Precipitation Stoichiometry

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www.mioy.org/chem161

Example

Say we mix an aqueous solution of nickel(II) chloride with an aqueous solution of sodium sulfide.

We can write the reactants of this chemical reaction:

$NiCl_2$ (*aq*) + Na_2S (*aq*) →

We can predict the products by exchanging the ions (we call these double exchange reactions):

$NiCl_2$ (*aq*) + Na_2S (*aq*) → NiS (*s*) + NaCl (*aq*)

Determine the products first.

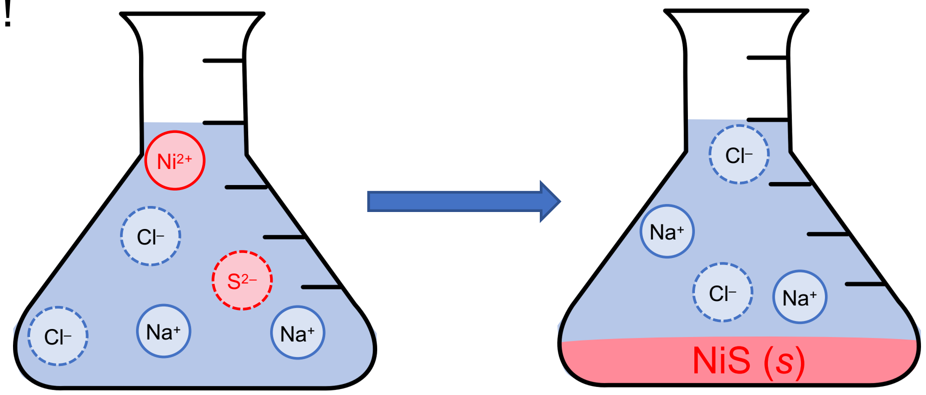
Then determine if each new product is soluble (*aq*) or insoluble (*s*)!

Don't forget to balance the chemical reaction!

$\underline{1}NiCl_2$ (*aq*) + $\underline{1}Na_2S$ (*aq*) → $\underline{1}NiS$ (*s*) + $\underline{2}NaCl$ (*aq*)

Always go back to the microscopic picture of what's happening!

→ NiS precipitate (solid) forms & other ions float in solution



Writing Chemical Equations

This is called a molecular equation because we keep the salt as neutral "molecules":

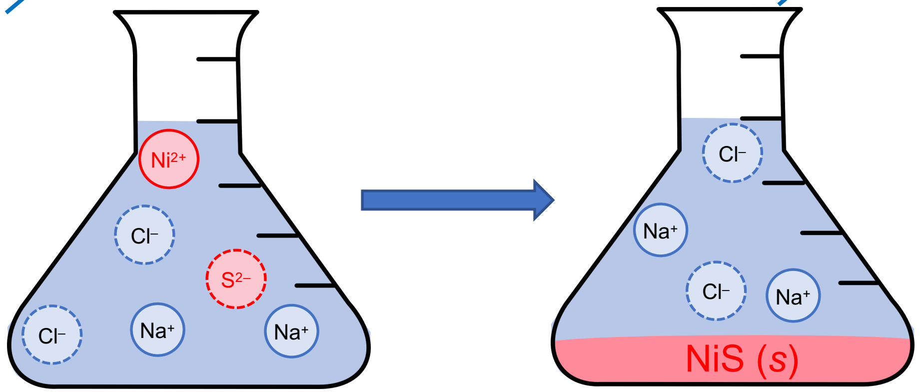
$$\underline{1}NiCl_2\ (aq) + \underline{1}Na_2S\ (aq) \rightarrow \underline{1}NiS\ (s) + \underline{2}NaCl\ (aq)$$

But we know that soluble salts (aqueous solutions) exist as dissociated ions!

It would be more accurate to dissociate the soluble salts.

We can transform the molecular equation into an complete ionic equation:

$$\underline{1}Ni^{+}\ (aq) + \underline{2}\cancel{Cl^{-}}\ (aq) + \underline{2}\cancel{Na^{+}}\ (aq) + \underline{1}S^{2-}\ (aq) \rightarrow \underline{1}NiS\ (s) + \underline{2}\cancel{Na^{+}}\ (aq) + \cancel{Cl^{-}}\ (aq)$$



SPECTACTOR IONS

You may notice that some of these ions (Na^{+} and Cl^{-}) don't actually chemically react.

We can ignore these!

We can write a net ionic equation by eliminating all the spectator ions: $\underline{1}Ni^{+}\ (aq) + \underline{1}S^{2-}\ (aq) \rightarrow \underline{1}NiS\ (s)$

