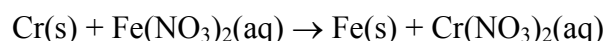


**Chapter 8 Applications of Stoichiometry****Solutions to Practice Problems****1.****Problem**

What would be the limiting reactant if 4.0 g of Cr(s) and 6.0 g of Fe(NO<sub>3</sub>)<sub>2</sub>(aq) were used in the reaction below?

**What is Required?**

You must determine which of the reactants is the limiting reactant.

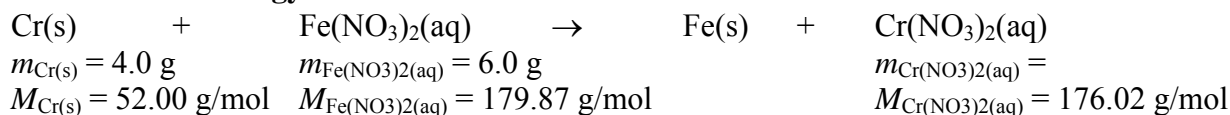
**What is Given?**

The balanced reaction is given.

The amounts of each of the reactants are given: 4.0 g of Cr(s) and 6.0 g of Fe(NO<sub>3</sub>)<sub>2</sub>(aq).

**Plan Your Strategy**

Complete a gravimetric stoichiometry calculation for each reactant. Whichever reactant produces the lowest mass of Cr(NO<sub>3</sub>)<sub>2</sub>(aq) is the limiting reactant.

**Act on Your Strategy**

Using Cr(s):

$$m_{\text{Cr(NO}_3)_2(\text{aq})} = 4.0 \text{ g Cr(s)} \times \frac{1 \text{ mol Cr(s)}}{52.00 \text{ g Cr(s)}} \times \frac{1 \text{ mol Cr(NO}_3)_2(\text{aq})}{1 \text{ mol Cr(s)}} \times \frac{176.02 \text{ g Cr(NO}_3)_2(\text{aq})}{1 \text{ mol Cr(NO}_3)_2(\text{aq})}$$

$$= 13.54 \text{ g}$$

Using Fe(NO<sub>3</sub>)<sub>2</sub>(aq):

$$m_{\text{Cr(NO}_3)_2(\text{aq})} = 6.0 \text{ g Fe(NO}_3)_2(\text{aq}) \times \frac{1 \text{ mol Fe(NO}_3)_2(\text{aq})}{179.87 \text{ g Fe(NO}_3)_2(\text{aq})} \times \frac{1 \text{ mol Cr(NO}_3)_2(\text{aq})}{1 \text{ mol Fe(NO}_3)_2(\text{aq})} \times \frac{176.02 \text{ g Cr(NO}_3)_2(\text{aq})}{1 \text{ mol Cr(NO}_3)_2(\text{aq})}$$

$$= 5.87 \text{ g}$$

The mass of product from 6.0 g of Fe(NO<sub>3</sub>)<sub>2</sub>(aq) is less than the mass of product from 4.0 g of Cr(s), so Fe(NO<sub>3</sub>)<sub>2</sub>(aq) is the limiting reactant.

**Check Your Solution**

Although there is more Fe(NO<sub>3</sub>)<sub>2</sub>(aq) available to react, its molecular weight is greater and it seems reasonable that it is the limiting reactant.

## 2.

### Problem

What would be the expected mass of Fe(s) formed in the reaction described in Question 1.

### What is Required?

You must calculate the expected mass of Fe(s).

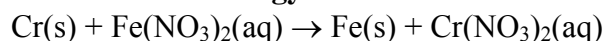
### What is Given?

The balanced equation is given. The mass of each of the reactants is known. Fe(NO<sub>3</sub>)<sub>2</sub>(aq) is the limiting reactant.

### Plan Your Strategy

Complete a gravimetric stoichiometry calculation using 6.0 g of Fe(NO<sub>3</sub>)<sub>2</sub>(aq) because it is the limiting reactant.

### Act on Your Strategy



$$M_{\text{Fe(s)}} = 55.85 \text{ g/mol}$$

$$\begin{aligned} m_{\text{Fe(s)}} &= 6.0 \text{ g Fe(NO}_3)_2(\text{aq}) \times \frac{1 \text{ mol Fe(NO}_3)_2(\text{aq})}{179.87 \text{ g Fe(NO}_3)_2(\text{aq})} \times \frac{1 \text{ mol Fe(s)}}{1 \text{ mol Fe(NO}_3)_2(\text{aq})} \times \frac{55.85 \text{ g Fe(s)}}{1 \text{ mol Fe(s)}} \\ &= 1.9 \text{ g} \end{aligned}$$

### Check Your Solution

The answer seems reasonable, has the correct units (g), and the correct number of significant digits (2).

## 3.

### Problem

Predict the mass of zinc sulfide produced from the reaction of 15 g of zinc with 15 g of yellow sulfur (S<sub>8</sub>(s)).

### What is Required?

You must determine which reagent, zinc or sulfur, is limiting in the reaction and calculate the mass of product, zinc sulfide.

### What is Given?

You know that 15 g of both zinc and sulfur are used in a formation reaction to produce zinc sulfide.

### Plan Your Strategy

Write the balanced equation for this reaction.

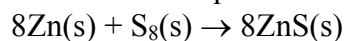
Determine the limiting reagent by calculating which reactant produces the lower mass of zinc sulfide. Calculate the molar mass,  $M$ , of each reactant and of the product. Use the formula  $n =$

$\frac{m}{M}$  to calculate the number of moles of each reactant. Use the mole ratios in the balanced

equation to calculate the number of moles of zinc sulfide that forms from each reactant. Convert the lower number of moles of ZnS(s) to grams using the formula  $m = n \times M$ . The mass of product can be calculated in multiple steps or as a one step calculation.

### Act on Your Strategy

The balanced equation for the reaction is



$$M(\text{Zn(s)}) = 65.39 \text{ g/mol}$$

$$M(\text{S}_8\text{(s)}) = 256.56 \text{ g/mol}$$

$$M(\text{ZnS(s)}) = 97.46 \text{ g/mol}$$

$$n(\text{Zn(s)}) = \frac{15 \text{ g}}{65.39 \frac{\text{g}}{\text{mol}}} = 0.229 \text{ mol}$$

$$n(\text{S}_8\text{(s)}) = \frac{15 \text{ g}}{256.56 \frac{\text{g}}{\text{mol}}} = 0.0585 \text{ mol}$$

Using Zn(s):

$$n\text{ZnS(s) produced} = \frac{0.229 \text{ mol Zn(s)} \times 8 \text{ mol ZnS(s)}}{8 \text{ mol Zn(s)}} = 0.229 \text{ mol ZnS(s)}$$

Using S<sub>8</sub>(s):

$$n\text{ZnS(s) produced} = \frac{0.0585 \text{ mol S}_8\text{(s)} \times 8 \text{ mol ZnS(s)}}{1 \text{ mol S}_8\text{(s)}} = 0.468 \text{ mol ZnS(s)}$$

Since the Zn(s) reactant produced a smaller amount of ZnS(s), Zn(s) is the limiting reagent. Based on this, the expected mass of ZnS(s) that will be produced is  $m = n \times M_{\text{ZnS(s)}} = 0.229 \text{ mol ZnS(s)} \times 97.46 \text{ g/mol} = 22 \text{ g ZnS(s)}$

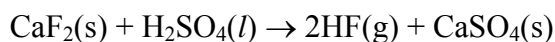
### Check Your Solution

The answer seems reasonable, has the correct units(g) and the correct number of significant digits (2).

## 4.

### Problem

Hydrogen fluoride, HF(g), is produced when concentrated sulfuric acid, H<sub>2</sub>SO<sub>4</sub>(l), reacts with the mineral fluorspar, CaF<sub>2</sub>(s):



Calculate the mass of HF(g) produced when 15 g of calcium fluoride is mixed with 20 g of H<sub>2</sub>SO<sub>4</sub>(l).

### What is Required?

You must determine which reagent, calcium fluoride or sulfuric acid, is limiting in the reaction and calculate the mass of product, hydrogen fluoride.

**What is Given?**

You know the mass of calcium fluoride is 15 g and the mass of sulfuric acid is 20 g. The balanced equation is given.

**Plan Your Strategy**

Determine the limiting reagent by calculating which reactant produces the lowest mass of hydrogen fluoride. Calculate the molar mass,  $M$ , of each reactant and of the product. Use the formula  $n = \frac{m}{M}$  to calculate the number of moles of each reactant. Use the mole ratios in the

balanced equation to calculate the number of moles of hydrogen fluoride that forms from each reactant. Convert the lower number of moles of HF(g) to grams using the formula  $m = n \times M$ . The mass of product can be calculated in multiple steps or as a one step calculation.

