

Fractional Precipitation

A solution contains two of the following cations: Pb^{2+} , Fe^{3+} , Ca^{2+} . Based upon the following experiments, determine which two cations are present!

You add 1M HCl(aq) to the solution.

No Pb^{2+} b/c Cl^- forms

PbCl_2 precipitate

You add 1M NaOH(aq) to the solution.

Fe(OH)_3 precipitate

No precipitate forms



You add $(\text{NH}_4)_2\text{HPO}_4/\text{NH}_3$ to the solution

Precipitate forms

$\text{Ca}_3(\text{PO}_4)_2$ precipitate

Solubility Rules for Ionic Compounds in Water

Compounds Containing the Following Ions Are Generally <u>Soluble</u>	Exceptions
Li^+ , Na^+ , K^+ , and NH_4^+	None
NO_3^- , CH_3COO^-	None
Cl^- , Br^- , and I^-	When these ions pair with Ag^+ , Hg_2^{2+} , or Pb^{2+} => insoluble compounds
SO_4^{2-}	When SO_4^{2-} pairs with Ca^{2+} , Sr^{2+} , Ba^{2+} , Pb^{2+} , or Ag^+ => insoluble

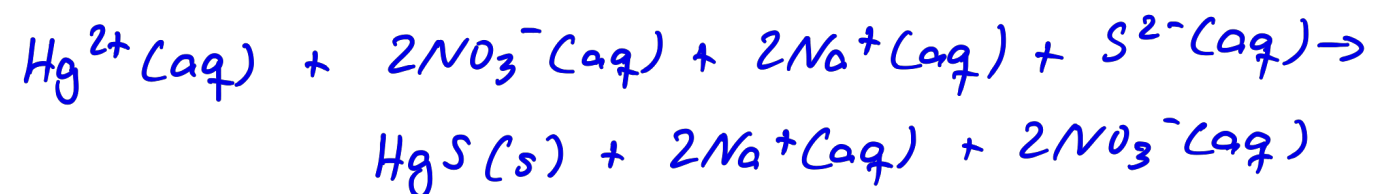
Compounds Containing the Following Ions Are Generally <u>Insoluble</u>	Exceptions
OH^- and S^{2-}	When these ions pair with Li^+ , Na^+ , K^+ , or NH_4^+ => soluble
	When S^{2-} pairs with Ca^{2+} , Sr^{2+} or Ba^{2+} => soluble
	When OH^- pairs with Ca^{2+} , Sr^{2+} or Ba^{2+} => slightly soluble
CO_3^{2-} and PO_4^{3-}	When these ions pair with Li^+ , Na^+ , K^+ , or NH_4^+ => soluble

Precipitation Calculations

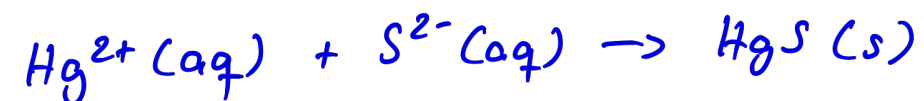
What mass of mercury(II) sulfide can be produced when 50 mL of 0.0150 M mercury(II) nitrate is mixed with a solution containing excess of sodium sulfide

Given: $\text{Hg}(\text{NO}_3)_2$ 50 mL 0.0150 M
 Na_2S excess

Complete Ionic Equation



Net Ionic



Calculation

$$\# \text{ moles } (\text{Hg}^{2+}) = 0.0150 \text{ M} \cdot 0.0500 \text{ L} = 0.00075 \text{ mol}$$

$$\# \text{ moles } (\text{HgS}) = 0.00075 \text{ mol } (\text{Hg}^{2+}) \cdot \frac{1 \text{ mol } (\text{HgS})}{1 \text{ mol } (\text{Hg}^{2+})}$$

$$\# \text{ moles } (\text{HgS}) = 0.00075 \text{ mol}$$

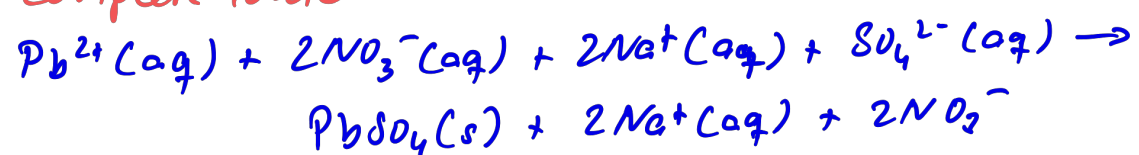
$$\text{mass } (\text{HgS}) = 0.00075 \text{ mol} \cdot \frac{232.7 \text{ g}}{1 \text{ mol}} = 0.175 \text{ g}$$