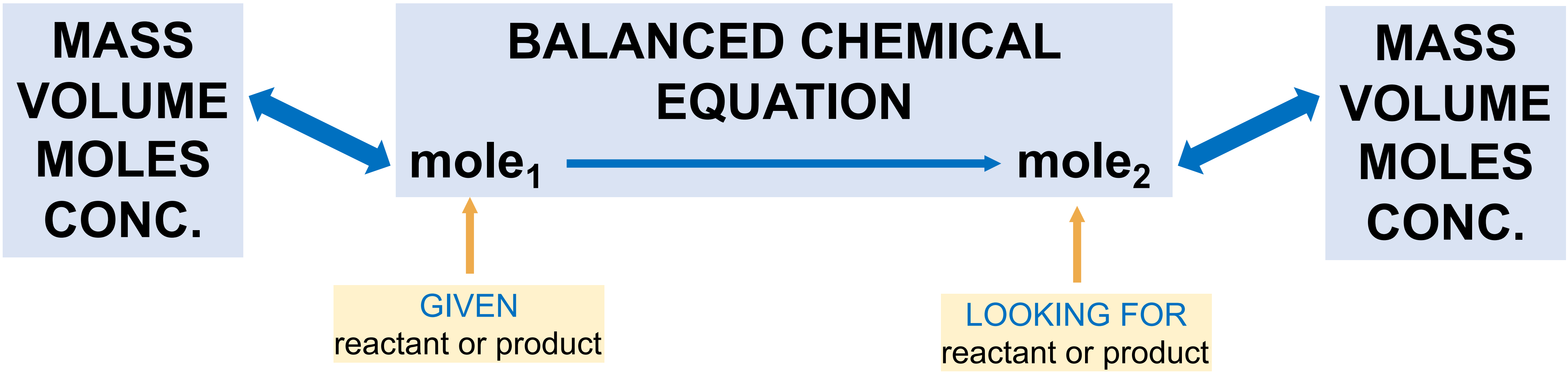
Everything so far has been qualitative.

But we can also understand everything quantitatively!

After all, these are still chemical reactions.

SUMMARIZING STOICHIOMETRY RELATIONSHIPS

We can add concentration now!



*I hope now you understand why I say to convert to moles before you do anything else. It's because a balanced chemical equation gives us **mole-to-mole ratios** that we can use to convert between one reactant/product to another reactant/product.*

A Guided Example

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$Pb(NO_3)_2$ (*aq*) + 2KI (*aq*) → PbI_2 (*s*) + 2KNO_3 (*aq*) *molecular*

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We mix an aqueous solution of lead(II) nitrate with an aqueous solution of potassium iodide.

$Pb(NO_3)_2\ (aq) + 2KI\ (aq) \rightarrow PbI_2\ (s) + 2KNO_3\ (aq)$ *molecular*

$Pb^{2+}\ (aq) + 2NO_3^-\ (aq) + 2K^+\ (aq) + 2I^-\ (aq) \rightarrow PbI_2\ (s) + 2K^+\ (aq) + 2NO_3^-\ (aq)$ *complete ionic*

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Pb^{2+} (*aq*) + $2I^-$ (*aq*) → PbI_2 (*s*)	*net ionic*

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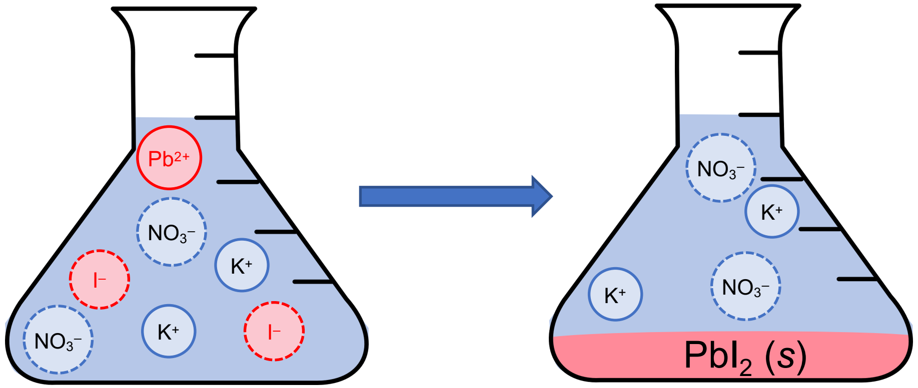
Revisit the microscopic picture!



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Let's say we mixed 100.0 mL of a 0.100 M $Pb(NO_3)_2$ solution with 250.0 mL of a 0.200 M KI solution.
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Let's say we mixed 100.0 mL of a 0.100 M $Pb(NO_3)_2$ solution with 250.0 mL of a 0.200 M KI solution. What is the mass of the solid that forms?

We can use the molecular equation to solve this problem. Let's find the moles of each reactant:

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Q: Which is the limiting reactant? *I'll leave it up to you to remember how to do this!*

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A: $Pb(NO_3)_2$ (*aq*) is the limiting reactant!

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We can use the net ionic equation to solve this problem.

This is the same as the previous slide but even less work! Convince yourself that you can do it!

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$$0.0500 \text{ mol KI} \times \frac{1 \text{ mol } K^+}{1 \text{ mol KI}} = 0.0500 \text{ mol } K^+$$

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Because Pb^{2+} was the limiting reactant, $[Pb^{2+}]$ = 0 M.

A Guided Example

We mix an aqueous solution of lead(II) nitrate with an aqueous solution of potassium iodide.

$Pb(NO_3)_2$ (*aq*) + 2KI (*aq*) → PbI_2 (*s*) + $2KNO_3$ (*aq*) *molecular*

Pb^{2+} (*aq*) + $2NO_3^-$ (*aq*) + $2K^+$ (*aq*) + $2I^-$ (*aq*) → PbI_2 (*s*) + $2K^+$ (*aq*) + $2NO_3^-$ (*aq*) *complete ionic*

Pb^{2+} (*aq*) + $2I^-$ (*aq*) → PbI_2 (*s*) *net ionic*

Let's say we mixed 100.0 mL of a 0.100 M $Pb(NO_3)_2$ solution with 250.0 mL of a 0.200 M KI solution.

