

Write out a balanced equation that represents the combination of the following aqueous salt, and then complete the stoichiometry problem. Write both the molecular equation and the net ionic equation.

1. If 25.0 ml of a 0.34 M solution of lead(II) nitrate is combined with 25.0 ml of 0.34 M sodium iodide solution,
  - a. calculate the mass of the precipitate that should form.
  - b. If 1.50 g of precipitate did form in the laboratory, calculate the % yield.
  - c. Determine the concentration (molarity) of each ion still floating in the solution.
2. If 25.0 ml of a 0.34 M solution of aluminum chloride is combined with 25.0 ml of 0.34 M copper (II) sulfate solution, calculate the mass of the precipitate that should form.
3. If 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.
4. If 40.0 ml of 0.87 M solution of silver nitrate is combined with 55.0 ml of 0.57 M of potassium chromate,
  - a. calculate the mass of precipitate that should form.
  - b. If 4.96 g of the precipitate forms in the laboratory, calculate the % yield of ppt formed.
  - c. Determine the concentration (molarity) of each ion still floating in the solution.
5. If an excess of cobalt (II) chloride is combined with 35.0 ml of a potassium carbonate solution, and 8.11 g of the precipitate forms, determine the original molarity of the potassium carbonate solution.
6. If 15.0 ml of potassium chromate is combined with an excess of a aluminum chloride solution, and 1.54 g of the precipitate forms, determine the original molarity of the potassium chromate solution.
7. If a volume of 0.50 M cobalt (II) chloride is combined with an excess of potassium ferrocyanide solution, and 4.37 g of the precipitate forms, determine the original volume of the cobalt (II) chloride solution.  
NOTE: ferrocyanide ion is  $\text{Fe}(\text{CN})_6^{4-}$
8. If a 0.64 M solution of copper (II) sulfate solution is combined with an excess of potassium chromate, and 12.3 g of the precipitate forms, determine the original volume of the copper (II) sulfate solution used.
9. If an excess of cobalt (II) chloride is combined with a 0.35 M solution of potassium chromate, and 3.86 g of the precipitate forms, determine the original volume of the potassium chromate solution used.



First, be sure you are working in moles – change both reactants to mole quantities

$$(0.025 \text{ L})(0.34 \text{ M}) = 0.0085 \text{ moles Pb}(\text{NO}_3)_2 \quad (0.025 \text{ L})(0.34 \text{ M}) = 0.0085 \text{ moles of NaI}$$

a. determine that the sodium iodide limits then calculate the mass of ppt that can form

$$0.0085 \text{ mol NaI} \times \left( \frac{1 \text{ PbI}_2}{2 \text{ NaI}} \right) = 0.00425 \text{ moles of PbI}_2(\text{ppt}) \text{ can be formed} \times \left( \frac{461 \text{ g}}{1 \text{ mol}} \right) = 1.96 \text{ g of ppt theor}$$

b. then determine the % yield

$$\left( \frac{1.50 \text{ g Experimental}}{1.96 \text{ g Theoretical}} \right) \times 100 = 76.6\%$$

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

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

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

$$0.0085 \text{ mol Pb}(\text{NO}_3)_2 \times \left( \frac{2 \text{ NO}_3^-}{1 \text{ Pb}(\text{NO}_3)_2} \right) = 0.017 \text{ moles NO}_3^-, \text{ next calculate molarity } \left( \frac{0.017 \text{ moles NO}_3^-}{0.05 \text{ L (total Vol)}} \right) = 0.34 \text{ M NO}_3^-$$

$$0.0085 \text{ mol NaI} \times \left( \frac{1 \text{ Na}^+}{1 \text{ NaI}} \right) = 0.0085 \text{ moles Na}^+ \text{ ions, next calculate molarity } \left( \frac{0.0085 \text{ Na}^+}{0.05 \text{ L (total Vol)}} \right) = 0.17 \text{ M Na}^+$$

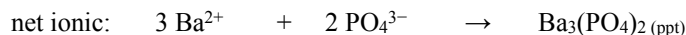
Then determine the moles of lead that are needed to go with the limiting reactant by “doing stoichiometry”

$$0.0085 \text{ mol NaI} \times \left( \frac{1 \text{ Pb}(\text{NO}_3)_2}{2 \text{ NaI}} \right) = 0.00425 \text{ moles of Pb}(\text{NO}_3)_2 \text{ required to go with all the I}^- \text{ ions}$$

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

Thus  $(0.0085 \text{ moles of Pb}^{2+} \text{ started with}) - (0.00425 \text{ moles of Pb}^{2+} \text{ needed}) = 0.00425 \text{ moles Pb}^{2+} \text{ left over}$  and then calculate molarity  $\left( \frac{0.00425 \text{ Pb}^{2+}}{0.05 \text{ L (total Vol)}} \right) = 0.085 \text{ M Pb}^{2+} \text{ ions remaining in solution.}$

