

### Gravimetric Analysis, Redox and Titration Problems

- 1) When a sample of impure potassium chloride (0.4500g) was dissolved in water and treated with an excess of silver nitrate, 0.8402 g of silver chloride was precipitated. Calculate the percentage KCl in the original sample.

	NEED (+)		HAVE (-)	
	<b>KCl<sub>(aq)</sub></b>	<b>+ Ag(NO<sub>3</sub>)<sub>(aq)</sub></b>	<b>→</b>	<b>AgCl<sub>(s)</sub></b> <b>K(NO<sub>3</sub>)<sub>(aq)</sub></b>
Molarity or MM	<b>74.5 g/mol</b>			<b>143 g/mol</b>
Amount				<b>0.8402g</b>
Moles Gram/ MM or Molarity x L				<b>0.0059 mol</b>
Moles/rxn (divide moles by SC)				<b>0.0059 mol</b>
React (Least Mol/ Rxn)	<b>+0.0059 mol</b>			<b>-0.0059 mol</b>
Final Mole/ Rxn	<b>0.0059 mol</b>			<b>0</b>
Final Moles (SC x final mol/ rxn)	<b>0.0059mol</b>			<b>0</b>
Final Amt (final moles x MM) or Concentration (final mols/ total volume)	<b>0.0059 mol x 74.5 g/mol = 0.440g</b>			

$$\text{Percent Purity} = \frac{\text{Pure}}{\text{Impure}} \times 100 = \frac{0.440\text{g}}{0.450\text{ g}} \times 100 = 97.78\%$$

- 2) Calculate the percent purity of a sample of Mg(OH)<sub>2</sub> if titration of 2.568 g of the sample required 38.45 mL of 0.6695 M H<sub>3</sub>PO<sub>4</sub>.

	NEED (+)	HAVE (-)		
	<b>3 Mg(OH)<sub>2(aq)</sub></b>	<b>+ 2 H<sub>3</sub>PO<sub>4(aq)</sub></b>	<b>→</b>	<b>Mg<sub>3</sub>(PO<sub>4</sub>)<sub>2(s)</sub></b> <b>+ 6H<sub>2</sub>O<sub>(l)</sub></b>
Molarity or MM	<b>58 g/mol</b>	<b>0.6695 mol/ L</b>		
Amount	<b>?</b>	<b>0.03845 L</b>		
Moles Gram/ MM or Molarity x L		<b>0.0257 mol</b>		
Moles/rxn (divide moles by SC)		<b>0.0129 mol</b>		
React (Least Mol/ Rxn)	<b>+0.0129 mol</b>	<b>-0.0129 mol</b>		
Final Mole/ Rxn	<b>0.00129</b>	<b>0</b>		
Final Moles (SC x final mol/ rxn)	<b>0.0386 mol</b>			
Final Amt (final moles x MM) or Concentration (final mols/ total volume)	<b>2.24g</b>			

$$\text{Percent Purity} = \frac{\text{Pure}}{\text{Impure}} \times 100 = \frac{2.24\text{g}}{2.568\text{ g}} \times 100 = 87.23\%$$

3) If 19g of BaCl<sub>2</sub>(aq) is mixed with 42g of Na<sub>2</sub>SO<sub>4</sub>(aq) in 250 mL of water, what is the resulting precipitate and the mass of the precipitate and what is the concentration of all ions left in solution?

	BaCl <sub>2</sub> (aq)	+ Na <sub>2</sub> (SO <sub>4</sub> )(aq)	→	Ba(SO <sub>4</sub> )(s)	+ 2NaCl(aq)
Molarity or MM	208.3 g/mol	142 g/mol		233.4 g/mol	58.5 g/mol
Amount	19	42			
Moles Gram/ MM or Molarity x L	0.0912 mol	0.296 mol			
Moles/rxn (divide moles by SC)	0.0912 mol	0.296 mol			
React (Least Mol/ Rxn)	-0.0912 mol	-0.0912 mol		+0.0912 mol	+0.0912 mol
Final Mole/ Rxn	0	0.205 mol		0.0912 mol	0.0912 mol
Final Moles (SC x final mol/ rxn)	0	0.205 mol		0.0912 mol	0.182 mol
Final Amt (final moles x MM) or Concentration (final mols/ total volume)	0	<u>0.205 mol</u> 0.25 L = 0.82 mol Na <sub>2</sub> SO <sub>4</sub> / L		21.29g	<u>0.182 mol</u> 0.25 L = 0.73 mol NaCl / L