What are the concentrations of the ions left in the solution after the reaction is complete?

Remember from before:

$$0.0100 \text{ mol Pb(NO}_3)_2 \times \frac{1 \text{ mol Pb}^{2+}}{1 \text{ mol Pb(NO}_3)_2} = 0.0100 \text{ mol Pb}^{2+}$$

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Because Pb^{2+} was the limiting reactant, $[Pb^{2+}] = 0$ M. What about $[I^-]$?

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$Pb(NO_3)_2$ (*aq*) + 2KI (*aq*) → PbI_2 (*s*) + $2KNO_3$ (*aq*) *molecular*

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Because Pb^{2+} was the limiting reactant, $[Pb^{2+}]$ = 0 M. What about $[I^-]$?

$$0.0100 \text{ mol Pb}^{2+} \times \frac{2 \text{ mol I}^-}{1 \text{ mol Pb}^{2+}} = 0.0200 \text{ mol I}^- \text{ used}$$

A Guided Example

We mix an aqueous solution of lead(II) nitrate with an aqueous solution of potassium iodide.

$Pb(NO_3)_2\ (aq) + 2KI\ (aq) \rightarrow PbI_2\ (s) + 2KNO_3\ (aq)$ *molecular*

$Pb^{2+}\ (aq) + 2NO_3^-\ (aq) + 2K^+\ (aq) + 2I^-\ (aq) \rightarrow PbI_2\ (s) + 2K^+\ (aq) + 2NO_3^-\ (aq)$ *complete ionic*

$Pb^{2+}\ (aq) + 2I^-\ (aq) \rightarrow PbI_2\ (s)$ *net ionic*

Let's say we mixed 100.0 mL of a 0.100 M $Pb(NO_3)_2$ solution with 250.0 mL of a 0.200 M KI solution.

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Because Pb^{2+} was the limiting reactant, $[Pb^{2+}] = 0$ M. What about $[I^-]$?

$$0.0100\ \text{mol Pb}^{2+} \times \frac{2\ \text{mol I}^-}{1\ \text{mol Pb}^{2+}} = 0.0200\ \text{mol I}^-\ \text{used}$$

$$0.0500\ \text{mol I}^- - 0.0200\ \text{mol I}^- = 0.0300\ \text{mol I}^-\ \text{left over}$$

A Guided Example

We mix an aqueous solution of lead(II) nitrate with an aqueous solution of potassium iodide.

$Pb(NO_3)_2$ (*aq*) + 2KI (*aq*) → PbI_2 (*s*) + $2KNO_3$ (*aq*) *molecular*

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$$0.0500 \text{ mol KI} \times \frac{1 \text{ mol } \mathrm{K^+}}{1 \text{ mol KI}} = 0.0500 \text{ mol } \mathrm{K^+}$$

$$0.0100 \text{ mol } \mathrm{Pb(NO_3)_2} \times \frac{2 \text{ mol } \mathrm{NO_3^-}}{1 \text{ mol } \mathrm{Pb(NO_3)_2}} = 0.0200 \text{ mol } \mathrm{NO_3^-}$$

$$0.0500 \text{ mol KI} \times \frac{1 \text{ mol } \mathrm{I^-}}{1 \text{ mol KI}} = 0.0500 \text{ mol } \mathrm{I^-}$$

Because Pb^{2+} was the limiting reactant, $[Pb^{2+}] = 0$ M. What about $[I^-]$?

$$0.0500 \text{ mol } \mathrm{I^-} - 0.0200 \text{ mol } \mathrm{I^-} = 0.0300 \text{ mol } \mathrm{I^-} \text{ left over}$$

$$V_{total} = 100.0 \text{ mL} + 250.0 \text{ mL} = 0.3500 \text{ L}$$

A Guided Example

We mix an aqueous solution of lead(II) nitrate with an aqueous solution of potassium iodide.

$Pb(NO_3)_2$ (*aq*) + 2KI (*aq*) → PbI_2 (*s*) + 2KNO_3 (*aq*) *molecular*

Pb^{2+} (*aq*) + 2NO_3^- (*aq*) + 2K^+ (*aq*) + 2I^- (*aq*) → PbI_2 (*s*) + 2K^+ (*aq*) + 2NO_3^- (*aq*) *complete ionic*

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Let's say we mixed 100.0 mL of a 0.100 M $Pb(NO_3)_2$ solution with 250.0 mL of a 0.200 M KI solution.

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Because Pb^{2+} was the limiting reactant, $[Pb^{2+}] = 0$ M. What about $[I^-]$?

$$[\text{I}^-] = \frac{0.0300\ \text{mol I}^-}{0.3500\ \text{L}} = 0.0857\ \text{M}$$

A Guided Example

We mix an aqueous solution of lead(II) nitrate with an aqueous solution of potassium iodide.

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Pb^{2+} (*aq*) + $2NO_3^-$ (*aq*) + $2K^+$ (*aq*) + $2I^-$ (*aq*) → PbI_2 (*s*) + $2K^+$ (*aq*) + $2NO_3^-$ (*aq*) *complete ionic*

Pb^{2+} (*aq*) + $2I^-$ (*aq*) → PbI_2 (*s*) *net ionic*

Let's say we mixed 100.0 mL of a 0.100 M $Pb(NO_3)_2$ solution with 250.0 mL of a 0.200 M KI solution.

