Name ____________________________ Section __________ Date __________

# Chapter 9: Basic Review Worksheet

1. Considering the reaction represented by the (unbalanced) equation

$$N_2(g) + H_2(g) \rightarrow NH_3(g)$$

determine the number of moles of $NH_3(g)$ that can be produced from the following:
a. 0.20 mol $N_2(g)$ reacts completely with $H_2(g)$.
b. 0.30 mol $H_2(g)$ reacts completely with $N_2(g)$.

2. Considering the reaction represented by the (unbalanced) equation

$$Mg(s) + HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

determine the mass of $H_2(g)$ that can be produced from the following:
a. 10.0 g Mg(s) reacts completely with HCl(aq).
b. 20.0 g HCl (aq) reacts completely with Mg(s).

3. What is meant by a *limiting reactant* in a particular reaction? What does it mean to say that one or more of the reactants are present *in excess*?

4. Considering the reaction represented by the (unbalanced) equation

$$H_2(g) + O_2(g) \rightarrow H_2O(l)$$

determine the limiting reactant in each case.
a. 4.0 mol $H_2(g)$ reacts with 3.0 mol $O_2(g)$
b. 10.0 g $H_2(g)$ reacts with 10.0 g $O_2(g)$
c. 10.0 mol $H_2(g)$ reacts with 10.0 mol $O_2(g)$
d. 5.0 g $H_2(g)$ reacts with 30.0 g $O_2(g)$

5. What do we mean by the *theoretical yield* for a reaction? What is meant by the *actual yield*?

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# Chapter 9: Review Worksheet

1. Balanced chemical equations give us information on the molecular level (individual molecules reacting in the proportions indicated by the coefficients), and also on the macroscopic level (moles). Write a balanced chemical equation of your choice, and interpret in words the meaning of the equation on the molecular and macroscopic levels.

2. Consider the *unbalanced* equation for the combustion of ethyl alcohol, $C_2H_5OH$:

$$C_2H_5OH(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$

For a given amount of ethyl alcohol, write the mole ratios that would enable you to calculate the number of moles of each product, as well as the number of moles of $O_2$ that would be required. Show how these mole ratios would be applied if 0.65 mol of ethyl alcohol is combusted.

3. When a limiting reactant is present, in what way is the reaction "limited"? What happens to a reaction when the limiting reactant is used up?

4. For each of the following balanced equations, calculate the mass of each product formed if 25.0 g of the reactant listed first reacts completely with the second.
   a. $2AgNO_3(aq) + CaSO_4(aq) \rightarrow Ag_2SO_4(s) + Ca(NO_3)_2(aq)$
   b. $2Al(s) + 6HNO_3(aq) \rightarrow 2Al(NO_3)_3(aq) + 3H_2(g)$
   c. $H_3PO_4(aq) + 3NaOH(aq) \rightarrow Na_3PO_4(aq) + 3H_2O(l)$
   d. $CaO(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l)$

5. For the reactions in Question 4, calculate the mass of each product formed if 12.5 g of the first reactant is combined with 10.0 g of the second reactant. Indicate which substance is the limiting reactant for each case.

6. Look at your answers to question 5. Is there a pattern to which reactant is limiting? That is, is the limiting reactant always the one that is present with the lowest mass in grams? Is it always the one that is present with the least number of moles? Explain.

7. In a problem, how do we determine the *theoretical yield*? Where do we get the *actual yield*? How do we use these to calculate the *percent yield*?

8. You have calculated the theoretical yield for a reaction to be 4.0 g Cu(s). You collect 2.8 g Cu(s) in the lab. Determine your percent yield.

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# Chapter 9: Challenge Review Worksheet

1. In the practice of chemistry one of the most important calculations concerns the masses of products expected when particular masses of reactants are used in an experiment. For example, chemists judge the practicality and efficiency of a reaction by seeing how close the amount of product actually obtained is to the expected amount. Using a balanced chemical equation and an amount of starting material of your choice, summarize and illustrate the various steps needed in such a calculation for the expected amount of product.