**Act on Your Strategy**

$$M(\text{CaF}_2(\text{s})) = 78.08 \text{ g/mol}$$

$$M(\text{H}_2\text{SO}_4(\text{l})) = 98.09 \text{ g/mol}$$

$$M(\text{HF}(\text{g})) = 20.01 \text{ g/mol}$$

$$n(\text{CaF}_2(\text{s})) = \frac{15 \text{ g}}{78.08 \frac{\text{g}}{\text{mol}}} = 0.192 \text{ mol}$$

$$n(\text{H}_2\text{SO}_4(\text{l})) = \frac{20 \text{ g}}{98.09 \frac{\text{g}}{\text{mol}}} = 0.204 \text{ mol}$$

Using  $\text{CaF}_2(\text{s})$ :

$$n\text{HF}(\text{g}) \text{ produced} = \frac{0.192 \text{ mol CaF}_2(\text{s}) \times 2 \text{ mol HF}(\text{g})}{1 \text{ mol CaF}_2(\text{s})} = 0.384 \text{ mol HF}(\text{g})$$

Using the  $\text{H}_2\text{SO}_4(\text{l})$  reactant:

$$n\text{HF}(\text{g}) \text{ produced} = \frac{0.204 \text{ mol H}_2\text{SO}_4(\text{l}) \times 2 \text{ mol HF}(\text{g})}{1 \text{ mol H}_2\text{SO}_4(\text{l})} = 0.408 \text{ mol HF}(\text{g})$$

Since the  $\text{CaF}_2(\text{s})$  reactant produced a smaller amount of HF(g),  $\text{CaF}_2(\text{s})$  is the limiting reagent. Based on this, the expected mass of HF(g) that will be produced is

$$m = n \times M(\text{HF}(\text{g})) = 0.384 \text{ mol (HF}(\text{g})) \times 20.01 \text{ g/mol} = 7.7 \text{ g HF}(\text{g})$$

**Check Your Solution**

The answer seems reasonable, has the correct units(g) and the correct number of significant digits (2).

**5.****Problem**

If 15 g of zinc powder is stirred into an aqueous solution of copper(II) nitrate that contains 50 g of solute, does the zinc react completely?

### What is Required?

You must determine if 15 of zinc reacts completely with a given mass of copper(II) nitrate.

### What is Given?

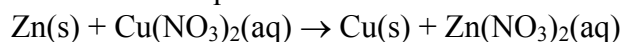
The reactants in a single replacement reaction are 15g of zinc solid and 50 g of copper(II) nitrate dissolved in aqueous solution.

### Plan Your Strategy

Write the balanced equation for this reaction. Calculate the molar mass,  $M$ , of each reactant. Use the formula  $n = \frac{m}{M}$  to calculate the number of moles of each reactant. Select one of the reactants, e.g., copper(II) nitrate, and determine how many moles of zinc it will react with. Compare this answer to the number of moles of zinc that were given. Based upon this information, determine if all of the zinc reacts with the copper(II) nitrate solution.

### Act on Your Strategy

The balanced equation for the reaction is



$$M(\text{Zn(s)}) = 65.39 \text{ g/mol}$$

$$M(\text{Cu(NO}_3)_2(\text{aq})) = 187.57 \text{ g/mol}$$

$$n(\text{Zn(s)}) = \frac{15 \text{ g}}{65.39 \frac{\text{g}}{\text{mol}}} = 0.229 \text{ mol}$$

$$n(\text{Cu(NO}_3)_2(\text{aq})) = \frac{50 \text{ g}}{187.57 \frac{\text{g}}{\text{mol}}} = 0.267 \text{ mol}$$

Using  $\text{Cu(NO}_3)_2(\text{aq})$ :

$$n\text{Zn(s) that react} = \frac{0.267 \text{ mol Cu(NO}_3)_2(\text{aq}) \times 1 \text{ mol Zn(s)}}{1 \text{ mol Cu(NO}_3)_2(\text{aq})} = 0.267 \text{ mol Zn(s)}$$

Since this is a larger number of moles of Zn(s) than is given, there is an excess of  $\text{Cu(NO}_3)_2(\text{aq})$  available to react with the Zn(s). All of 15 g of Zn(s) will react with the  $\text{Cu(NO}_3)_2(\text{aq})$  solution. Alternatively, if the calculation was based upon the 15 g (0.229 mol) of Zn(s),

$$n\text{Cu(NO}_3)_2(\text{aq) that react} = \frac{0.229 \text{ mol Zn(s)} \times 1 \text{ mol Cu(NO}_3)_2(\text{aq)}}{1 \text{ mol Zn(s)}} = 0.229 \text{ mol Cu(NO}_3)_2(\text{aq)}$$

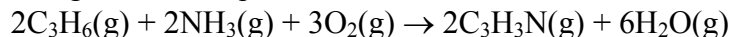
Since this is less than the amount of  $\text{Cu(NO}_3)_2(\text{aq})$  given (0.267 mol), the Zn(s) is limiting and would be completely used up.

### Check Your Solution

The answer seems reasonable and the same result is found using either reactant.

## 6. Problem

Acrylonitrile,  $C_3H_3N(g)$ , is the starting material for acrylic fibres. It is produced using the gas phase reaction of ammonia,  $NH_3(g)$ ; oxygen from the air,  $O_2(g)$ ; and propene,  $C_3H_6(g)$ , according to the following chemical equation:



What is the limiting reactant if 1.0 kg of  $C_3H_6(g)$  reacts with 600 g of  $NH_3(g)$ ?

### What is Required?

You must determine which reagent,  $C_3H_6(g)$  or  $NH_3(g)$ , is limiting in a reaction to produce acrylonitrile.

### What is Given?

You know the mass of  $C_3H_6(g)$  is 1.0 kg and the mass of  $NH_3(g)$  is 600 g. The balanced equation is given.

### Plan Your Strategy

Determine the limiting reagent by calculating which reactant produces the lower mass of  $C_3H_3N(g)$ . Calculate the molar mass,  $M$ , of each reactant and of the product.

Recall that 1.0 kg = 1000 g. Use the formula  $n = \frac{m}{M}$  to calculate the number of moles of each

reactant. Use the mole ratios in the balanced equation to calculate the number of moles of  $C_3H_3N(g)$  that forms from each reactant. Convert the lower number of moles of  $C_3H_3N(g)$  to grams using the formula  $m = n \times M$ .

The mass of product can be calculated in multiple steps or as a one step calculation.

### Act on Your Strategy

$$M(C_3H_6(g)) = 42.09 \text{ g/mol}$$

$$M(NH_3(g)) = 17.04 \text{ g/mol}$$

$$M(C_3H_3N(g)) = 53.07 \text{ g/mol}$$

$$n(C_3H_6(g)) = \frac{1000 \text{ g}}{42.09 \frac{\text{g}}{\text{mol}}} = 23.7586 \text{ mol}$$

$$n(NH_3(g)) = \frac{600 \text{ g}}{17.04 \frac{\text{g}}{\text{mol}}} = 35.2113 \text{ mol}$$

Using  $C_3H_6(g)$ :

$$n_{C_3H_3N(g) \text{ produced}} = \frac{23.7586 \text{ mol } C_3H_6(g) \times 1 \text{ mol } C_3H_3N(g)}{1 \text{ mol } C_3H_6(g)} = 23.7586 \text{ mol } C_3H_3N(g)$$

Using  $NH_3(g)$ :

$$n_{C_3H_3N(g) \text{ produced}} = \frac{35.2113 \text{ mol } NH_3(g) \times 2 \text{ mol } C_3H_3N(g)}{2 \text{ mol } NH_3(g)} = 35.2113 \text{ mol } C_3H_3N(g)$$

Since the  $C_3H_6(g)$  reactant produced a smaller amount of  $C_3H_3N(g)$ ,  $C_3H_6(g)$  is the limiting reagent.

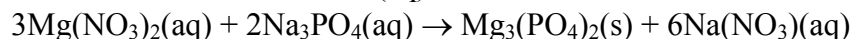
### Check Your Solution

The answer seems reasonable, and is consistent with the definition of limiting reagent.

7.

#### Problem

What would be the limiting reactant if 100.0 mL of a 0.5 mol/L solution of  $\text{Mg}(\text{NO}_3)_2(\text{aq})$  and 125.0 mL of a 1.2 mol/L solution of  $\text{Na}_3\text{PO}_4(\text{aq})$  were used in the reaction below?



#### What is Required?

You must determine which reagent,  $\text{Mg}(\text{NO}_3)_2(\text{aq})$  or  $\text{Na}_3\text{PO}_4(\text{aq})$ , is the limiting reactant.

#### What is Given?

The balanced equation is given.

The volumes and concentrations of each of the reactants are given: 100.0 mL of 0.5 mol/L  $\text{Mg}(\text{NO}_3)_2(\text{aq})$  and 125.0 mL of 1.2 mol/L  $\text{Na}_3\text{PO}_4(\text{aq})$ .

#### Plan Your Strategy

Calculate the number of moles of each reactant using the equation  $n = c \times V$ . Use the mole ratios from the balanced equation to calculate the number of moles of  $\text{Mg}_3(\text{PO}_4)_2(\text{s})$  that form from each reactant. The limiting reagent is the reactant that produces the lower amount of  $\text{Mg}_3(\text{PO}_4)_2(\text{s})$ .

#### Act on Your Strategy

For  $\text{Mg}(\text{NO}_3)_2(\text{aq})$ :

$$n\text{Mg}(\text{NO}_3)_2(\text{aq}) = c \times V = 0.5 \text{ mol/L} \times 100.0 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.0500 \text{ mol}$$

From the balanced equation:

$$\frac{n(\text{Mg}(\text{NO}_3)_2(\text{aq}))}{n(\text{Mg}_3(\text{PO}_4)_2(\text{s}))} = \frac{3}{1}$$

$$\frac{0.0500 \text{ mol Mg}(\text{NO}_3)_2(\text{aq})}{n(\text{Mg}_3(\text{PO}_4)_2(\text{s}))} = \frac{3}{1}$$

$$n\text{Mg}_3(\text{PO}_4)_2(\text{s}) \text{ produced} = \frac{0.0500 \text{ mol Mg}(\text{NO}_3)_2(\text{aq})}{3} = 0.0166 \text{ mol Mg}_3(\text{PO}_4)_2(\text{s})$$

For  $\text{Na}_3\text{PO}_4(\text{aq})$ :

$$n\text{Na}_3\text{PO}_4(\text{aq}) = c \times V = 1.2 \text{ mol/L} \times 125.0 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.150 \text{ mol}$$

From the balanced equation:

$$\frac{n(\text{Na}_3\text{PO}_4(\text{aq}))}{n(\text{Mg}_3(\text{PO}_4)_2(\text{s}))} = \frac{2}{1}$$

$$\frac{0.150 \text{ mol Na}_3\text{PO}_4(\text{aq})}{n(\text{Mg}_3(\text{PO}_4)_2(\text{s}))} = \frac{2}{1}$$

$$n\text{Mg}_3(\text{PO}_4)_2(\text{s}) \text{ produced} = \frac{0.150 \text{ mol Na}_3\text{PO}_4(\text{aq})}{2} = 0.0750 \text{ mol Mg}_3(\text{PO}_4)_2(\text{s})$$

$\text{Mg}(\text{NO}_3)_2(\text{aq})$  is the limiting reactant because it produces the lower amount of  $\text{Mg}_3(\text{PO}_4)_2(\text{s})$ .

