

2.6 LIMITING REACTANT AND EXCESS REACTANT

<https://subarnam.com.np/>

Learning outcomes

After the completion of the chapter, the students should be able to:

- Identify the **limiting reactant** and **excess reactant** in a given chemical reaction.
- Calculate the maximum amount of products formed based on stoichiometric principles.
- Solve numerical problems related to leftover reactants, gas volumes at NTP, and sequential neutralisation reactions.

Introduction

In a perfect laboratory setup, reactants would always be mixed in exact theoretical (stoichiometric) ratios. In real-world experiments, however, reactants are rarely mixed in perfect stoichiometric amounts.

Usually, one reactant is present in a smaller proportion than required, while the other is present in a larger amount.

- **Limiting Reactant (or Limiting Reagent):** The reactant that is completely consumed first in a chemical reaction. Because it finishes first, it limits the amount of product that can be formed and halts the reaction.
- **Excess Reactant:** The reactant that is left over after the reaction has completely stopped.

Why is the Limiting Reactant Essential in Stoichiometry?

The limiting reactant is the single most important factor in stoichiometric calculations because **it determines the theoretical yield of the product**. Once it is exhausted, the remaining excess reactant cannot magically transform into products on its own. Therefore, all product calculations must be based strictly on the limiting reactant.

How to Identify the Limiting Reactant?

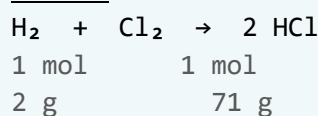
There are two common approaches used in NEB-level numericals:

- *Comparison method: Calculate how much of the second reactant is actually required by the given mass of the first reactant, then compare it with the amount actually available.*
- *Mole-ratio method: Convert both reactants to moles and divide each by its coefficient in the balanced equation — the smaller value identifies the limiting reactant.*

Worked Example 1

1 g of H_2 is mixed with 71 g of Cl_2 . Identify the limiting reactant.

Solution:



From the equation, 2 g of H_2 reacts exactly with 71 g of Cl_2 .

$\therefore$ 1 g of H_2 reacts with $71/2 = 35.5$ g of Cl_2 .

Cl_2 left unreacted = $71 - 35.5 = 35.5$ g.

Since Cl_2 remains left over, H_2 is the limiting reactant.

1. Given, $\text{CaCO}_3(\text{s}) + 2\text{HCl}(\text{aq.}) \rightarrow \text{CaCl}_2(\text{aq.}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$

If 10 grams of pure CaCO_3 are added to a solution containing 7.665 grams of HCl ,

i) Find the limiting reactant. (CaCO_3)

ii) Calculate the moles of excess reactant left over. (0.01)

iii) Calculate the volume of CO_2 gas produced at NTP. (2.24 litres)

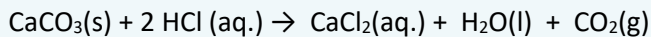
iv) Calculate the number of grams of NaOH required for absorbing the whole of the CO₂ gas as

$$\text{Na}_2\text{CO}_3. \text{ (8)}$$

5

Solution:

First, write down the balanced chemical equation and calculate molecular weights



1 mol 2 mol 1 mol 1 mol 1 mol

100g 2 x 36.5 g 22.4 l of CO₂ at NTP

$$[\text{CaCO}_3 = 40 + 12 + 16 \times 3 = 100 \text{ amu} \qquad \text{HCl} = 1 + 35.5 = 36.5 \text{ amu}]$$

i) For limiting reactant,

100 g of CaCO_3 reacts completely with 73 g of HCl.

10 g $73/100 \times 10 = 7.3 \text{ g HCl.}$

Comparison: We are given 7.665 g of HCl. Since the required amount of 7.3 g is **less** than the given amount 7.665 g, HCl is in excess.

Result: CaCO_3 is the limiting reactant

ii) For Excess reactant,

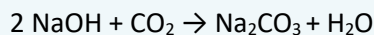
$$7.665 - 7.3g = 0.365g$$
$$36.5 \text{ g HCl} = 1 \text{ mol HCl}$$
$$0.365 \text{ g HCl} = 1/36.5 \times 0.365 = 0.01 \text{ moles.}$$

iii) **For Volume of CO_2**

100 g CaCO_3 gives 22.4 l CO_2 at NTP.

10 g 22.4/100 x 10 = 2.24 l of CO₂ at NTP.

iii) For mass of NaOH



2 mol 1 mol

2 x40g 22.4 l CO₂ at NTP

$$\text{NaOH} = 23 + 16 + 1 = 40 \text{ amu}$$

22.4 l CO₂ at NTP requires 80 g NaOH.

