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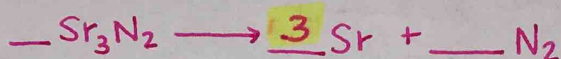
Per \_\_\_\_\_

## Percent Yield

Practice Sheet #

1. Strontium nitride can be decomposed into strontium and nitrogen.

a. Write a balanced equation for this reaction.



b. If 58.2 g of strontium nitride yields 45.0 g of strontium, determine the percent yield for the reaction.

$$58.2 \text{ g Sr}_3\text{N}_2 \times \frac{1 \text{ mol Sr}_3\text{N}_2}{290.88 \text{ g Sr}_3\text{N}_2} \times \frac{3 \text{ mol Sr}}{1 \text{ mol Sr}_3\text{N}_2} \times \frac{87.62 \text{ g Sr}}{1 \text{ mol Sr}} = \boxed{52.6 \text{ g Sr}}$$

theoretical yield.

$$\frac{\text{actual}}{\text{theoretical}} \times 100 = \% \text{ yield} \quad \frac{45.0 \text{ g}}{52.6 \text{ g}} \times 100 = \boxed{85.6\%}$$

c. What mass of nitrogen will actually be obtained?

$$45.0 \text{ g Sr} \times \frac{1 \text{ mol Sr}}{87.62 \text{ g Sr}} \times \frac{1 \text{ mol N}_2}{3 \text{ mol Sr}} \times \frac{28.02 \text{ g N}_2}{1 \text{ mol N}_2} = \boxed{4.80 \text{ g N}_2}$$

2. Mercury (II) oxide can be decomposed into mercury and oxygen.

a. Write a balanced chemical equation for this reaction.



b. If 54.15 g of mercury (II) oxide yields 3.250 g of oxygen, determine the percent yield for the reaction.

$$54.15 \text{ g HgO} \times \frac{1 \text{ mol HgO}}{216.6 \text{ g HgO}} \times \frac{1 \text{ mol O}_2}{2 \text{ mol HgO}} \times \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} = \boxed{4.000 \text{ g O}_2}$$

theoretical yield

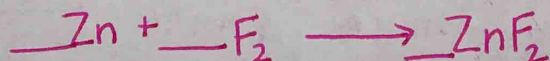
$$\frac{3.250 \text{ g}}{4.000 \text{ g}} \times 100 = \boxed{81.3\%}$$

c. What mass of mercury will actually be obtained?

$$3.250 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \times \frac{2 \text{ mol Hg}}{1 \text{ mol O}_2} \times \frac{200.6 \text{ g Hg}}{1 \text{ mol Hg}} = \boxed{40.75 \text{ g Hg}}$$

3. Zinc and fluorine can be combined to form zinc fluoride.

a. Write a balanced chemical equation for this reaction.



- b. If 16.35 g of zinc yields 16.50 g of zinc fluoride, determine the percent yield of this reaction.

$$16.35 \text{ g Zn} \times \frac{1 \text{ mol Zn}}{65.38 \text{ g Zn}} \times \frac{1 \text{ mol ZnF}_2}{1 \text{ mol Zn}} \times \frac{103.38 \text{ g ZnF}_2}{1 \text{ mol ZnF}_2} = \boxed{25.85 \text{ g ZnF}_2} \text{ theoretical yield}$$

$$\frac{16.50 \text{ g}}{25.85 \text{ g}} \times 100 = \boxed{63.8\%}$$

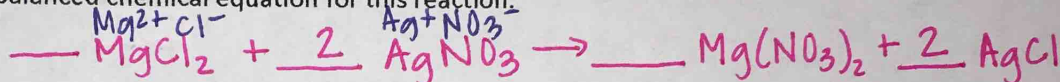
- c. If 0.380 g of fluorine reacts with a percent yield of 60.0%, what mass of zinc fluoride will actually be obtained?

$$0.380 \text{ g F}_2 \times \frac{1 \text{ mol F}_2}{38.00 \text{ g F}_2} \times \frac{1 \text{ mol ZnF}_2}{1 \text{ mol F}_2} \times \frac{103.38 \text{ g ZnF}_2}{1 \text{ mol ZnF}_2} = \boxed{1.03 \text{ g ZnF}_2}$$

$$\frac{x}{1.03 \text{ g}} = 0.60 \quad x = \text{actual mass of ZnF}_2 = \boxed{0.618 \text{ g}}$$

4. Magnesium chloride reacts with silver nitrate.

- a. Write a balanced chemical equation for this reaction.