### Check Your Solution

From the information provided, the answer seems reasonable.

## 8.

### Problem

What would be the expected mass of  $\text{Mg}_3\text{PO}_4(\text{s})$  formed in the reaction described in question 7?

### What is Required?

You must calculate the expected mass of  $\text{Mg}_3\text{PO}_4(\text{s})$  formed in the reaction described in Question 7.

### What is Given?

The balanced equation is given.

The volume and concentration of the limiting reactant as determined in question 7 is given 100.0 mL of 0.5 mol/L  $\text{Mg}(\text{NO}_3)_2(\text{aq})$ .

### Plan Your Strategy

Use the mole ratio in the balanced equation to calculate the number of moles of  $\text{Mg}_3(\text{PO}_4)_2(\text{s})$  that form from 100.0 mL of 0.5 mol/L  $\text{Mg}(\text{NO}_3)_2(\text{aq})$ . Calculate the mass of  $\text{Mg}_3(\text{PO}_4)_2(\text{s})$  that would be produced using the equation  $m = n \times M$ .

### Act on Your Strategy

From question 7 we know that:

$$n\text{Mg}(\text{NO}_3)_2(\text{aq}) = c \times V = 0.5 \text{ mol/L} \times 100.0 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.0500 \text{ mol}$$

From the balanced equation:

$$\frac{\text{Mg}(\text{NO}_3)_2(\text{aq})}{\text{Mg}_3(\text{PO}_4)_2(\text{s})} = \frac{3}{1}$$

$$\frac{0.0500 \text{ mol Mg}(\text{NO}_3)_2(\text{aq})}{n(\text{Mg}_3(\text{PO}_4)_2(\text{s}))} = \frac{3}{1}$$

$$n\text{Mg}_3\text{PO}_4(\text{s}) \text{ produced} = \frac{0.0500 \text{ mol Mg}(\text{NO}_3)_2(\text{aq})}{3} = 0.0166 \text{ mol}$$

$$M\text{Mg}_3\text{PO}_4(\text{s}) = 167.90 \text{ g/mol}$$

$$m\text{Mg}_3\text{PO}_4(\text{s}) = n \times M = 0.0166 \text{ mol} \times 167.90 \text{ g/mol} = 2.787 \text{ g} \approx 3 \text{ g}$$

### Check Your Solution

The answer seems reasonable, has the correct units (g), and has the correct number of significant digits (1).

## 9.

### Problem



Would 600 mL of 0.085 mol/L Na<sub>2</sub>S(s) be sufficient to remove all the mercury(II) ions as HgS(s) from 200 mL of B5 solution (0.221 mol/L HgCl<sub>2</sub>(aq))? Explain your answer.

### What is Required?

You must determine if a given volume of 0.085 mol/L Na<sub>2</sub>S(aq) can precipitate the Hg<sup>2+</sup>(aq) ions from a given volume of B5 solution having a concentration of 0.221 mol/L HgCl<sub>2</sub>(aq).

### What is Given?

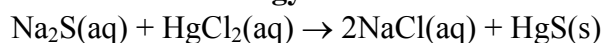
For Na<sub>2</sub>S(aq):  
 $V = 600 \text{ mL} = 0.600 \text{ L}$   
 $c = 0.085 \text{ mol/L}$

For HgCl<sub>2</sub>(aq):  
 $c = 0.221 \text{ mol/L}$

### Plan Your Strategy

Write the balanced equation for the reaction. Determine the limiting reagent by calculating which reactant produces the lower amount of mercury(II) sulfide, HgS(s). Calculate the number of moles of each reactant using the equation  $n = c \times V$ . Use the mole ratios in the balanced equation to calculate the number of moles of HgS(s) that form from each reactant. Use this information to explain if sufficient Na<sub>2</sub>S(aq) has been provided.

### Act on Your Strategy



Using Na<sub>2</sub>S(aq):

$$n_{\text{Na}_2\text{S}(\text{aq})} = 0.085 \text{ mol/L} \times 0.600 \text{ L} = 0.0510 \text{ mol Na}_2\text{S}(\text{aq})$$

$$n_{\text{HgS}(\text{s}) \text{ produced}} = 0.0510 \text{ mol Na}_2\text{S}(\text{aq}) \times \frac{1 \text{ mol HgS}(\text{s})}{1 \text{ mol Na}_2\text{S}(\text{aq})} = 0.0510 \text{ mol HgS}(\text{s})$$

Using HgCl<sub>2</sub>(aq):

$$n_{\text{HgCl}_2(\text{aq})} = 0.221 \text{ mol/L} \times 0.200 \text{ L} = 0.0442 \text{ mol HgCl}_2(\text{aq})$$

$$n_{\text{HgS}(\text{s}) \text{ produced}} = 0.0442 \text{ mol HgCl}_2(\text{aq}) \times \frac{1 \text{ mol HgS}(\text{s})}{1 \text{ mol HgCl}_2(\text{aq})} = 0.0442 \text{ mol HgS}(\text{s})$$

The HgCl<sub>2</sub>(aq) produces less HgS(s) and, therefore, is the limiting reagent. This means that the HgCl<sub>2</sub>(aq) in the B5 solution would be used up first. There is sufficient Na<sub>2</sub>S(aq) to precipitate all of the Hg<sup>2+</sup>(aq).

### Checking Your Solution

This seems to be a reasonable answer and the calculations have the correct units (mol) and number of significant digits.

## 10.

### Problem

If 250 mL of 0.400 mol/L aqueous lead(II) nitrate, Pb(NO<sub>3</sub>)<sub>2</sub>(aq), is mixed with 300 mL of 0.22 mol/L aqueous potassium iodide, KI(aq), what is the maximum mass of precipitate that would be formed in the resulting reaction?

### What is Required?

You must determine which reagent,  $\text{Pb}(\text{NO}_3)_2(\text{aq})$  or  $\text{KI}(\text{aq})$ , is limiting in the reaction and calculate the mass of product,  $\text{PbI}_2(\text{s})$ .

### What is Given?

For  $\text{Pb}(\text{NO}_3)_2(\text{aq})$

$c = 0.400 \text{ mol/L}$

$V = 250 \text{ ml} = 0.250 \text{ L}$

For  $\text{KI}(\text{aq})$

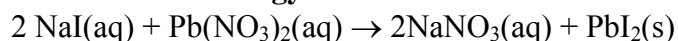
$c = 0.22 \text{ mol/L}$

$V = 300 \text{ mL} = 0.300 \text{ L}$

### Plan Your Strategy

Write the balanced equation for this reaction. Determine the limiting reagent by calculating which reactant produces the lower amount of precipitate  $\text{PbI}_2(\text{s})$ . Use the mole ratios from the balanced equation to calculate the number of moles of  $\text{PbI}_2(\text{s})$  that form from each reactant. Convert the lower number of moles of  $\text{PbI}_2(\text{s})$  to grams using the formula  $m = n \times M$ .

### Act on Your Strategy



Using  $\text{NaI}(\text{aq})$ :

$$n_{\text{NaI}(\text{aq})} = 0.22 \text{ mol/L} \times 0.300 \text{ L} = 0.0660 \text{ mol NaI}(\text{aq})$$

$$n_{\text{PbI}_2(\text{s}) \text{ produced}} = 0.0660 \text{ mol NaI}(\text{aq}) \times \frac{1 \text{ mol PbI}_2(\text{s})}{2 \text{ mol NaI}(\text{aq})} = 0.0330 \text{ mol PbI}_2(\text{s})$$

Using  $\text{Pb}(\text{NO}_3)_2(\text{aq})$ :

$$n_{\text{Pb}(\text{NO}_3)_2(\text{aq})} = 0.400 \text{ mol/L} \times 0.250 \text{ L} = 0.100 \text{ mol Pb}(\text{NO}_3)_2(\text{aq})$$

$$n_{\text{PbI}_2(\text{s}) \text{ produced}} = 0.100 \text{ mol Pb}(\text{NO}_3)_2(\text{aq}) \times \frac{1 \text{ mol PbI}_2(\text{s})}{1 \text{ mol Pb}(\text{NO}_3)_2(\text{aq})} = 0.100 \text{ mol PbI}_2(\text{s})$$

Since  $\text{NaI}(\text{aq})$  produced less  $\text{PbI}_2(\text{s})$ ,  $\text{NaI}(\text{aq})$  is the limiting reagent. Based upon this, the mass of  $\text{PbI}_2(\text{s})$  produced =  $0.0330 \text{ mol} \times 461.0 \text{ g/mol} = 15 \text{ g}$ .

