

AP CHEMISTRY

TOPIC 2: STOICHIOMETRY, PART E

(PART II)

Day 19:

Stoichiometry:

- Precipitation Reactions

- 1) Calculate the mass, in grams, of the precipitate formed when 58.3 mL of 3.58 M aluminum chlorate is mixed with an ammonium chromate solution (that is in excess). **SHOW ALL WORK**

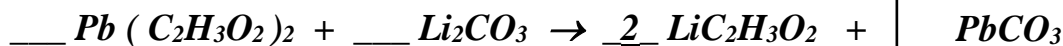
Answers:



$$\frac{58.3 \text{ mL Al}(\text{ClO}_3)_3}{1000 \text{ mL}} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{3.58 \text{ mol Al}(\text{ClO}_3)_3}{L} \times \frac{1 \text{ mol Al}_2(\text{CrO}_4)_3}{2 \text{ mol Al}(\text{ClO}_3)_3} \times \frac{401.96 \text{ g}}{1 \text{ mol Al}_2(\text{CrO}_4)_3} = 41.9 \text{ g Al}_2(\text{CrO}_4)_3$$

- 2) Calculate the moles, of the precipitate formed when 422 mL of a 13.5 M lead(II) acetate solution is mixed with a lithium carbonate solution (that is in excess). **SHOW ALL WORK**

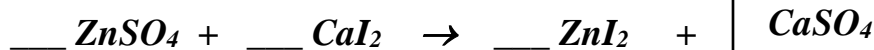
Answers:



$$\frac{422 \text{ mL}}{1000 \text{ mL}} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{13.5 \text{ mol Pb}(\text{C}_2\text{H}_3\text{O}_2)_2}{L} \times \frac{1 \text{ mol PbCO}_3}{1 \text{ mol Pb}(\text{C}_2\text{H}_3\text{O}_2)_2} = 5.70 \text{ mol PbCO}_3$$

- 3) Calculate the mass, in grams, of the precipitate formed when 3.22 liters of a 3.85 M zinc sulfate solution is mixed with 1550 mL of a 8.33 M calcium iodide solution. **SHOW ALL WORK**

Answers:



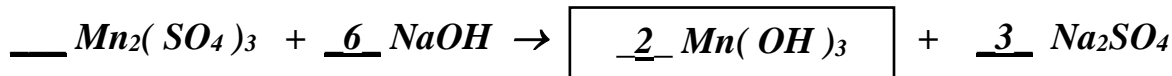
$$\frac{3.22 \text{ L}}{1 \text{ L}} \times \frac{3.85 \text{ mol ZnSO}_4}{L} \times \frac{1 \text{ mol CaI}_2}{1 \text{ mol ZnSO}_4} \times \frac{L}{8.33 \text{ mol CaI}_2} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 1488 \text{ mL CaI}_2$$

ZnSO₄ is the Limiting Reactant

$$\frac{3.22 \text{ L}}{1 \text{ L}} \times \frac{3.85 \text{ mol ZnSO}_4}{L} \times \frac{1 \text{ mol CaSO}_4}{1 \text{ mol ZnSO}_4} \times \frac{136.14 \text{ g}}{1 \text{ mol CaSO}_4} = 1.69 \times 10^3 \text{ g CaSO}_4$$

- 4) Calculate the mass, in milligrams, of the precipitate that is produced when 75.0 mL of a 9.75 M manganese(III) sulfate solution is added to 378 mL of a 6.75 M sodium hydroxide solution. **SHOW ALL WORK**

Answers:



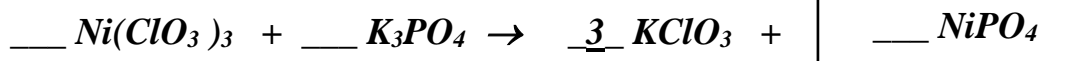
$$\frac{378 \text{ mL}}{1000 \text{ mL}} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{6.75 \text{ mol NaOH}}{\text{L}} \times \frac{1 \text{ mol Mn}_2(\text{SO}_4)_3}{6 \text{ mol NaOH}} \times \frac{\text{L}}{9.75 \text{ mol Mn}_2(\text{SO}_4)_3} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 43.6 \text{ mL Mn}_2(\text{SO}_4)_3$$

