When there is not enough of one reactant in a chemical reaction, the reaction stops abruptly. To figure out the amount of product produced, it must be determined which reactant will limit the chemical reaction (the limiting reagent) and which reactant is in excess (the excess reagent). One way of finding the limiting reagent is by calculating the amount of product that can be formed by each reactant; the one that produces less product is the limiting reagent.

**Introduction**

The following scenario illustrates the significance of limiting reagents. In order to assemble a car, 4 tires and 2 headlights are needed (among other things). In this example, imagine that the tires and headlights are reactants while the car is the product formed from the reaction of 4 tires and 2 headlights. If you have 20 tires and 14 headlights, how many cars can be made? With 20 tires, 5 cars can be produced because there are 4 tires to a car. With 14 headlights, 7 cars can be built (each car needs 2 headlights). Although more cars can be made from the headlights available, only 5 full cars are possible because of the limited number of tires available. In this case, the headlights are in excess. Because the number of cars formed by 20 tires is less than number of cars produced by 14 headlights, the tires are the limiting reagent (they limit the full completion of the reaction, in which all of the reactants are used up). This scenario is illustrated below:

$4 \text{ Tires} + 2 \text{ Headlights} = 1 \text{ Car}$
The initial condition is that there must be 4 tires to 2 headlights. The reactants must thus occur in that ratio; otherwise, one will limit the reaction. There are 20 tires and 14 headlights, so there are two ways of looking at this problem. For 20 tires, 10 headlights are required, whereas for 14 headlights, 28 tires are required. Because there are not enough tires (20 tires is less than the 28 required), tires are the limiting "reactant."

The limiting reagent is the reactant that is completely used up in a reaction, and thus determines when the reaction stops. From the reaction stoichiometry, the exact amount of reactant needed to react with another element can be calculated. If the reactants are not mixed in the correct stoichiometric proportions (as indicated by the balanced chemical equation), then one of the reactants will be entirely consumed while another will be left over. The limiting reagent is the one that is totally consumed; it limits the reaction from continuing because there is none left to react with the in-excess reactant.

There are two ways to determine the limiting reagent. One method is to find and compare the mole ratio of the reactants used in the reaction (approach 1). Another way is to calculate the grams of products produced from the given quantities of reactants; the reactant that produces the smallest amount of product is the limiting reagent (approach 2).

How to Find the Limiting Reagent: Approach 1

Find the limiting reagent by looking at the number of moles of each reactant.

1. Determine the balanced chemical equation for the chemical reaction.
2. Convert all given information into moles (most likely, through the use of molar mass as a conversion factor).
3. Calculate the mole ratio from the given information. Compare the calculated ratio to the actual ratio.
4. Use the amount of limiting reactant to calculate the amount of product produced.
5. If necessary, calculate how much is left in excess of the non-limiting reagent.

How to Find the Limiting Reagent: Approach 2

Find the limiting reagent by calculating and comparing the amount of product each reactant will produce.

1. Balance the chemical equation for the chemical reaction.
2. Convert the given information into moles.
3. Use stoichiometry for each individual reactant to find the mass of product produced.
4. The reactant that produces a lesser amount of product is the limiting reagent.
5. The reactant that produces a larger amount of product is the excess reagent.
6. To find the amount of remaining excess reactant, subtract the mass of excess reagent consumed from the total
mass of excess reagent given.

Example \(\PageIndex{1}\): Photosynthesis

Consider respiration, one of the most common chemical reactions on earth.

\[
\text{C}_6\text{H}_{12}\text{O}_6 + 6 \text{O}_2 \rightarrow 6 \text{CO}_2 + 6 \text{H}_2\text{O} + \text{energy}
\]

What mass of carbon dioxide forms in the reaction of 25 grams of glucose with 40 grams of oxygen?

**Solution**

When approaching this problem, observe that every 1 mole of glucose \((\text{C}_6\text{H}_{12}\text{O}_6)\) requires 6 moles of oxygen to obtain 6 moles of carbon dioxide and 6 moles of water.