What are the concentrations of the ions left in the solution after the reaction is complete?

Remember from before:

$$0.0100 \text{ mol Pb(NO}_3)_2 \times \frac{1 \text{ mol Pb}^{2+}}{1 \text{ mol Pb(NO}_3)_2} = 0.0100 \text{ mol Pb}^{2+}$$

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Because Pb^{2+} was the limiting reactant, $[Pb^{2+}]$ = 0 M. What about $[I^-]$? What about $[NO_3^-]$ and $[K^+]$?

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We mix an aqueous solution of lead(II) nitrate with an aqueous solution of potassium iodide.

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Pb^{2+} (*aq*) + 2NO_3^- (*aq*) + 2K^+ (*aq*) + 2I^- (*aq*) → PbI_2 (*s*) + 2K^+ (*aq*) + 2NO_3^- (*aq*) *complete ionic*

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Let's say we mixed 100.0 mL of a 0.100 M $Pb(NO_3)_2$ solution with 250.0 mL of a 0.200 M KI solution.

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Because Pb^{2+} was the limiting reactant, $[Pb^{2+}]$ = 0 M. What about $[I^-]$? What about $[NO_3^-]$ and $[K^+]$?

$$[\mathrm{I^-}] = \frac{0.0300\ \mathrm{mol\ I^-}}{0.3500\ \mathrm{L}} = 0.0857\ \mathrm{M} \qquad [\mathrm{NO_3^-}] = \frac{0.0200\ \mathrm{mol\ NO_3^-}}{0.3500\ \mathrm{L}} = 0.0571\ \mathrm{M}$$

A Guided Example

We mix an aqueous solution of lead(II) nitrate with an aqueous solution of potassium iodide.

$Pb(NO_3)_2\ (aq) + 2KI\ (aq) \rightarrow PbI_2\ (s) + 2KNO_3\ (aq)$ *molecular*

$Pb^{2+}\ (aq) + 2NO_3^-\ (aq) + 2K^+\ (aq) + 2I^-\ (aq) \rightarrow PbI_2\ (s) + 2K^+\ (aq) + 2NO_3^-\ (aq)$ *complete ionic*

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Let's say we mixed 100.0 mL of a 0.100 M $Pb(NO_3)_2$ solution with 250.0 mL of a 0.200 M KI solution.

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$$0.0500\ \text{mol KI} \times \frac{1\ \text{mol}\ I^-}{1\ \text{mol KI}} = 0.0500\ \text{mol}\ I^-$$

Because Pb^{2+} was the limiting reactant, $[Pb^{2+}] = 0$ M. What about $[I^-]$? What about $[NO_3^-]$ and $[K^+]$?

$$[I^-] = \frac{0.0300\ \text{mol}\ I^-}{0.3500\ \text{L}} = 0.0857\ \text{M} \qquad [NO_3^-] = \frac{0.0200\ \text{mol}\ NO_3^-}{0.3500\ \text{L}} = 0.0571\ \text{M} \qquad [K^+] = \frac{0.0500\ \text{mol}\ NO_3^-}{0.3500\ \text{L}} = 0.143\ \text{M}$$

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Let's say we mixed 100.0 mL of a 0.100 M $Pb(NO_3)_2$ solution with 250.0 mL of a 0.200 M KI solution.

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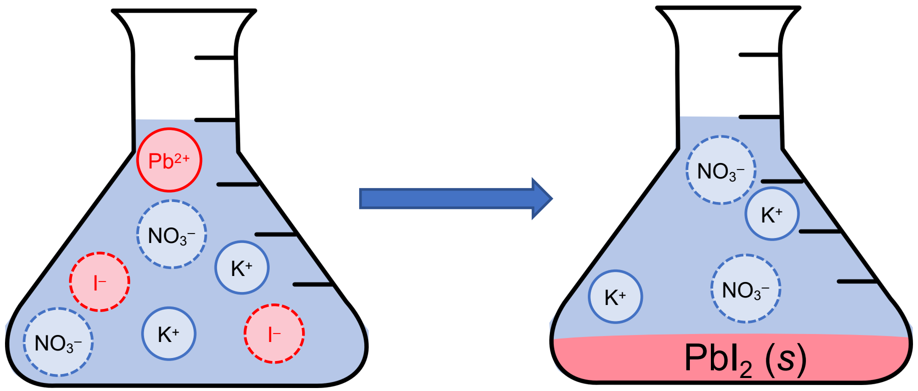
Does this all make sense though?

Go back to the microscopic picture of the solution!



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Initial concentrations

$[Pb^{2+}] = 0.100$ M
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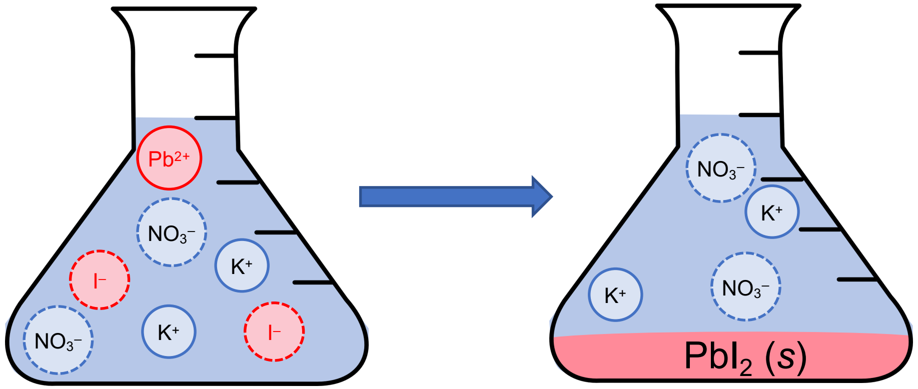