NaOH is the Limiting Reactant

$$\frac{378 \text{ mL}}{1000 \text{ mL}} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{6.75 \text{ mol NaOH}}{\text{L}} \times \frac{2 \text{ mol Mn}(\text{OH})_3}{6 \text{ mol NaOH}} \times \frac{105.9617 \text{ g}}{1 \text{ mol Mn}(\text{OH})_3} \times \frac{1000 \text{ mg}}{1 \text{ g}} = 9.01 \times 10^4 \text{ mg Mn}(\text{OH})_3$$

- 5) Calculate the mass, in grams, of the precipitate formed when 2.25 L of a nickel(III) chlorate solution is mixed with 865.3 grams potassium phosphate dissolved in enough water so that the total volume of the solution equals 3.50 liters. The density (at room temperature) for “this” nickel(III) chlorate solution is 1.236 g/cm³. **SHOW ALL WORK**

Answers:



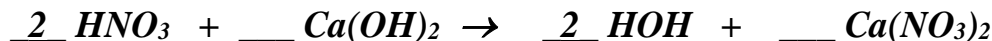
$$\frac{2.25 \text{ L}}{1 \text{ L}} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times \frac{1 \text{ cm}^3}{1 \text{ mL}} \times \frac{1.236 \text{ g Ni}(\text{ClO}_3)_3}{\text{cm}^3} \times \frac{1 \text{ mol Ni}(\text{ClO}_3)_3}{309.049} \times \frac{1 \text{ mol K}_3\text{PO}_4}{1 \text{ mol Ni}(\text{ClO}_3)_3} \times \frac{212.274 \text{ g}}{1 \text{ mol K}_3\text{PO}_4} = 1.91 \times 10^3 \text{ g K}_3\text{PO}_4$$

K₃PO₄ is the Limiting Reactant

$$\frac{865.3 \text{ g K}_3\text{PO}_4}{212.274 \text{ g}} \times \frac{1 \text{ mol K}_3\text{PO}_4}{1 \text{ mol K}_3\text{PO}_4} \times \frac{1 \text{ mol NiPO}_4}{1 \text{ mol K}_3\text{PO}_4} \times \frac{153.664 \text{ g}}{1 \text{ mol NiPO}_4} = 626 \text{ g NiPO}_4$$

- 6) Nitric acid (75.3 mL of 2.35 M) is added to 225.0 mL of 0.755 M calcium hydroxide solution. What is the concentration of the excess H⁺ or OH⁻ ions left in this solution?

Answers:



$$\frac{75.3 \text{ mL}}{1000 \text{ mL}} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{2.35 \text{ mol HNO}_3}{\text{L}} \times \frac{1 \text{ mol Ca}(\text{OH})_2}{2 \text{ mol HNO}_3} \times \frac{\text{L}}{0.755 \text{ mol Ca}(\text{OH})_2} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 117.189 \text{ mL}$$

HNO₃ is the Limiting Reactant

$$\text{Ca}(\text{OH})_2 \text{ used} = 117.189 \text{ mL} ; \text{Ca}(\text{OH})_2 \text{ remaining} = 225 \text{ mL} - 117.189 \text{ mL} = 107.81 \text{ mL}$$

$$\frac{107.81 \text{ mL}}{1000 \text{ mL}} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.755 \text{ mol Ca}(\text{OH})_2}{\text{L}} \times \frac{2 \text{ mol OH}^-}{1 \text{ mol Ca}(\text{OH})_2} = 0.162795 \text{ mol OH}^-$$

$$[\text{OH}^-] = \frac{0.162795 \text{ mol OH}^-}{0.3003 \text{ L}} = 0.542 \text{ mol OH}^-$$