

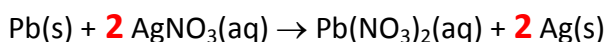
# Homework #9, Graded Answers

## Chem20, Elementary Chemistry

---

8.29) Consider the unbalanced equation for the reaction of solid lead with silver nitrate.

a.) Balance the equation.



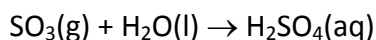
b.) How many moles of silver nitrate are required to completely react with 9.3 moles of lead?

$$9.3 \text{ mols Pb} \times \frac{2 \text{ moles AgNO}_3}{1 \text{ mole Pb}} = 18.6 \rightarrow \mathbf{19 \text{ moles AgNO}_3}$$

c.) How many moles of Ag are formed by the complete reaction of 28.4 moles Pb?

$$28.4 \text{ moles Pb} \times \frac{2 \text{ moles Ag}}{1 \text{ mol Pb}} = 56.8 \rightarrow \mathbf{56.8 \text{ moles Ag}}$$

8.62) Consider the reaction between sulfur trioxide and water. If 61.5 g SO<sub>3</sub> and 11.2 g of H<sub>2</sub>O react and 54.9 g H<sub>2</sub>SO<sub>4</sub> is collected, calculate the **limiting reactant**, **theoretical yield**, and **percent yield** for the reaction.



Calculate the molar masses for SO<sub>3</sub>, H<sub>2</sub>O, and H<sub>2</sub>SO<sub>4</sub>.

$$\text{SO}_3 = (1)(32.07 \text{ g/mol}) + (3)(16.00 \text{ g/mol}) = 80.07 \text{ g/mol}$$

$$\text{H}_2\text{O} = (2)(1.008 \text{ g/mol}) + (1)(16.00 \text{ g/mol}) = 18.016 \text{ g/mol}$$

$$\text{H}_2\text{SO}_4 = (2)(1.008 \text{ g/mol}) + (1)(32.07 \text{ g/mol}) + (4)(16.00 \text{ g/mol}) = 98.016 \text{ g/mol}$$

Convert g SO<sub>3</sub> → g H<sub>2</sub>SO<sub>4</sub>.

$$61.5 \text{ g SO}_3 \times \frac{1 \text{ mole SO}_3}{80.07 \text{ g SO}_3} \times \frac{1 \text{ mole H}_2\text{SO}_4}{1 \text{ mole SO}_3} \times \frac{98.016 \text{ g H}_2\text{SO}_4}{1 \text{ mole H}_2\text{SO}_4} = 75.28 \text{ g H}_2\text{SO}_4$$

Convert g H<sub>2</sub>O → g H<sub>2</sub>SO<sub>4</sub>

$$11.2 \text{ g H}_2\text{O} \times \frac{1 \text{ mole H}_2\text{O}}{18.016 \text{ g H}_2\text{O}} \times \frac{1 \text{ mole H}_2\text{SO}_4}{1 \text{ mole H}_2\text{O}} \times \frac{98.016 \text{ g H}_2\text{SO}_4}{1 \text{ mole H}_2\text{SO}_4} = 60.98 \text{ g H}_2\text{SO}_4$$

Since H<sub>2</sub>O produces fewer g product

**Limiting Reactant: H<sub>2</sub>O, Theoretical Yield: 61.0 g H<sub>2</sub>SO<sub>4</sub>**

$$\text{percent yield} = \frac{54.9 \text{ g}}{61.0 \text{ g}} \times 100 = 90.0 \rightarrow \mathbf{90.0 \% \text{ yield}}$$

8.69) Classify each process as exothermic or endothermic, and indicate the sign of ΔH.

a.) Butane gas burning in a lighter → **exothermic, -ΔH** (releases heat)

b.) The reaction that occurs in the chemical cold packs used to ice athletic injuries – **endothermic, +ΔH** (absorbs heat)

c.) The burning of wax in a candle → **exothermic, -ΔH** (releases heat)

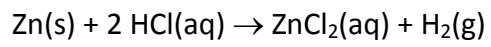
8.76) The evaporation of water is endothermic:



What minimum mass of water (in g) has to evaporate to absorb 175 kJ of heat?

$$175 \text{ kJ} \times \frac{1 \text{ mole H}_2\text{O}}{44.01 \text{ kJ}} \times \frac{18.016 \text{ g H}_2\text{O}}{1 \text{ mole H}_2\text{O}} = 71.63 \rightarrow \mathbf{71.6 \text{ g H}_2\text{O}}$$

8.87) Hydrogen gas can be prepared in the laboratory by a single-displacement reaction in which solid zinc reacts with hydrochloric acid. How much zinc in grams is required to make 14.5 g of hydrogen gas through this reaction?



$$14.5 \text{ g H}_2 \times \frac{1 \text{ mole H}_2}{2.016 \text{ g H}_2} \times \frac{1 \text{ mole Zn}}{1 \text{ mole H}_2} \times \frac{65.39 \text{ g Zn}}{1 \text{ mole Zn}} = 470.3 \rightarrow \mathbf{4.70 \times 10^2 \text{ g Zn}}$$