2. For a balanced chemical equation of your choice, and using 25.0 g of each of the reactants in your equation, illustrate and explain how you would determine which reactant is the limiting reactant. Indicate *clearly* in your discussion how the choice of limiting reactant follows from your calculations.

3. Chlorine gas is a very reactive substance and will combine with most metals. For example,

   $2K(s) + Cl_2(g) \rightarrow 2KCl(s)$
   $Ca(s) + Cl_2(g) \rightarrow CaCl_2(s)$
   $2Al(s) + 3Cl_2(g) \rightarrow 2AlCl_3(s)$

   Suppose individual 25.0-g samples of these three metals are reacted with separate 50.0-g samples of $Cl_2(g)$. In each case, determine whether the metal or chlorine is the limiting reactant, and calculate the theoretical yield of metal chloride for each process.

4. Suppose you run the reaction between potassium and chlorine (with the amounts given in problem 3) and you collect 31.2 g of potassium chloride. Determine your percent yield.

5. Your teacher gives you 5.00 g of a mixture of the two salts silver nitrate and potassium nitrate and asks you to determine the percent silver nitrate by mass in the mixture. You dissolve the mixture in water and add an excess of aqueous sodium chloride. You collect and dry the white solid that precipitates and find that it has a mass of 1.48 g.

   Provide balanced equations for all reactions that occur in this process and determine the percent silver nitrate by mass in the original mixture.

# Chapter 9: Basic Review Worksheet

1. The balanced equation is $N_2 + 3H_2 \rightarrow 2NH_3$

   a. $0.20 \text{ mol } N_2 \times \dfrac{2 \text{ mol } NH_3}{1 \text{ mol } N_2} = 0.40 \text{ mol } NH_3$

   b. $0.30 \text{ mol } H_2 \times \dfrac{2 \text{ mol } NH_3}{3 \text{ mol } H_2} = 0.20 \text{ mol } NH_3$

2. The balanced equation is $Mg + 2HCl \rightarrow MgCl_2 + H_2$

   a. $10.0\text{g } Mg \times \dfrac{1 \text{ mol } Mg}{24.31\text{g } Mg} \times \dfrac{1 \text{ mol } H_2}{1 \text{ mol } Mg} \times \dfrac{2.016 \text{ } H_2}{1 \text{ mol } H_2} = 0.829\text{g } H_2$

   b. $20.0\text{g } HCl \times \dfrac{1 \text{ mol } HCl}{36.458\text{g } HCl} \times \dfrac{1 \text{ mol } H_2}{2 \text{ mol } HCl} \times \dfrac{2.016 \text{ } H_2}{1 \text{ mol } H_2} = 0.553\text{g } H_2$

3. Although we can calculate specifically the exact amounts of each reactant needed for a chemical reaction, oftentimes reaction mixtures are prepared using more or less arbitrary amounts of the reagents. However, regardless of how much of each reagent may be used for a reaction, the substances still react stoichiometrically, according to the mole ratios derived from the balanced chemical equation for the reaction. When arbitrary amounts of reactants are used, there will be one reactant which, stoichiometrically, is present in the least amount. This substance is called the *limiting* reactant for the experiment. We say that the other reactants in the experiment are present in the excess, which means that a portion of these reactants will still be present unchanged after the reaction has ended and the limiting reactant has been used up completely.

4. The balanced equation is $2H_2 + O_2 \rightarrow 2H_2O$

   a. $4.0 \text{ mol } H_2 \times \dfrac{1 \text{ mol } O_2}{2 \text{ mol } H_2} = 2.0 \text{ mol } O_2$; $H_2$ is limiting

   b. $10.0\text{g } H_2 \times \dfrac{1 \text{ mol } H_2}{2.016\text{g } H_2} \times \dfrac{1 \text{ mol } O_2}{2 \text{ mol } H_2} \times \dfrac{32.00\text{g } O_2}{1 \text{ mol } O_2} = 79.4\text{g } O_2$; $O_2$ is limiting

   c. $10.0 \text{ mol } H_2 \times \dfrac{1 \text{ mol } O_2}{2 \text{ mol } H_2} = 5.00 \text{ mol } O_2$; $H_2$ is limiting

   d. $5.0\text{g } H_2 \times \dfrac{1 \text{ mol } H_2}{2.016\text{g } H_2} \times \dfrac{1 \text{ mol } O_2}{2 \text{ mol } H_2} \times \dfrac{32.00\text{g } O_2}{1 \text{ mol } O_2} = 40.\text{g } O_2$; $O_2$ is limiting

