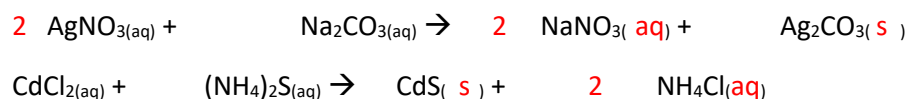
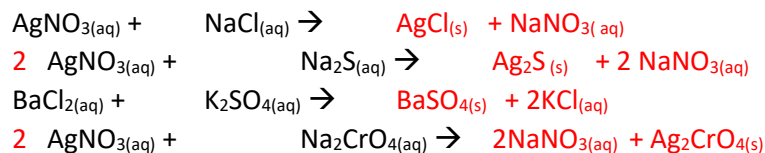


**Questions:**

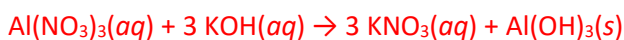
1. Balance and predict the solubility of each product for the following two reactions:



2. Complete the following reactions, balance, and predict the states of the products:



3. What mass of solid aluminum hydroxide can be produced when 50.0 mL of 0.200 M of aluminum nitrate is added to 200.0 mL of 0.100 M of potassium hydroxide?  
What is the concentration of aluminum ions left over in solution after the reaction is complete?



$C = n/V$ , so

$0.200 \text{ M Al}(\text{NO}_3)_3 \times 0.0500 \text{ L} = 0.0100 \text{ mol Al}(\text{NO}_3)_3$ , which will produce  $0.0100 \text{ mol Al}(\text{OH})_3$

$0.100 \text{ M KOH} \times 0.2000 \text{ L} = 0.0200 \text{ mol KOH}$ , which will produce  $(0.0200 \text{ mol}/3 =) 0.00667 \text{ mol Al}(\text{OH})_3$

The limiting reactant is KOH.

$$0.0200 \text{ mol KOH} \times \frac{1 \text{ mol Al}(\text{OH})_3}{3 \text{ mol KOH}} \times \frac{78.003 \text{ g Al}(\text{OH})_3}{1 \text{ mol Al}(\text{OH})_3} = 0.520 \text{ g Al}(\text{OH})_3$$

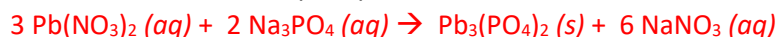
The excess reagent is  $\text{Al}(\text{NO}_3)_3$ .

$0.0200 \text{ mol KOH}$  requires  $(0.0200 \text{ mol KOH} \times 1 \text{ mol Al}(\text{NO}_3)_3/3 \text{ mol KOH} =) 0.00667 \text{ mol Al}(\text{NO}_3)_3$

Excess moles =  $0.0100 - 0.00667 = 0.00333 \text{ mol}$

Excess concentration =  $n/V = 0.00333 \text{ mol} / (0.050 \text{ L} + 0.200 \text{ L}) = 0.0133 \text{ M}$

4. If 10.0 mL of a 0.30 mol/L sodium phosphate solution reacts with 20.0 mL of a 0.20 mol/L lead(II) nitrate solution, what is the mass of the precipitate?



**NOTE: Your work does not need to match the format below; however, it does have to properly show your logic. Additionally, it should include properly labelled units throughout the problem.**

$$10.0 \text{ mL Na}_3\text{PO}_4 \left( \frac{1 \text{ L}}{1000 \text{ mL}} \right) \left( \frac{0.30 \text{ mol Na}_3\text{PO}_4}{1 \text{ L}} \right) = 0.0030 \text{ mol Na}_3\text{PO}_4$$

$$20.0 \text{ mL} \left( \frac{1 \text{ L}}{1000 \text{ mL}} \right) \left( \frac{0.20 \text{ mol Pb}(\text{NO}_3)_2}{1 \text{ L}} \right) = 0.0040 \text{ mol Pb}(\text{NO}_3)_2$$

Determine limiting reagent.

Assuming  $\text{Na}_3\text{PO}_4$  is limiting:

$$0.0030 \text{ mol Na}_3\text{PO}_4 \left( \frac{1 \text{ mol Pb}_3(\text{PO}_4)_2}{2 \text{ mol Na}_3\text{PO}_4} \right) = 0.0015 \text{ mol Pb}_3(\text{PO}_4)_2$$

Assuming  $Pb(NO_3)_2$  is limiting:

$$0.0040 \text{ mol } Pb(NO_3)_2 \left( \frac{1 \text{ mol } Pb_3(PO_4)_2}{3 \text{ mol } Pb(NO_3)_2} \right) = 0.0013333 \text{ mol } Pb_3(PO_4)_2 \text{ (should be 2 s. f.)}$$

Lead(II) nitrate leads to less product and is the limiting reagent.