### Check Your Solution

The answer seems to be reasonable given this data, has the correct units (g) and the correct number of significant digits (2).

## 11.

### Problem

Priti mixed 25.0 mL of 0.320 mol/L aqueous copper(II) sulfate,  $\text{CuSO}_4(\text{aq})$ , with 29.7 mL of 0.270 mol/L aqueous strontium nitrate,  $\text{Sr}(\text{NO}_3)_2(\text{aq})$ .

a) Write the balanced reaction equation for this precipitation reaction.

b) The precipitate that Priti collected by filtration through paper and dried in an oven was a white, powdery solid. What colour was the filtrate after the reaction mixture was filtered? Explain your answer.

c) Calculate the mass of precipitate that Priti can expect from this reaction?

d) Priti collected 1.432 g of precipitate. Calculate the percentage yield of precipitate.

### What is Required?

- You must write the balanced equation for the reaction between  $\text{CuSO}_4(\text{aq})$  and  $\text{Sr}(\text{NO}_3)_2(\text{aq})$ .
- You must determine which ion was in excess to determine the colour of the filtrate.
- You must use the balanced equation to calculate the mass of precipitate that is produced in the reaction.
- You must calculate the percentage yield of precipitate in the reaction.

### What is Given?

- The chemical formulas of the reactants are given.

**b)**

For  $\text{CuSO}_4(\text{aq})$

$$c = 0.320 \text{ mol/L}$$

$$V = 25.0 \text{ ml} = 0.0250 \text{ L}$$

For  $\text{Sr}(\text{NO}_3)_2(\text{aq})$

$$c = 0.270 \text{ mol/L}$$

$$V = 29.7 \text{ mL} = 0.0297 \text{ L}$$

- Use the balanced equation from **a**).

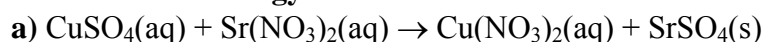
- The experimental yield in the reaction is 1.432 g.

### Plan Your Strategy

- Write the balanced equation for this double replacement reaction.
- Determine the limiting reagent in the reaction by calculating which reactant produces the lower amount of  $\text{SrSO}_4(\text{s})$ . Use the mole ratios in the balanced equation to calculate the number of moles of  $\text{SrSO}_4(\text{s})$  that form from each reactant.
- Convert the lower number of moles of  $\text{SrSO}_4(\text{s})$  to grams using the formula  $m = n \times M$ .
- Calculate the percentage yield using the equation

$$\text{percentage yield} = \frac{\text{experimental yield}}{\text{predicted yield}} \times 100\%$$

### Act on Your Strategy



The precipitate is strontium sulphate,  $\text{SrSO}_4(\text{s})$ .

**b)** Using  $\text{CuSO}_4(\text{aq})$ :

$$n\text{CuSO}_4(\text{aq}) = 0.320 \text{ mol/L} \times 0.0250 \text{ L} = 0.00800 \text{ mol CuSO}_4(\text{aq})$$

$$n\text{SrSO}_4(\text{s}) \text{ produced} = 0.00800 \text{ mol CuSO}_4(\text{aq}) \times \frac{1 \text{ mol SrSO}_4(\text{s})}{1 \text{ mol CuSO}_4(\text{aq})} = 0.00800 \text{ mol SrSO}_4(\text{s})$$

Using  $\text{Sr}(\text{NO}_3)_2(\text{aq})$ :

$$n \text{Sr}(\text{NO}_3)_2(\text{aq}) = 0.270 \text{ mol/L} \times 0.0297 \text{ L} = 0.008019 \text{ mol Sr}(\text{NO}_3)_2(\text{aq})$$

$$n\text{SrSO}_4(\text{s}) \text{ produced} = 0.008019 \text{ mol Sr}(\text{NO}_3)_2(\text{aq}) \times \frac{1 \text{ mol SrSO}_4(\text{s})}{1 \text{ mol Sr}(\text{NO}_3)_2(\text{aq})} = 0.008019 \text{ mol}$$

$\text{SrSO}_4(\text{s})$

Since  $\text{CuSO}_4(\text{aq})$  produced less  $\text{SrSO}_4(\text{s})$ ,  $\text{CuSO}_4(\text{aq})$  is the limiting reagent. Since the  $\text{Sr}(\text{NO}_3)_2(\text{aq})$  is in excess, the filtrate will be red.

**c)**  $M\text{SrSO}_4(\text{s}) = 183.69 \text{ g/mol}$

Based upon  $\text{CuSO}_4(\text{aq})$  as the limiting reagent, the mass of  $\text{SrSO}_4(\text{s})$  produced is  $0.008000 \text{ mol} \times 183.69 \text{ g/mol} = 1.4695 \text{ g} \approx 1.47 \text{ g}$

$$\text{d) Percentage yield} = \frac{\text{experimental yield}}{\text{predicted yield}} \times 100\% = \frac{1.432 \text{ g}}{1.4695 \text{ g}} = 97.4\%$$

### Check Your Solution

The answer seems to be reasonable given this data, have the correct units (g and %) and the correct number of significant digits (3).

## 12.

### Problem

Silicate salts of most transition metal ions, many of which are too toxic for disposal down the drain, have very low aqueous solubilities. It is common laboratory practice to add  $\text{Na}_2\text{SiO}_3(\text{aq})$  to remove these ions for disposal as solid waste. Suppose 150 mL of 0.250 mol/L aqueous sodium silicate,  $\text{Na}_2\text{SiO}_3(\text{aq})$ , is mixed with 950 mL of a 0.035 mol/L silver nitrate solution,  $\text{AgNO}_3(\text{aq})$ .

- Write the balanced chemical equation for this precipitation reaction.
- Which reactant is the limiting reactant? Explain your answer.
- Calculate the predicted yield of precipitate.
- Are all the silver ions removed from the solution? Explain briefly.

### What is Required?

- You must write the balanced equation for the reaction between  $\text{Na}_2\text{SiO}_3(\text{aq})$  and  $\text{AgNO}_3(\text{aq})$ .
- You must determine which reagent is limiting in this reaction.
- You must use the balanced equation to calculate the mass of precipitate (the predicted yield) that is produced in the reaction.
- You must determine if all of the  $\text{Ag}^+(\text{aq})$  is precipitated in the reaction.

### What is Given?

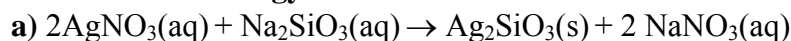
- The chemical formulas of the reactants are given.
- For  $\text{Na}_2\text{SiO}_3(\text{aq})$ 

$c = 0.250 \text{ mol/L}$	For $\text{AgNO}_3(\text{aq})$
$V = 150 \text{ mL} = 0.150 \text{ L}$	$c = 0.035 \text{ mol/L}$
	$V = 950 \text{ mL} = 0.950 \text{ L}$

### Plan Your Strategy

- Write the balanced equation for this double replacement reaction.
- Determine the limiting reagent in the reaction by calculating which reactant produces the lower amount of  $\text{Ag}_2\text{SiO}_3(\text{s})$ . Use the mole ratios in the balanced equation to calculate the number of moles of  $\text{Ag}_2\text{SiO}_3(\text{s})$  that forms from each reactant.
- Convert the lower number of moles of  $\text{Ag}_2\text{SiO}_3(\text{s})$  to grams using the formula  $m = n \times M$ .

### Act on Your Strategy



The precipitate is silver silicate,  $\text{Ag}_2\text{SiO}_3(\text{s})$ .

b) Using  $\text{AgNO}_3(\text{aq})$ :

$$n_{\text{AgNO}_3(\text{aq})} = 0.035 \text{ mol/L} \times 0.950 \text{ L} = 0.0332 \text{ mol}$$

$$n_{\text{Ag}_2\text{SiO}_3(\text{s}) \text{ produced}} = 0.0332 \text{ mol AgNO}_3(\text{aq}) \times \frac{1 \text{ mol Ag}_2\text{SiO}_3(\text{s})}{2 \text{ mol AgNO}_3(\text{aq})} = 0.0166 \text{ mol Ag}_2\text{SiO}_3(\text{s})$$

Using  $\text{Na}_2\text{SiO}_3(\text{aq})$ :

$$n_{\text{Na}_2\text{SiO}_3(\text{aq})} = 0.250 \text{ mol/L} \times 0.150 \text{ L} = 0.0375 \text{ mol Na}_2\text{SiO}_3(\text{aq})$$

$$n_{\text{Ag}_2\text{SiO}_3(\text{s}) \text{ produced}} = 0.0375 \text{ mol Na}_2\text{SiO}_3(\text{aq}) \times \frac{1 \text{ mol Ag}_2\text{SiO}_3(\text{s})}{1 \text{ mol Na}_2\text{SiO}_3(\text{aq})} = 0.0375 \text{ mol}$$

$\text{Ag}_2\text{SiO}_3(\text{s})$

Since the  $\text{AgNO}_3(\text{aq})$  produced less  $\text{Ag}_2\text{SiO}_3(\text{s})$ ,  $\text{AgNO}_3(\text{aq})$  is the limiting reagent.

c)  $M_{\text{Ag}_2\text{SiO}_3(\text{s})} = 291.83 \text{ g/mol}$

Based upon this,

$$m_{\text{Ag}_2\text{SiO}_3(\text{s}) \text{ produced}} = 0.0166 \text{ mol} \times 291.83 \text{ g/mol} = 4.9 \text{ g}$$

d) Yes, all of the  $\text{Ag}^+(\text{aq})$  are removed because the  $\text{AgNO}_3(\text{aq})$  is the limiting reagent.