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$[Pb^{2+}] = 0$ M
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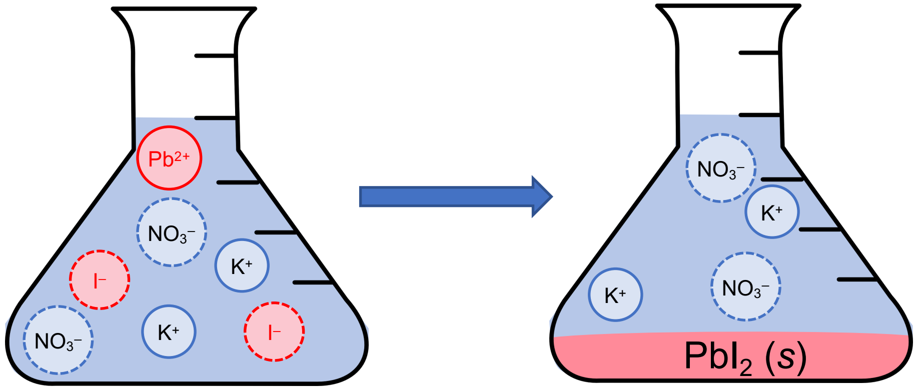

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$$[I^-] = 0.0857 \text{ M}$$

$[Pb^{2+}]$ = 0 M; everything else is diluted: I^- since there are less moles and K^+/NO_3^- since volume is larger.

What volume of 0.750 M $FeCl_2$ is needed to completely react with 1.00×10^2 mL of 0.25 M $KMnO_4$ in the reaction:

$$5FeCl_2\ (aq) + KMnO_4\ (aq) + 8HCl\ (aq) \rightarrow 5FeCl_3\ (aq) + MnCl_2\ (aq) + KCl\ (aq) + 4H_2O\ (l)$$

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Start by finding the number of moles of $KMnO_4$:

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$$0.25\ \text{M} = \frac{\text{x mol}}{1.00\times10^2\ \text{mL}\times\frac{1\ \text{L}}{1000\ \text{mL}}}$$

$$\text{x} = 0.025\ \text{mol KMnO}_4$$

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Now, use the balanced chemical equation to find out how many moles of $FeCl_2$ are needed:

$$0.025\ \text{mol KMnO}_4\times\frac{5\ \text{mol FeCl}_2}{1\ \text{mol KMnO}_4} = 0.12_5\ \text{mol FeCl}_2\ \text{needed}$$

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Finally, find the volume of the 0.750 M $FeCl_2$ solution that would have this many moles:

$$0.750\ \mathrm{M} = \frac{0.12_5\ \mathrm{mol\ FeCl_2}}{\mathrm{V}}$$

$$\mathrm{V} = 0.17\ \mathrm{L}$$

Consider the reaction between 200.0 mL of 0.10 M Na_2CrO_4 and 150.0 mL of 0.10 M $AgNO_3$.

What mass of precipitate would form?

Start by writing the equations for this reaction:

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Start by writing the equations for this reaction:

$$Na_2CrO_4\ (aq) + 2AgNO_3\ (aq) \rightarrow 2NaNO_3\ (aq) + Ag_2CrO_4\ (s)$$

Consider the reaction between 200.0 mL of 0.10 M Na_2CrO_4 and 150.0 mL of 0.10 M $AgNO_3$.

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$$2Na^+\ (aq) + CrO_4^{2-}\ (aq) + 2Ag^+\ (aq) + 2NO_3^-\ (aq) \rightarrow 2Na^+\ (aq) + 2NO_3^-\ (aq) + Ag_2CrO_4\ (s)$$

Consider the reaction between 200.0 mL of 0.10 M Na_2CrO_4 and 150.0 mL of 0.10 M $AgNO_3$.

What mass of precipitate would form?

Start by writing the equations for this reaction:

$Na_2CrO_4\ (aq) + 2AgNO_3\ (aq) \rightarrow 2NaNO_3\ (aq) + Ag_2CrO_4\ (s)$

$2Na^+\ (aq) + CrO_4^{2-}\ (aq) + 2Ag^+\ (aq) + 2NO_3^-\ (aq) \rightarrow 2Na^+\ (aq) + 2NO_3^-\ (aq) + Ag_2CrO_4\ (s)$

$CrO_4^{2-}\ (aq) + 2Ag^+\ (aq) \rightarrow Ag_2CrO_4\ (s)$

Consider the reaction between 200.0 mL of 0.10 M Na_2CrO_4 and 150.0 mL of 0.10 M $AgNO_3$.

What mass of precipitate would form?

