(_2\text{O}\) is required and 5.55 mol H\(_2\text{O}\) is supplied, there is an excess of H\(_2\text{O}\). So CaC\(_2\)(s) is the limiting reactant and all of it can react.

**Why is it always important to work in the unit of moles when determining limiting reactants? Why not grams or milliliters?**
There’s another way to determine the limiting reactant using the number of moles of a product that can be made from each reactant.

- How many moles of \( \text{C}_2\text{H}_2 \) is formed from one mole of \( \text{CaC}_2 \)? How many moles of \( \text{C}_2\text{H}_2 \) is formed from one mole of \( \text{H}_2\text{O} \)? Use this information to determine the number of moles of product that can be made from our starting quantities.

\[
\text{Moles of } \text{C}_2\text{H}_2 \text{ (aq)} \text{ from } \text{CaC}_2(s) = 1.56 \text{ mol } \text{CaC}_2(s) \times \frac{1 \text{ mol } \text{C}_2\text{H}_2 \text{ (aq)}}{1 \text{ mol } \text{CaC}_2(s)} = 1.56 \text{ mol } \text{C}_2\text{H}_2
\]

\[
\text{Moles of } \text{C}_2\text{H}_2 \text{ (aq)} \text{ from } \text{H}_2\text{O} \text{ (l)} = 5.55 \text{ mol } \text{H}_2\text{O} \text{ (l)} \times \frac{1 \text{ mol } \text{C}_2\text{H}_2 \text{ (aq)}}{2 \text{ mol } \text{H}_2\text{O} \text{ (l)}} = 2.78 \text{ mol } \text{C}_2\text{H}_2
\]

- From this calculation, which reactant is the limiting reactant? Why? Is it the limiting reactant that was determined previously?

- Determine which reactant is the limiting reactant when 40 g of magnesium and 20 g of nitrogen react in the following reaction:

\[
3 \text{ Mg (s)} + \text{N}_2 \text{ (g)} \longrightarrow \text{Mg}_3\text{N}_2 \text{ (s)}
\]

- How many moles of the limiting reactant are consumed by the reaction? How many grams of the excess reactant are left after the reaction?
YIELDS

The theoretical yield is the maximum product that can be obtained from the amount (mass, moles, volume) of reactant(s) used. Calculate the maximum number of moles of product that can be obtained from the following reaction, when 13.45 g of N\textsubscript{2} reacts with 35 g of Mg.

*Always remember to check for limiting reactants!

\[
3 \text{ Mg (s)} + \text{ N}_2 (g) \rightarrow \text{ Mg}_3\text{N}_2 (s)
\]

The actual yield of a reaction is the amount (moles, volume, mass) of product obtained at the end of the reaction.

The percentage yield can be calculated by:

\[
\text{Percentage Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%
\]

► In the reaction that forms magnesium nitride, the actual yield was 39.8 g. The theoretical yield was not obtained because the Mg was impure, meaning that when 35 g was weighed out, it was not all Mg (s). What is the percentage yield for this reaction?

► Name three reasons that a chemist could obtain a yield less than the theoretical yield.

► Determine the theoretical yield for the reaction magnesium nitride reaction when 35 g of Mg reacts with 16 g of N\textsubscript{2}.
YIELDS

Mark each set of pictures and each ChemLog as either showing a reaction that proceeded to the theoretical yield, the reaction which shows an actual yield, or the reaction which shows there was a limiting reactant.

Connect the reaction pictures above with the ChemLog that each refers to.
MOLARITY

A solution is a homogeneous (uniform in composition) mixture of two chemicals. The solute in a solution is the substance that is being dissolved. The solvent in a solution is the substance doing the dissolving.

Chemists talk about the concentrations of solutions. The concentration of solution is the amount of solute per solvent. One way to state the concentration of a liquid solution is to state its molarity. Molarity is defined as the number of moles of solute per liter of solvent.

\[
\text{Molarity (M)} = \frac{\text{moles of solute}}{\text{liters of solvent}}
\]

Supposed we poured sugar into water to form a solution.

Which of the following pictures would be an accurate representation of the sugar solution? Circle the solution.