- b. If 21.2 g of silver nitrate react with a percent yield of 80.0%, what mass would actually be obtained for each of the products?

$$21.2 \text{ g AgNO}_3 \times \frac{1 \text{ mol AgNO}_3}{169.9 \text{ g AgNO}_3} \times \frac{1 \text{ mol Mg(NO}_3)_2}{2 \text{ mol AgNO}_3} \times \frac{148.31 \text{ g Mg(NO}_3)_2}{1 \text{ mol Mg(NO}_3)_2} = 9.25 \text{ g}$$

$$\text{"} \times \frac{2 \text{ mol AgCl}}{2 \text{ mol AgNO}_3} \times \frac{143.35 \text{ g AgCl}}{1 \text{ mol AgCl}} = 17.9 \text{ g}$$

$$\text{"} \times 0.8 = \boxed{7.4 \text{ g}}$$

- c. If 476 g of magnesium chloride yields 400 g of magnesium nitrate, determine the percent yield of the reaction. What mass of silver chloride would actually be obtained?

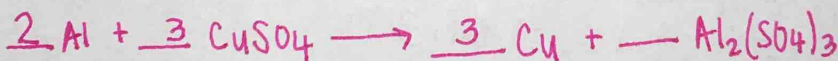
$$476 \text{ g MgCl}_2 \times \frac{1 \text{ mol MgCl}_2}{95.21 \text{ g MgCl}_2} \times \frac{1 \text{ mol Mg(NO}_3)_2}{1 \text{ mol MgCl}_2} \times \frac{148.33 \text{ g Mg(NO}_3)_2}{1 \text{ mol Mg(NO}_3)_2} = 742 \text{ g theoretical}$$

$$\frac{400 \text{ g}}{742 \text{ g}} \times 100 = \boxed{53.9\%}$$

$$400 \text{ Mg(NO}_3)_2 \times \frac{1 \text{ mol}}{148.33 \text{ g}} \times \frac{2 \text{ mol AgCl}}{1 \text{ mol}} \times \frac{143.35 \text{ g AgCl}}{1 \text{ mol AgCl}} = 773 \text{ g AgCl}$$

5. Aluminum is reacted with copper (II) sulfate.

- a. Write a balanced chemical equation for this reaction.



- b. If 5.60 g of aluminum react, determine the mass of copper (II) sulfate required.

$$5.60 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \times \frac{3 \text{ mol CuSO}_4}{2 \text{ mol Al}} \times \frac{159.61 \text{ g CuSO}_4}{1 \text{ mol CuSO}_4} = \boxed{49.7 \text{ g CuSO}_4}$$

- c. If 18.0 g of copper are actually produced, determine the percent yield of the reaction. What mass of aluminum sulfate will actually be obtained?

$$5.60 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \times \frac{3 \text{ mol Cu}}{2 \text{ mol Al}} \times \frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}} = \boxed{19.8 \text{ g Cu theoretical}}$$

$$\rightarrow \frac{18.0 \text{ g}}{19.8 \text{ g}} \times 100 = \boxed{90.9\% \text{ yield}}$$

$$18.0 \text{ g Cu} \times \frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}} \times \frac{1 \text{ mol Al}_2(\text{SO}_4)_3}{3 \text{ mol Cu}} \times \frac{342.14 \text{ g Al}_2(\text{SO}_4)_3}{1 \text{ mol Al}_2(\text{SO}_4)_3} = \boxed{32.3 \text{ g Al}_2(\text{SO}_4)_3}$$