5. The theoretical yield for an experiment is the mass of product calculated based on the limiting reactant for the experiment being completely consumed. The actual yield for an experiment is the mass of product actually collected by the experimenter.

# Chapter 9: Review Worksheet

1. Answers will vary. An example is included below.

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$

which describes the decomposition reaction of hydrogen peroxide.

Microscopic: Two molecules of hydrogen peroxide (in aqueous solution) decompose to produce two molecules of liquid water and one molecule of oxygen gas.

Macroscopic: Two moles of hydrogen peroxide (present in aqueous solution) decompose to produce two moles of liquid water and one mole of oxygen gas.

2. The mole ratios for a reaction are based on the coefficients of the balanced chemical equation for the reaction. These coefficients show in what proportions molecules (or moles of molecules) combine. From the balanced equation

$$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$$

(and assuming a given amount of $C_2H_5OH$) various mole ratios can be constructed.

to calculate mol $CO_2$ produced: $\dfrac{2\text{ mol } CO_2}{1\text{ mol } C_2H_5OH}$

to calculate mol $H_2O$ produced: $\dfrac{3\text{ mol } H_2O}{1\text{ mol } C_2H_5OH}$

to calculate mol $O_2$ required: $\dfrac{3\text{ mol } O_2}{1\text{ mol } C_2H_5OH}$

We could calculate the number of moles of the other substances if 0.65 mol of $C_2H_5OH$ were to be combusted as follows:

$$\text{mol } CO_2 \text{ produced} = (0.65\text{ mol } C_2H_5OH) \times \frac{2\text{ mol } CO_2}{1\text{ mol } C_2H_5OH} = 1.3\text{ mol } CO_2$$

$$\text{mol } H_2O \text{ produced} = (0.65\text{ mol } C_2H_5OH) \times \frac{3\text{ mol } H_2O}{1\text{ mol } C_2H_5OH} = 1.95 = 2.0\text{ mol } H_2O$$

$$\text{mol } O_2 \text{ required} = (0.65\text{ mol } C_2H_5OH) \times \frac{3\text{ mol } O_2}{1\text{ mol } C_2H_5OH} = 1.95 = 2.0\text{ mol } O_2$$

3. It is the limiting reactant that controls how much product is formed, regardless of how much of the other reactants are present. The limiting reactant limits the amount of product that can form in the experiment, because once the limiting reactant has reacted completely, the reaction must stop.

4. a. $2AgNO_3(aq) + CaSO_4(aq) \rightarrow Ag_2SO_4(s) + Ca(NO_3)_2(aq)$

Molar masses: $AgNO_3$, 169.9 g; $Ag_2SO_4$, 311.9 g; $Ca(NO_3)_2$, 164.1 g

$$25.0\text{ g } AgNO_3 \times \frac{1\text{ mol}}{169.9\text{ g}} = 0.147\text{ mol } AgNO_3$$

$$0.147\text{ mol } AgNO_3 \times \frac{1\text{ mol } Ag_2SO_4}{2\text{ mol } AgNO_3} = 0.0735\text{ mol } Ag_2SO_4$$

$$0.0735\text{ mol } Ag_2SO_4 \times \frac{311.9\text{ g}}{1\text{ mol}} = 22.9\text{ g } Ag_2SO_4$$

$$0.147\text{ mol } AgNO_3 \times \frac{1\text{ mol } Ca(NO_3)_2}{2\text{ mol } AgNO_3} = 0.0735\text{ mol } Ca(NO_3)_2$$

$$0.0735\text{ mol } Ca(NO_3)_2 \times \frac{164.1\text{ g}}{1\text{ mol}} = 12.1\text{ g } Ca(NO_3)_2$$