Mass = 21.29 g BaSO<sub>4</sub>

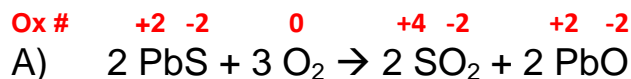
### Concentrations

Ba<sup>2+</sup> = 0 is part of the limiting reagent

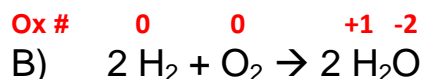
$$\text{Na}^+ = \frac{0.82 \text{ mol Na}_2\text{SO}_4}{1 \text{ L}} \times \frac{2 \text{ mol Na}^+}{1 \text{ mol Na}_2\text{SO}_4} = 0.164 \text{ M} + \frac{0.73 \text{ mol NaCl}}{1 \text{ L}} \times \frac{1 \text{ mol Na}^+}{1 \text{ mol NaCl}} = 0.073 \text{ M} = 0.237 \text{ M}$$

$$\text{Cl}^- = \frac{0.73 \text{ mol NaCl}}{1 \text{ L}} \times \frac{1 \text{ mol Cl}^-}{1 \text{ mol NaCl}} = 0.073 \text{ M}$$

4) In each of the following equations, indicate the element that has been oxidized and the one that has been reduced. You should also label the oxidation state of each before and after the process:



S goes from a -2 to a +4 which means it's getting more positive or being oxidized and is the "reducing agent"  
O goes from a 0 to a -2 which means it's getting more negative or being reduced and is the "oxidizing agent"



H goes from a 0 to a +1 which means it's getting more positive or being oxidized and is the "reducing agent"  
O goes from a 0 to a -2 which means it's getting more negative or being reduced and is the "oxidizing agent"

5) When aqueous solutions, like AgF and Sr(NO<sub>3</sub>)<sub>2</sub> combine, the precipitate, SrF<sub>2</sub> forms. Calculate the mass of the precipitate formed if 3.0L of 6.0M AgF and 12.0L of 0.10M Sr(NO<sub>3</sub>)<sub>2</sub> are mixed. Also, give the concentration of all ions in solution.

	<b>2 AgF<sub>(aq)</sub></b>	<b>+ Sr(NO<sub>3</sub>)<sub>2(aq)</sub></b>	<b>→</b>	<b>SrF<sub>2(s)</sub></b>	<b>+ 2Ag(NO<sub>3</sub>)<sub>(aq)</sub></b>
Molarity or MM	<b>6 mol/ L</b>	<b>1 mol/ L</b>		<b>125.6 g/mol</b>	
Amount	<b>3 L</b>	<b>12</b>			
Moles Gram/ MM or Molarity x L	<b>18 mol</b>	<b>12 mol</b>			
Moles/rxn (divide moles by SC)	<b>9 mol</b>	<b>12 mol</b>			
React (Least Mol/ Rxn)	<b>-9</b>	<b>-9 mol</b>		<b>+9 mol</b>	<b>+9 mol</b>
Final Mole/ Rxn	<b>0</b>	<b>3 mol</b>		<b>9 mol</b>	<b>9 mol</b>
Final Moles (SC x final mol/ rxn)	<b>0</b>	<b>3 mol</b>		<b>9 mol</b>	<b>18 mol</b>
Final Amt (final moles x MM) or Concentration (final mols/ total volume)	<b>0</b>	<b><u>3 mol</u> 15 L = 0.2 mol Sr(NO<sub>3</sub>)<sub>2</sub> / L</b>		<b>1130.4 g</b>	<b><u>18 mol</u> 15 L = 1.2 mol Ag(NO<sub>3</sub>) / L</b>

**Mass = 1130.4 g SrF<sub>2</sub>**

**Concentrations**

**F<sup>-</sup> = 0 is part of the limiting reagent**

$$\text{NO}_3^- = \frac{0.2 \text{ mol Sr(NO}_3)_2}{1 \text{ L}} \times \frac{2 \text{ mol NO}_3^-}{1 \text{ mol Sr(NO}_3)_2} = 0.4 \text{ M} + \frac{1.2 \text{ mol Ag(NO}_3)}{1 \text{ L}} \times \frac{1 \text{ mol NO}_3^-}{1 \text{ mol Ag(NO}_3)} = 1.2 \text{ M} = 1.6 \text{ M}$$

$$\text{Ag}^+ = \frac{1.2 \text{ mol Ag(NO}_3)}{1 \text{ L}} \times \frac{1 \text{ mol Ag}^+}{1 \text{ mol Ag(NO}_3)} = 1.2 \text{ M}$$