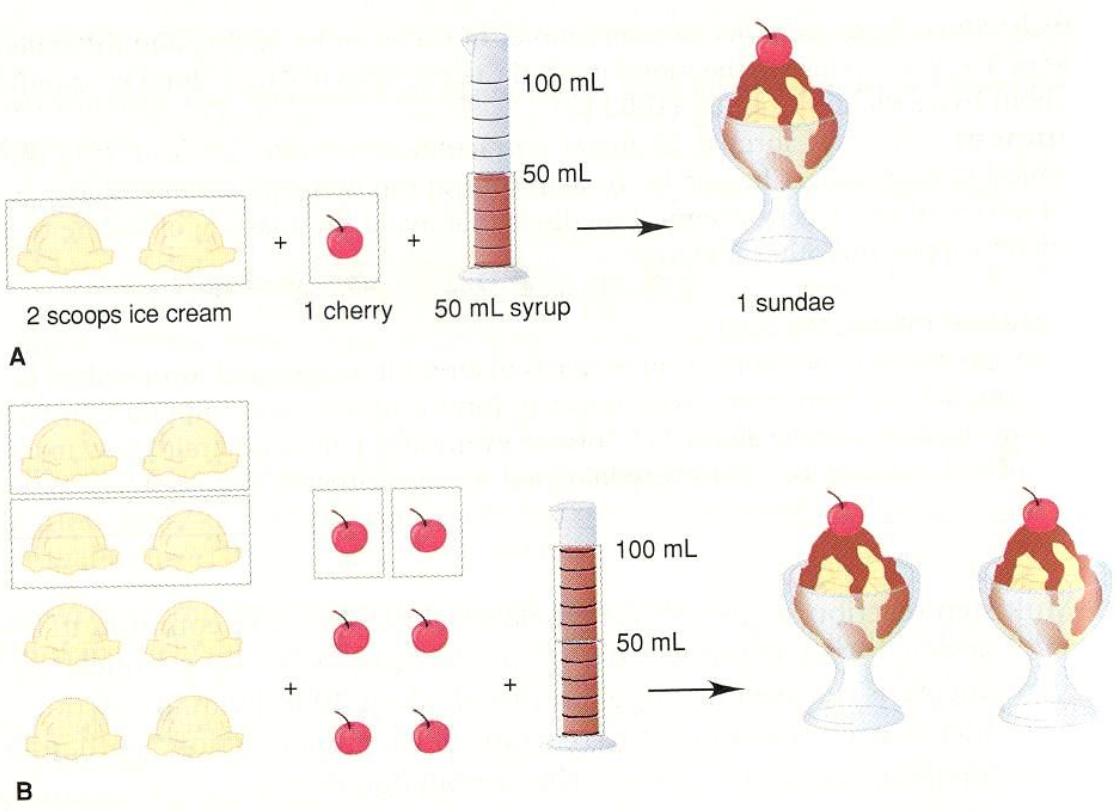
LIMITING REACTANT

- When ***2 or more reactants*** are combined in ***non-stoichiometric*** ratios, the amount of ***product*** produced is ***limited*** by the reactant that is ***not in excess*** (***limiting reactant).***

Analogy:



The number of sundaes possible is ***limited*** by the amount of syrup, the ***limiting reactant***.

Limiting Reactant (Reagent) Problems always involve 2 steps:

1. **Identify the Limiting Reactant (LR)**
   - convert all masses to moles
   -compare actual mole ratio to mole ratio given by the the balanced chemical equation

   OR

   - calculate the number of moles obtained from each reactant in turn.
   - The reactant that gives the smaller amount of product is the Limiting R.eactant.

