3.9 Stoichiometric Calcs: Amounts of Reactants and Products

Equation: \[ \text{C}(s) + \text{O}_2(g) \overset{\Delta}{\rightarrow} \text{CO}_2(g) \]
Law of Conservation of Mass

The Law of Conservation of Mass indicates that in an ordinary chemical reaction,

- Matter cannot be created nor destroyed.
- No change in total mass occurs in a reaction.
- Mass of products is equal to mass of reactants.
Conservation of Mass

2 moles Ag + 1 moles S = 1 mole Ag₂S
2 (107.9 g) + 1 (32.1 g) = 1 (247.9 g)

247.9 g reactants = 247.9 g product
Consider the following equation:

\[ 4 \text{ Fe}(s) + 3 \text{ O}_2(g) \rightarrow 2 \text{ Fe}_2\text{O}_3(s) \]

This equation can be read in “moles” by placing the word “moles” between each coefficient and formula.

\[ 4 \text{ moles Fe} + 3 \text{ moles O}_2 \rightarrow 2 \text{ moles Fe}_2\text{O}_3 \]
A mole-mole factor is a ratio of the moles for any two substances in an equation.

\[ 4\text{Fe}(s) + 3\text{O}_2(g) \rightarrow 2\text{Fe}_2\text{O}_3(s) \]

**Fe and O\textsubscript{2}**
- \(4 \text{ moles Fe}\) and \(3 \text{ moles O}_2\)
- \(3 \text{ moles O}_2\) and \(4 \text{ moles Fe}\)

**Fe and Fe\textsubscript{2}O\textsubscript{3}**
- \(4 \text{ moles Fe}\) and \(2 \text{ moles Fe}_2\text{O}_3\)
- \(2 \text{ moles Fe}_2\text{O}_3\) and \(4 \text{ moles Fe}\)

**O\textsubscript{2} and Fe\textsubscript{2}O\textsubscript{3}**
- \(3 \text{ moles O}_2\) and \(2 \text{ moles Fe}_2\text{O}_3\)
- \(2 \text{ moles Fe}_2\text{O}_3\) and \(3 \text{ moles O}_2\)
Consider the following equation:

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

A. A mole-mole factor for H\(_2\) and N\(_2\) is

1) 3 moles N\(_2\) 
2) 1 mole N\(_2\) 
3) 1 mole N\(_2\)

1 mole H\(_2\) 
3 moles H\(_2\) 
2 moles H\(_2\)

B. A mole-mole factor for NH\(_3\) and H\(_2\) is

1) 1 mole H\(_2\) 
2) 2 moles NH\(_3\) 
3) 3 moles N\(_2\)

2 moles NH\(_3\) 
3 moles H\(_2\) 
2 moles NH\(_3\)
Solution

$$3\text{H}_2(g) + \text{N}_2(g) \rightarrow 2\text{NH}_3(g)$$

A. A mole-mole factor for $\text{H}_2$ and $\text{N}_2$ is

2) $\frac{1 \text{ mole } \text{N}_2}{3 \text{ moles } \text{H}_2}$

B. A mole-mole factor for $\text{NH}_3$ and $\text{H}_2$ is

2) $\frac{2 \text{ moles } \text{NH}_3}{3 \text{ moles } \text{H}_2}$
How many moles of Fe₂O₃ can form from 6.0 mole O₂?

\[ 4\text{Fe(s)} + 3\text{O}_2(g) \rightarrow 2\text{Fe}_2\text{O}_3(s) \]

Relationship: \[ 3 \text{ mole O}_2 = 2 \text{ mole Fe}_2\text{O}_3 \]

Write a mole-mole factor to determine the moles of Fe₂O₃.

\[ 6.0 \text{ mole O}_2 \times \frac{2 \text{ mole Fe}_2\text{O}_3}{3 \text{ mole O}_2} = 4.0 \text{ mole Fe}_2\text{O}_3 \]
Guide to Using Mole Factors

STEP 1
Write the given and needed moles.

STEP 2
Write a plan to convert the given to the needed moles.

STEP 3
Use coefficients to write relationships and mole–mole factors.

STEP 4
Set up problem using the mole factor that cancels given moles.

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Learning Check

How many moles of Fe are needed for the reaction of 12.0 moles $O_2$?

$$4 \text{Fe}(s) + 3 \text{O}_2(g) \rightarrow 2 \text{Fe}_2\text{O}_3(s)$$

1) 3.00 moles Fe
2) 9.00 moles Fe
3) 16.0 moles Fe
Solution

3) 16.0 moles Fe

\[
\frac{12.0 \text{ moles } O_2}{3 \text{ moles } O_2} \times \frac{4 \text{ moles } Fe}{3 \text{ moles } O_2} = 16.0 \text{ moles } Fe
\]
Steps in Finding the Moles and Masses in a Chemical Reaction

Step 1  If necessary convert the mass of substance A to moles using its molar mass.
grams of A → moles of A

Step 2  Convert the moles of A to moles of the desired substance B using the mole factor derived from the coefficients of the balanced equation.
moles of A → moles of B

Step 3  If needed, convert the moles of B to grams using its molar mass.
moles of B → grams of B

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Moles to Grams

Suppose we want to determine the mass (g) of NH₃ that can form from 2.50 moles N₂.