Start by writing the equations for this reaction:

Na_2CrO_4 (*aq*) + $2AgNO_3$ (*aq*) → $2NaNO_3$ (*aq*) + Ag_2CrO_4 (*s*)

$2Na^+$ (*aq*) + CrO_4^{2-} (*aq*) + $2Ag^+$ (*aq*) + $2NO_3^-$ (*aq*) → $2Na^+$ (*aq*) + $2NO_3^-$ (*aq*) + Ag_2CrO_4 (*s*)

CrO_4^{2-} (*aq*) + $2Ag^+$ (*aq*) → Ag_2CrO_4 (*s*)

Then, determine the number of moles of each reactant and the limiting reactant:

Consider the reaction between 200.0 mL of 0.10 M Na_2CrO_4 and 150.0 mL of 0.10 M $AgNO_3$.

What mass of precipitate would form?

Start by writing the equations for this reaction:

Na_2CrO_4 (*aq*) + $2AgNO_3$ (*aq*) → $2NaNO_3$ (*aq*) + Ag_2CrO_4 (*s*)

$2Na^+$ (*aq*) + CrO_4^{2-} (*aq*) + $2Ag^+$ (*aq*) + $2NO_3^-$ (*aq*) → $2Na^+$ (*aq*) + $2NO_3^-$ (*aq*) + Ag_2CrO_4 (*s*)

CrO_4^{2-} (*aq*) + $2Ag^+$ (*aq*) → Ag_2CrO_4 (*s*)

Then, determine the number of moles of each reactant and the limiting reactant:

$$0.10\ \mathrm{M\ Na_2CrO_4} = \frac{\mathrm{x\ mol}}{0.2000\ \mathrm{L}}$$
$$\mathrm{x} = 0.020\ \mathrm{mol\ Na_2CrO_4}$$

$$0.10\ \mathrm{M\ AgNO_3} = \frac{\mathrm{x\ mol}}{0.1500\ \mathrm{L}}$$
$$\mathrm{x} = 0.015\ \mathrm{mol\ AgNO_3}$$

Consider the reaction between 200.0 mL of 0.10 M Na_2CrO_4 and 150.0 mL of 0.10 M $AgNO_3$.

What mass of precipitate would form?

Start by writing the equations for this reaction:

$Na_2CrO_4\ (aq) + 2AgNO_3\ (aq) \rightarrow 2NaNO_3\ (aq) + Ag_2CrO_4\ (s)$

$2Na^+\ (aq) + CrO_4^{2-}\ (aq) + 2Ag^+\ (aq) + 2NO_3^-\ (aq) \rightarrow 2Na^+\ (aq) + 2NO_3^-\ (aq) + Ag_2CrO_4\ (s)$

$CrO_4^{2-}\ (aq) + 2Ag^+\ (aq) \rightarrow Ag_2CrO_4\ (s)$

Then, determine the number of moles of each reactant and the limiting reactant:

$$0.10\text{ M Na}_2\text{CrO}_4 = \frac{\text{x mol}}{0.2000\text{ L}}$$
$$\text{x} = 0.020\text{ mol Na}_2\text{CrO}_4$$

$$0.10\text{ M AgNO}_3 = \frac{\text{x mol}}{0.1500\text{ L}}$$
$$\text{x} = 0.015\text{ mol AgNO}_3$$

→ The limiting reactant is $AgNO_3$ (or Ag^+). *I'll leave it to you to be able to solve this part!*

Consider the reaction between 200.0 mL of 0.10 M Na_2CrO_4 and 150.0 mL of 0.10 M $AgNO_3$.

What mass of precipitate would form?

Start by writing the equations for this reaction:

$Na_2CrO_4\ (aq) + 2AgNO_3\ (aq) \rightarrow 2NaNO_3\ (aq) + Ag_2CrO_4\ (s)$

$2Na^+\ (aq) + CrO_4^{2-}\ (aq) + 2Ag^+\ (aq) + 2NO_3^-\ (aq) \rightarrow 2Na^+\ (aq) + 2NO_3^-\ (aq) + Ag_2CrO_4\ (s)$

$CrO_4^{2-}\ (aq) + 2Ag^+\ (aq) \rightarrow Ag_2CrO_4\ (s)$

Then, determine the number of moles of each reactant and the limiting reactant:

$$0.10\ \mathrm{M\ Na_2CrO_4} = \frac{\mathrm{x\ mol}}{0.2000\ \mathrm{L}}$$
$$\mathrm{x} = 0.020\ \mathrm{mol\ Na_2CrO_4}$$

$$0.10\ \mathrm{M\ AgNO_3} = \frac{\mathrm{x\ mol}}{0.1500\ \mathrm{L}}$$
$$\mathrm{x} = 0.015\ \mathrm{mol\ AgNO_3}$$

→ The limiting reactant is $AgNO_3$ (or Ag^+). *I'll leave it to you to be able to solve this part!*

Finally, find the mass of Ag_2CrO_4 formed from the starting amount of the limiting reactant:

Consider the reaction between 200.0 mL of 0.10 M Na_2CrO_4 and 150.0 mL of 0.10 M $AgNO_3$.

What mass of precipitate would form?