**Step 1:** Determine the balanced chemical equation for the chemical reaction.

The balanced chemical equation is already given.

**Step 2:** Convert all given information into moles (most likely, through the use of molar mass as a conversion factor).

\[
\text{g} \times \frac{\text{mol}}{\text{g}} = \text{mol}
\]

\[
\begin{align*}
25 \text{ g} \times \frac{1 \text{ mol}}{180.06 \text{ g}} &= 0.1388 \text{ mol} \text{ C}_6\text{H}_{12}\text{O}_6 \\
40 \text{ g} \times \frac{1 \text{ mol}}{32 \text{ g}} &= 1.25 \text{ mol} \text{ O}_2
\end{align*}
\]

**Step 3:** Calculate the mole ratio from the given information. Compare the calculated ratio to the actual ratio.

a. If all of the 1.25 moles of oxygen were to be used up, there would need to be \(\frac{1 \text{ mol} \times \frac{1}{6}}{\text{mol} \text{ C}_6\text{H}_{12}\text{O}_6}\) or 0.208 moles of glucose. There is only 0.1388 moles of glucose available which makes it the limiting reactant.

\[
\begin{align*}
1.25 \text{ mol } \text{O}_2 \times \frac{1 \text{ mol} \text{ C}_6\text{H}_{12}\text{O}_6}{6 \text{ mol } \text{O}_2} &= 0.208 \text{ mol} \text{ C}_6\text{H}_{12}\text{O}_6 \\
0.1388 \text{ mol} \text{ C}_6\text{H}_{12}\text{O}_6 \times \frac{6 \text{ mol } \text{O}_2}{1 \text{ mol} \text{ C}_6\text{H}_{12}\text{O}_6} &= 0.8328 \text{ mol } \text{O}_2
\end{align*}
\]

b. If all of the 0.1388 moles of glucose were used up, there would need to be 0.1388 x 6 or 0.8328 moles of oxygen. Because there is an excess of oxygen, the glucose amount is used to calculate the amount of the products in the reaction.

\[
\begin{align*}
0.8328 \text{ mol } \text{O}_2 \times \frac{6 \text{ mol} \text{ C}_6\text{H}_{12}\text{O}_6}{1 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6} &= 4.9968 \text{ mol} \text{ C}_6\text{H}_{12}\text{O}_6 \\
0.1388 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6 \times \frac{6 \text{ mol } \text{O}_2}{1 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6} &= 0.8328 \text{ mol } \text{O}_2
\end{align*}
\]

If more than 6 moles of O\(_2\) are available per mole of C\(_6\)H\(_{12}\)O\(_6\), the oxygen is in excess and glucose is the limiting reactant. If less than 6 moles of oxygen are available per mole of glucose, oxygen is the limiting reactant. The ratio is 6 mole oxygen per 1 mole glucose, OR 1 mole oxygen per 1/6 mole glucose. This means: 6 mol O\(_2\) / 1 mol C\(_6\)H\(_{12}\)O\(_6\).

Therefore, the mole ratio is: \((0.8328 \text{ mol } \text{O}_2)/(0.208 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6)\)

This gives a 4.004 ratio of O\(_2\) to C\(_6\)H\(_{12}\)O\(_6\).

**Step 4:** Use the amount of limiting reactant to calculate the amount of CO\(_2\) or H\(_2\)O produced.
For carbon dioxide produced: \(\text{0.1388 moles glucose} \times \frac{6}{1} = \text{0.8328 moles carbon dioxide}\).

