Name: KEY! Per____

Percent Yield

Practice Sheet #

1. Strontium nitride can be decomposed into strontium and nitrogen.
   a. Write a balanced equation for this reaction.

$$\_\,Sr_3N_2 \longrightarrow \underline{3}\,Sr + \_\_\,N_2$$

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

$$58.2\,g\ Sr_3N_2 \times \frac{1\,mol\ Sr_3N_2}{290.88\,g\ Sr_3N_2} \times \frac{3\,mol\ Sr}{1\,mol\ Sr_3N_2} \times \frac{87.62\,g\ Sr}{1\,mol\ Sr} = \boxed{52.6\,g\ Sr}$$

theoretical yield.

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

   c. What mass of nitrogen will actually be obtained?

$$45.0\,g\ Sr \times \frac{1\,mol\ Sr}{87.62\,g\ Sr} \times \frac{1\,mol\ N_2}{3\,mol\ Sr} \times \frac{28.02\,g\ N_2}{1\,mol\ N_2} = \boxed{4.80\,g\ N_2}$$

2. Mercury (II) oxide can be decomposed into mercury and oxygen.
   a. Write a balanced chemical equation for this reaction.

$$\underline{2}\,HgO \longrightarrow \underline{2}\,Hg + \_\_\,O_2$$

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

$$54.15\,g\ HgO \times \frac{1\,mol\ HgO}{216.6\,g\ HgO} \times \frac{1\,mol\ O_2}{2\,mol\ HgO} \times \frac{32.00\,g\ O_2}{1\,mol\ O_2} = \boxed{4.000\,g\ O_2}$$

theoretical yield

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

   c. What mass of mercury will actually be obtained?

$$3.250\,g\ O_2 \times \frac{1\,mol\ O_2}{32.00\,g\ O_2} \times \frac{2\,mol\ Hg}{1\,mol\ O_2} \times \frac{200.6\,g\ Hg}{1\,mol\ Hg} = \boxed{40.75\,g\ Hg}$$

3. Zinc and fluorine can be combined to form zinc fluoride.
   a. Write a balanced chemical equation for this reaction.

$$\_\_\,Zn + \_\_\,F_2 \longrightarrow \_\_\,ZnF_2$$

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

$$16.35 g\ Zn \times \frac{1 mol\ Zn}{65.38 g\ Zn} \times \frac{1 mol\ ZnF_2}{1 mol\ Zn} \times \frac{103.38 g\ ZnF_2}{1 mol\ ZnF_2} = \boxed{25.85 g\ ZnF_2}$$ theoretical yield

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

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

$$0.380 g\ F_2 \times \frac{1 mol\ F_2}{38.00 g\ F_2} \times \frac{1 mol\ ZnF_2}{1 mol\ F_2} \times \frac{103.38 g\ ZnF_2}{1 mol\ ZnF_2} = \boxed{1.03 g\ ZnF_2}$$

$$\frac{x}{1.03 g} = 0.60$$ x = actual mass of ZnF_2 = $\boxed{0.618 g}$

4. Magnesium chloride reacts with silver nitrate.

a. Write a balanced chemical equation for this reaction.

$Mg^{2+}\ Cl^-$ $Ag^+\ NO_3^-$

$$\_\_ MgCl_2 + \underline{2}\ AgNO_3 \rightarrow \_\_ Mg(NO_3)_2 + \underline{2}\ AgCl$$

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?

$$21.2 g\ AgNO_3 \times \frac{1 mol\ AgNO_3}{169.9 g\ AgNO_3} \times \frac{1 mol\ Mg(NO_3)_2}{2 mol\ AgNO_3} \times \frac{148.31 g\ Mg(NO_3)_2}{1 mol\ Mg(NO_3)_2} = 9.25 g$$

× 0.8 = [illegible]

" " $$\times \frac{2 mol\ AgCl}{2 mol\ AgNO_3} \times \frac{143.35 g\ AgCl}{1 mol\ AgCl} = 17.9 g$$

× 0.80 = $\boxed{14.3 g}$

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

$$476 g\ MgCl_2 \times \frac{1 mol\ MgCl_2}{95.21 g\ MgCl_2} \times \frac{1 mol\ Mg(NO_3)_2}{1 mol\ MgCl_2} \times \frac{148.33 g\ Mg(NO_3)_2}{1 mol\ Mg(NO_3)_2} = 742 g$$ theoretical

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

$$400\ Mg(NO_3)_2 \times \frac{1 mol}{148.33 g} \times \frac{2 mol\ AgCl}{1 mol} \times \frac{143.35 g\ AgCl}{1 mol\ AgCl} = \boxed{773 g\ AgCl}$$

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

a. Write a balanced chemical equation for this reaction.

$$\underline{2}\ Al + \underline{3}\ CuSO_4 \rightarrow \underline{3}\ Cu + \_\_ Al_2(SO_4)_3$$

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

$$5.60 g\ Al \times \frac{1 mol\ Al}{26.98 g\ Al} \times \frac{3 mol\ CuSO_4}{2 mol\ Al} \times \frac{159.61 g\ CuSO_4}{1 mol\ CuSO_4} = \boxed{49.7 g\ CuSO_4}$$

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?

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

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

$$18.0 g\ Cu \times \frac{1 mol\ Cu}{63.55 g\ Cu} \times \frac{1 mol\ Al_2(SO_4)_3}{3 mol\ Cu} \times \frac{342.14 g\ Al_2(SO_4)_3}{1 mol\ Al_2(SO_4)_3} = \boxed{32.3 g\ Al_2(SO_4)_3}$$