Start by writing the equations for this reaction:

$Na_2CrO_4\ (aq) + 2AgNO_3\ (aq) \rightarrow 2NaNO_3\ (aq) + Ag_2CrO_4\ (s)$

$2Na^+\ (aq) + CrO_4^{2-}\ (aq) + 2Ag^+\ (aq) + 2NO_3^-\ (aq) \rightarrow 2Na^+\ (aq) + 2NO_3^-\ (aq) + Ag_2CrO_4\ (s)$

$CrO_4^{2-}\ (aq) + 2Ag^+\ (aq) \rightarrow Ag_2CrO_4\ (s)$

Then, determine the number of moles of each reactant and the limiting reactant:

$$0.10\ \text{M}\ Na_2CrO_4 = \frac{\text{x mol}}{0.2000\ \text{L}}$$
$$\text{x} = 0.020\ \text{mol}\ Na_2CrO_4$$

$$0.10\ \text{M}\ AgNO_3 = \frac{\text{x mol}}{0.1500\ \text{L}}$$
$$\text{x} = 0.015\ \text{mol}\ AgNO_3$$

→ The limiting reactant is $AgNO_3$ (or Ag^+). *I'll leave it to you to be able to solve this part!*

Finally, find the mass of Ag_2CrO_4 formed from the starting amount of the limiting reactant:

$$0.015\ \text{mol}\ AgNO_3 \times \frac{1\ \text{mol}\ Ag_2CrO_4}{2\ \text{mol}\ AgNO_3} \times \frac{331.8\ \text{g}\ Ag_2CrO_4}{1\ \text{mol}\ Ag_2CrO_4} = 2.5\ \text{g}\ Ag_2CrO_4$$

Consider the reaction between 200.0 mL of 0.10 M Na_2CrO_4 and 150.0 mL of 0.10 M $AgNO_3$.

What is the concentration of the chromate ions left in the solution after the reaction occurs?

Consider the reaction between 200.0 mL of 0.10 M Na_2CrO_4 and 150.0 mL of 0.10 M $AgNO_3$.

What is the concentration of the chromate ions left in the solution after the reaction occurs?

Start by writing the equations for this reaction:

$Na_2CrO_4\ (aq) + 2AgNO_3\ (aq) \rightarrow 2NaNO_3\ (aq) + Ag_2CrO_4\ (s)$

$2Na^+\ (aq) + CrO_4^{2-}\ (aq) + 2Ag^+\ (aq) + 2NO_3^-\ (aq) \rightarrow 2Na^+\ (aq) + 2NO_3^-\ (aq) + Ag_2CrO_4\ (s)$

$CrO_4^{2-}\ (aq) + 2Ag^+\ (aq) \rightarrow Ag_2CrO_4\ (s)$

→ The limiting reactant is $AgNO_3$ (or Ag^+).

Now find the amount of CrO_4^{2-} needed to react with the Ag^+ ions to form the precipitate:

Consider the reaction between 200.0 mL of 0.10 M Na_2CrO_4 and 150.0 mL of 0.10 M $AgNO_3$.

What is the concentration of the chromate ions left in the solution after the reaction occurs?

Start by writing the equations for this reaction:

$Na_2CrO_4\ (aq) + 2AgNO_3\ (aq) \rightarrow 2NaNO_3\ (aq) + Ag_2CrO_4\ (s)$

$2Na^+\ (aq) + CrO_4^{2-}\ (aq) + 2Ag^+\ (aq) + 2NO_3^-\ (aq) \rightarrow 2Na^+\ (aq) + 2NO_3^-\ (aq) + Ag_2CrO_4\ (s)$

$CrO_4^{2-}\ (aq) + 2Ag^+\ (aq) \rightarrow Ag_2CrO_4\ (s)$

→ The limiting reactant is $AgNO_3$ (or Ag^+).

Now find the amount of CrO_4^{2-} needed to react with the Ag^+ ions to form the precipitate:

$$0.10\ \text{M AgNO}_3 = \frac{\text{x mol}}{0.1500\ \text{L}}$$

$$\text{x} = 0.015\ \text{mol AgNO}_3$$

$$0.015\ \text{mol AgNO}_3 \times \frac{1\ \text{mol Ag}^+}{1\ \text{mol AgNO}_3} = 0.015\ \text{mol Ag}^+$$

Consider the reaction between 200.0 mL of 0.10 M Na_2CrO_4 and 150.0 mL of 0.10 M $AgNO_3$.

What is the concentration of the chromate ions left in the solution after the reaction occurs?

Start by writing the equations for this reaction:

$Na_2CrO_4\ (aq) + 2AgNO_3\ (aq) \rightarrow 2NaNO_3\ (aq) + Ag_2CrO_4\ (s)$

$2Na^+\ (aq) + CrO_4^{2-}\ (aq) + 2Ag^+\ (aq) + 2NO_3^-\ (aq) \rightarrow 2Na^+\ (aq) + 2NO_3^-\ (aq) + Ag_2CrO_4\ (s)$

$CrO_4^{2-}\ (aq) + 2Ag^+\ (aq) \rightarrow Ag_2CrO_4\ (s)$

→ The limiting reactant is $AgNO_3$ (or Ag^+).

Now find the amount of CrO_4^{2-} needed to react with the Ag^+ ions to form the precipitate:

$$0.10\ \text{M AgNO}_3 = \frac{\text{x mol}}{0.1500\ \text{L}}$$

$$\text{x} = 0.015\ \text{mol AgNO}_3$$

$$0.015\ \text{mol AgNO}_3 \times \frac{1\ \text{mol Ag}^+}{1\ \text{mol AgNO}_3} = 0.015\ \text{mol Ag}^+$$

$$0.015\ \text{mol Ag}^+ \times \frac{1\ \text{mol CrO}_4^{2-}}{2\ \text{mol Ag}^+} = 0.0075\ \text{mol CrO}_4^{2-}\ \text{reacted}$$

Consider the reaction between 200.0 mL of 0.10 M Na_2CrO_4 and 150.0 mL of 0.10 M $AgNO_3$.