b. $2Al(s) + 6HNO_3(aq) \rightarrow 2Al(NO_3)_3(aq) + 3H_2(g)$

Molar masses: Al, 26.98 g; $Al(NO_3)_3$, 213.0 g; $H_2$, 2.016 g

$$25.0\text{ g Al} \times \frac{1\text{ mol}}{26.98\text{ g}} = 0.927\text{ mol Al}$$

$$0.927\text{ mol Al} \times \frac{2\text{ mol } Al(NO_3)_3}{2\text{ mol Al}} = 0.927\text{ mol } Al(NO_3)_3$$

$$0.927\text{ mol } Al(NO_3)_3 \times \frac{213.0\text{ g}}{1\text{ mol}} = 197\text{ g } Al(NO_3)_3$$

$$0.927\text{ mol Al} \times \frac{3\text{ mol } H_2}{2\text{ mol Al}} = 1.39\text{ mol } H_2$$

$$1.39\text{ mol } H_2 \times \frac{2.016\text{ g}}{1\text{ mol}} = 2.80\text{ g } H_2$$

c. $H_3PO_4(aq) + 3NaOH(aq) \rightarrow Na_3PO_4(aq) + 3H_2O(l)$

Molar masses: $H_3PO_4$, 97.99 g; $Na_3PO_4$, 163.9 g; $H_2O$, 18.02 g

$$25.0 \text{ g } H_3PO_4 \times \frac{1 \text{ mol}}{97.99 \text{ g}} = 0.255 \text{ mol } H_3PO_4$$

$$0.255 \text{ mol } H_3PO_4 \times \frac{1 \text{ mol } Na_3PO_4}{1 \text{ mol } H_3PO_4} = 0.255 \text{ mol } Na_3PO_4$$

$$0.255 \text{ mol } Na_3PO_4 \times \frac{163.9 \text{ g}}{1 \text{ mol}} = 41.8 \text{ g } Na_3PO_4$$

$$0.255 \text{ mol } H_3PO_4 \times \frac{3 \text{ mol } H_2O}{1 \text{ mol } H_3PO_4} = 0.765 \text{ mol } H_2O$$

$$0.765 \text{ mol } H_2O \times \frac{18.02 \text{ g}}{1 \text{ mol}} = 13.8 \text{ g } H_2O$$

d. $CaO(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l)$

Molar masses: CaO, 56.08 g; $CaCl_2$, 111.0 g; $H_2O$, 18.02 g

$$25.0 \text{ g CaO} \times \frac{1 \text{ mol}}{56.08 \text{ g}} = 0.446 \text{ mol CaO}$$

$$0.446 \text{ mol CaO} \times \frac{1 \text{ mol } CaCl_2}{1 \text{ mol CaO}} = 0.446 \text{ mol } CaCl_2$$

$$0.446 \text{ mol } CaCl_2 \times \frac{111.0 \text{ g}}{1 \text{ mol}} = 49.5 \text{ g } CaCl_2$$

$$0.446 \text{ mol CaO} \times \frac{1 \text{ mol } H_2O}{1 \text{ mol CaO}} = 0.446 \text{ mol } H_2O$$

$$0.446 \text{ mol } H_2O \times \frac{18.02 \text{ g}}{1 \text{ mol}} = 8.04 \text{ g } H_2O$$

5. a. $2AgNO_3(aq) + CaSO_4(aq) \rightarrow Ag_2SO_4(s) + Ca(NO_3)_2(aq)$

Molar masses: $AgNO_3$, 169.9 g; $CaSO_4$, 136.2 g
$Ag_2SO_4$, 311.9 g; $Ca(NO_3)_2$, 164.1 g

$$12.5 \text{ g } AgNO_3 \times \frac{1 \text{ mol}}{169.9 \text{ g}} = 0.0736 \text{ mol } AgNO_3$$

$$10.0 \text{ g } CaSO_4 \times \frac{1 \text{ mol}}{136.2 \text{ g}} = 0.0734 \text{ mol } CaSO_4$$

$AgNO_3$ is the limiting reactant

$$0.0736 \text{ mol } AgNO_3 \times \frac{1 \text{ mol } Ag_2SO_4}{2 \text{ mol } AgNO_3} \times \frac{311.9 \text{ g } Ag_2SO_4}{1 \text{ mol } Ag_2SO_4} = 11.5 \text{ g } Ag_2SO_4$$

$$0.0736 \text{ mol } AgNO_3 \times \frac{1 \text{ mol } Ca(NO_3)_2}{2 \text{ mol } AgNO_3} \times \frac{164.1 \text{ g}}{1 \text{ mol}} = 6.04 \text{ g } Ca(NO_3)_2$$

b. $2Al(s) + 6HNO_3(aq) \rightarrow Al(NO_3)_3(aq) + 3H_2(g)$

Molar masses: Al, 26.98 g; $HNO_3$, 63.02 g, $Al(NO_3)_3$, 213.0 g; $H_2$, 2.016 g

$$12.5 \text{ g Al} \times \frac{1 \text{ mol}}{26.98 \text{ g}} = 0.463 \text{ mol Al}$$

$$10.0 \text{ g } HNO_3 \times \frac{1 \text{ mol}}{63.02 \text{ g}} = 0.159 \text{ mol } HNO_3$$

$HNO_3$ is the limiting reactant

$$0.159 \text{ mol } HNO_3 \times \frac{2 \text{ mol } Al(NO_3)_3}{6 \text{ mol } HNO_3} \times \frac{213.0 \text{ g}}{1 \text{ mol}} = 11.3 \text{ g } Al(NO_3)_3$$

$$0.159 \text{ mol } HNO_3 \times \frac{3 \text{ mol } H_2}{6 \text{ mol } HNO_3} \times \frac{2.016 \text{ g}}{1 \text{ mol}} = 0.160 \text{ g } H_2$$

c. $H_3PO_4(aq) + 3NaOH(aq) \rightarrow Na_3PO_4(aq) + 3H_2O(l)$

Molar masses: $H_3PO_4$, 97.99 g; NaOH, 40.00 g, $Na_3PO_4$, 163.9 g; $H_2O$, 18.02 g

$$12.5 \text{ g } H_3PO_4 \times \frac{1 \text{ mol}}{97.99 \text{ g}} = 0.128 \text{ mol } H_3PO_4$$

$$10.0 \text{ g NaOH} \times \frac{1 \text{ mol}}{40.00 \text{ g}} = 0.250 \text{ mol NaOH}$$

NaOH is the limiting reactant

$$0.250 \text{ mol NaOH} \times \frac{1 \text{ mol } Na_3PO_4}{3 \text{ mol NaOH}} \times \frac{163.9 \text{ g}}{1 \text{ mol}} = 13.7 \text{ g } Na_3PO_4$$

$$0.250 \text{ mol NaOH} \times \frac{3 \text{ mol } H_2O}{3 \text{ mol NaOH}} \times \frac{18.02 \text{ g}}{1 \text{ mol}} = 4.51 \text{ g } H_2O$$

d. $CaO(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l)$

Molar masses: $CaO$, 56.08 g; $HCl$, 36.46 g, $CaCl_2$, 111.0 g; $H_2O$, 18.02 g

$$12.5 \text{ g CaO} \times \frac{1 \text{ mol}}{56.08 \text{ g}} = 0.222 \text{ mol CaO}$$

$$10.0 \text{ g HCl} \times \frac{1 \text{ mol}}{36.458 \text{ g}} = 0.274 \text{ mol HCl}$$

Since twice as many moles of HCl (compared to CaO) are required, HCl is the limiting reactant.

$$0.274 \text{ mol HCl} \times \frac{1 \text{ mol } CaCl_2}{2 \text{ mol HCl}} \times \frac{111.0 \text{ g}}{1 \text{ mol}} = 15.2 \text{ g } CaCl_2$$

$$0.274 \text{ mol HCl} \times \frac{1 \text{ mol } H_2O}{2 \text{ mol HCl}} \times \frac{18.02 \text{ g}}{1 \text{ mol}} = 2.47 \text{ g } H_2O$$

6. There is no pattern. For example, in part "a", $AgNO_3$ is present with the highest mass and greatest number of moles and it is the limiting reactant. In part "b", $HNO_3$ limits the reaction and is present in the least amount (of mass and moles). Students need to understand that they must figure out the limiting reactant, not just memorize an incorrect short cut such as "the limiting reactant is present in least amount".