To determine mass of the precipitate:

$$0.0013333 \text{ mol } Pb_3(PO_4)_2 \left( \frac{811.5 \text{ g } Pb_3(PO_4)_2}{1 \text{ mol } Pb_3(PO_4)_2} \right) = 1.082026 \text{ g } Pb_3(PO_4)_2$$

**1.1 g  $Pb_3(PO_4)_2$  (with sig. figs.)**

Alternatively, you could use an IRF table to help determine more information at the same time.

	$2Na_3PO_4$	+	$3Pb(NO_3)_2$	→	$Pb_3(PO_4)_2$	+	$6NaNO_3$
<b>Initial (mol)</b>	0.0030	X	0.0040	X	0	X	0
<b>Reaction (mol)</b>	-2x	X	-3x	X	+x	X	+6x
<b>Final (mol)</b>	0.0030-2x	X	0.0040-3x	X	x	X	6x

To solve for reaction value, the limiting reagent must be completely consumed.

$$0.0040 - 3x = 0$$

$$x = 0.0013333$$

Therefore, the IRF table can be completed (values displayed to correct sig. figs.; however unrounded values will be used in future calculations):

	$2Na_3PO_4$	+	$3Pb(NO_3)_2$	→	$Pb_3(PO_4)_2$	+	$6NaNO_3$
<b>I (mol)</b>	0.0030	X	0.0040	X	0	X	0
<b>R (mol)</b>	-2x	X	-3x	X	+x	X	+6x
<b>F (mol)</b>	0.0004	X	0	X	0.0013	X	0.0080

To determine mass of the precipitate:

$$0.0013333 \text{ mol } Pb_3(PO_4)_2 \left( \frac{811.5 \text{ g } Pb_3(PO_4)_2}{1 \text{ mol } Pb_3(PO_4)_2} \right) = 1.082026 \text{ g } Pb_3(PO_4)_2$$

**1.1 g  $Pb_3(PO_4)_2$  (with sig. figs.)**

5. How many grams of precipitate are formed with 100 mL of 0.100M potassium sulfate and 100 mL of 0.200M barium nitrate? Note – it's a double displacement reaction.

Write out the complete and net ionic equations after the calculations.



1) Moles of  $K_2SO_4$ :

$$100 \text{ mL } K_2SO_4 \times \frac{0.100 \text{ mol}}{1L} \times \frac{1L}{1000 \text{ mL}} = 0.0100 \text{ mol } K_2SO_4$$

2) Moles of  $BaSO_4$  from  $K_2SO_4$ :

$$0.0100 \text{ mol } K_2SO_4 \times \frac{1 \text{ mol } BaSO_4}{1 \text{ mol } K_2SO_4} = \boxed{0.0100 \text{ mol } BaSO_4} \quad \text{Limiting reactant}$$

3) Moles of  $\text{Ba}(\text{NO}_3)_2$  :

$$100 \text{ mL Ba}(\text{NO}_3)_2 \times \frac{0.200 \text{ mol}}{1\text{L}} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.0200 \text{ mol Ba}(\text{NO}_3)_2$$

4) Moles of  $\text{BaSO}_4$  from  $\text{Ba}(\text{NO}_3)_2$  :

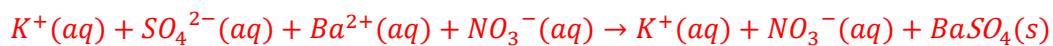
$$0.0200 \text{ mol Ba}(\text{NO}_3)_2 \times \frac{1 \text{ mol BaSO}_4}{1 \text{ mol Ba}(\text{NO}_3)_2} = 0.0200 \text{ mol BaSO}_4$$

5) Grams of  $\text{BaSO}_4$  from limiting reactant ( $\text{K}_2\text{SO}_4$ ):

$$0.0100 \text{ mol BaSO}_4 \times \frac{233.390 \text{ g BaSO}_4}{\text{mol BaSO}_4} = 2.3339 \text{ g BaSO}_4$$

Therefore 2.33 g  $\text{BaSO}_4$

Complete ionic:



Net ionic:

