

Combining aqueous solutions to produce a precipitate and making quantitative predictions.

Assume the reaction goes to completion when 35.0 ml of a 0.27 M solution of copper(II) nitrate is combined with 25.0 ml of 0.36 M sodium hydroxide solution. Assuming the volumes are additive, we can write an overall reaction and then calculate several items:

- (A) calculate the mass of the precipitate that should form.  
 (B) If 0.38 g of precipitate did form in the laboratory, calculate the % yield.  
 (C) Determine the concentration (molarity) of each ion still in the solution.



Change both reactants to millimole quantities

$$0.27 \text{ M} \times 35 \text{ ml} = \mathbf{9.45 \text{ millimoles Cu}(\text{NO}_3)_2} \quad 0.36 \text{ M} \times 25 \text{ ml} = \mathbf{9 \text{ millimoles of NaOH}}$$

(A) Use the coefficients to determine which limits:  $\frac{9.45 \text{ mmol Cu}(\text{NO}_3)_2}{1} > \frac{9 \text{ mmol NaOH}}{2}$

Thus the **sodium hydroxide limits** then calculate the mass of ppt that can form

$$0.009 \text{ mol NaOH} \times \frac{1 \text{ mol Cu}(\text{OH})_2}{2 \text{ NaOH}} \times \frac{97.57 \text{ g}}{1 \text{ mol}} = 0.439 \text{ g Cu}(\text{OH})_2$$

(B) Determine the % yield  $\frac{0.38 \text{ g Experimental}}{0.439 \text{ g Theoretical}} \times 100 = 86.6\% \text{ yield}$

(C) To determine the amount of ions left in solution,

First, realize that we will assume that there will be **no hydroxide ions left in solution** since they are the limiting reactant.

Then determine the amount of **spectator ions** that will be floating in the solution

$$9.45 \text{ mmol Cu}(\text{NO}_3)_2 \times \frac{2 \text{ NO}_3^-}{1 \text{ Cu}(\text{NO}_3)_2} = 18.9 \text{ mmol NO}_3^- \quad \frac{18.9 \text{ mmol NO}_3^-}{60 \text{ ml}} = 0.315 \text{ M NO}_3^-$$

$$9 \text{ mmol NaOH} \times \frac{\text{OH}^-}{1 \text{ NaOH}} = 9 \text{ mmol OH}^- \quad \frac{9 \text{ mmol OH}^-}{60 \text{ ml}} = 0.15 \text{ M OH}^-$$

Then determine the moles of copper ions that are needed to go with the limiting hydroxide ions by “doing stoichiometry”

$$9 \text{ mmol NaOH} \times \frac{1 \text{ OH}^-}{1 \text{ NaOH}} \times \frac{1 \text{ Cu}^{2+}}{2 \text{ OH}^-} = 4.5 \text{ mmol Cu}^{2+} \text{ needed for reaction}$$

then calculate the amount of copper ion that will be left over

Thus 9.45 mmol  $\text{Cu}^{2+}$  started with – 4.5 moles of  $\text{Cu}^{2+}$  needed = 4.95 mmol  $\text{Cu}^{2+}$  left over and then calculate molarity

$$\frac{4.95 \text{ mmol Cu}^{2+}}{60 \text{ ml (total)}} = 0.0825 \text{ M Cu}^{2+} \text{ Left Over}$$

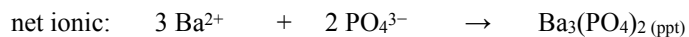
## P B1      Solution Stoichiometry Practice Problem

Name \_\_\_\_\_ Per \_\_\_\_\_

*Show your work clearly in the space below.*

Assume the reaction goes to completion, and the volumes are additive, when 45.0 ml of a 0.54 M solution of barium nitrate is combined with 35.0 ml of 0.54 M sodium phosphate solution,

- (A) calculate the mass of the precipitate that should form.
- (B) If 3.32 g of precipitate did form in the laboratory, calculate the % yield.
- (C) Determine the concentration (molarity) of each ion still floating in the solution.



