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 $Fe(NO_3)_2$(aq) were used in the reaction below?

$$Cr(s) + Fe(NO_3)_2(aq) \rightarrow Fe(s) + Cr(NO_3)_2(aq)$$

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 $Fe(NO_3)_2$(aq).

Plan Your Strategy
Complete a gravimetric stoichiometry calculation for each reactant. Whichever reactant produces the lowest mass of $Cr(NO_3)_2$(aq) is the limiting reactant.

Act on Your Strategy

Cr(s)	+	$Fe(NO_3)_2$(aq)	→	Fe(s)	+	$Cr(NO_3)_2$(aq)
$m_{Cr(s)}$ = 4.0 g		$m_{Fe(NO3)2(aq)}$ = 6.0 g				$m_{Cr(NO3)2(aq)}$ =
$M_{Cr(s)}$ = 52.00 g/mol		$M_{Fe(NO3)2(aq)}$ = 179.87 g/mol				$M_{Cr(NO3)2(aq)}$ = 176.02 g/mol

Using Cr(s):

$$m_{Cr(NO3)2(aq)} = 4.0 \text{ g Cr(s)} \times \frac{1 \text{ mol Cr(s)}}{52.00 \text{ g Cr(s)}} \times \frac{1 \text{ mol Cr(NO}_3)_2\text{(aq)}}{1 \text{ mol Cr(s)}} \times \frac{176.02 \text{ g Cr(NO}_3)_2\text{(aq)}}{1 \text{ mol Cr(NO}_3)_2\text{(aq)}}$$
$$= 13.54 \text{ g}$$

Using $Fe(NO_3)_2$(aq):

$$m_{Cr(NO3)2(aq)} = 6.0 \text{ g Fe(NO}_3)_2\text{(aq)} \times \frac{1 \text{ mol Fe(NO}_3)_2\text{(aq)}}{179.87 \text{ g Fe(NO}_3)_2\text{(aq)}} \times \frac{1 \text{ mol Cr(NO}_3)_2\text{(aq)}}{1 \text{ mol Fe(NO}_3)_2\text{(aq)}} \times$$
$$\frac{176.02 \text{ g Cr(NO}_3)_2\text{(aq)}}{1 \text{ mol Cr(NO}_3)_2\text{(aq)}}$$
$$= 5.87 \text{ g}$$

The mass of product from 6.0 g of $Fe(NO_3)_2$(aq) is less than the mass of product from 4.0 g of Cr(s), so $Fe(NO_3)_2$(aq) is the limiting reactant.

Check Your Solution
Although there is more $Fe(NO_3)_2$(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. $Fe(NO_3)_2(aq)$ is the limiting reactant.

Plan Your Strategy
Complete a gravimetric stoichiometry calculation using 6.0 g of $Fe(NO_3)_2(aq)$ because it is the limiting reactant.

Act on Your Strategy
$Cr(s) + Fe(NO_3)_2(aq) \rightarrow Fe(s) + Cr(NO_3)_2(aq)$
$M_{Fe(s)} = 55.85$ g/mol

$$m_{Fe(s)} = 6.0\text{ g }Fe(NO_3)_2(aq) \times \frac{1\text{ mol }Fe(NO_3)_2(aq)}{179.87\text{ g }Fe(NO_3)_2(aq)} \times \frac{1\text{ mol }Fe(s)}{1\text{ mol }Fe(NO_3)_2(aq)} \times \frac{55.85\text{ g }Fe(s)}{1\text{ mol }Fe(s)}$$
$$= 1.9\text{ g}$$

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 ($S_8(s)$).

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
$8Zn(s) + S_8(s) \rightarrow 8ZnS(s)$
$M(Zn(s)) = 65.39$ g/mol
$M(S_8(s)) = 256.56$ g/mol
$M(ZnS(s)) = 97.46$ g/mol

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

$$n(S_8(s)) = \frac{15\text{ g}}{256.56\,\frac{\text{g}}{\text{mol}}} = 0.0585\text{ mol}$$

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_8(s):

$$n\text{ZnS(s) produced} = \frac{0.0585\text{ mol } S_8(s) \times 8\text{ mol ZnS(s)}}{1\text{ mol } S_8(s)} = 0.468\text{ mol ZnS(s)}$$

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
$m = n \times M$ZnS(s) = 0.229 mol ZnS(s) × 97.46 g/mol = 22 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, $H_2SO_4(l)$, reacts with the mineral fluorspar, $CaF_2(s)$:

$$CaF_2(s) + H_2SO_4(l) \rightarrow 2HF(g) + CaSO_4(s)$$

Calculate the mass of HF(g) produced when 15 g of calcium fluoride is mixed with 20 g of $H_2SO_4(l)$.

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
$M(CaF_2(s)) = 78.08$ g/mol
$M(H_2SO_4(l) = 98.09$ g/mol
$M(HF(g)) = 20.01$ g/mol

$$n(CaF_2(s)) = \frac{15\ \text{g}}{78.08\ \frac{\text{g}}{\text{mol}}} = 0.192\ \text{mol}$$

$$n(H_2SO_4(l)) = \frac{20\ \text{g}}{98.09\ \frac{\text{g}}{\text{mol}}} = 0.204\ \text{mol}$$

Using $CaF_2(s)$:

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

Using the $H_2SO_4(l)$ reactant:

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

Since the $CaF_2(s)$ reactant produced a smaller amount of HF(g), $CaF_2(s)$ is the limiting reagent. Based on this, the expected mass of HF(g) that will be produced is
$m = n \times M(\text{HF(g)}) = 0.384\ \text{mol (HF(g))} \times 20.01\ \text{g/mol} = 7.7\ \text{g HF(g)}$

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
$Zn(s) + Cu(NO_3)_2(aq) \rightarrow Cu(s) + Zn(NO_3)_2(aq)$
$M(Zn(s)) = 65.39$ g/mol
$M(Cu(NO_3)_2(aq)) = 187.57$ g/mol

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

$$n(Cu(NO_3)_2(aq)) = \frac{50\,g}{187.57\,\frac{g}{mol}} = 0.267 \text{ mol}$$

Using $Cu(NO_3)_2(aq)$:

$$nZn(s) \text{ that react} = \frac{0.267\,mol\,Cu(NO_3)_2(aq) \times 1\,mol\,Zn(s)}{1\,mol\,Cu(NO_3)_2(aq)} = 0.267 \text{ mol Zn(s)}$$

Since this is a larger number of moles of Zn(s) than is given, there is an excess of $Cu(NO_3)_2(aq)$ available to react with the Zn(s). All of 15 g of Zn(s) will react with the $Cu(NO_3)_2(aq)$ solution. Alternatively, if the calculation was based upon the 15 g (0.229 mol) of Zn(s),

$$nCu(NO_3)_2(aq) \text{ that react} = \frac{0.229\ mol\,Zn(s) \times 1\ mol\,Cu(NO_3)_2(aq)}{1\,mol\,Zn(s)} = 0.229 \text{ mol } Cu(NO_3)_2(aq)$$

Since this is less than the amount of $Cu(NO_3)_2(aq)$ 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:

$$2C_3H_6(g) + 2NH_3(g) + 3O_2(g) \rightarrow 2C_3H_3N(g) + 6H_2O(g)$$

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
$M(C_3H_6(g)) = 42.09$ g/mol
$M(NH_3(g)) = 17.04$ g/mol
$M(C_3H_3N(g)) = 53.07$ 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 $C_3H_6(g)$:

$$nC_3H_3N(g)\ \text{produced} = \frac{23.7586\ \text{mol}\ C_3H_6(g) \times 1\ \text{mol}\ C_3H_6N(g)}{1\ \text{mol}\ C_3H_6(g)} = 23.7586\ \text{mol}\ C_3H_3N(g)$$

Using $NH_3(g)$:

$$nC_3H_3N(g)\ \text{produced} = \frac{35.2113\ \text{mol}\ NH_3(g) \times 2\ \text{mol}\ C_3H_3N(g)}{2\ \text{mol}\ NH_3(g)} = 35.2113\ \text{mol}\ C_3H_3N(g)$$

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 $Mg(NO_3)_2(aq)$ and 125.0 mL of a 1.2 mol/L solution of $Na_3PO_4(aq)$ were used in the reaction below?

$$3Mg(NO_3)_2(aq) + 2Na_3PO_4(aq) \rightarrow Mg_3(PO_4)_2(s) + 6Na(NO_3)(aq)$$

What is Required?
You must determine which reagent, $Mg(NO_3)_2(aq)$ or $Na_3PO_4(aq)$, 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 $Mg(NO_3)_2(aq)$ and 125.0 mL of 1.2 mol/L $Na_3PO_4(aq)$.

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 $Mg_3(PO_4)_2(s)$ that form from each reactant. The limiting reagent is the reactant that produces the lower amount of $Mg_3(PO_4)_2(s)$.

Act on Your Strategy
For $Mg(NO_3)_2(aq)$:

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

From the balanced equation:

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

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

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

For $Na_3PO_4(aq)$:

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

From the balanced equation:

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

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

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

$Mg(NO_3)_2$(aq) is the limiting reactant because it produces the lower amount of $Mg_3(PO_4)_2$(s).

Check Your Solution
From the information provided, the answer seems reasonable.

8.
Problem
What would be the expected mass of Mg_3PO_4(s) formed in the reaction described in question 7?

What is Required?
You must calculate the expected mass of Mg_3PO_4(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 $Mg(NO_3)_2$(aq).

Plan Your Strategy
Use the mole ratio in the balanced equation to calculate the number of moles of $Mg_3(PO_4)_2$(s) that form from 100.0 mL of 0.5 mol/L $Mg(NO_3)_2$(aq). Calculate the mass of $Mg_3(PO_4)_2$(s) that would be produced using the equation $m = n \times M$.

Act on Your Strategy
From question 7 we know that:

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

From the balanced equation:

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

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

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

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

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 $Na_2S(s)$ be sufficient to remove all the mercury(II) ions as HgS(s) from 200 ml of B5 solution (0.221 mol/L $HgCl_2(aq)$)? Explain your answer.

What is Required?
You must determine if a given volume of 0.085 mol/L $Na_2S(aq)$ can precipitate the $Hg^{2+}(aq)$ ions from a given volume of B5 solution having a concentration of 0.221 mol/L $HgCl_2(aq)$.

What is Given?

For $Na_2S(aq)$:
V = 600 mL = 0.600 L
c = 0.085 mol/L

For $HgCl_2(aq)$:
c = 0.221 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 $Na_2S(aq)$ has been provided.

Act on Your Strategy
$Na_2S(aq) + HgCl_2(aq) \rightarrow 2NaCl(aq) + HgS(s)$
Using $Na_2S(aq)$:
$n Na_2S(aq) = 0.085 \text{ mol/L} \times 0.600 \text{ L} = 0.0510 \text{ mol } Na_2S(aq)$

$$n HgS(s) \text{ produced} = 0.0510 \text{ mol } Na_2S(aq) \times \frac{1 \text{ mol } HgS(s)}{1 \text{ mol } Na_2S(aq)} = 0.0510 \text{ mol } HgS(s)$$

Using $HgCl_2(aq)$:
$n HgCl_2(aq) = 0.221 \text{ mol/L} \times 0.200 \text{ L} = 0.0442 \text{ mol } HgCl_2(aq)$

$$n HgS(s) \text{ produced} = 0.0442 \text{ mol } HgCl_2(aq) \times \frac{1 \text{ mol } HgS(s)}{1 \text{ mol } HgCl_2(aq)} = 0.0442 \text{ mol } HgS(s)$$

The $HgCl_2(aq)$ produces less HgS(s) and, therefore, is the limiting reagent. This means that the $HgCl_2(aq)$ in the B5 solution would be used up first. There is sufficient $Na_2S(aq)$ to precipitate all of the $Hg^{2+}(aq)$.

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, $Pb(NO_3)_2(aq)$, 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, $Pb(NO_3)_2(aq)$ or KI(aq), is limiting in the reaction and calculate the mass of product, $PbI_2(s)$.

What is Given?

For $Pb(NO_3)_2(aq)$
c = 0.400 mol/L
V = 250 ml = 0.250 L

For KI(aq)
c = 0.22 mol/L
V = 300 mL = 0.300 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 $PbI_2(s)$. Use the mole ratios from the balanced equation to calculate the number of moles of PbS(s) that form from each reactant. Convert the lower number of moles of $PbI_2(s)$ to grams using the formula $m = n \times M$.

Act on Your Strategy
$2\ NaI(aq) + Pb(NO_3)_2(aq) \rightarrow 2NaNO_3(aq) + PbI_2(s)$
Using NaI(aq):
nNaI(aq) = 0.22 mol/L × 0.300 L = 0.0660 mol NaI(aq)

$$nPbI_2(s) \text{ produced} = 0.0660 \text{ mol NaI(aq)} \times \frac{1\,\text{mol}\,PbI_2(s)}{2\,\text{mol}\,NaI(aq)} = 0.0330 \text{ mol } PbI_2(s)$$

Using $Pb(NO_3)_2(aq)$:
$nPb(NO_3)_2(aq)$ = 0.400 mol/L × 0.250 L = 0.100 mol $Pb(NO_3)_2(aq)$

$$nPbI_2(s) \text{ produced} = 0.100 \text{ mol } Pb(NO_3)_2(aq) \times \frac{1\,\text{mol}\,PbI_2(s)}{1\,\text{mol}\,Pb(NO_3)_2(aq)} = 0.100 \text{ mol } PbI_2(s)$$