When aqueous solutions of Na_2SO_4 and $\text{Pb}(\text{NO}_3)_2$ are mixed PbSO_4 precipitates. Calculate the mass of PbSO_4 formed when 1.25 L of 0.0500 M $\text{Pb}(\text{NO}_3)_2$ and 2.00 L of 0.0250 M Na_2SO_4 are mixed

Given:

1.25 L of 0.0500 M $\text{Pb}(\text{NO}_3)_2$
2.00 L of 0.0250 M Na_2SO_4

Complete Ionic



Net Ionic



What is the limiting reagent?

$$\begin{aligned} \text{moles}(\text{Pb}^{2+}) &= 1.25 \text{ L} \cdot \frac{0.0500 \text{ mol}}{1 \text{ L}} = 0.0625 \text{ mol} \\ \text{moles}(\text{PbSO}_4) &= 0.0625 \text{ mol}(\text{Pb}^{2+}) \cdot \frac{1 \text{ mol}(\text{PbSO}_4)}{1 \text{ mol} \text{ Pb}^{2+}} \\ &= 0.0625 \text{ mol} \end{aligned}$$

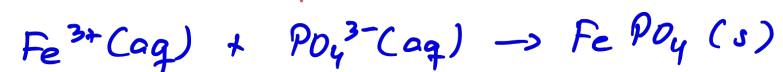
$$\begin{aligned} \text{moles}(\text{SO}_4^{2-}) &= 2.00 \text{ L} \cdot \frac{0.0250 \text{ mol}}{1 \text{ L}} = 0.0500 \text{ mol} \\ \text{moles}(\text{PbSO}_4) &= 0.0500 \text{ mol}(\text{SO}_4^{2-}) \cdot \frac{1 \text{ mol}(\text{PbSO}_4)}{1 \text{ mol}(\text{SO}_4^{2-})} \\ &= 0.0500 \text{ mol} \end{aligned}$$

SO_4^{2-} is the limiting reagent

$$\text{mass}(\text{PbSO}_4) = 0.0500 \text{ mol} \cdot \frac{303.3 \text{ g}}{1 \text{ mol}} = 15.2 \text{ g}$$

Example 3: You mix 200 mL of 0.20 M $\text{Fe}_2(\text{SO}_4)_3$ solution with 250 mL of 0.20 M Na_3PO_4 solution. What is the precipitate that will form? What is its mass? What is the conc. of the ion that is in excess?

1. Net Ionic Equation



2. Determine limiting reagent

$$\begin{aligned} \# \text{ moles } (\text{Fe}^{3+}) &= 200 \text{ mL} \cdot \frac{1 \text{ L}}{1000 \text{ mL}} \cdot 0.20 \text{ M } (\text{Fe}_2(\text{SO}_4)_3) \\ &\quad \cdot \frac{2 \text{ mol Fe}^{3+}}{1 \text{ mol Fe}_2(\text{SO}_4)_3} \\ &= 0.080 \text{ mol} \end{aligned}$$

$$\# \text{ moles } (\text{PO}_4^{3-}) = \dots = 0.050 \text{ mol}$$

limiting reagent PO_4^{3-}

3. mass of FePO_4

$$\begin{aligned} \text{mass } (\text{FePO}_4) &= 0.050 \text{ mol} \cdot \frac{150.82 \text{ g}}{1 \text{ mol}} \\ &= 7.54 \text{ g} \end{aligned}$$

4. What is the conc. of the ion in excess?

$$\begin{aligned} \# \text{ moles } (\text{Fe}^{3+})_{\text{remaining}} &= 0.080 \text{ mol} - 0.050 \text{ mol} \\ &\quad \text{(initial)} \quad \text{consumed} \\ &= 0.030 \text{ mol} \end{aligned}$$

$$\begin{aligned} \text{total volume: } 200 \text{ mL} + 250 \text{ mL} &= 450 \text{ mL} \\ &= 0.45 \text{ L} \end{aligned}$$

$$M(\text{Fe}^{3+}) = \frac{0.030 \text{ mol}}{0.45 \text{ L}} = 0.067 \text{ M} = 67 \text{ mM}$$