*change both reactants to mole quantities*

$$(45 \text{ ml})(0.54 \text{ M}) = 24.3 \text{ mmol Ba}(\text{NO}_3)_2$$

$$(35 \text{ ml})(0.54 \text{ M}) = 18.9 \text{ mmol of Na}_3\text{PO}_4$$

a. *determine that the barium nitrate limits then calculate the mass of ppt that can form*

$$24.3 \text{ mmol Ba}(\text{NO}_3)_2 \times \left( \frac{1 \text{ Ba}_3(\text{PO}_4)_2}{3 \text{ Ba}(\text{NO}_3)_2} \right) = 8.1 \text{ mmol of Ba}_3(\text{PO}_4)_2 (\text{ppt}) \text{ can be formed} \times \left( \frac{602 \text{ g}}{1 \text{ mol}} \right) = \mathbf{4,880 \text{ mg (4.88 g) of ppt}}$$

b. *determine the % yield*

$$\left( \frac{3.32 \text{ g Experimental}}{4.88 \text{ g Theoretical}} \right) \times 100 = 68.1\%$$

c. To determine the amount of ions left in solution,

first, realize that we will assume there will be **no barium ions left in solution** since they are the limiting reactant.

Then determine the amount of **spectator ions** that will be floating in the solution

$$24.3 \text{ mmol Ba(NO}_3)_2 \times \left( \frac{2 \text{ NO}_3^-}{1 \text{ Ba(NO}_3)_2} \right) = 48.6 \text{ mmol NO}_3^-, \text{ next calculate molarity } \left( \frac{48.6 \text{ mmol NO}_3^-}{80 \text{ ml (total Vol)}} \right) = \mathbf{0.61 \text{ M NO}_3^-}$$

$$18.9 \text{ mmol Na}_3\text{PO}_4 \times \left( \frac{3 \text{ Na}^+}{1 \text{ Na}_3\text{PO}_4} \right) = 56.7 \text{ mmol Na}^+ \text{ ions, next calculate molarity } \left( \frac{56.7 \text{ mmol Na}^+}{80 \text{ ml (total Vol)}} \right) = \mathbf{0.71 \text{ M Na}^+}$$

Then determine the moles of phosphate that are needed to go with the limiting reactant by “doing stoichiometry”

$$24.3 \text{ mmol Ba(NO}_3)_2 \times \left( \frac{2 \text{ Na}_3\text{PO}_4}{3 \text{ Ba(NO}_3)_2} \right) = 16.2 \text{ mmol of Na}_3\text{PO}_4 = 16.2 \text{ mmol PO}_4^{3-} \text{ required to go with all the Ba}^{2+}$$

then calculate the amount of phosphate ion that will be left over (Phosphate is the excess reactant so “some” of it will remain.)

$$(18.9 \text{ mmol of PO}_4^{3-} \text{ started with}) - (16.2 \text{ mmol of PO}_4^{3-} \text{ needed}) = 2.7 \text{ mmol PO}_4^{3-} \text{ left over}$$

$$\text{and then calculate the molarity } \left( \frac{2.7 \text{ mmol PO}_4^{3-}}{80 \text{ ml (total Vol)}} \right) = \mathbf{0.034 \text{ M PO}_4^{3-}}$$



change both reactants to mole quantities

$$(0.040 \text{ L})(0.87 \text{ M}) = 0.0348 \text{ moles AgNO}_3 \quad (0.055 \text{ L})(0.57 \text{ M}) = 0.03135 \text{ moles of K}_2\text{CrO}_4$$

a. determine that the silver nitrate limits then calculate the mass of ppt that can form

$$0.0348 \text{ mol AgNO}_3 \times \left( \frac{1 \text{ Ag}_2\text{CrO}_4}{2 \text{ AgNO}_3} \right) = 0.0174 \text{ moles of Ag}_2\text{CrO}_{4(\text{ppt})} \text{ can be formed} \times \left( \frac{331.8 \text{ g}}{1 \text{ mol}} \right) = \mathbf{5.77 \text{ g of ppt}}$$

b. determine the % yield

$$\left( \frac{4.96 \text{ g Experimental}}{5.77 \text{ g Theoretical}} \right) \times 100 = 86.0\%$$

c. To determine the amount of ions left in solution,

first, realize that there will be **no silver ions left in solution** since they are the limiting reactant.