Since NaI(aq) produced less $PbI_2(s)$, NaI(aq) is the limiting reagent. Based upon this, the mass of $PbI_2(s)$ produced = 0.0330 mol × 461.0 g/mol = 15 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, $CuSO_4(aq)$, with 29.7 mL of 0.270 mol/L aqueous strontium nitrate, $Sr(NO_3)_2(aq)$.
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?
a) You must write the balanced equation for the reaction between $CuSO_4(aq)$ and $Sr(NO_3)_2(aq)$.
b) You must determine which ion was in excess to determine the colour of the filtrate.
c) You must use the balanced equation to calculate the mass of precipitate that is produced in the reaction.
d) You must calculate the percentage yield of precipitate in the reaction.

What is Given?
a) The chemical formulas of the reactants are given.
b)

For $CuSO_4(aq)$	For $Sr(NO_3)_2(aq)$
$c = 0.320$ mol/L	$c = 0.270$ mol/L
$V = 25.0$ ml = 0.0250 L	$V = 29.7$ mL = 0.0297 L

c) Use the balanced equation from **a)**.
d) The experimental yield in the reaction is 1.432 g.

Plan Your Strategy
a) Write the balanced equation for this double replacement reaction.
b) Determine the limiting reagent in the reaction by calculating which reactant produces the lower amount of $SrSO_4(s)$. Use the mole ratios in the balanced equation to calculate the number of moles of $SrSO_4(s)$ that form from each reactant.
c) Convert the lower number of moles of $SrSO_4(s)$ to grams using the formula $m = n \times M$.
d) Calculate the percentage yield using the equation

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

Act on Your Strategy
a) $CuSO_4(aq) + Sr(NO_3)_2(aq) \rightarrow Cu(NO_3)_2(aq) + SrSO_4(s)$
The precipitate is strontium sulphate, $SrSO_4(s)$.
b) Using $CuSO_4(aq)$:
$n CuSO_4(aq) = 0.320 \text{ mol/L} \times 0.0250 \text{ L} = 0.00800 \text{ mol } CuSO_4(aq)$

$$n SrSO_4(s) \text{ produced} = 0.00800 \text{ mol } CuSO_4(aq) \times \frac{1 \text{ mol } SrSO_4(s)}{1 \text{ mol } CuSO_4(aq)} = 0.00800 \text{ mol } SrSO_4(s)$$

Using $Sr(NO_3)_2(aq)$:
$n\ Sr(NO_3)_2(aq) = 0.270 \text{ mol/L} \times 0.0297 \text{ L} = 0.008019 \text{ mol } Sr(NO_3)_2(aq)$

$$n SrSO_4(s) \text{ produced} = 0.008019 \text{ mol } Sr(NO_3)_2(aq) \times \frac{1 \text{ mol } SrSO_4(s)}{1 \text{ mol } Sr(NO_3)_2(aq)} = 0.008019 \text{ mol}$$

$SrSO_4(s)$

Since $CuSO_4(aq)$ produced less $SrSO_4(s)$, $CuSO_4(aq)$ is the limiting reagent. Since the $Sr(NO_3)_2(aq)$ is in excess, the filtrate will be red.
c) $M SrSO_4(s) = 183.69$ g/mol
Based upon $CuSO_4(aq)$ as the limiting reagent, the mass of $SrSO_4(s)$ produced is 0.008000 mol × 183.69 g/mol = 1.4695 g ≈ 1.47 g

d) Percentage yield = $\dfrac{\text{experimental yield}}{\text{predicted yield}} \times 100\% = \dfrac{1.432\text{ g}}{1.4695\text{ g}}$ =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 Na_2SiO_3(aq) to remove these ions for disposal as solid waste. Suppose 150 mL of 0.250 mol/L aqueous sodium silicate, Na_2SiO_3(aq), is mixed with 950 mL of a 0.035 mol/L silver nitrate solution, $AgNO_3$(aq).
a) Write the balanced chemical equation for this precipitation reaction.
b) Which reactant is the limiting reactant? Explain your answer.
c) Calculate the predicted yield of precipitate.
d) Are all the silver ions removed from the solution? Explain briefly.

What is Required?
a) You must write the balanced equation for the reaction between Na_2SiO_3(aq) and $AgNO_3$(aq).
b) You must determine which reagent is limiting in this reaction.
c) You must use the balanced equation to calculate the mass of precipitate (the predicted yield) that is produced in the reaction.
d) You must determine if all of the Ag^+(aq) is precipitated in the reaction.