Of the following pairs of pictures. Circle the picture which is most concentrated (the picture with the highest molarity).
**ChemLog STOICHIOMETRY**

**BALANCING CHEMICAL REACTIONS**

The numbers to the left of entire chemical formulas in a reaction are called the stoichiometric coefficients. A coefficient of one, as seen in the reaction below, is not written explicitly but is implied. When the number of atoms of each element on each side of the arrow are the same, the reaction is said to be balanced.

\[
\text{N}_2 + \text{H}_2 \rightarrow \text{NH}_3 \quad 100 \text{ C}
\]

Notice the number in front of the product; this is called a stoichiometric coefficient. Since we doubled the number amount of product, we must show this in the chemical reaction using coefficients.

\[
\text{N}_2 + 3 \text{H}_2 \rightarrow 2 \text{NH}_3 \quad 100 \text{ C}
\]

In order to balance the chemical reaction, we must make the number of blue elements before the reaction equal to the number of blue elements after the reaction. The same must be done for the red elements. Let's first begin with the products of the reaction. In order to make the number of blue blocks before and after the reaction equal, we need to double the blue blocks after the reaction. In doing that, we double the red blocks too.

Before Reaction | After Reaction
---|---
2 | 1
2 | 3

Notice that the amount of product has been doubled. Complete the following table using the ChemLog below.

\[
\text{N}_2 + \text{H}_2 \rightarrow 2 \text{NH}_3 \quad 100 \text{ C}
\]

Before Reaction | After Reaction
---|---
2 | 2 x 1
2 | 6
MOLE-TO-MOLE CALCULATIONS

Using the chemical reaction we balanced previously, let’s do some simple calculations with that information.

\[ \text{N}_2 + 3 \text{H}_2 \rightarrow 2 \text{NH}_3 \]

Here are a few ways to think about and visualize this reaction.

► How many moles of product can come from one mole of \( \text{N}_2 \)?

To answer, just look at the fact given in the reaction. One mole of \( \text{N}_2 \) (along with something not important to answer the question) goes to two moles of \( \text{NH}_3 \), the product.

So, 2 moles of product can come from 1 mole of \( \text{N}_2 \).

► How many moles of product can come from three moles of \( \text{H}_2 \)?

Since this question is more difficult and cannot be answered just by looking at the reaction, we’ll use some simple math to figure it out.

\[ X \begin{array}{c} \square \square \square \square \square \square \end{array} = 1 \times \frac{2}{3} \text{ moles NH}_3 \]

\[ = \frac{2}{3} \text{ moles NH}_3 \]

Hint: Remember that the ratio we use has the number of moles of product (the thing we need to answer the question) on the top.

► How many moles of \( \text{NH}_3 \) can be made from four moles of \( \text{N}_2 \)?

► Why must moles be used instead of molecules when answering this question? Try to figure out the number of molecules of product you would make when starting with one molecule of \( \text{H}_2 \). Explain your answer.
Avogadro’s number is the number of particles of substance in a mole. Just as 1 dozen means 12 of something, regardless of what it is - eggs, diamonds, molecules, 1 mole means $6.022 \times 10^{23}$ of something. There are Avogadro’s number, $6.022 \times 10^{23}$, of particles in every mole. That’s a lot!

► How many atoms are in a mole of atoms? How many molecules are in a mole of molecules?

► Since there are the same number of items in a mole, does a mole have a specific mass, say 5 grams? Why or why not?

Let’s look at the following reaction of carbon oxygen to form carbon dioxide.

\[
C + O_2 \rightarrow CO_2
\]

6.022 x $10^{23}$ atoms of C = 1 mole of C

6.022 x $10^{23}$ molecules of O$_2$ = 1 mole of O$_2$

6.022 x $10^{23}$ molecules of CO$_2$ = 1 mole of CO$_2$

What is the mass of 2 moles of C? What is the mass of 0.5 moles of CO$_2$? What is the mass of 1 mole of ozone, O$_3$?

You may be wondering, how scientists came up with Avogadro’s number - $6.022 \times 10^{23}$ isn’t a number one normally thinks of off the top of one’s head! Well, Amadeo Avogadro was not the first scientist who realized this number. However, he was the first scientist to sense the significance of the mole, so the number is named after him. Technically, a mole is an amount of substance that contains as many elementary entities as there are atoms in exactly 12 g of the carbon - 12 isotope.

Avogadro’s Number $= N_A = 6.022 \times 10^{23}$