2. **Calculate the amount of product obtained from the Limiting Reactant**

Example 1

Sodium hydrogen carbonate is prepared from NaCl and ammonium hydrogen carbonate, according to the equation:

$$NH_4HCO_3(aq) + NaCl(aq) \longrightarrow NaHCO_3(aq) + NH_4Cl(aq)$$

If 0.300 moles of NH_4HCO_3 are reacted with 0.2567 moles of NaCl, how many grams of $NaHCO_3$ are obtained ?

$$\mathbf{1}\ NH_4HCO_3(aq) + \mathbf{1}\ NaCl(aq) \longrightarrow \mathbf{1}\ NaHCO_3(aq) + \mathbf{1}\ NH_4Cl(aq)$$

0.300 moles 0.2567 moles ? g

L.R.

$$? \text{ g } NaHCO_3 = \mathbf{0.2567 \text{ moles NaCl}} \times \frac{\mathbf{1} \text{ mole } NaHCO_3}{\mathbf{1} \text{ mole NaCl}} \times \frac{84.01 \text{ g } NaHCO_3}{1 \text{ mole } NaHCO_3}$$

$$= \mathbf{21.57 \text{ g } NaHCO_3}$$

Example 2

A 1.4 g sample of magnesium is treated with 8.1 g of hydrochloric acid to produce magnesium chloride and hydrogen gas. How many grams of hydrogen are produced ?

$$\underset{1.4\ g}{Mg(s)} + \underset{8.1\ g}{2\ HCl(aq)} \longrightarrow MgCl_2(aq) + \underset{?\ g}{H_2(g)}$$

Change masses of reactants in moles:

$1\ Mg(s) + 2\ HCl(aq) \longrightarrow 1\ MgCl_2(aq) + 1\ H_2(g)$

$$1.4\ g \times \frac{1\ mole}{24.31\ g} = 0.0576\ moles \qquad 8.1\ g \times \frac{1\ mole}{36.46\ g} = 0.222\ moles$$

L.R.

requires → 2 x 0.0576 moles HCl = 0.115 moles HCl

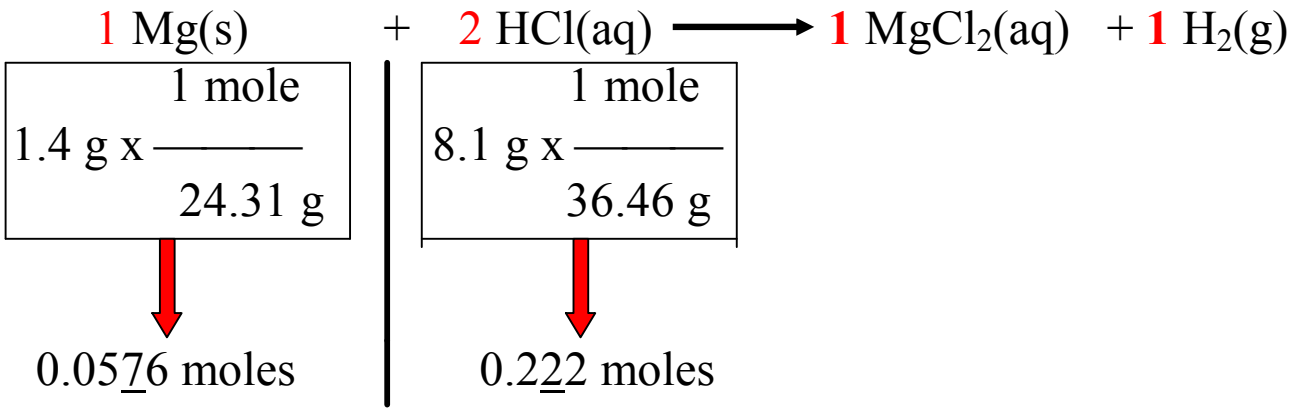
(0.222 moles HCl) available

HCl is an excess!

$$?\ g\ H_2 = 0.0576\ moles\ Mg \times \frac{1\ mole\ H_2}{1\ mole\ Mg} \times \frac{2.02\ g\ H_2}{1\ mole\ H_2} = \mathbf{0.12\ g\ H_2}$$

L.R.