What is Given?
a) The chemical formulas of the reactants are given.
b) For Na_2SiO_3(aq)
c = 0.250 mol/L
V = 150 ml = 0.150 L

For $AgNO_3$(aq)
c = 0.035 mol/L
V = 950 mL = 0.950 L

Plan Your Strategy
a) Write the balanced equation for this double replacement reaction.
b) Determine the limiting reagent in the reaction by calculating which reactant produces the lower amount of Ag_2SiO_3(s). Use the mole ratios in the balanced equation to calculate the number of moles of Ag_2SiO_3(s) that forms from each reactant.
c) Convert the lower number of moles of Ag_2SiO_3(s) to grams using the formula $m = n \times M$.

Act on Your Strategy
a) $2AgNO_3(aq) + Na_2SiO_3(aq) \rightarrow Ag_2SiO_3(s) + 2\ NaNO_3(aq)$
The precipitate is silver silicate, Ag_2SiO_3(s).
b) Using $AgNO_3$(aq):
$n$$AgNO_3$(aq) = 0.035 mol/L × 0.950 L = 0.0332 mol

$n\text{Ag}_2\text{SiO}_3\text{(s)}$ produced = 0.0332 mol $\text{AgNO}_3\text{(aq)}$ × $\dfrac{1\,\text{mol}\,\text{Ag}_2\text{SiO}_3\text{(s)}}{2\,\text{mol}\,\text{AgNO}_3\text{(aq)}}$ = 0.0166 mol $\text{Ag}_2\text{SiO}_3\text{(s)}$

Using $\text{Na}_2\text{SiO}_3\text{(aq)}$:
$n\text{Na}_2\text{SiO}_3\text{(aq)}$ = 0.250 mol/L × 0.150 L = 0.0375 mol $\text{Na}_2\text{SiO}_3\text{(aq)}$
$n\text{Ag}_2\text{SiO}_3\text{(s)}$ produced = 0.0375 mol $\text{Na}_2\text{SiO}_3\text{(aq)}$ × $\dfrac{1\,\text{mol}\,\text{Ag}_2\text{SiO}_3\text{(s)}}{1\,\text{mol}\,\text{Na}_2\text{SiO}_3\text{(aq)}}$ = 0.0375 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 g/mol
Based upon this,
$m\text{Ag}_2\text{SiO}_3\text{(s)}$ produced = 0.0166 mol × 291.83 g/mol = 4.9 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:

$$\text{Fe(s)} + \text{CuCl}_2\text{(aq)} \rightarrow \text{FeCl}_2\text{(aq)} + \text{Cu(s)}$$

a) What is the predicted yield if 10.0 g of Fe(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 Cu(s) is actually obtained?

What is Required?
a) Determine the predicted yield of Cu(s).
b) Determine the percentage yield of Cu(s).

What is Given?
a) The balanced equation is given. The amount of Fe(s) is 10.0 g.
b) The balanced equation is given. The actual yield of Cu(s) is 9.0 g.

Plan Your Strategy
a) Calculate the number of g of Fe(s) using the equation $n \times M = m$. Using the mole ratio from the balanced equation, calculate the predicted yield of Cu(s) if 10.0 g of Fe(s) is reacted.
b) Calculate the number of moles of Cu(s) obtained using the equation $n = \dfrac{m}{M}$. Calculate the percentage yield is 9.0 g of Cu(s) is actually obtained using

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

Act on Your Strategy
a) $M\text{Fe(s)}$ = 55.85 g/mol
$M\text{cu(s)}$ = 63.55 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 mol Cu(s) × 63.55 g/mol = 11.3786 g = 11.4 g Cu(s)

b) MCu(s) = 63.55 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 SO_2(s) is produced using 50 g of S_8(s).

$$S_8(s) + 8O_2(g) \rightarrow 8SO_2(s)$$

What is Required?

You must calculate the percentage yield for the reaction if 60 g of SO_2(s) is produced using 50 g of S_8(s).

What is Given?

The balanced equation is given.
60 g of SO_2(s) is produced using 50 g of S_8(s).

Plan Your Strategy

Calculate the number of moles of S_8(s) used with the equation $n = \frac{m}{M}$. Then using the molar ratios from the balanced equation, calculate the predicted yield of SO_2(s) using 50 g of S_8(s). Determine the number of moles of SO_2(s) that were actually produced. Use these values to determine the percentage yield with the equation percentage yield = $\frac{\text{experimental yield}}{\text{predicted yield}} \times 100\%$.

Act on Your Strategy
$MS_8(s) = 256.56$ g/mol
$MSO(s) = 64.07$ g/mol

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

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

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

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:

$$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$$

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 trail 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

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

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

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

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:

$$Cl_2(l) + 2NaOH(aq) \rightarrow NaOCl(aq) + NaCl(aq) + H_2O(l)$$

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,$Cl_2(l)$, 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 $Cl_2(l)$ and for NaOCl(aq). Use the equation $n = \frac{m}{M}$ to calculate the moles of $Cl_2(l)$.

