Quantity relationships: *How much?*  *How many?*

How many dragonflies’ wings are required for 4 dragonflies to be able to fly?

1 Dragonfly  →  4 wings
\[
\frac{1 \ DF}{5 \ DFs} = \frac{4 \ wings}{?}
\]
Quantity Relationships in Chemical Reactions:
How to solve stoichiometry problems:

What we need to know are two subjects as:
1) A balanced chemical reaction.
2) Relationship between mass and mole.

EXAMPLE:
How many moles of oxygen are required to burn 2.40 moles of ethane gas?

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

Given

\[
\begin{align*}
&2 \text{ mol C}_2\text{H}_6 \\
\end{align*}
\]

Wanted

\[
\begin{align*}
&7 \text{ mol O}_2 \\
&2.40 \text{ mol C}_2\text{H}_6
\end{align*}
\]
\[
\frac{2 \text{ mol}}{2.4 \text{ mol}} = \frac{7 \text{ mol}}{?} 
\]

\[2 \text{ mol (C}_2\text{H}_6) \times ? = 7 \text{ mol (O}_2\text{)} \times 2.4 \text{ mol (C}_2\text{H}_6)\]

\[? = \frac{7 \text{ mol}}{2 \text{ mol}} \times 2.4 \text{ mol}\]

EXAMPLE: Calculate the number of grams of oxygen that are required to burn 155 g of ethane.

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

Given \hspace{1cm} Wanted

\[\frac{2 \text{ mol}}{155 \text{ g}} = \frac{7 \text{ mol}}{\text{? g}}\]

\[\frac{155 \text{ g}}{30.06 \left(\frac{\text{g}}{\text{mol}}\right)}\]

\[\frac{155 \text{ g}}{30.06 \left(\frac{\text{g}}{\text{mol}}\right)}\]

Same Units \hspace{1cm} Same Units
\[
\frac{2 \text{ mol}}{5.15 \text{ mol}} = \frac{7 \text{ mol}}{?} \]

\(? = 5.15 \times \frac{7}{2}\)

\(? = 18.03 \text{ mol (oxygen)}\)

Wanted in gram!

\[
mole = \frac{mass}{molar \ mass}
\]

molar mass of oxygen = \(2 \times 16.00 \text{ (g/mol)}\)

\(? = 18.03 \text{ mol} \times 32.00 \text{ (g/mol)}\)

\(? = 577 \text{ g } \text{O}_2\)
EXAMPLE:
How many milligrams of Nickel Chloride are in a solution if 503 mg of Silver Chloride is precipitated in the reaction of Silver Nitrate and Nickel Chloride solution?

\[
2 \text{ AgNO}_3 (aq) + \text{NiCl}_2 (aq) \rightarrow 2 \text{ AgCl} (s) + \text{Ni(NO}_3)_2 (aq)
\]

Wanted: 2 mol AgCl
Given: 1 mol NiCl₂

➢ (mg) to (g):
\[
503 \text{ mg} \times \left(\frac{1 \text{ g}}{1000 \text{ mg}}\right) = 0.503 \text{ g}
\]
(mass) to (mole)

\[
mole = \frac{mass}{molar\ mass}
\]

mole of AgCl = \(0.503\ g \div 143.35\ (g/mol)\)
mole of AgCl = \(3.51 \times 10^{-3}\ mol\)

\[
\frac{1\ mol}{?} = \frac{2\ mol}{72.1\ mol}
\]

\(= 3.51 \times 10^{-3}\ mol\ AgCl \times \frac{1\ mol\ NiCl_2}{2\ mol\ AgCl}\)

\(? = 1.75 \times 10^{-3}\ mol\ NiCl_2\)

Wanted in milligram!

\[
mole = \frac{mass}{molar\ mass}
\]

molar mass of NiCl₂ = 129.59 (g/mol)

Mass of NiCl₂ = \(1.75 \times 10^{-3}\ mol \times 129.59\ (g/mol) = 0.227\ g\)
Mass of NiCl₂ = 227 mg
Gas Stoichiometry at Standard Temperature and Pressure:

One mole of any gas at STP occupies 22.4 L

EXAMPLE:
What volume of hydrogen, at STP, can be released by 42.7 g of zinc as it reacts with hydrochloric acid?

\[
\text{Zn(s) + 2 HCl (aq) } \rightarrow \text{ H}_2 (g) + \text{ ZnCl}_2 (aq)
\]

Given

- 1 mol Zn

Wanted

- 22.4 L

\[
\frac{42.7 \text{ g}}{65.39 \left( \frac{\text{g}}{\text{mol}} \right)} = \frac{1 \text{ mol}}{0.653 \text{ mol}} \rightarrow \frac{22.4 \text{ L}}{?}
\]
EXAMPLE:

1.30L of gaseous ethylene is burned. What volume of oxygen is required if both gas volume are measured at STP?

\[ \text{C}_2\text{H}_4 \ (g) + 3\text{O}_2 \ (g) \rightarrow 2\text{CO}_2 \ (g) + 2\ \text{H}_2\text{O} \ (l) \]

<table>
<thead>
<tr>
<th>Given</th>
<th>Wanted</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 mol C\textsubscript{2}H\textsubscript{4}</td>
<td>3 mol O\textsubscript{2}</td>
</tr>
</tbody>
</table>

\[ \frac{22.4 \ L}{1.30 \ L} = \frac{67.2}{?} \]
Gas Stoichiometry at nonstandard condition:

In STP condition: \( T = 273 \) K, and \( P = 760 \) torr

\[
\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}
\]

STP = nonstandard

In the previous example for the reaction:

\[
\text{Zn (s) + 2 HCl (aq) \rightarrow H}_2 \text{ (g) + ZnCl}_2 \text{ (aq)}
\]

We calculated the volume of hydrogen as 14.6 L, now let’s change the problem. Instead of measuring the hydrogen volume at STP, let’s use \( t = 21 \) °C and \( P = 748 \) torr.

\[
T_k = 21 + 273 = 294
\]

\[
\frac{760 \times 14.6}{273_1} = \frac{748 V_2}{294}
\]

STP = nonstandard

\( V_2 = 16.0 \) L
How many of the above electronic circuits (ec) can you assemble from 22 resistors, 18 capacitors, and 14 transistors?

2 resistors + 4 transistors + 3 capacitors →

Calculate the theoretical yield of ec:

\[ 14 \times T \times \frac{1}{2} \times \frac{ec}{T} = 7 \text{ ec} \]

\[ 22 \times R \times \frac{1}{4} \times \frac{ec}{T} = 5.5 \text{ ec} \approx 5 \text{ ec} \]

\[ 18 \times C \times \frac{1}{3} \times \frac{ec}{C} = 6 \text{ ec} \]

Resistor is the limiting item, because 22 of resistors give 5 ec.
Example:

The fertilizer ammonium nitrate can be made by direct combination of ammonia with nitric acid. If 74.4 g of ammonia is reacted with 159 g of nitric acid, how many grams of ammonium nitrate can be produced? Also calculate the mass of unreacted reactant that is in excess.

\[
\text{NH}_3 + \text{HNO}_3 \rightarrow \text{NH}_4\text{NO}_3
\]

Convert the number of grams of each reactant to moles.

\[
\text{NH}_3: \quad \frac{74.4 \text{ g}}{17.03 \text{ g/mol}} = 4.37 \text{ mol} \quad \text{HNO}_3: \quad \frac{159 \text{ g}}{63.02 \text{ g/mol}} = 2.52 \text{ mol}
\]

\[
4.37 \text{ mol } \text{NH}_3 \times \frac{1 \text{ mol } \text{NH}_4\text{NO}_3}{1 \text{ mol } \text{NH}_3} = 4.37 \text{ mol } \text{NH}_4\text{NO}_3
\]

\[
2.52 \text{ mol } \text{HNO}_3 \times \frac{1 \text{ mol } \text{NH}_4\text{NO}_3}{1 \text{ mol } \text{HNO}_3} = 2.52 \text{ mol } \text{NH}_4\text{NO}_3
\]

\[
\therefore \text{ The reactant that yields the smaller amount of product is limiting reactant.}
\]

1 mol of NH₃ ≡ 1 mol of HNO₃

2.52 mol ≡ ?