Then determine the amount of **spectator ions** that will be floating in the solution

$$0.0348 \text{ mol AgNO}_3 \times \left( \frac{1 \text{ NO}_3^-}{1 \text{ AgNO}_3} \right) = 0.0348 \text{ moles NO}_3^-, \text{ next calculate molarity } \left( \frac{0.0348 \text{ moles NO}_3^-}{0.095 \text{ L (total Vol)}} \right) = \mathbf{0.37 \text{ M NO}_3^-}$$

$$0.03135 \text{ mol K}_2\text{CrO}_4 \times \left( \frac{2 \text{ K}^+}{1 \text{ K}_2\text{CrO}_4} \right) = 0.0627 \text{ moles K}^+ \text{ ions, next calculate molarity } \left( \frac{0.0627 \text{ K}^+}{0.095 \text{ L (total Vol)}} \right) = \mathbf{0.66 \text{ M K}^+}$$

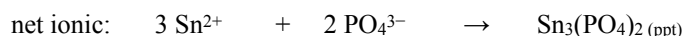
Then determine the moles of chromate that are needed to go with the limiting reactant by “doing stoichiometry”

$$0.0348 \text{ mol AgNO}_3 \times \left( \frac{1 \text{ K}_2\text{CrO}_4}{2 \text{ AgNO}_3} \right) = 0.0174 \text{ moles of CrO}_4^{2-} \text{ required to go with all the Ag}^+$$

then calculate the amount of chromate ion that will be left over (Chromate is the excess reactant so “some” of it will remain.)

(0.03135 moles of  $\text{CrO}_4^{2-}$  started with) – (0.0174 moles of  $\text{CrO}_4^{2-}$  needed) = 0.01395 moles  $\text{CrO}_4^{2-}$  left over and then

$$\text{calculate the molarity } \left( \frac{0.01395 \text{ mol CrO}_4^{2-}}{0.095 \text{ L (total Vol)}} \right) = \mathbf{0.15 \text{ M CrO}_4^{2-}}$$



change both reactants to mole quantities

$$(35 \text{ ml})(0.25 \text{ M}) = \mathbf{8.75 \text{ mmol Sn(NO}_3)_2} \quad (45 \text{ ml})(0.30 \text{ M}) = \mathbf{13.5 \text{ mmol of Na}_3\text{PO}_4}$$

a. determine that the barium nitrate limits then calculate the mass of ppt that can form

$$8.75 \text{ mmol Sn(NO}_3)_2 \times \left( \frac{1 \text{ Ba}_3(\text{PO}_4)_2}{3 \text{ Sn(NO}_3)_2} \right) = 2.91 \text{ mmol of Sn}_3(\text{PO}_4)_2 \text{ (ppt) can be formed} \times \left( \frac{546 \text{ g}}{1 \text{ mol}} \right) = \mathbf{1.59 \text{ g of ppt}}$$

b. determine the % yield

$$\left( \frac{1.00 \text{ g Experimental}}{1.59 \text{ g Theoretical}} \right) \times 100 = \mathbf{62.9\%}$$

c. To determine the amount of ions left in solution,

first, realize that we will assume there will be **no barium ions left in solution** since they are the limiting reactant.

Then determine the amount of **spectator ions** that will be floating in the solution

$$8.75 \text{ mmol Sn(NO}_3)_2 \times \left( \frac{2 \text{ NO}_3^-}{1 \text{ Sn(NO}_3)_2} \right) = 17.5 \text{ moles NO}_3^-, \text{ next calculate molarity } \left( \frac{17.5 \text{ moles NO}_3^-}{80 \text{ mL (total Vol)}} \right) = \mathbf{0.219 \text{ M NO}_3^-}$$

$$13.5 \text{ mol K}_3\text{PO}_4 \times \left( \frac{3 \text{ Na}^+}{1 \text{ Na}_3\text{PO}_4} \right) = 40.5 \text{ moles Na}^+ \text{ ions, next calculate molarity } \left( \frac{40.5 \text{ Na}^+}{80 \text{ mL (total Vol)}} \right) = \mathbf{0.506 \text{ M Na}^+}$$

Then determine the moles of phosphate that are needed to go with the limiting reactant by “doing stoichiometry”

$$13.5 \text{ mmol Sn(NO}_3)_2 \times \left( \frac{2 \text{ K}_3\text{PO}_4}{3 \text{ Sn(NO}_3)_2} \right) = 9.0 \text{ mmol of K}_3\text{PO}_4 = 9.0 \text{ mmol of PO}_4^{3-} \text{ required to go with all the Sn}^{2+}$$

then calculate the amount of phosphate ion that will be left over (Phosphate is the excess reactant so “some” of it will remain.)