Solution recommended by textbook:



Calculate the number of moles obtained from each reactant in turn. The reactant that gives the smaller amount of product is the Limiting Reactant.

$$\text{? g } H_2 = \textbf{0.222 moles HCl} \times \frac{1 \text{ mole } H_2}{2 \text{ moles HCl}} \times \frac{2.02 \text{ g } H_2}{1 \text{ mole } H_2} = \textbf{0.22 g } \mathbf{H_2}$$

$$\text{? g } H_2 = \textbf{0.0576 moles Mg} \times \frac{1 \text{ mole } H_2}{1 \text{ mole Mg}} \times \frac{2.02 \text{ g } H_2}{1 \text{ mole } H_2} = \textbf{0.12 g } \mathbf{H_2}$$

smaller ! (correct answer)

Since Mg produces the smaller amount of product, Mg is the L.R.

THE YIELD CONCEPT

- Quantities of product calculated represent the maximum amount obtainable (100 % yield)
- Most chemical reactions do not give 100 % yield of product because of:

  - side reactions (unwanted reactions)
  - reversible reactions (reactants ⟵ products)
  - losses in handling and transferring

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100$$

Actual Yield: Amount of product actually obtained (experimental)

Theoretical Yield: Maximum amount of product obtainable (calculated from equation)

Example 1

A 35.0 g sample of calcium hydroxide is reacted with excess phosphoric acid, according to the following balanced chemical equation:

$$\mathbf{3}\ Ca(OH)_2(aq) + \mathbf{2}\ H_3PO_4(aq) \longrightarrow \mathbf{1}\ Ca_3(PO_4)_2(s) + \mathbf{6}\ H_2O(l)$$

(a) How many grams of calcium phosphate can be produced ?

$$?\ g\ Ca_3(PO_4)_2 = 35.0\ g\ Ca(OH)_2 \times \frac{1\ mole\ Ca(OH)_2}{74.10\ g\ Ca(OH)_2} \times \frac{\mathbf{1}\ mole\ Ca_3(PO_4)_2}{\mathbf{3}\ mole\ Ca(OH)_2} \times \frac{310.3\ g\ Ca_3(PO_4)_2}{1 mole\ Ca_3(PO_4)_2} = \mathbf{48.9\ g\ Ca_3(PO_4)_2}$$

(b) If **45.2 grams** of calcium phosphate are actually obtained in a **laboratory experiment**, what is the percent yield ?

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100 = \frac{45.2\ g}{48.9\ g} \times 100 = \mathbf{92.4\ \%}$$

Example 2

Sodium hydrogen carbonate is prepared from NaCl and ammonium hydrogen carbonate, according to the equation:

$$NH_4HCO_3(aq) + NaCl(aq) \longrightarrow NaHCO_3(aq) + NH_4Cl(aq)$$

If 0.300 moles of NH_4HCO_3 are reacted with 0.2567 moles of NaCl, and 10.45 g of $NaHCO_3$ are obtained, what is the percent yield?

1. First calculate the maximum amount obtainable (theoretical yield) from the given quantities (theoretical yield)

$$\mathbf{1}\ NH_4HCO_3(aq) + \mathbf{1}\ NaCl(aq) \longrightarrow \mathbf{1}\ NaHCO_3(aq) + \mathbf{1}\ NH_4Cl(aq)$$

0.300 moles 0.2567 moles ? g

L.R.

$$?\ g\ NaHCO_3 = \mathbf{0.2567\ moles\ NaCl} \times \frac{\mathbf{1}\ \text{mole}\ NaHCO_3}{\mathbf{1}\ \text{mole NaCl}} \times \frac{84.01\ g\ NaHCO_3}{1\ \text{mole}\ NaHCO_3} = \mathbf{21.57\ g\ NaHCO_3\ (theoretical\ yield)}$$

2. Second, calculate % yield from actual and theoretical yield

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100 = \frac{10.45\ g}{21.57\ g} \times 100 = \mathbf{48.45\ \%}$$