? = 2.52 mol of HNO₃

4.37 mol – 2.52 mol = 1.85 mol NH₃ left
A mixture of 5.0 g of H_2 (g) and 10.0 g of O_2(g) is ignited. Water forms according to the following combination reaction:

\[ 2H_2(g) + O_2(g) \rightarrow 2H_2O(g) \]

Which reactant is limiting? How much water will the reaction produce?
Percent Yield:
The calculated amount of product if it is based on the assumption that all of the reactant is converted into product is called the **theoretical yield**. In laboratory or in industrial production, the actual amount of product isolated from a reaction is usually less than the theoretical yield, and it is called **actual yield**.

**General solution:**

\[ \alpha A + \beta B \rightarrow \gamma C \]

**Given:** n gram of A  
**Given:** actual yield  
**Wanted:** %yield

Step 1: Convert mass of A into the mole of A.

Step 2: Calculate the **theoretical yield**:

\[ \text{Number of moles } A \times \frac{\gamma \text{ mol } C}{\alpha \text{ mol } A} \]

Step 3: Convert mole of C into the mass of C.
Step 4:

\[
\%Yield = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100
\]

Example:

A general chemistry student, preparing copper metal by the reaction of 1.274 g of copper sulfate with zinc metal, obtained a yield of 0.392 g copper. What was the percent yield?

To calculate %yield, we need both the theoretical yield of copper and the actual yield.

\[
\text{CuSO}_4 (aq) + \text{Zn(s)} \rightarrow \text{Cu(s)} + \text{ZnSO}_4 (aq)
\]

**Mass to Mole:**

\[
\frac{1.274 \text{ g CuSO}_4}{159.6 \text{ g/mol}} = 7.982 \times 10^{-3} \text{ mol CuSO}_4
\]

The theoretical yield:

\[
7.982 \times 10^{-3} \text{ mol CuSO}_4 \times \frac{1 \text{ mol Cu}}{1 \text{ mol CuSO}_4} = 7.982 \times 10^{-3} \text{ mol Cu}
\]

**Mole to Mass:**

\[
7.982 \times 10^{-3} \text{ mol Cu} \times 63.55 \text{ g/mol} = 0.5072 \text{ g Cu}
\]

\[
\%Yield = \frac{0.392 \text{ g}}{0.5072 \text{ g}} \times 100 = 77.3 \%
\]
Liquid Oxygen

Liquid Hydrogen

Main Engines
Thermochemical Equations:

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

Gives off heat  
Exothermic  
\(-\Delta H\)

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

Absorbs heat  
Endothermic  
\(+\Delta H\)

The SI unit for energy: joule (J)

How many kilojoule of energy are released when 73.0 g C2H6 burns according to:

\[ 2\text{C}_2\text{H}_6 (g) + 7 \text{O}_2 (g) \rightarrow 4 \text{CO}_2(g) + 6\text{H}_2\text{O} (l) + 3119\text{KJ} \]

\[
\frac{73.0 \text{ g}}{30.05 \left( \frac{\text{g}}{\text{mol}} \right)} = 2 \text{ mol C}_2\text{H}_6
\]

\[
\frac{3119 \text{ KJ}}{?}
\]
ΔH = 2.82 × 10³ KJ for the photosynthesis reaction by which plants use energy from the sun to form one mole sugar from carbon dioxide and water. How much energy is required to form 454g (1lb) of simple sugar C₆H₁₂O₆?

\[ 6 \text{ CO}_2 (\text{g}) + 6 \text{ H}_2\text{O (g)} + \text{sun} \rightarrow \text{C}_6\text{H}_{12}\text{O}_6 (s) + 6\text{O}_2(\text{g}) \]

\[
\begin{align*}
454 \text{ g} & \rightarrow 454 \text{ g} \\
\frac{1 \text{ mol}}{180 \left( \frac{\text{g}}{\text{mol}} \right)} & \cdot 2.82 \times 10^3 \text{ KJ} \\
\therefore \ 7.11 \times 10^3 \text{ KJ}
\end{align*}
\]