Step 5: **If necessary, calculate how much is left in excess.**

1.25 mol - 0.8328 mol = **0.4172 moles of oxygen left over**

Example \((\text{PageIndex}[2])\): Oxidation of Magnesium

Calculate the mass of magnesium oxide possible if 2.40 g \((\text{Mg})\) reacts with 10.0 g \((\text{O}_2)\)

\[
\text{Mg} + \text{O}_2 \rightarrow \text{MgO}
\]

**Solution**

Step 1: **Balance equation**

\[
\text{2 Mg} + \text{O}_2 \rightarrow 2 \text{MgO}
\]

Step 2 and Step 3: **Converting mass to moles and stoichiometry**

\[
\text{2.40 g Mg} \times \frac{1.00 \text{ mol Mg}}{24.31 \text{ g Mg}} \times \frac{2.00 \text{ mol MgO}}{2.00 \text{ mol Mg}} \times \frac{40.31 \text{ g MgO}}{1.00 \text{ mol MgO}} = \text{3.98 g MgO}
\]

\[
\text{10.0 g O}_2 \times \frac{1 \text{ mol O}_2}{32.0 \text{ g O}_2} \times \frac{2 \text{ mol MgO}}{1 \text{ mol O}_2} \times \frac{40.31 \text{ g MgO}}{1 \text{ mol MgO}} = \text{25.2 g MgO}
\]

Step 4: **The reactant that produces a smaller amount of product is the limiting reagent**

Mg produces less MgO than does O\(_2\) (3.98 g MgO vs. 25.2 g MgO), therefore Mg is the limiting reagent in this reaction.

Step 5: **The reactant that produces a larger amount of product is the excess reagent**

O\(_2\) produces more amount of MgO than Mg (25.2 g MgO vs. 3.98 g MgO), therefore O\(_2\) is the excess reagent in this reaction.

Step 6: **Find the amount of remaining excess reactant by subtracting the mass of the excess reagent consumed from the total mass of excess reagent given.**

Mass of excess reagent calculated using the limiting reagent:

\[
\text{Mass of excess reagent calculated using the mass of the product:}
\]

\[
\text{OR Mass of excess reagent calculated using the mass of the product:}
\]

Example \((\text{PageIndex}[3])\): Limiting Reagent

What is the limiting reagent if 76.4 grams of \((\text{C}_2\text{H}_3\text{Br}_3)\) were reacted with 49.1 grams of \((\text{O}_2)\)?

\[
\text{4 C}_2\text{H}_3\text{Br}_3 + 11 \text{O}_2 \rightarrow 8 \text{CO}_2 + 6 \text{H}_2\text{O} + 6 \text{Br}_2
\]

**Solution**
Using Approach 1:

A. \(76.4 \text{g} \times \frac{1 \text{ mole}}{266.72 \text{g}} = 0.286 \text{ moles of C}_2\text{H}_3\text{Br}_3\)

\(49.1 \text{g} \times \frac{1 \text{ mole}}{32 \text{g}} = 1.53 \text{ moles of O}_2\)

B. Assuming that all of the oxygen is used up, \(1.53 \times \frac{4}{11}\) or 0.556 moles of C$_2$H$_3$Br$_3$ are required. Because there are only 0.286 moles of C$_2$H$_3$Br$_3$ available, C$_2$H$_3$Br$_3$ is the limiting reagent.

Using Approach 2:

\(76.4 \text{g of C}_2\text{H}_3\text{Br}_3 \times \frac{1 \text{ mol of C}_2\text{H}_3\text{Br}_3}{266.72 \text{g of C}_2\text{H}_3\text{Br}_3} \times \frac{8 \text{ mol of CO}_2}{4 \text{ mol of C}_2\text{H}_3\text{Br}_3} \times \frac{44.01 \text{g of CO}_2}{1 \text{ mol of CO}_2} = 25.2 \text{g of CO}_2\)

\(49.1 \text{g of O}_2 \times \frac{1 \text{ mol of O}_2}{32 \text{g of O}_2} \times \frac{8 \text{ mol of CO}_2}{11 \text{ mol of O}_2} \times \frac{44.01 \text{g of CO}_2}{1 \text{ mol of CO}_2} = 49.1 \text{g of CO}_2\)

Therefore, by either method, C$_2$H$_3$Br$_3$ is the limiting reagent.

Example $(\PageIndex{4})$: Limiting Reagent

What is the limiting reagent if 78 grams of Na$_2$O$_2$ were reacted with 29.4 grams of H$_2$O?