7. We determine the theoretical yield by stoichiometric calculations. The actual yield is determined by experiment. The percent yield is calculated by taking the actual yield, dividing by the theoretical yield, and multiplying this number by 100%.

8. (2.8 g)/(4.0g) x 100% = 70.% yield.

## Chapter 9: Challenge Review Worksheet

1. Answers will vary. An example is provided showing the decomposition of calcium carbonate, producing calcium oxide and carbon dioxide.

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

Let's suppose that 50.0 g of $CaCO_3$ is to be decomposed.

 

Molar masses: $CaCO_3$, 100.09 g; CaO, 56.08; $CO_2$, 44.01 g

$$\text{mol } CaCO_3 = 50.0 \text{ g} \times \frac{1 \text{ mol}}{100.09 \text{ g}} = 0.4995 \text{ mol } CaCO_3$$

$$\text{mol CaO} = 0.4995 \text{ mol } CaCO_3 \times \frac{1 \text{ mol CaO}}{1 \text{ mol } CaCO_3} = 0.4995 \text{ mol CaO}$$

$$\text{mass CaO} = 0.4995 \text{ mol CaO} \times \frac{56.08 \text{ g}}{1 \text{ mol}} = 28.0 \text{ g CaO}$$

$$\text{mol } CO_2 = 0.4995 \text{ mol } CaCO_3 \times \frac{1 \text{ mol } CO_2}{1 \text{ mol } CaCO_3} = 0.4995 \text{ mol } CO_2$$

$$\text{mass } CO_2 = 0.4995 \text{ mol } CO_2 \times \frac{44.01 \text{ g}}{1 \text{ mol}} = 22.0 \text{ g } CO_2$$

The results illustrate an important point: the sum of the masses of the two products (28.0 g + 22.0 g) equals the mass of the reactant (50.0 g).

2. Answers will vary. An example of hydrogen and oxygen reacting to form water is provided.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

Molar masses: $H_2$, 2.016 g; $O_2$, 32.00 g

To determine which reactant is limiting, we first need to realize that the masses of the reactants (25.0 g of each) tell us nothing: we need to calculate how many moles of each reactant is present.

$$\text{mol } H_2 = 25.0 \text{ g } H_2 \times \frac{1 \text{ mol } H_2}{2.016 \text{ g } H_2} = 12.0 \text{ mol } H_2$$

$$\text{mol } O_2 = 25.0 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} = 0.781 \text{ mol } O_2$$

Considering now these numbers of moles, it is clear that there is considerably more hydrogen present than oxygen. Chances are, the hydrogen is present in excess, and oxygen is the limiting reactant. We need to prove this, however, by calculation. If we consider that the 0.7813 mol of oxygen may be the limiting reactant, we can calculate how much hydrogen would be needed for complete reaction. This requires the mole ratio as determined by the coefficients of the balanced chemical equation.

$$0.781 \text{ mol } O_2 \times \frac{2 \text{ mol } H_2}{1 \text{ mol } O_2} = 1.56 \text{ mol } H_2 \text{ required for reaction}$$

Since only 1.56 mol of $H_2$ is required to react with 0.781 mol of $O_2$, and since we have considerably more hydrogen present in our sample than this amount, clearly hydrogen is present in excess and oxygen is, indeed, the limiting reactant.