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 $Cl_2(l)$. Calculate the mass of NaOCl(aq) using the equation $m = n \times M$.

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

Act on Your Strategy

0.150 kg = 150 g

$M Cl_2(l)$ = 70.90 g/mol

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

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

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

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

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

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, $C_6H_{12}O_6$(aq), to ethanol, C_2H_5OH(aq), and carbon dioxide, CO_2(g):

$$C_6H_{12}O_6(aq) \xrightarrow{\text{baker's yeast}} 2C_2H_5OH(aq) + 2CO_2(g)$$

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 kg = 1 000g and convert 1.63 kg to grams. Determine the molar mass, *M*, for $C_6H_{12}O_6$(aq) and for C_2H_5OH(aq). Use the equation $n = \dfrac{m}{M}$ to calculate the moles of $C_6H_{12}O_6$(aq). Use the mole ratio in the balanced equation to calculate the moles C_2H_5OH(aq) of that are predicted to form from this number of moles of $C_6H_{12}O_6$(aq). Calculate the mass of C_2H_5OH(aq) using the equation $m = n \times M$.

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

Act on Your Strategy

1.63 kg = 1630 g

$MC_6H_{12}O_6(aq) = 180.16$ g/mol

$MC_2H_5OH(aq) = 46.08$ g/mol

$$nC_6H_{12}O_6(aq) = \frac{1630\text{ g}}{180.16\frac{\text{g}}{\text{mol}}} = 9.048\text{ mol } C_6H_{12}O_6(aq)$$

$$nC_2H_5OH(aq)\text{ predicted} = \frac{9.048\text{ mol } C_6H_{12}O_6(aq) \times 2\text{ mol } C_2H_5OH(aq)}{1\text{ mol } C_6H_{12}O_6(aq)} = 18.10\text{ mol}$$

$C_2H_5OH(aq)$

$mC_2H_5OH(aq)$ predicted = 18.10 mol × 46.08 g/mol = 833.8 g $C_2H_5OH(aq)$

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

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, $Cu_2(CO_3)\cdot(OH_2)(s)$, is heated vigorously over a Bunsen flame for about 20 minutes, it is transformed into black copper(II) oxide:

$$Cu_2(CO_3)\cdot(OH_2)(s) \rightarrow 2CuO(s) + CO_2(g) + H_2O(g)$$

a) Calculate the predicted yield of CuO(s) from vigorous heating of 4.00 g of malachite.
b) If 2.80 g of CuO(s) remains after the $CO_2(g)$ and $H_2O(g)$ have been burned off, what is the percentage yield of the solid product?

What is Required?

You must calculate the predicted yield of CuO(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, $CuCO_3\cdot Cu(OH)_2(s)$, is 4.00 g.
b) The experimental yield of CuO(s) is 2.80 g.

Plan Your Strategy

a) Determine the molar mass, *M*, for $CuCO_3\cdot Cu(OH)_2(s)$ and of CuO(s). Use the equation n = $\frac{m}{M}$ to calculate the moles of malachite, $CuCO_3\cdot Cu(OH)_2(s)$. Use the mole ratio in the balanced equation to calculate the moles CuO(s) that are predicted to form from this number of moles of $CuCO_3\cdot Cu(OH)_2(s)$. Calculate the mass of CuO(s) using the equation $m = n \times M$.
b) Predict the percentage yield using percentage yield = $\frac{\text{experimental yield}}{\text{predicted yield}} \times 100\%$.

Act on Your Strategy

a) $MCuCO_3\cdot Cu(OH)_2(s)$ = 221.13 g/mol

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

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

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

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

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, $C_7H_8(l)$, with potassium permanganate, KMnO4(aq), proceeds with significantly less than 100% yield under most conditions:

$$C_7H_8(l) + 2KMnO_4(aq) \rightarrow KC_7H_5O_2(aq) + 2MnO_2(s) + KOH(aq) + 2H_2O(l)$$

a) If 8.60 g toluene reacts with excess potassium permanganate, what is the predicted yield, in grams, of potassium benzoate, $KC_7H_5O_2(aq)$?
b) If the percentage yield is 70.0%, what mass of potassium benzoate would you expect to be produced?
c) What mass of toluene is needed to produce 13.4 g of potassium benzoate if the percentage yield is 60.0%?

What is Required?
a) You must calculate a predicted yield of potassium benzoate that can be produced from a given mass of toluene.
b) You must determine the experimental yield, in grams, when the percentage yield is given.
c) 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?
a) The balanced equation is given and the mass of the reactant toluene is 8.60 g.
b) The percentage yield for the reaction in **a)** is given.
c) The experimental yield of potassium benzoate is 13.4 g and the percentage yield for the reaction is 60.0%.

Plan Your Strategy
a) Calculate the molar mass of toluene and of potassium benzoate.

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

Use the mole ratio from the balanced equation to calculate the number of moles of $KC_7H_5O_2(aq)$ that will be produced.
Calculate the predicted mass of $KC_7H_5O_2(aq)$ that will be produced using $m = n \times M$.

b) Calculate the experimental yield of $KC_7H_5O_2$(aq) using

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

c) Use the equation for experimental yield and the given percentage yield to calculate the predicted mass of $KC_7H_5O_2$(aq). Calculate the number of moles of $KC_7H_5O_2$(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 $C_7H_8(l)$ using the equation $n = m \times M$.

Act on Your Strategy

a) $MC_7H_8(l)$ = 92.15 g/mol

$MKC_7H_5O_2$(aq) = 160.22 g/mol

$nC_7H_8(l) = \dfrac{8.60\text{ g}}{92.15\dfrac{\text{g}}{\text{mol}}} = 0.09333$ mol

0.09333 mol $C_7H_8(l) \times \dfrac{1\text{ mol } KC_7H_5O_2(aq)}{1\text{ mol } C_7H_8([el])} = 0.09333$ mol $KC_7H_5O_2$(aq)

$mKC_7H_5O_2$(aq) = 0.09333 mol $KC_7H_5O_2$(aq) × 160.22 g/mol = 15.0 g

b) Percentage yield $KC_7H_5O_2$(aq) = $\dfrac{\text{experimental yield}}{\text{predicted yield}} \times 100\%$ = 70.0% × 15.0 g = 10.5 g

c) Predicted yield of $KC_7H_5O_2$(aq) = $\dfrac{\text{experimental yield}}{\text{percentage yield}} \times 100\% = \dfrac{13.4}{60.0} \times 100\% = 22.33$ g

$nKC_7H_5O_2(aq) = \dfrac{22.33\text{ g}}{160.22\dfrac{\text{g}}{\text{mol}}} = 0.1394$ mol

$nC_7H_8(l) = \dfrac{0.1394\text{ mol } KC_7H_5O_2(aq) \times 1\text{ mol } C_7H_8([el])}{1\text{ mol } KC_7H_8O_2(aq)}$ = 0.1394 mol $C_7H_8(l)$

$mC_7H_8(l)$ = 0.1394 mol × 92.15 g/mol = 12.8 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 NaOH(aq) reaches 7 after you add 36.2 mL of a 1.5 mol/L solution of acetic acid, CH_3COOH(aq), what is the concentration of the NaOH(aq) solution?

What is Required?

You must calculate the concentration of the NaOH(aq) solution.

What is Given?

V(NaOH(aq)) = 50 mL = 0.050 L

c(CH_3COOH(aq)) = 1.50 mol/L

$V(CH_3COOH(aq))$ = 36.2 mL = 0.0362 L

Plan Your Strategy
Write the balanced chemical equation for the reaction.
Calculate the moles of $CH_3COOH(aq)$ using the equation $n = c \times V$.
Use the mole ratio from the balanced equation to determine the moles of NaOH(aq) that react.
Calculate the concentration of NaOH(aq) using the equation $c = \frac{n}{V}$.

Act on Your Strategy
$NaOH(aq) + CH_3COOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$
$n\mathrm{CH_3COOH(aq)} = c \times V = 1.50\ \mathrm{mol/L} \times 0.0362\ \mathrm{L} = 0.0543\ \mathrm{mol}$
From the balanced equation:

$$\frac{n\mathrm{CH3COOH(aq)}}{n\mathrm{NaOH(aq)}} = \frac{1}{1}$$

$$n\mathrm{NaOH(aq)} = \frac{0.0543\ \mathrm{mol\ CH3COOH(aq)}}{1} = 0.0543\ \mathrm{mol\ NaOH(aq)}$$

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

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 HCl(aq) with 0.150 mol/L NaOH(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 HCl(aq)?

What is Required?
You must calculate the concentration of the hydrochloric acid solution.

What is Given?
$V(\mathrm{HCl_{initial}})$ = 1.50 mL
$\mathrm{V(HCl_{final})}$ = 19.35 mL
c(NaOH(aq)) = 0.150 mol/L
V(NaOH(aq)) = 25.00 mL = 0.02500 L

Plan Your Strategy
Write the balanced equation for the reaction.
Calculate the volume of HCl(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 HCl(aq) that react.
Calculate the concentration of HCl(aq) using the equation $c = \frac{n}{V}$.