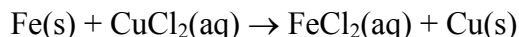
### Check Your Solution

The answer seems to be reasonable given this data, have the correct units (g) and the correct number of significant digits (2).

## 13.

### Problem

For the following reaction:



a) What is the predicted yield if 10.0 g of  $\text{Fe}(\text{s})$  is reacted with an excess amount of  $\text{CuCl}_2(\text{aq})$ ?

b) What would be the percentage yield if 9.0 g of  $\text{Cu}(\text{s})$  is actually obtained?

### What is Required?

a) Determine the predicted yield of  $\text{Cu}(\text{s})$ .

b) Determine the percentage yield of  $\text{Cu}(\text{s})$ .

### What is Given?

a) The balanced equation is given. The amount of  $\text{Fe}(\text{s})$  is 10.0 g.

b) The balanced equation is given. The actual yield of  $\text{Cu}(\text{s})$  is 9.0 g.

### Plan Your Strategy

a) Calculate the number of g of  $\text{Fe}(\text{s})$  using the equation  $n \times M = m$ . Using the mole ratio from the balanced equation, calculate the predicted yield of  $\text{Cu}(\text{s})$  if 10.0 g of  $\text{Fe}(\text{s})$  is reacted.

b) Calculate the number of moles of  $\text{Cu}(\text{s})$  obtained using the equation  $n = \frac{m}{M}$ . Calculate the

percentage yield if 9.0 g of  $\text{Cu}(\text{s})$  is actually obtained using

$$\text{percentage yield} = \frac{\text{experimental yield}}{\text{predicted yield}} \times 100\%$$

### Act on Your Strategy

a)  $M_{\text{Fe}(\text{s})} = 55.85 \text{ g/mol}$

$M_{\text{Cu}(\text{s})} = 63.55 \text{ g/mol}$

$$n_{\text{Fe(s)}} = \frac{m}{M} = \frac{10.0 \text{ g}}{55.85 \frac{\text{g}}{\text{mol}}} = 0.17905 \text{ mol}$$

From the balanced equation:

$$\frac{n_{\text{Fe(s)}}}{n_{\text{Cu(s)}}} = \frac{1}{1}$$

$$n_{\text{Cu(s)}} = \frac{0.17905 \text{ mol Fe(s)}}{1} = 0.17905 \text{ mol Cu(s)}$$

$$0.17905 \text{ mol Cu(s)} \times 63.55 \text{ g/mol} = 11.3786 \text{ g} = 11.4 \text{ g Cu(s)}$$

**b)**  $M_{\text{Cu(s)}} = 63.55 \text{ g/mol}$

$$n_{\text{Cu(s)}} = \frac{m}{M} = \frac{9.0 \text{ g}}{63.55 \frac{\text{g}}{\text{mol}}} = 0.1416 \text{ mol Cu(s)}$$

$$\text{Percentage yield} = \frac{\text{experimental yield}}{\text{predicted yield}} \times 100\% = \frac{0.1416 \text{ mol Cu(s)}}{0.17905 \text{ mol Cu(s)}} \times 100\% = 79\%$$

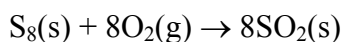
### Check Your Solution

The answers seem reasonable, have the correct units (mol and %), and have the correct number of significant digits (2).

## 14.

### Problem

Calculate the percentage yield for the following reaction if 60 g of  $\text{SO}_2(\text{s})$  is produced using 50 g of  $\text{S}_8(\text{s})$ .



### What is Required?

You must calculate the percentage yield for the reaction if 60 g of  $\text{SO}_2(\text{s})$  is produced using 50 g of  $\text{S}_8(\text{s})$ .

### What is Given?

The balanced equation is given.

60 g of  $\text{SO}_2(\text{s})$  is produced using 50 g of  $\text{S}_8(\text{s})$ .

### Plan Your Strategy

Calculate the number of moles of  $\text{S}_8(\text{s})$  used with the equation  $n = \frac{m}{M}$ . Then using the molar ratios from the balanced equation, calculate the predicted yield of  $\text{SO}_2(\text{s})$  using 50 g of  $\text{S}_8(\text{s})$ . Determine the number of moles of  $\text{SO}_2(\text{s})$  that were actually produced. Use these values to determine the percentage yield with the equation  $\text{percentage yield} = \frac{\text{experimental yield}}{\text{predicted yield}} \times 100\%$ .

**Act on Your Strategy**

$$MS_8(s) = 256.56 \text{ g/mol}$$

$$MSO(s) = 64.07 \text{ g/mol}$$

Predicted yield of  $SO_2(s)$  with 50 g of  $S_8(s)$ :

$$mSO_2(s) = 50 \text{ g } S_8(s) \times \frac{1 \text{ mol } S_8(s)}{256.56 \text{ g}} \times \frac{8 \text{ mol } SO_2(s)}{1 \text{ mol } S_8(s)} \times \frac{64.07 \text{ g } SO_2(s)}{1 \text{ mol } SO_2(s)} = 99.890 \text{ g } SO_2(s)$$

$$\begin{aligned} \text{Percentage yield} &= \frac{\text{experimental yield}}{\text{predicted yield}} \times 100\% = \frac{60 \text{ g } SO_2(s)}{99.890 \text{ g } SO_2(s)} \times 100\% \\ &= 60\% \end{aligned}$$

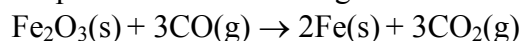
**Check Your Solution**

The answer seems to be reasonable given this data, has the correct units (%), and has the correct number of significant digits (2).

**15.**

**Problem**

A chemical engineering student performs the following reaction:



The student carries out three trials with the same mass of  $Fe_2O_3(s)$  and obtains the following masses of  $Fe(s)$ .

Trial 1: 23.5 g

Trial 2: 23.2 g

Trial 3: 23.9 g

The predicted yield of  $Fe(s)$  is 24.6 g. Calculate the percentage yield for each trial and the average percentage yield for this series of trials.

**What is Required?**

You must calculate the percentage yield for each trial and average these results.

**What is Given?**

The predicted yield for the reaction is 24.6 g. The experimental yields for three trials are:

Trial 1: 23.5 g

Trial 2: 23.2 g

Trial 3: 23.9 g

**Plan Your Strategy**

Calculate the percentage yield for each trial using the formula:

$$\text{Percentage yield} = \frac{\text{experimental yield}}{\text{predicted yield}} \times 100\%$$

Average the results.

### Act on Your Strategy

$$\text{Trial 1: Percentage yield} = \frac{23.5 \text{ g}}{24.6 \text{ g}} \times 100\% = 95.5\%$$

$$\text{Trial 2: Percentage yield} = \frac{23.2 \text{ g}}{24.6 \text{ g}} \times 100\% = 94.3\%$$

$$\text{Trial 3: Percentage yield} = \frac{23.9 \text{ g}}{24.6 \text{ g}} \times 100\% = 97.2\%$$

$$\text{Average percentage yield} = \frac{95.5\% + 94.3\% + 97.2\%}{3} = 95.7\%$$

### Check Your Solution

The average seems reasonable and has the correct number of significant digits (3).

## 16.

### Problem

Prairie Chemical in northeast Edmonton manufactures household and industrial bleach from liquid chlorine using the following process:



For every 0.150 kg of liquid chlorine consumed, 0.150 kg of NaOCl(aq) is produced. Calculate the percentage yield of NaOCl(aq).

### What is Required?

You must calculate the percentage yield of NaOCl(aq) for the reaction in which liquid bleach is prepared.

### What is Given?

The balanced equation is given. The mass of the reactant chlorine,  $\text{Cl}_2(l)$ , is 0.150 kg and the experimental yield of NaOCl(aq) is 0.150 kg.

### Plan Your Strategy

Recall that 1.0 kg = 1 000g and convert 0.150 kg to grams. Determine the molar mass,  $M$ , for  $\text{Cl}_2(l)$  and for NaOCl(aq). Use the equation  $n = \frac{m}{M}$  to calculate the moles of  $\text{Cl}_2(l)$ .

Use the mole ratio in the balanced equation to calculate the moles of NaOCl(aq) that are predicted to form from this number of moles of  $\text{Cl}_2(l)$ . Calculate the mass of NaOCl(aq) using the equation  $m = n \times M$ .

Calculate the percentage yield using  $\text{percentage yield} = \frac{\text{experimental yield}}{\text{predicted yield}} \times 100\%$ .

### Act on Your Strategy

$$0.150 \text{ kg} = 150 \text{ g}$$

$$M_{\text{Cl}_2(l)} = 70.90 \text{ g/mol}$$

$$M_{\text{NaOCl}(aq)} = 74.44 \frac{\text{g}}{\text{mol}}$$



$$n_{\text{Cl}_2(l)} = \frac{150 \text{ g}}{70.90 \frac{\text{g}}{\text{mol}}} = 2.116 \text{ mol Cl}_2(l)$$

$$n_{\text{NaOCl(aq)}} \text{ predicted} = \frac{2.116 \text{ mol Cl}_2(l) \times 1 \text{ mol NaOCl}}{1 \text{ mol Cl}_2(l)} = 2.116 \text{ mol NaOCl(aq)}$$

$$m_{\text{NaOCl(aq)}} \text{ predicted} = 2.116 \text{ mol} \times 74.44 \text{ g/mol} = 157.5 \text{ g NaOCl(aq)}$$

$$\text{Percentage yield} = \frac{150 \text{ g}}{157.5 \text{ g}} \times 100\% = 95.2\%$$

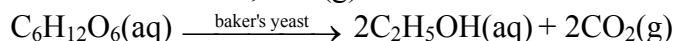
### Check Your Solution

The answer seems to be reasonable, has the correct units (%), and has the correct number of significant digits (3).

