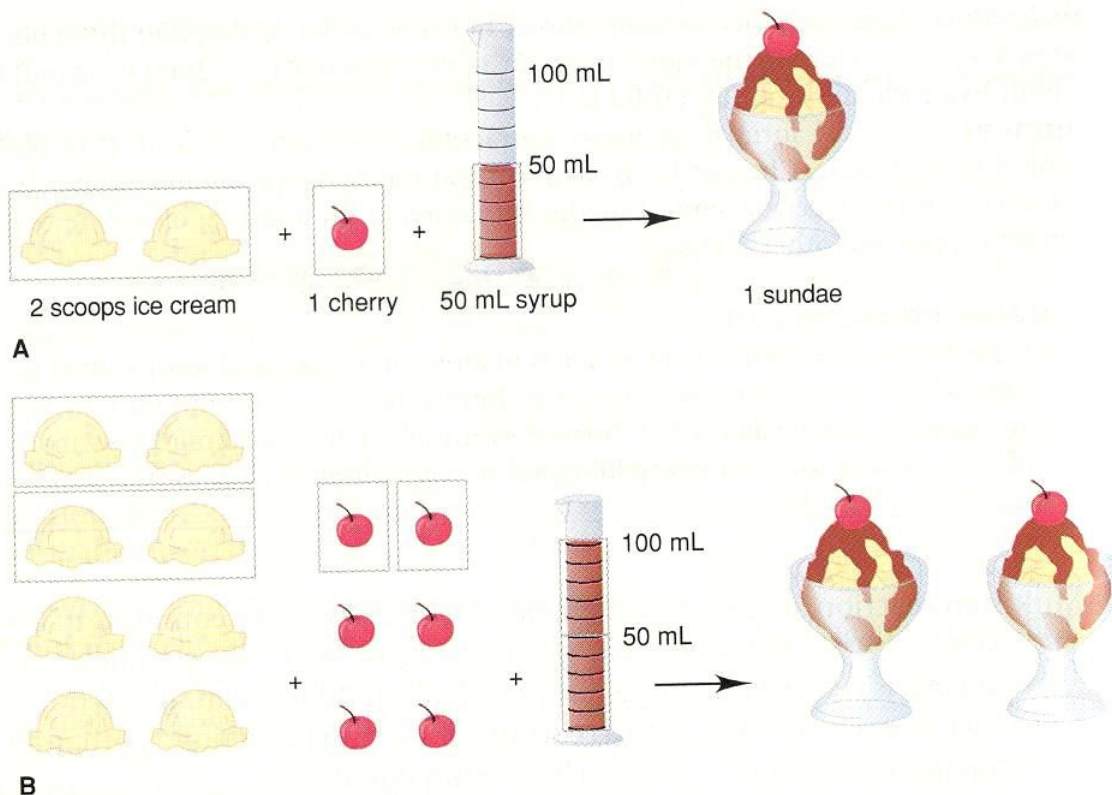


LIMITING REACTANT

- 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.

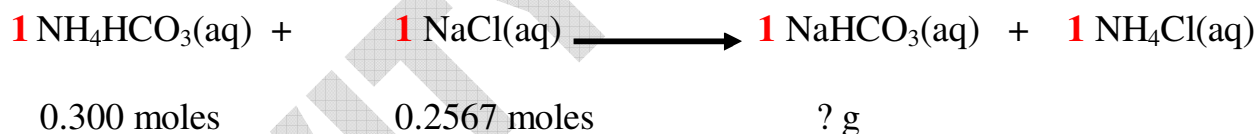
2. **Calculate the amount of product obtained from the Limiting Reactant**

Example1

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



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

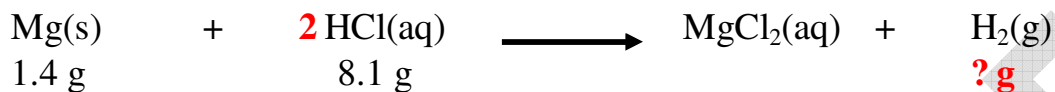


L.R.

$$\begin{aligned} ? \text{ g NaHCO}_3 &= \mathbf{0.2567 \text{ moles NaCl}} \times \frac{\mathbf{1 \text{ mole NaHCO}_3}}{\mathbf{1 \text{ mole NaCl}}} \times \frac{84.01 \text{ g NaHCO}_3}{1 \text{ mole NaHCO}_3} \\ &= \mathbf{21.57 \text{ g NaHCO}_3} \end{aligned}$$

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 ?



Change masses of reactants in moles:



$$\begin{array}{c} \boxed{\begin{array}{c} 1 \text{ mole} \\ 1.4 \text{ g} \times \frac{\quad}{24.31 \text{ g}} \end{array}} \end{array}$$

0.0576 moles

L.R.

requires

$$\boxed{\begin{array}{c} 1 \text{ mole} \\ 8.1 \text{ g} \times \frac{\quad}{36.46 \text{ g}} \end{array}}$$

0.222 moles

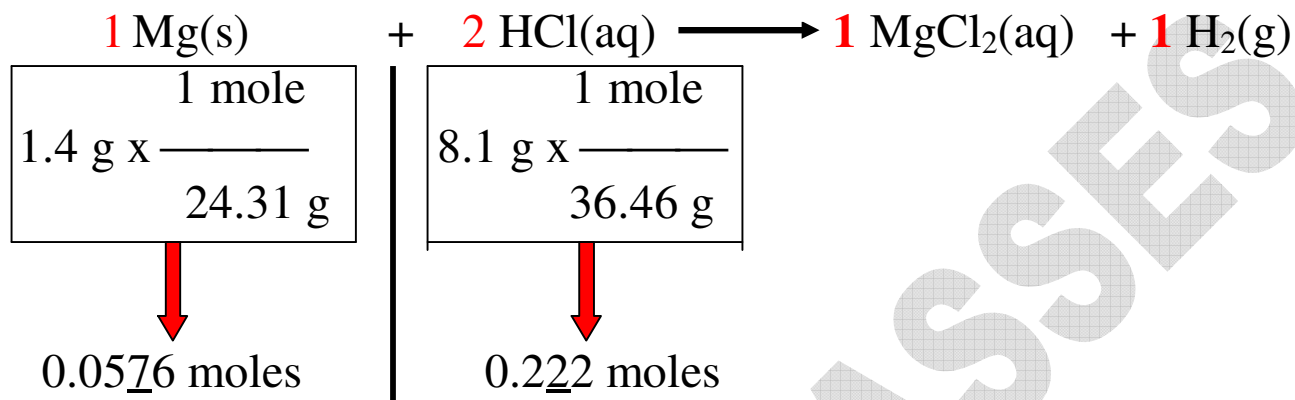
HCl is an excess!

**$2 \times 0.0576 \text{ moles HCl} = 0.115 \text{ moles HCl}$
 $(0.222 \text{ moles HCl}) \text{ available}$**

$$? \text{ g H}_2 = 0.0576 \text{ moles Mg} \times \frac{1 \text{ mole H}_2}{1 \text{ mole Mg}} \times \frac{2.02 \text{ g H}_2}{1 \text{ mole H}_2} = \mathbf{0.12 \text{ g H}_2}$$

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.

$$? \text{ g H}_2 = \text{0.222 moles HCl} \times \frac{\text{1 mole H}_2}{\text{2 moles HCl}} \times \frac{\text{2.02 g H}_2}{\text{1 mole H}_2} = \text{0.22 g H}_2$$

$$? \text{ g H}_2 = \text{0.0576 moles Mg} \times \frac{\text{1 mole H}_2}{\text{1 mole Mg}} \times \frac{\text{2.02 g H}_2}{\text{1 mole H}_2} = \text{0.12 g H}_2$$

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:



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

$$? \text{ g Ca}_3(\text{PO}_4)_2 = 35.0 \text{ g Ca(OH)}_2 \times \frac{1 \text{ mole Ca(OH)}_2}{74.10 \text{ g Ca(OH)}_2} \times \frac{1 \text{ mole Ca}_3(\text{PO}_4)_2}{3 \text{ mole Ca(OH)}_2} \times \frac{310.3 \text{ g Ca}_3(\text{PO}_4)_2}{1 \text{ mole Ca}_3(\text{PO}_4)_2} = 48.9 \text{ g Ca}_3(\text{PO}_4)_2$$

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

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

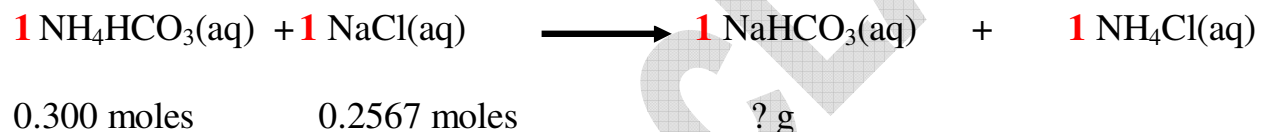
Example 2

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



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)



L.R.

$$? \text{ g NaHCO}_3 = \mathbf{0.2567 \text{ moles NaCl}} \times \frac{\mathbf{1 \text{ mole NaHCO}_3}}{\mathbf{1 \text{ mole NaCl}}} \times \frac{84.01 \text{ g NaHCO}_3}{1 \text{ mole NaHCO}_3} = \mathbf{21.57 \text{ g NaHCO}_3 \text{ (theoretical yield)}}$$

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

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