2.24 | "80/22.4 x 2.24= 8 g NaOH

Worked Example 3

2 g of magnesium is burnt in a closed vessel containing 1.2 g of oxygen, producing magnesium oxide.

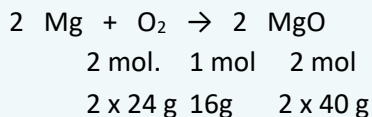
i) Find the limiting reagent.

ii) Calculate the number of molecules of unreacted reagent left over. (5.019×10^{21})

iii) What mass of MgO is produced? (3g)

iv) How many grams of pure HCl are required to neutralise the whole MgO produced? (5.475)

Solution:



i. 48 g of Mg reacts with 32 g of oxygen

$$\begin{aligned} 2 \text{ g of Mg reacts with } & \frac{32 \times 2}{48} \text{ g of oxygen} \\ & = 1.34 \text{ g of oxygen} \end{aligned}$$

Since the amount of oxygen required by calculation is higher than the given, **the limiting reagent is oxygen.**

ii. 32 g of oxygen reacts with 48 g of Mg

$$1.2 \text{ g of oxygen reacts with } \frac{48 \times 1.2}{32} \text{ g Mg} = 1.8 \text{ g of Mg}$$

$$\therefore \text{The leftover magnesium} = (32 - 1.8) \text{ g} = 0.2 \text{ g}$$

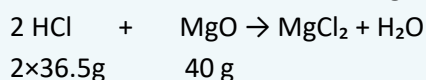
Now,

$$\begin{aligned} \text{No of moles of leftover magnesium} &= \frac{0.2}{24} \text{ moles} \\ &= \frac{0.2}{24} \times 6.023 \times 10^{23} \\ &= \underline{5.01 \times 10^{21} \text{ Mg atoms}} \end{aligned}$$

iii. 32 g of oxygen produces 80 g of MgO

$$\begin{aligned} 1.2 \text{ g of oxygen produces } & \frac{80 \times 1.2}{32} \text{ g of MgO} \\ & = \underline{3 \text{ g of MgO}} \end{aligned}$$

iv. The reaction between HCl and MgO is as follows:



40 g of MgO is completely neutralised by 2×36.5 g of HCl

$$\begin{aligned} 3 \text{ g of MgO is completely neutralised by } & \frac{2 \times 36.5}{40} \text{ g of HCl} \\ & = \underline{5.475 \text{ g of HCl}} \end{aligned}$$

Worked Example 4

2g of Magnesium is burnt in a closed vessel containing 3g of oxygen.

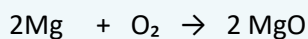
i) Which one is the limiting reagent?

ii) Calculate the moles of reactant left over. (0.052)

iii) How many grams of MgO are produced? (3.33)

iv) What is the mass of H₂SO₄ required to neutralise MgO formed in the reaction? (8.16g) 5

Solution:



i. From the balanced chemical equation,

48 g of Mg reacts with 32 g of oxygen

2 g of Mg reacts with $32 \times 2 / 48$ g of oxygen

$$= 1.34 \text{ g of oxygen}$$

Since the amount of oxygen required by calculation is smaller than the amount given, Mg is the limiting reagent.

ii. Mass of the reactant left over = $(3 - 1.34) \text{ g} = 1.66 \text{ g}$

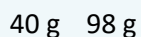
$$\text{No. of moles} = 1.66 / 32 = \underline{0.051 \text{ moles}}$$

iii. Since magnesium is the limiting reagent,

48 g of Mg produces 80 g of MgO

2 g of Mg produces $80 \times 2 / 48$ g of MgO = 3.33 g of MgO

iv. Now, the MgO produced is reacted with H₂SO₄



40 g of MgO reacts with 98 g of H₂SO₄

3.33 g of MgO reacts with $98 \times 3.33 / 40$ g of H₂SO₄

$$= \underline{8.15 \text{ g of H}_2\text{SO}_4}$$

☑ **Exam Tips (NEB Style)**

- Always start by writing the balanced chemical equation with mole ratios and gram masses written below each substance — this earns method marks even if the final answer is wrong.
- State clearly which reactant is limiting and which is in excess before proceeding — examiners specifically look for this line.
- Keep units consistent (g, mol, L) throughout, and box or underline your final numerical answer with correct units.

Key Points to Remember
The limiting reactant is completely consumed first and decides the theoretical yield of product.
The excess reactant is left over, partially unreacted, once the reaction stops.
Product calculations are always based on the limiting reactant — never the excess one.
Two methods to find it: the comparison method (mass required vs mass available) and the mole-ratio method.

*** ... ***

<https://subarnam.com.np/>