## 17.

### Problem

The fermentation enzymes of baker's yeast convert a solution of glucose,  $\text{C}_6\text{H}_{12}\text{O}_6(\text{aq})$ , to ethanol,  $\text{C}_2\text{H}_5\text{OH}(\text{aq})$ , and carbon dioxide,  $\text{CO}_2(\text{g})$ :



If 223 g of ethanol obtained from the fermentation of 1.63 kg of glucose, what is the percentage yield of the reaction?

### What is Required?

You must calculate the percentage yield of ethanol from a known mass of glucose in a fermentation reaction.

### What is Given?

The balanced equation is given. The mass of the reactant glucose is 223 g and the experimental yield of ethanol is 1.63 kg.

### Plan Your Strategy

Recall that  $1.0 \text{ kg} = 1\,000 \text{ g}$  and convert 1.63 kg to grams. Determine the molar mass,  $M$ , for

$\text{C}_6\text{H}_{12}\text{O}_6(\text{aq})$  and for  $\text{C}_2\text{H}_5\text{OH}(\text{aq})$ . Use the equation  $n = \frac{m}{M}$  to calculate the moles of

$\text{C}_6\text{H}_{12}\text{O}_6(\text{aq})$ . Use the mole ratio in the balanced equation to calculate the moles  $\text{C}_2\text{H}_5\text{OH}(\text{aq})$  of that are predicted to form from this number of moles of  $\text{C}_6\text{H}_{12}\text{O}_6(\text{aq})$ . Calculate the mass of  $\text{C}_2\text{H}_5\text{OH}(\text{aq})$  using the equation  $m = n \times M$ .

Calculate the percentage yield using  $\text{percentage yield} = \frac{\text{experimental yield}}{\text{predicted yield}} \times 100\%$ .

### Act on Your Strategy

$$1.63 \text{ kg} = 1630 \text{ g}$$

$$M_{\text{C}_6\text{H}_{12}\text{O}_6(\text{aq})} = 180.16 \text{ g/mol}$$

$$M_{\text{C}_2\text{H}_5\text{OH}(\text{aq})} = 46.08 \text{ g/mol}$$

$$n_{\text{C}_6\text{H}_{12}\text{O}_6(\text{aq})} = \frac{1630 \text{ g}}{180.16 \frac{\text{g}}{\text{mol}}} = 9.048 \text{ mol C}_6\text{H}_{12}\text{O}_6(\text{aq})$$

$$n_{\text{C}_2\text{H}_5\text{OH}(\text{aq})} \text{ predicted} = \frac{9.048 \text{ mol C}_6\text{H}_{12}\text{O}_6(\text{aq}) \times 2 \text{ mol C}_2\text{H}_5\text{OH}(\text{aq})}{1 \text{ mol C}_6\text{H}_{12}\text{O}_6(\text{aq})} = 18.10 \text{ mol}$$

$\text{C}_2\text{H}_5\text{OH}(\text{aq})$

$$m_{\text{C}_2\text{H}_5\text{OH}(\text{aq})} \text{ predicted} = 18.10 \text{ mol} \times 46.08 \text{ g/mol} = 833.8 \text{ g C}_2\text{H}_5\text{OH}(\text{aq})$$

$$\text{Percentage yield} = \frac{223 \text{ g}}{833.8 \text{ g}} \times 100\% = 26.7\%$$

### Check Your Solution

The answer is reasonable, has the correct units (%), and has the correct number of significant digits (3).

## 18.

### Problem

When a 1 g to 5 g sample of the pale-green mineral malachite,  $\text{Cu}_2(\text{CO}_3) \cdot (\text{OH})_2(\text{s})$ , is heated vigorously over a Bunsen flame for about 20 minutes, it is transformed into black copper(II) oxide:



a) Calculate the predicted yield of  $\text{CuO}(\text{s})$  from vigorous heating of 4.00 g of malachite.

b) If 2.80 g of  $\text{CuO}(\text{s})$  remains after the  $\text{CO}_2(\text{g})$  and  $\text{H}_2\text{O}(\text{g})$  have been burned off, what is the percentage yield of the solid product?

### What is Required?

You must calculate the predicted yield of  $\text{CuO}(\text{s})$  that will form when a known mass of malachite is heated.

### What is Given?

a) The balanced equation is given. The mass of the reactant malachite,  $\text{CuCO}_3 \cdot \text{Cu}(\text{OH})_2(\text{s})$ , is 4.00 g.

b) The experimental yield of  $\text{CuO}(\text{s})$  is 2.80 g.

### Plan Your Strategy

a) Determine the molar mass,  $M$ , for  $\text{CuCO}_3 \cdot \text{Cu}(\text{OH})_2(\text{s})$  and of  $\text{CuO}(\text{s})$ . Use the equation  $n = \frac{m}{M}$  to calculate the moles of malachite,  $\text{CuCO}_3 \cdot \text{Cu}(\text{OH})_2(\text{s})$ . Use the mole ratio in the balanced equation to calculate the moles  $\text{CuO}(\text{s})$  that are predicted to form from this number of moles of  $\text{CuCO}_3 \cdot \text{Cu}(\text{OH})_2(\text{s})$ . Calculate the mass of  $\text{CuO}(\text{s})$  using the equation  $m = n \times M$ .

b) Predict the percentage yield using  $\text{percentage yield} = \frac{\text{experimental yield}}{\text{predicted yield}} \times 100\%$ .

### Act on Your Strategy

a)  $M_{\text{CuCO}_3 \cdot \text{Cu}(\text{OH})_2(\text{s})} = 221.13 \text{ g/mol}$

$$n\text{CuCO}_3 \cdot \text{Cu}(\text{OH})_2(\text{s}) = \frac{4.00 \text{ g}}{221.13 \frac{\text{g}}{\text{mol}}} = 0.180889 \text{ mol CuCO}_3 \cdot \text{Cu}(\text{OH})_2(\text{s})$$

$$n\text{CuO}(\text{s}) \text{ predicted} = \frac{0.180889 \text{ mol CuCO}_3 \cdot \text{Cu}(\text{OH})_2(\text{s}) \times 2 \text{ mol CuO}(\text{s})}{1 \text{ mol CuCO}_3 \cdot \text{Cu}(\text{OH})_2(\text{s})} = 0.36178 \text{ mol CuO}(\text{s})$$

$$m\text{CuO}(\text{s}) \text{ predicted} = 0.3618 \text{ mol} \times 79.55 \text{ g/mol} = 2.8779 \text{ g CuO}(\text{s})$$

$$\text{b) Percentage yield} = \frac{2.80 \text{ g}}{2.88 \text{ g}} \times 100\% = 97.29\% = 97.3\%$$

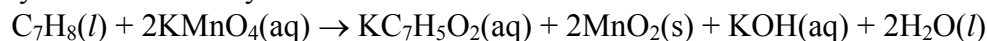
### Check Your Solution

The answer is reasonable, has the correct units (%), and has the correct number of significant digits (3).

## 19.

### Problem

The reaction of toluene,  $\text{C}_7\text{H}_8(\text{l})$ , with potassium permanganate,  $\text{KMnO}_4(\text{aq})$ , proceeds with significantly less than 100% yield under most conditions:



- If 8.60 g toluene reacts with excess potassium permanganate, what is the predicted yield, in grams, of potassium benzoate,  $\text{KC}_7\text{H}_5\text{O}_2(\text{aq})$ ?
- If the percentage yield is 70.0%, what mass of potassium benzoate would you expect to be produced?
- What mass of toluene is needed to produce 13.4 g of potassium benzoate if the percentage yield is 60.0%?

### What is Required?

- You must calculate a predicted yield of potassium benzoate that can be produced from a given mass of toluene.
- You must determine the experimental yield, in grams, when the percentage yield is given.
- You must calculate the mass of toluene needed to produce a given mass of potassium benzoate when the percentage yield is known.

### What is Given?

- The balanced equation is given and the mass of the reactant toluene is 8.60 g.
- The percentage yield for the reaction in **a)** is given.
- The experimental yield of potassium benzoate is 13.4 g and the percentage yield for the reaction is 60.0%.

### Plan Your Strategy

- Calculate the molar mass of toluene and of potassium benzoate.

Determine the number of moles of toluene using the equation  $n = \frac{m}{M}$ .

Use the mole ratio from the balanced equation to calculate the number of moles of  $\text{KC}_7\text{H}_5\text{O}_2(\text{aq})$  that will be produced.

Calculate the predicted mass of  $\text{KC}_7\text{H}_5\text{O}_2(\text{aq})$  that will be produced using  $m = n \times M$ .

b) Calculate the experimental yield of  $\text{KC}_7\text{H}_5\text{O}_2(\text{aq})$  using

$$\text{percentage yield} = \frac{\text{experimental yield}}{\text{predicted yield}} \times 100\%$$

c) Use the equation for experimental yield and the given percentage yield to calculate the predicted mass of  $\text{KC}_7\text{H}_5\text{O}_2(\text{aq})$ . Calculate the number of moles of  $\text{KC}_7\text{H}_5\text{O}_2(\text{aq})$  and use the mole ratio in the balanced equation to determine the number of moles of toluene that are required. Calculate the mass of  $\text{C}_7\text{H}_8(\text{l})$  using the equation  $n = m \times M$ .