Suppose we had not initially considered that oxygen was the limiting reactant (because there is so much less oxygen present on a mole basis) and had wondered if $H_2$ was the limiting reactant. For the given amount of $H_2$ (12.4 mol), we could calculate how much oxygen would be required to react

$$12.4 \text{ mol } H_2 \times \frac{1 \text{ mol } O_2}{2 \text{ mol } H_2} = 6.20 \text{ mol } O_2 \text{ would be required}$$

Since we do not have 6.20 mol of $O_2$ (we have only 0.781 mol $O_2$), clearly there is not enough oxygen present to react with all the hydrogen, and we would conclude again that oxygen must be the limiting reactant.

3. For potassium: $2K(s) + Cl_2(g) \rightarrow 2KCl(s)$

Molar masses: K, 39.10 g; $Cl_2$, 70.90 g; KCl, 74.55 g

$$25.0 \text{ g K} \times \frac{1 \text{ mol}}{39.10 \text{ g}} = 0.639 \text{ mol K}$$

$$50.0 \text{ g } Cl_2 \times \frac{1 \text{ mol}}{70.90 \text{ g}} = 0.705 \text{ mol } Cl_2$$

K is the limiting reactant

$$0.639 \text{ mol K} \times \frac{2 \text{ mol KCl}}{2 \text{ mol K}} \times \frac{74.55 \text{ g}}{1 \text{ mol}} = 47.6 \text{ g KCl}$$

For calcium: $Ca(s) + Cl_2 \rightarrow CaCl_2(s)$

Molar Masses: Ca, 40.08 g; $Cl_2$, 70.90 g; $CaCl_2$, 111.0 g

$$25.0 \text{ g Ca} \times \frac{1 \text{ mol}}{40.08 \text{ g}} = 0.624 \text{ mol Ca}$$

$$50.0 \text{ g } Cl_2 \times \frac{1 \text{ mol}}{70.90 \text{ g}} = 0.705 \text{ mol } Cl_2$$

Ca is the limiting reactant

$$0.624 \text{ mol Ca} \times \frac{1 \text{ mol } CaCl_2}{1 \text{ mol Ca}} \times \frac{111.0 \text{ g}}{1 \text{ mol}} = 69.3 \text{ g } CaCl_2$$

For aluminum: $2\,Al(s) + 3Cl_2(g) \rightarrow 2AlCl_3(s)$

Molar masses: Al, 26.98 g; $Cl_2$ 70.90 g; $AlCl_3$, 133.3 g

$$25.0\ gAl \times \frac{1\ mol}{26.98\ g} = 0.927\ mol\ Al$$

$$50.0\ g\ Cl_2 \times \frac{1\ mol}{70.90\ g} = 0.705\ mol\ Cl_2$$

$Cl_2$ is the limiting reactant

$$0.705\ mol\ Cl_2 \times \frac{2\ mol\ AlCl_3}{3\ mol\ Cl_2} \times \frac{133.3\ g}{1\ mol} = 62.7\ g\ AlCl_3$$

4. (31.2 g KCl / 47.6 g KCl) x 100% = 65.6% yield

5. The relevant reaction is $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

$$1.48\ AgCl \times \frac{1\ mol\ AgCl}{143.35\ g\ AgCl} \times \frac{1\ mol\ Ag^+}{1\ mol\ AgCl} = 0.0103\ mol\ Ag^+$$

The 0.0103 mol $Ag^+$ comes from $AgNO_3$, thus we started with 0.0103 mol $AgNO_3$ in the mixture.

$$0.0103\ mol\ AgNO_3 \times \frac{169.91 g}{1\ mol\ AgNO_3} = 1.75 g\ AgNO_3$$

$$\frac{1.75 g\ AgNO_3}{5.00 g\ mixture} \times 100\% = 35.0\%.$$ The mixture is 35.0% $AgNO_3$

## Chapter 10: Basic Review Worksheet

1. Scientists define energy as "the capacity to do work or to produce heat". As with "matter", energy is such a fundamental concept that it is hard to define.

2. potential; kinetic; total

3. A state function is a property that is independent of pathway. Energy and elevation are examples of state functions. Heat and work are not state functions.

4. Temperature is a measure of the average kinetic energy of the particles of a system.

5. An exothermic reaction or process is one in which energy as heat is released to the surroundings; an endothermic reaction or process is one the system absorbs energy as heat from the surroundings.

 