What is the concentration of the chromate ions left in the solution after the reaction occurs?

Start by writing the equations for this reaction:

$Na_2CrO_4\ (aq) + 2AgNO_3\ (aq) \rightarrow 2NaNO_3\ (aq) + Ag_2CrO_4\ (s)$

$2Na^+\ (aq) + CrO_4^{2-}\ (aq) + 2Ag^+\ (aq) + 2NO_3^-\ (aq) \rightarrow 2Na^+\ (aq) + 2NO_3^-\ (aq) + Ag_2CrO_4\ (s)$

$CrO_4^{2-}\ (aq) + 2Ag^+\ (aq) \rightarrow Ag_2CrO_4\ (s)$

→ The limiting reactant is $AgNO_3$ (or Ag^+).

Now find the amount of CrO_4^{2-} needed to react with the Ag^+ ions to form the precipitate:

$$0.10\ \text{M AgNO}_3 = \frac{\text{x mol}}{0.1500\ \text{L}}$$

$$\text{x} = 0.015\ \text{mol AgNO}_3$$

$$0.015\ \text{mol AgNO}_3 \times \frac{1\ \text{mol Ag}^+}{1\ \text{mol AgNO}_3} = 0.015\ \text{mol Ag}^+$$

$$0.015\ \text{mol Ag}^+ \times \frac{1\ \text{mol CrO}_4^{2-}}{2\ \text{mol Ag}^+} = 0.0075\ \text{mol CrO}_4^{2-}\ \text{reacted}$$

Finally, find the final concentration:

Consider the reaction between 200.0 mL of 0.10 M Na_2CrO_4 and 150.0 mL of 0.10 M $AgNO_3$.

What is the concentration of the chromate ions left in the solution after the reaction occurs?

Start by writing the equations for this reaction:

$Na_2CrO_4\ (aq) + 2AgNO_3\ (aq) \rightarrow 2NaNO_3\ (aq) + Ag_2CrO_4\ (s)$

$2Na^+\ (aq) + CrO_4^{2-}\ (aq) + 2Ag^+\ (aq) + 2NO_3^-\ (aq) \rightarrow 2Na^+\ (aq) + 2NO_3^-\ (aq) + Ag_2CrO_4\ (s)$

$CrO_4^{2-}\ (aq) + 2Ag^+\ (aq) \rightarrow Ag_2CrO_4\ (s)$

→ The limiting reactant is $AgNO_3$ (or Ag^+).

Now find the amount of CrO_4^{2-} needed to react with the Ag^+ ions to form the precipitate:

$$0.10\ \mathrm{M\ AgNO_3} = \frac{\mathrm{x\ mol}}{0.1500\ \mathrm{L}}$$

$$\mathrm{x} = 0.015\ \mathrm{mol\ AgNO_3}$$

$$0.015\ \mathrm{mol\ AgNO_3} \times \frac{1\ \mathrm{mol\ Ag^+}}{1\ \mathrm{mol\ AgNO_3}} = 0.015\ \mathrm{mol\ Ag^+}$$

$$0.015\ \mathrm{mol\ Ag^+} \times \frac{1\ \mathrm{mol\ CrO_4^{2-}}}{2\ \mathrm{mol\ Ag^+}} = 0.0075\ \mathrm{mol\ CrO_4^{2-}\ reacted}$$

Finally, find the final concentration:

Consider the reaction between 200.0 mL of 0.10 M Na_2CrO_4 and 150.0 mL of 0.10 M $AgNO_3$.

What is the concentration of the chromate ions left in the solution after the reaction occurs?

Start by writing the equations for this reaction:

Na_2CrO_4 (*aq*) + 2$AgNO_3$ (*aq*) → 2$NaNO_3$ (*aq*) + Ag_2CrO_4 (*s*)

2Na^+ (*aq*) + CrO_4^{2-} (*aq*) + 2Ag^+ (*aq*) + 2NO_3^- (*aq*) → 2Na^+ (*aq*) + 2NO_3^- (*aq*) + Ag_2CrO_4 (*s*)

CrO_4^{2-} (*aq*) + 2Ag^+ (*aq*) → Ag_2CrO_4 (*s*)

→ The limiting reactant is $AgNO_3$ (or Ag^+).

Now find the amount of CrO_4^{2-} needed to react with the Ag^+ ions to form the precipitate:

$$0.10\text{ M AgNO}_3 = \frac{\text{x mol}}{0.1500\text{ L}}$$

$$\text{x} = 0.015\text{ mol AgNO}_3$$

$$0.015\text{ mol AgNO}_3 \times \frac{1\text{ mol Ag}^+}{1\text{ mol AgNO}_3} = 0.015\text{ mol Ag}^+$$

$$0.015\text{ mol Ag}^+ \times \frac{1\text{ mol CrO}_4^{2-}}{2\text{ mol Ag}^+} = 0.0075\text{ mol CrO}_4^{2-}\text{ reacted}$$

Finally, find the final concentration: $$[\text{CrO}_4^{2-}] = \frac{0.020\text{ mol CrO}_4^{2-} - 0.0075\text{ mol CrO}_4^{2-}}{0.2000\text{ L} + 0.1500\text{ L}} = 0.036\text{ M CrO}_4^{2-}$$