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Limiting Reactants



CALCULATIONS FROM CHEMICAL EQUATIONS

From the previous lesson, do you remember what does a chemical reaction tells us?

Below is an example of an equation:

Methane + Oxygen \rightarrow Carbon dioxide + Water



Reactants

Products

1 molecule

2 molecules

1 molecule

2 molecules

1 mole

2 mole

1 mole


2 mole

Ratio of the moles of reactants and products




Reacting Masses





How to calculate reacting masses?

- ☒ Mass of one substance must be given.
 - ☒ Balanced equation.
 - ☒ Ratio of moles from the equation.
 - ☒ No. of moles.
- 

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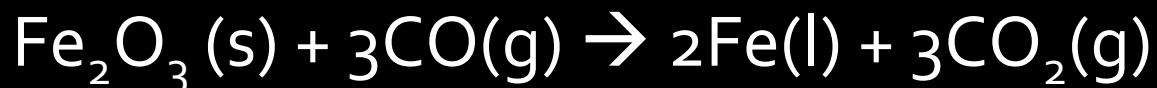
Step 4: Find mass

Mass of H_2O :

$$\text{No. of moles} = \frac{\text{Mass (g)}}{\text{Molar mass (g) or Mr}}$$
$$\text{Mass (g)} = \text{No. of moles} \times \text{Molar mass or Mr}$$
$$= 0.5 \text{ mol} \times 18 \text{ (g)}$$
$$= 9 \text{ g.}$$

- Example 2:

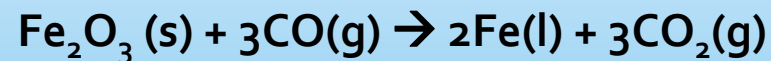
- To obtain iron, iron(III) oxide reacts with carbon monoxide according to the equation:



What mass of iron is produced from 40g of iron (II) oxide?

Solution:

Step 1: Balanced equation



Step 2: Mole ratio

1 : 2

Step 2.5*: Calculate no. of moles

Mass of $\text{Fe}_2\text{O}_3 = 40\text{g}$

Molar mass = $(2 \times 56) + (3 \times 16) = 160\text{g}$

No. of moles of Fe_2O_3 :

$$\frac{\text{Mass (g)}}{\text{Molar mass}}$$

$$\frac{40}{160}$$

$$= 0.25 \text{ mol}$$

Step 3: Find no. of moles

1 : 2

0.25mol ~~:~~ x

$$x = (0.25 \times 2) / 1$$

$$= 0.50 \text{ mol}$$


❖ No. of moles of Fe = 0.50 mol.

Step 4: Find mass


Mass of Fe:

$$\text{No. of moles} = \frac{\text{Mass (g)}}{\text{Molar mass or Mr}}$$

$$\begin{aligned} \text{Mass (g)} &= \text{No. of moles} \times \text{Molar mass or Mr} \\ &= 0.50 \text{ mol} \times 56\text{g} \\ &= 28 \text{ g.} \end{aligned}$$



How about calculating volumes of gases produced or used in reactions.

- Same procedure!
 - Must be there:
 - ☑ Mass /volume of one substance must be given.
 - ☑ Balanced equation.
 - ☑ Ratio of moles from the equation.
 - ☑ No. of moles.
- 

■ Example 1:

Magnesium reacts with hydrochloric acid according to the equation:



Calculate the **volume** of hydrogen gas, measured at room conditions, produced from the reaction of **14.6g** of hydrochloric acid.

■ Solution:

Step 1: Balanced equation



Step 2: Mole ratio

2 : 1

Step 2.5*: Calculate no. of moles

Mass of HCl = 14.6g

Molar mass = 1 + 35.5 = 36.5g

No. of moles of HCl :

$$\frac{\text{Mass (g)}}{\text{Molar mass}}$$
$$\frac{14.6}{36.5}$$
$$= 0.4 \text{ mol}$$

Step 3: Find no. of moles

$$\begin{array}{ccc} 2 & & 1 \\ 0.4 \text{ mol} & \begin{array}{c} \diagdown \quad : \quad \diagup \\ : \quad \quad \end{array} & x \end{array}$$
$$x = (0.4 \times 1) / 2$$
$$= 0.2 \text{ mol}$$

❖ No. of moles of $\text{H}_2 = 0.2 \text{ mol}$.