Act on Your Strategy
VHCl(aq) used = $V_{final} - V_{initial}$ = 19.35 mL – 1.50 mL = 17.85 mL = 0.01785 L
The balanced equation for the reaction is
$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$
nNaOH(aq) = $c \times V$ = 0.150 mol/L × 0.02500 L = 0.003750 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_3(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_3(aq)?

What is Required?
You must calculate the concentration of the nitric acid solution.

What is Given?
$V(HNO_{3initial})$ = 0.05 mL
$V(HNO_{3final})$ = 38.00 mL
c(NaOH(aq)) = 0.2030 mol/L
V(NaOH(aq)) = 40.00 mL = 0.04000 L

Plan Your Strategy
Write the balanced equation for the reaction.
Calculate the volume of HNO_3(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_3(aq) that react. Calculate the concentration of HNO_3(aq) using $c = \frac{n}{V}$.

Act on Your Strategy
VHNO$_3$(aq) = $V_{final} - V_{initial}$ = 38.00 mL – 0.050 mL = 37.95 mL = 0.03795 L
The balanced equation for the reaction is
$HNO_3(aq) + NaOH(aq) \rightarrow NaNO_3(aq) + H_2O(l)$
nNaOH(aq) = $c \times V$ = 0.2030 mol/L × 0.04000 L = 0.008120 mol

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

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

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_3(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(HNO_3(aq)) = 0.5 mol/L
Trial 1: Final burette volume = 5.35 mL
Trial 2: Final burette volume = 5.42 mL
Trial 3: Final burette volume = 6.58 mL
Trial 1: Initial burette volume = 0.15 mL
Trial 2: Initial burette volume = 0.20 mL
Trial 3: Initial burette volume = 1.43 mL
Trial 1: V(KOH(aq)) = 10.0 mL
Trial 2: V(KOH(aq)) = 10.0 mL
Trial 3: V(KOH(aq)) = 10.0 mL

Plan Your Strategy

Write the balanced equation for the reaction.
Calculate the volume of HNO_3(aq) used in the titration.
Calculate the moles of HNO_3(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 $c = \dfrac{n}{V}$.

Act on Your Strategy

$KOH(aq) + HNO_3(aq) \rightarrow KNO_3(aq) + HO(l)$

$V HNO_3(aq)$ used in the titration:

Trial 1: 5.35 mL – 0.15 mL = 5.20 mL
Trial 2: 5.42 mL – 0.20 mL = 5.22 mL
Trial 3: 6.58 mL – 1.43 mL = 5.15 mL

$n HNO_3(aq) = c \times V$ for each trial in the titration:

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

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

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

$n KOH(aq)$ that react for each trial in the titration, from the balanced equation $\dfrac{1 \text{ mol } HNO_3(aq)}{1 \text{ mol } KOH(aq)}$:

Trial 1: 0.00260 mol $HNO_3(aq)$ = 0.00260 mol KOH(aq)
Trial 2: 0.00261 mol $HNO_3(aq)$ = 0.00261 mol KOH(aq)
Trial 3: 0.00258 mol $HNO_3(aq)$ = 0.00258 mol KOH(aq)

$c KOH(aq) = \dfrac{n}{V}$ for each trial in the titration:

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

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

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

Average $c(KOH(aq))$ 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 HCl(aq) using 0.1500 mol/L 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 $V = \dfrac{n}{c}$.)

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
cNaOH(aq) = 0.1500 mol/L
VNaOH(aq) = 25.00 mL = 0.02500 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
$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

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

nNaOH(aq) needed = $c \times V$ = 0.1500 mol/L × 0.02500 L = 0.00375 mol

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

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

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:
$CaCO_3(s) + 2\ HCl(aq) \rightarrow CO_2(g) + H_2O(l) + CaCl_2(aq)$
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.
$V(\text{HCl(aq)}) = 125 \text{ mL} = 0.125 \text{ L}$
$m(\text{CaCO}_3\text{(s)}) = 3.28 \text{ g}$

Plan Your Strategy
Determine the molar mass, M, of CaCO_3(s) and calculate the number of moles of CaCO_3(s) using the equation $n = \frac{m}{M}$.
Use the mole ratio in the balanced equation to calculate the number of moles of HCl(aq).
Calculate the volume of HCl(aq) using the formula $V = \frac{n}{c}$.

Act on Your Strategy

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

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

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

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

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).