### Act on Your Strategy

a)  $M_{\text{C}_7\text{H}_8(\text{l})} = 92.15 \text{ g/mol}$

$M_{\text{KC}_7\text{H}_5\text{O}_2(\text{aq})} = 160.22 \text{ g/mol}$

$$n_{\text{C}_7\text{H}_8(\text{l})} = \frac{8.60 \text{ g}}{92.15 \frac{\text{g}}{\text{mol}}} = 0.09333 \text{ mol}$$

$$0.09333 \text{ mol C}_7\text{H}_8(\text{l}) \times \frac{1 \text{ mol KC}_7\text{H}_5\text{O}_2(\text{aq})}{1 \text{ mol C}_7\text{H}_8(\text{l})} = 0.09333 \text{ mol KC}_7\text{H}_5\text{O}_2(\text{aq})$$

$$m_{\text{KC}_7\text{H}_5\text{O}_2(\text{aq})} = 0.09333 \text{ mol KC}_7\text{H}_5\text{O}_2(\text{aq}) \times 160.22 \text{ g/mol} = 15.0 \text{ g}$$

b)  $\text{Percentage yield KC}_7\text{H}_5\text{O}_2(\text{aq}) = \frac{\text{experimental yield}}{\text{predicted yield}} \times 100\% = 70.0\% \times 15.0 \text{ g} = 10.5 \text{ g}$

c)  $\text{Predicted yield of KC}_7\text{H}_5\text{O}_2(\text{aq}) = \frac{\text{experimental yield}}{\text{percentage yield}} \times 100\% = \frac{13.4}{60.0} \times 100\% = 22.33 \text{ g}$

$$n_{\text{KC}_7\text{H}_5\text{O}_2(\text{aq})} = \frac{22.33 \text{ g}}{160.22 \frac{\text{g}}{\text{mol}}} = 0.1394 \text{ mol}$$

$$n_{\text{C}_7\text{H}_8(\text{l})} = \frac{0.1394 \text{ mol KC}_7\text{H}_5\text{O}_2(\text{aq}) \times 1 \text{ mol C}_7\text{H}_8(\text{l})}{1 \text{ mol KC}_7\text{H}_5\text{O}_2(\text{aq})} = 0.1394 \text{ mol C}_7\text{H}_8(\text{l})$$

$$m_{\text{C}_7\text{H}_8(\text{l})} = 0.1394 \text{ mol} \times 92.15 \text{ g/mol} = 12.8 \text{ g}$$

### Check Your Solution

The answers are reasonable, have the appropriate units and the correct number of significant digits (3).

## 20.

### Problem

If the pH of a 50 mL solution of  $\text{NaOH}(\text{aq})$  reaches 7 after you add 36.2 mL of a 1.5 mol/L solution of acetic acid,  $\text{CH}_3\text{COOH}(\text{aq})$ , what is the concentration of the  $\text{NaOH}(\text{aq})$  solution?

### What is Required?

You must calculate the concentration of the  $\text{NaOH}(\text{aq})$  solution.

### What is Given?

$$V(\text{NaOH}(\text{aq})) = 50 \text{ mL} = 0.050 \text{ L}$$

$$c(\text{CH}_3\text{COOH}(\text{aq})) = 1.50 \text{ mol/L}$$

$$V(\text{CH}_3\text{COOH}(\text{aq})) = 36.2 \text{ mL} = 0.0362 \text{ L}$$

### Plan Your Strategy

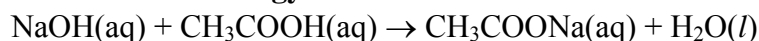
Write the balanced chemical equation for the reaction.

Calculate the moles of  $\text{CH}_3\text{COOH}(\text{aq})$  using the equation  $n = c \times V$ .

Use the mole ratio from the balanced equation to determine the moles of  $\text{NaOH}(\text{aq})$  that react.

Calculate the concentration of  $\text{NaOH}(\text{aq})$  using the equation  $c = \frac{n}{V}$ .

### Act on Your Strategy



$$n\text{CH}_3\text{COOH}(\text{aq}) = c \times V = 1.50 \text{ mol/L} \times 0.0362 \text{ L} = 0.0543 \text{ mol}$$

From the balanced equation:

$$\frac{n\text{CH}_3\text{COOH}(\text{aq})}{n\text{NaOH}(\text{aq})} = \frac{1}{1}$$

$$n\text{NaOH}(\text{aq}) = \frac{0.0543 \text{ mol CH}_3\text{COOH}(\text{aq})}{1} = 0.0543 \text{ mol NaOH}(\text{aq})$$

$$c\text{NaOH}(\text{aq}) = \frac{n}{V} = \frac{0.0543 \text{ mol NaOH}(\text{aq})}{0.050 \text{ L}} = 1.086 = 1.1 \text{ mol/L}$$

### Check Your Solution

The answer seems reasonable, has the appropriate units (mol/L), and the correct number of significant digits (2).

## 21.

### Problem

A student titrates 20.00 mL of  $\text{HCl}(\text{aq})$  with 0.150 mol/L  $\text{NaOH}(\text{aq})$ . The initial reading on the burette is 1.50 mL. The final reading on the burette is 29.51 mL. What is the concentration of the  $\text{HCl}(\text{aq})$ ?

### What is Required?

You must calculate the concentration of the hydrochloric acid solution.

### What is Given?

$$V(\text{HCl}_{\text{initial}}) = 1.50 \text{ mL}$$

$$V(\text{HCl}_{\text{final}}) = 19.35 \text{ mL}$$

$$c(\text{NaOH}(\text{aq})) = 0.150 \text{ mol/L}$$

$$V(\text{NaOH}(\text{aq})) = 25.00 \text{ mL} = 0.02500 \text{ L}$$

### Plan Your Strategy

Write the balanced equation for the reaction.

Calculate the volume of  $\text{HCl}(\text{aq})$  used in the titration.

Calculate the moles of  $\text{NaOH}(\text{aq})$  that are given using the equation  $n = c \times V$ .

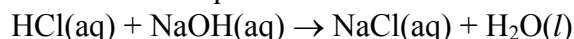
Use the mole ratio from the balanced equation to determine the moles of HCl(aq) that react.

Calculate the concentration of HCl(aq) using the equation  $c = \frac{n}{V}$ .

### Act on Your Strategy

$$V_{\text{HCl(aq) used}} = V_{\text{final}} - V_{\text{initial}} = 19.35 \text{ mL} - 1.50 \text{ mL} = 17.85 \text{ mL} = 0.01785 \text{ L}$$

The balanced equation for the reaction is



$$n_{\text{NaOH(aq)}} = c \times V = 0.150 \text{ mol/L} \times 0.02500 \text{ L} = 0.003750 \text{ mol}$$

$$n_{\text{HCl(aq)}} = 0.003750 \text{ mol NaOH(aq)} \times \frac{1 \text{ mol HCl(aq)}}{1 \text{ mol NaOH(aq)}} = 0.003750 \text{ mol}$$

$$c_{\text{HCl(aq)}} = \frac{n}{V} = \frac{0.003750 \text{ mol}}{0.01785 \text{ L}} = 0.210 \text{ mol/L}$$

### Check Your Solution

The answer seems reasonable based on the given data. The unit is correct (mol/L) and the number of significant digits is correct (3).

## 22.

### Problem

A student uses 0.2030 mol/L NaOH(aq) to titrate 40.00 mL of HNO<sub>3</sub>(aq). The initial reading on the burette is 0.05 mL. The final reading on the burette at endpoint is 38.00 mL. What is the concentration of the HNO<sub>3</sub>(aq)?

### What is Required?

You must calculate the concentration of the nitric acid solution.

### What is Given?

$$V(\text{HNO}_3_{\text{initial}}) = 0.05 \text{ mL}$$

$$V(\text{HNO}_3_{\text{final}}) = 38.00 \text{ mL}$$

$$c(\text{NaOH(aq)}) = 0.2030 \text{ mol/L}$$

$$V(\text{NaOH(aq)}) = 40.00 \text{ mL} = 0.04000 \text{ L}$$

### Plan Your Strategy

Write the balanced equation for the reaction.

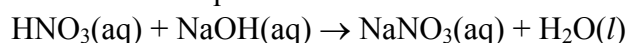
Calculate the volume of HNO<sub>3</sub>(aq) used in the titration. Calculate the moles of NaOH(aq) that are given using the equation  $n = c \times V$ . Use the mole ratio from the balanced equation to determine

the moles of HNO<sub>3</sub>(aq) that react. Calculate the concentration of HNO<sub>3</sub>(aq) using  $c = \frac{n}{V}$ .

### Act on Your Strategy

$$V_{\text{HNO}_3(\text{aq})} = V_{\text{final}} - V_{\text{initial}} = 38.00 \text{ mL} - 0.050 \text{ mL} = 37.95 \text{ mL} = 0.03795 \text{ L}$$

The balanced equation for the reaction is



$$n_{\text{NaOH(aq)}} = c \times V = 0.2030 \text{ mol/L} \times 0.04000 \text{ L} = 0.008120 \text{ mol}$$

$$n\text{HNO}_3(\text{aq}) = 0.008120 \text{ mol NaOH}(\text{aq}) \times \frac{1 \text{ mol HNO}_3(\text{aq})}{1 \text{ mol NaOH}(\text{aq})} = 0.008120 \text{ mol}$$

$$c\text{HCl}(\text{aq}) = \frac{n}{V} = \frac{0.008120 \text{ mol}}{0.03795 \text{ L}} = 0.21392 \text{ mol/L} = 0.2 \text{ mol/L}$$

**Check Your Solution**

The answer seems reasonable based on the given data. The unit is correct (mol/L) and the number of significant digits is correct (4).

**23.****Problem**

A student performs a series of titrations in order to determine the concentration of a KOH(aq) solution. The titrant used is a 0.5 mol/L solution of HNO<sub>3</sub>(aq). The following is a table summarizing the student's experimental data for the 3 trials that were performed:

Trial #	1	2	3
Final burette volume	5.35	5.42	6.58
Initial burette volume	0.15	0.20	1.43
Volume of KOH(aq)	10.0 mL	10.0 mL	10.0 mL

Calculate the concentration of the KOH(aq) solution for each trial. What is the average value for the concentration of KOH(aq)?

**What is Required?**

You must calculate the concentration of KOH(aq) for each trial. You must calculate the average concentration value of KOH(aq).

**What is Given?**

$$c(\text{HNO}_3(\text{aq})) = 0.5 \text{ mol/L}$$

$$\text{Trial 1: Final burette volume} = 5.35 \text{ mL}$$

$$\text{Trial 2: Final burette volume} = 5.42 \text{ mL}$$

$$\text{Trial 3: Final burette volume} = 6.58 \text{ mL}$$

$$\text{Trial 1: Initial burette volume} = 0.15 \text{ mL}$$

$$\text{Trial 2: Initial burette volume} = 0.20 \text{ mL}$$

$$\text{Trial 3: Initial burette volume} = 1.43 \text{ mL}$$

$$\text{Trial 1: } V(\text{KOH}(\text{aq})) = 10.0 \text{ mL}$$

$$\text{Trial 2: } V(\text{KOH}(\text{aq})) = 10.0 \text{ mL}$$

$$\text{Trial 3: } V(\text{KOH}(\text{aq})) = 10.0 \text{ mL}$$

**Plan Your Strategy**

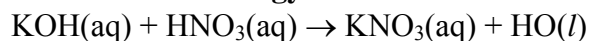
Write the balanced equation for the reaction.

Calculate the volume of HNO<sub>3</sub>(aq) used in the titration.

Calculate the moles of HNO<sub>3</sub>(aq) that are given using the equation  $n = c \times V$ .

Use the mole ratio from the balanced equation to determine the moles of KOH(aq) that react.

Calculate the concentration of KOH(aq) using the equation  $c = \frac{n}{V}$ .

**Act on Your Strategy**

$V_{\text{HNO}_3(\text{aq})}$  used in the titration:

$$\text{Trial 1: } 5.35 \text{ mL} - 0.15 \text{ mL} = 5.20 \text{ mL}$$

$$\text{Trial 2: } 5.42 \text{ mL} - 0.20 \text{ mL} = 5.22 \text{ mL}$$

$$\text{Trial 3: } 6.58 \text{ mL} - 1.43 \text{ mL} = 5.15 \text{ mL}$$

$n_{\text{HNO}_3(\text{aq})} = c \times V$  for each trial in the titration:

$$\text{Trial 1: } 0.5 \text{ mol/L} \times 5.20 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.00260 \text{ mol}$$

$$\text{Trial 2: } 0.5 \text{ mol/L} \times 5.22 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.00261 \text{ mol}$$

$$\text{Trial 3: } 0.5 \text{ mol/L} \times 5.15 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.00258 \text{ mol}$$

$n_{\text{KOH(aq)}}$  that react for each trial in the titration, from the balanced equation  $\frac{1 \text{ mol HNO}_3(\text{aq})}{1 \text{ mol KOH(aq)}}$ :

$$\text{Trial 1: } 0.00260 \text{ mol HNO}_3(\text{aq}) = 0.00260 \text{ mol KOH(aq)}$$

$$\text{Trial 2: } 0.00261 \text{ mol HNO}_3(\text{aq}) = 0.00261 \text{ mol KOH(aq)}$$

$$\text{Trial 3: } 0.00258 \text{ mol HNO}_3(\text{aq}) = 0.00258 \text{ mol KOH(aq)}$$

$c_{\text{KOH(aq)}} = \frac{n}{V}$  for each trial in the titration:

$$\text{Trial 1: } \frac{0.00260 \text{ mol KOH(aq)}}{10.0 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.260 = 0.26 \text{ mol/L}$$

$$\text{Trial 2: } \frac{0.00261 \text{ mol KOH(aq)}}{10.0 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.261 = 0.26 \text{ mol/L}$$

$$\text{Trial 3: } \frac{0.00258 \text{ mol KOH(aq)}}{10.0 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.258 = 0.26 \text{ mol/L}$$

Average  $c_{\text{KOH(aq)}}$  from all 3 trials:

$$\frac{0.260 + 0.261 + 0.258}{3} = 0.2596 \approx 0.26 \text{ mol/L}$$

**Check Your Solution**

The answers seem reasonable, have the appropriate units (mol/L), and the correct number of significant digits (2).

**24.****Problem**

A student is about to titrate 25.00 mL of  $\text{HCl(aq)}$  using 0.1500 mol/L  $\text{NaOH(aq)}$ . The teacher tells the class that the concentration of the acid is approximately 0.2 mol/L.

**a)** Using this information, estimate what volume of sodium hydroxide will be required to titrate the hydrochloric acid. (**Hint:** Use  $V = \frac{n}{c}$ .)

**b)** Why might it be useful to perform this type of calculation before a titration?



### What is Required?

a) You must calculate the approximate volume of HCl(aq) needed to titrate a standard solution of NaOH(aq).

### What is Given?

The approximate concentration of the HCl(aq) is 0.2 mol/L

$c_{\text{NaOH(aq)}} = 0.1500 \text{ mol/L}$

$V_{\text{NaOH(aq)}} = 25.00 \text{ mL} = 0.02500 \text{ L}$

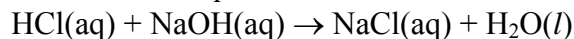
### Plan Your Strategy

Write the balanced equation for the reaction. Calculate the number of moles of NaOH(aq) using the equation  $n = c \times V$ . Use the molar ratio from the balanced equation to calculate the number of moles of HCl(aq) that react with the NaOH(aq). Estimate the volume of 0.2 mol/L HCl(aq)

required in the titration using the equation  $V = \frac{n}{c}$ .

### Act on Your Strategy

The balanced equation for the reaction is



$$n_{\text{HCl(aq)}} = 25.00 \text{ mL} \times 0.2 \text{ mol/L} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.005 \text{ mol}$$

$$n_{\text{NaOH(aq) needed}} = c \times V = 0.1500 \text{ mol/L} \times 0.02500 \text{ L} = 0.00375 \text{ mol}$$

$$n_{\text{NaOH(aq)}} = \frac{0.005 \text{ mol HCl(aq)}}{1} = 0.005 \text{ mol NaOH(aq)}$$

$$V_{\text{NaOH(aq)}} = \frac{n}{c} = \frac{0.005 \text{ mol NaOH(aq)}}{\frac{0.15 \text{ mol}}{\text{L}} \text{ NaOH(aq)}} = 0.0333 \text{ L} \approx 33.3 \text{ mL} \text{ or } 3 \times 10^1 \text{ mL}$$

### Check Your Solution

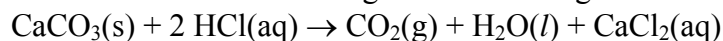
a) The answer seems to be reasonable based upon the given data. The unit is correct (mL) and the number of significant digits is correct (1).

b) By performing this calculation before the titration, you know approximately the volume of titrant that will be used. The HCl(aq) can be added quickly up to a few mL of the endpoint (perhaps 15 mL), then proceed to titrate slowly toward the endpoint. Therefore, time is saved in performing the titration.

## 25.

### Problem

Geologists carry small bottles of hydrochloric acid in the field with them to identify carbonate rocks and minerals. When a few drops of the acid are applied to a carbonate containing rock or mineral, the geologist observes fizzing. The test geologists perform can be recreated on a larger scale in the chemistry laboratory. It takes 125.0 mL of hydrochloric acid to react with 3.28 g of calcium carbonate according to the following reaction equation:



What is the molar concentration of the hydrochloric acid?

### What is Required?

You must calculate the concentration of a hydrochloric acid solution.

### What is Given?

A balanced equation is given.

$$V(\text{HCl}(\text{aq})) = 125 \text{ mL} = 0.125 \text{ L}$$

$$m(\text{CaCO}_3(\text{s})) = 3.28 \text{ g}$$

### Plan Your Strategy

Determine the molar mass,  $M$ , of  $\text{CaCO}_3(\text{s})$  and calculate the number of moles of  $\text{CaCO}_3(\text{s})$  using

$$\text{the equation } n = \frac{m}{M}.$$

Use the mole ratio in the balanced equation to calculate the number of moles of  $\text{HCl}(\text{aq})$ .

$$\text{Calculate the volume of } \text{HCl}(\text{aq}) \text{ using the formula } V = \frac{n}{c}.$$

### Act on Your Strategy

$$M_{\text{CaCO}_3(\text{s})} = 100.09 \frac{\text{g}}{\text{mol}}$$

$$n_{\text{CaCO}_3(\text{s})} = \frac{m}{M} = \frac{3.28 \text{ g}}{100.09 \frac{\text{g}}{\text{mol}}} = 0.03277 \text{ mol}$$

$$n_{\text{HCl}(\text{aq})} = 0.03277 \text{ mol CaCO}_3(\text{s}) \times \frac{2 \text{ mol HCl}(\text{aq})}{1 \text{ mol CaCO}_3(\text{s})} = 0.06554 \text{ mol}$$

$$c_{\text{HCl}(\text{aq})} = \frac{n}{V} = \frac{0.06554 \text{ mol}}{0.125 \text{ L}} = 0.524 \text{ mol/L}$$

### Check Your Solution

The answer seems reasonable based on the given data. The unit is correct (mol/L) and the number of significant digits is correct (3).