Step 4: Find volume

Volume of H_2 :

$$\text{No. of moles} = \frac{\text{Volume (dm}^3\text{)}}{\text{Molar volume (dm}^3\text{)}}$$

$$\text{Volume (dm}^3\text{)} = \text{No. of moles} \times \text{Molar volume}$$
$$= 0.2 \text{ mol} \times 24 \text{ dm}^3$$
$$= 4.8 \text{ dm}^3.$$

Example 2:

Lead (IV) oxide, PbO_2 , decomposes when heated strongly. The equation for the reaction is:



Calculate the **mass** of lead (IV) oxide that decomposes to produce **1.2dm³** of oxygen, measured at room conditions.

Solution:

Step 1: Balanced equation



Step 2: Mole ratio

2 : 1

Step 2.5*: Calculate no. of moles

Volume of $\text{O}_2 = 1.2 \text{ dm}^3$

Molar volume = 24 dm^3

No. of moles of O_2 :

$$\frac{\text{Volume (dm}^3\text{)}}{\text{Molar volume (dm}^3\text{)}} \\ \frac{1.2}{24} \\ = \underline{0.05 \text{ mol}}$$

Step 3: Find no. of moles

$\begin{array}{ccc} 2 & : & 1 \\ \times & & 0.05 \text{ mol} \end{array}$

$$x = (0.05 \times 2) / 1$$

$$= 0.10 \text{ mol}$$

❖ No. of moles of $\text{PbO}_2 = 0.10 \text{ mol}$.

Step 4: Find mass

Mass of PbO_2 :

Molar mass of $\text{PbO}_2 = 207 + (2 \times 16) = 239$

No. of moles = $\frac{\text{Mass (g)}}{\text{Molar Mass (g) or Mr}}$

Mass (g) = No. of moles \times Molar Mass or Mr

$$= 0.10 \text{ mol} \times 239$$

$$= 23.9 \text{ g}$$

Limiting reactants/reagents

- It is the chemical that is completely consumed in a reaction.
- The reaction stops as soon as the limiting reactant is totally consumed.
- The limiting reactant limits the amount of products formed.
- Check for limiting reactant when the quantities of **two** (or **more**) reactants are given.
- Calculation: Always involves no. of moles.

Example 1:

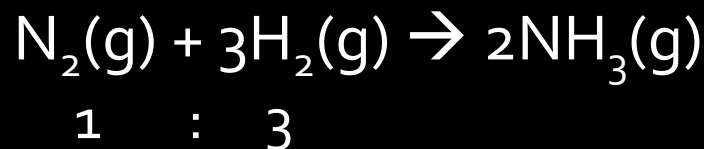
In the Haber process, ammonia is manufactured from nitrogen and hydrogen. A mixture of 5 moles of nitrogen and 10 moles of hydrogen were allowed to react according to the following equation:



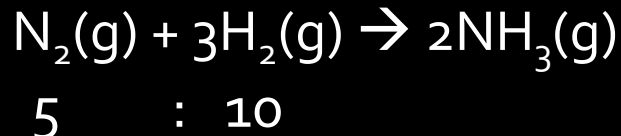
Which substance is the limiting reagent?

■ **Solution:**

Mole ratio from equation:



From the quantities given:



No. of moles of NH_3 (Products):

Using mole ratio $\text{N}_2 = (5 \times 2) / 1 = 10 \text{ mol.}$

Using mole ratio $\text{H}_2 = (10 \times 2) / 3 = 6.67 \text{ mol.}$

- ❖ The lowest no. of moles produced is the limiting reactants.
- ❖ Excess of N_2 , limiting reagent/reactant is H_2 .