(13.5 mmoles of  $\text{PO}_4^{3-}$  started with) – (9.0 mmoles of  $\text{PO}_4^{3-}$  needed) = 4.5 mmol  $\text{PO}_4^{3-}$  left over and then calculate the

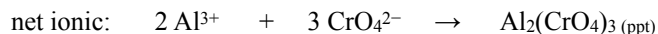
$$\text{molarity } \left( \frac{4.5 \text{ mmol PO}_4^{3-}}{80 \text{ mL (total Vol)}} \right) = \mathbf{0.056 \text{ M PO}_4^{3-}}$$



*the mass of precipitate will determine the original molarity of the potassium chromate solution*

$$8.11 \text{ g} \left( \frac{1 \text{ mol}}{118.9 \text{ g}} \right) = 0.0682 \text{ mol of ppt which is the same mol of } \text{K}_2\text{CO}_3 \text{ because of the 1:1 ratio}$$

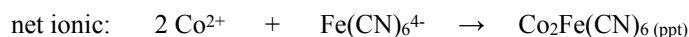
$$\left( \frac{0.0682 \text{ mol } \text{K}_2\text{CO}_3}{0.035 \text{ L (total Vol)}} \right) = \mathbf{1.95 \text{ M of } \text{K}_2\text{CO}_3 \text{ solution}}$$



*the mass of precipitate will determine the original molarity of the potassium chromate solution*

$$1.54 \text{ g} \left( \frac{1 \text{ mol}}{402 \text{ g}} \right) = 0.00383 \text{ mol ppt} \times \left( \frac{3 \text{K}_2\text{CrO}_4}{1 \text{Al}_2(\text{CrO}_4)_3} \right) = 0.0115 \text{ mol of } \text{K}_2\text{CrO}_4$$

$$\left( \frac{0.0115 \text{ mol } \text{K}_2\text{CrO}_4}{0.015 \text{ L (total Vol)}} \right) = \mathbf{0.77 \text{ M of } \text{K}_2\text{CrO}_4 \text{ solution}}$$



*the mass of precipitate will determine the original volume of the cobalt(II) chloride solution*

$$4.37 \text{ g} \left( \frac{1 \text{ mol}}{331 \text{ g}} \right) = 0.0132 \text{ mol ppt} \times \left( \frac{2 \text{CoCl}_2}{1 \text{Co}_2\text{Fe}(\text{CN})_6} \right) = 0.0264 \text{ mol of } \text{CoCl}_2$$

$$\left( \frac{0.0264 \text{ mol } \text{CoCl}_2}{0.05 \text{ M}} \right) = \mathbf{0.0528 \text{ L of } \text{CoCl}_2 \text{ solution, which is 52.8 ml } \text{CoCl}_2 \text{ solution}}$$



*the mass of precipitate will determine the original volume of the copper(II) sulfate solution required*

$$12.3 \text{ g} \left( \frac{1 \text{ mol}}{179.55 \text{ g}} \right) = 0.0685 \text{ mol of ppt of } \text{CuCrO}_4 \times \left( \frac{1 \text{CuSO}_4}{1 \text{CuCrO}_4} \right) = 0.0685 \text{ mol of } \text{CuSO}_4$$

$$\left( \frac{0.0685 \text{ mol } \text{CuSO}_4}{0.64 \text{ M}} \right) = \mathbf{0.107 \text{ L of } \text{CuSO}_4 \text{ solution, which is 107 ml } \text{CuSO}_4 \text{ solution}}$$



*the mass of precipitate will determine the original volume of the cobalt(II) chloride solution*

$$3.86 \text{ g} \left( \frac{1 \text{ mol}}{175 \text{ g}} \right) = 0.022 \text{ moles of } \text{CoCrO}_4 \times \left( \frac{1 \text{K}_2\text{CrO}_4}{1 \text{CoCrO}_4} \right) = 0.022 \text{ moles of } \text{K}_2\text{CrO}_4$$

$$\left( \frac{0.022 \text{ mol } \text{K}_2\text{CrO}_4}{0.35 \text{ M}} \right) = \mathbf{0.0630 \text{ L of } \text{K}_2\text{CrO}_4 \text{ solution, which is 63.0 ml } \text{K}_2\text{CrO}_4 \text{ solution}}$$