2. There is no ppt because both the products are soluble therefore no reaction occurs, so there is no need to write a reaction.



change both reactants to mole quantities

$$(0.045 \text{ L})(0.54 \text{ M}) = 0.0243 \text{ moles Ba}(\text{NO}_3)_2$$

$$(0.035 \text{ L})(0.54 \text{ M}) = 0.0189 \text{ moles of Na}_3\text{PO}_4$$

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

$$0.0243 \text{ mol Ba}(\text{NO}_3)_2 \times \left( \frac{1 \text{ Ba}_3(\text{PO}_4)_2}{3 \text{ Ba}(\text{NO}_3)_2} \right) = 0.0081 \text{ moles of Ba}_3(\text{PO}_4)_2 (\text{ppt}) \text{ can be formed} \times \left( \frac{602 \text{ g}}{1 \text{ mol}} \right) = 4.88 \text{ g of ppt}$$

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 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

$$0.0243 \text{ mol Ba}(\text{NO}_3)_2 \times \left( \frac{2 \text{ NO}_3^-}{1 \text{ Ba}(\text{NO}_3)_2} \right) = 0.0486 \text{ moles NO}_3^-, \text{ next calculate molarity } \left( \frac{0.0486 \text{ moles NO}_3^-}{0.08 \text{ L (total Vol)}} \right) = 0.61 \text{ M NO}_3^-$$

$$0.0189 \text{ mol Na}_3\text{PO}_4 \times \left( \frac{3 \text{ Na}^+}{1 \text{ Na}_3\text{PO}_4} \right) = 0.0567 \text{ moles Na}^+ \text{ ions, next calculate molarity } \left( \frac{0.0567 \text{ Na}^+}{0.08 \text{ L (total Vol)}} \right) = 0.71 \text{ M Na}^+$$

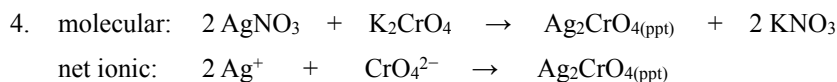
Then determine the moles of phosphate that are needed to go with the limiting reactant by “doing stoichiometry”

$$0.0243 \text{ mol Ba}(\text{NO}_3)_2 \times \left( \frac{2 \text{ Na}_3\text{PO}_4}{3 \text{ Ba}(\text{NO}_3)_2} \right) = 0.0162 \text{ moles of PO}_4^{3-} \text{ required to go with all the Ba}^{2+}$$

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

$$(0.0189 \text{ moles of PO}_4^{3-} \text{ started with}) - (0.0162 \text{ moles of PO}_4^{3-} \text{ needed}) = 0.0027 \text{ mol PO}_4^{3-} \text{ left over and then calculate the}$$

$$\text{molarity } \left( \frac{0.0027 \text{ mol PO}_4^{3-}}{0.08 \text{ L (total Vol)}} \right) = 0.034 \text{ M PO}_4^{3-}$$



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

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

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

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

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

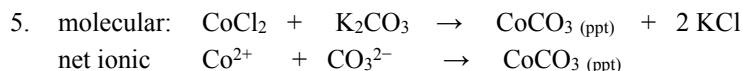
Then determine the moles of chromate that are needed to go with the limiting reactant by “doing stoichiometry”

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

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

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

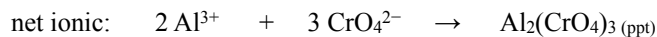
$$\text{calculate the molarity} \left( \frac{0.01395 \text{ mol CrO}_4^{2-}}{0.095 \text{ L (total Vol)}} \right) = 0.15 \text{ M CrO}_4^{2-}$$



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

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

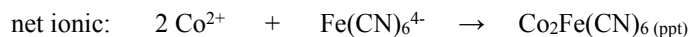
$$\left( \frac{0.0682 \text{ mol K}_2\text{CO}_3}{0.035 \text{ L (total Vol)}} \right) = 1.95 \text{ M of K}_2\text{CO}_3 \text{ solution}$$



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

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

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



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

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

$$\left( \frac{0.0264 \text{ mol CoCl}_2}{0.05 \text{ M}} \right) = 0.528 \text{ L of CoCl}_2 \text{ solution, which is 52.8 ml CoCl}_2 \text{ solution}$$



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

$$12.3 \text{ g} \left( \frac{1 \text{ mol}}{179.5 \text{ g}} \right) = 0.0685 \text{ moles ppt of CuCrO}_4 \times \left( \frac{1 \text{K}_2\text{CrO}_4}{1 \text{CuSO}_4} \right) = 0.0685 \text{ moles of CuSO}_4$$

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



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

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

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