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

The plan needed would be

moles N₂ → moles NH₃ → grams NH₃

The factors needed would be:
mole factor NH₃/N₂ and the molar mass NH₃
Moles to Grams

The setup for the solution would be:

\[
2.50 \text{ mole } N_2 \times \frac{2 \text{ moles } NH_3}{1 \text{ mole } N_2} \times \frac{17.0 \text{ g } NH_3}{1 \text{ mole } NH_3}
\]

given           mole-mole factor         molar mass

= 85.0 g NH_3
Learning Check

How many grams of $O_2$ are needed to produce 0.400 mole $Fe_2O_3$ in the following reaction?

$$4 \text{ Fe}(s) + 3 \text{ O}_2(g) \rightarrow 2 \text{ Fe}_2\text{O}_3(s)$$

1) 38.4 g $O_2$
2) 19.2 g $O_2$
3) 1.90 g $O_2$
Solution

2) 19.2 g $O_2$

$$0.400 \text{ mole } Fe_2O_3 \times \frac{3 \text{ mole } O_2}{2 \text{ mole } Fe_2O_3} \times \frac{32.0 \text{ g } O_2}{1 \text{ mole } O_2} = 19.2 \text{ g } O_2$$

mole factor  molar mass
The reaction between H₂ and O₂ produces 13.1 g water. How many grams of O₂ reacted?

\[ 2 \text{H}_2(g) + \text{O}_2(g) \rightarrow 2 \text{H}_2\text{O}(g) \]

\[ \text{? g} \quad 13.1 \text{ g} \]

The plan and factors would be:

\[
\text{g H}_2\text{O} \quad \rightarrow \quad \text{mole H}_2\text{O} \quad \rightarrow \quad \text{mole O}_2 \quad \rightarrow \quad \text{g O}_2
\]

- molar
- mole-mole
- molar
- mass H₂O
- factor
- mass O₂
Calculating the Mass of a Reactant

The setup would be:

\[
\begin{align*}
13.1 \text{ g H}_2\text{O} \times \frac{1 \text{ mole H}_2\text{O}}{18.0 \text{ g H}_2\text{O}} \times \frac{1 \text{ mole O}_2}{2 \text{ moles H}_2\text{O}} \times \frac{32.0 \text{ g O}_2}{1 \text{ mole O}_2}
\end{align*}
\]

\[
= 11.6 \text{ g O}_2
\]
Learning Check

Acetylene gas $\text{C}_2\text{H}_2$ burns in the oxyacetylene torch for welding. How many grams of $\text{C}_2\text{H}_2$ are burned if the reaction produces 75.0 g $\text{CO}_2$?

$$2 \text{ C}_2\text{H}_2(g) + 5 \text{ O}_2(g) \rightarrow 4 \text{ CO}_2(g) + 2 \text{ H}_2\text{O}(g)$$

1) 88.6 g $\text{C}_2\text{H}_2$
2) 44.3 g $\text{C}_2\text{H}_2$
3) 22.2 g $\text{C}_2\text{H}_2$
3) 22.2 g C₂H₂

\[ 2 \text{C}_2\text{H}_2(g) + 5 \text{O}_2(g) \rightarrow 4 \text{CO}_2(g) + 2 \text{H}_2\text{O}(g) \]

\[
\begin{align*}
75.0 \text{ g CO}_2 & \times \frac{1 \text{ mole CO}_2}{44.0 \text{ g CO}_2} \times \frac{2 \text{ moles C}_2\text{H}_2}{4 \text{ moles CO}_2} \times \frac{26.0 \text{ g C}_2\text{H}_2}{1 \text{ mole C}_2\text{H}_2} \\
& \text{molar mass CO}_2 & \text{mole-mole factor} & \text{molar mass C}_2\text{H}_2 \\
& = 22.2 \text{ g C}_2\text{H}_2
\end{align*}
\]
Calculating the Mass of Product

When 18.6 g ethane gas $C_2H_6$ burns in oxygen, how many grams of $CO_2$ are produced?

$$2 \text{C}_2\text{H}_6(g) + 7 \text{O}_2(g) \rightarrow 4 \text{CO}_2(g) + 6 \text{H}_2\text{O}(g)$$

18.6 g \quad ? \ g

The plan and factors would be

\[ \text{g C}_2\text{H}_6 \rightarrow \text{mole C}_2\text{H}_6 \rightarrow \text{mole CO}_2 \rightarrow \text{g CO}_2 \]

- molar
- mole-mole
- molar

mass $C_2H_6$ factor mass $CO_2$
Calculating the Mass of Product

\[2 \text{C}_2\text{H}_6(g) + 7 \text{O}_2(g) \rightarrow 4 \text{CO}_2(g) + 6 \text{H}_2\text{O}(g)\]

The setup would be

\[
\frac{18.6 \text{ g C}_2\text{H}_6}{30.1 \text{ g C}_2\text{H}_6} \times \frac{1 \text{ mole C}_2\text{H}_6}{2 \text{ moles C}_2\text{H}_6} \times \frac{4 \text{ moles CO}_2}{1 \text{ mole CO}_2} \times \frac{44.0 \text{ g CO}_2}{
\]

= 54.4 g CO\(_2\)
Learning Check

How many grams H₂O are produced when 35.8 g C₃H₈ react by the following equation?

C₃H₈(g) + 5 O₂(g) → 3 CO₂(g) + 4 H₂O(g)

1) 14.6 g H₂O
2) 58.4 g H₂O
3) 117 g H₂O
Solution

2) 58.4 g H₂O

\[ C_3H_8(g) + 5 O_2(g) \rightarrow 3 CO_2(g) + 4 H_2O(g) \]

\[
\frac{35.8 \text{ g } C_3H_8 \times 1 \text{ mole } C_3H_8}{44.1 \text{ g } C_3H_8} \times \frac{4 \text{ mole } H_2O}{1 \text{ mole } C_3H_8} \times \frac{18.0 \text{ g } H_2O}{1 \text{ mole } H_2O} = 58.4 \text{ g } H_2O
\]