Solution

Using Approach 1:

A. \(78 \text{g of Na}_2\text{O}_2 \times \frac{1 \text{ mole}}{77.96 \text{g}} = 1.001 \text{ moles of Na}_2\text{O}_2\)

\(29.4 \text{g of H}_2\text{O} \times \frac{1 \text{ mole}}{18 \text{g}}= 1.633 \text{ moles of H}_2\text{O}\)

B. Assume that all of the water is consumed, \(1.633 \times \frac{2}{2}\) or 1.633 moles of Na$_2$O$_2$ are required. Because there are only 1.001 moles of Na$_2$O$_2$, it is the limiting reactant.

Using Approach 2:

\(78 \text{g of Na}_2\text{O}_2 \times \frac{1 \text{ mol of Na}_2\text{O}_2}{77.96 \text{g of Na}_2\text{O}_2} \times \frac{4 \text{ mol of NaOH}}{2 \text{ mol of Na}_2\text{O}_2} \times \frac{40 \text{g of NaOH}}{1 \text{ mol of NaOH}} = 80.04 \text{g of NaOH}\)

Using either approach gives Na$_2$O$_2$ as the limiting reagent.

Example $(\PageIndex{5})$: Excess Reagent

How much the excess reagent remains if 24.5 grams of CoO is reacted with 2.58 grams of O$_2$?

\[4 \text{CoO} + \text{O}_2 \rightarrow 2 \text{Co}_2\text{O}_3\]

Solution

Using Approach 1:

A. \(24.5 \text{g of CoO} \times \frac{1 \text{ mole}}{74.9 \text{g}}= 0.327 \text{ moles of CoO}\)
\[ \text{mass of } \text{O}_2 = 2.58 \text{ g} \times \frac{1 \text{ mole}}{32 \text{ g}} = 0.0806 \text{ moles of } \text{O}_2 \]

B. Assuming that all of the oxygen is used up, \[ \text{mass of } \text{CoO} = 0.0806 \times \frac{4}{1} = 0.3224 \text{ moles of } \text{CoO} \]

Because there are 0.327 moles of \text{CoO}, \text{CoO} is in excess and thus \text{O}_2 is the limiting reactant.

C. \[ 0.327 \text{ mol} - 0.3224 \text{ mol} = 0.0046 \text{ moles left in excess.} \]

Example (PageIndex{6})): Identifying the Limiting Reagent

Will 28.7 grams of \(\text{SiO}_2\) react completely with 22.6 grams of \(\text{H}_2\text{F}_2\)? If not, identify the limiting reagent.

\[ \text{SiO}_2 + 2 \text{H}_2\text{F}_2 \rightarrow \text{SiF}_4 + 2 \text{H}_2\text{O} \]

**Solution**

A. \[ \text{mass of } \text{SiO}_2 = 28.7 \text{ g} \times \frac{1 \text{ mole}}{60.08 \text{ g}} = 0.478 \text{ moles of } \text{SiO}_2 \]

\[ \text{mass of } \text{H}_2\text{F}_2 = 22.6 \text{ g} \times \frac{1 \text{ mole}}{39.8 \text{ g}} = 0.568 \text{ moles of } \text{H}_2\text{F}_2 \]

B. There must be 1 mole of \text{SiO}_2 for every 2 moles of \text{H}_2\text{F}_2 consumed. Because the ratio is 0.478 to 0.568, 28.7 grams of \text{SiO}_2 do not react with the \text{H}_2\text{F}_2.

C. Assuming that all of the silicon dioxide is used up, \[ \text{mass of } \text{H}_2\text{F}_2 = 0.478 \times \frac{2}{1} = 0.956 \text{ moles of } \text{H}_2\text{F}_2 \]

Because there are only 0.568 moles of \text{H}_2\text{F}_2, it is the *limiting reagent*.

---

**References**


---

**Contributors and Attributions**

- Sarick Shah (UCD)