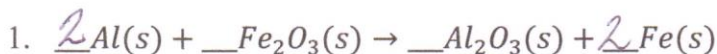


Percentage Purity and Percentage Yield**Percentage Purity:**

Virtually impossible to obtain a chemical in a pure form. Purity of a chemical is determined by:

$$\text{Percentage purity (by mass)} = \frac{\text{mass of pure chemical}}{\text{mass of impure sample}} \times 100\%$$

**Common questions:**

a. Balance the equation

b. Determine the charge on iron based on  $\text{Fe}_2\text{O}_3$   $\{ \text{no overall charge} \therefore \text{Fe} = 6+$ 

c. If 2.44g of 95% pure aluminum is reacted, how many grams of aluminum oxide can be produced?

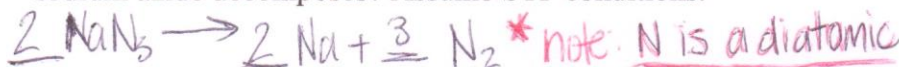
imp(g)  
and %p

$$\%p = \frac{\text{p(g)}}{\text{imp(g)}} \times 100\%$$

$$2.31\text{g of Al} \times \frac{1\text{ mol Al}}{26.98\text{g Al}} \times \frac{1\text{ mol Al}_2\text{O}_3}{2\text{ mol Al}} \times \frac{101.96\text{g}}{1\text{ mol Al}_2\text{O}_3} = 4.36\text{g of Al}_2\text{O}_3$$

$$\text{p(g)} = \frac{\%p}{100\%} \times \text{imp(g)} = \frac{95\%}{100\%} \times 2.44\text{g} = 2.31\text{g of Al (pure)}$$

2. Automotive air bags inflate when solid sodium azide ( $\text{NaN}_3$ ) decomposes explosively into its constituent elements. What volume of nitrogen gas is formed if 120g of 85% pure sodium azide decomposes? Assume STP conditions.



\* note: N is a diatomic

$$\text{p(g)} = \frac{\%p}{100\%} \times \text{imp(g)} = \frac{85\%}{100\%} \times 120\text{g of impure NaN}_3 = 102\text{g of pure NaN}_3$$

$$102\text{g of NaN}_3 \times \frac{1\text{ mol NaN}_3}{65.02\text{g}} \times \frac{3\text{ mol N}_2}{2\text{ mol NaN}_3} \times \frac{22.4\text{L}}{1\text{ mol}} = 53\text{L}$$

3. When dilute hydrochloric acid is added to 5.73g of contaminated calcium carbonate, 2.49g of  $\text{CO}_2$  is produced. What was the percentage purity of the calcium carbonate?



$$2.49\text{g of CO}_2 \times \frac{1\text{ mol CO}_2}{44.01\text{g}} \times \frac{1\text{ mol CaCO}_3}{1\text{ mol CO}_2} \times \frac{100.09\text{g}}{1\text{ mol CaCO}_3} = 5.66\text{g of pure CaCO}_3$$

$$\%p = \frac{5.66\text{g}}{5.73\text{g}} \times 100\% = 98.8\%$$



a. Balance the equation

b. Calculate the percent purity of a sample of  $\text{Mg}(\text{OH})_2$  if titration of 2.568g of the sample required 38.45mL of 0.6995M  $\text{H}_3\text{PO}_4$ 

$$\%p = \frac{2.353\text{g}}{2.568\text{g}} \times 100\% = 91.63\%$$

$$\text{moles} = CV$$

$$= 0.6995\text{M} \times \left( 38.45\text{mL} \times \frac{1\text{L}}{1000\text{mL}} \right)$$

$$= 0.02689\text{ moles of H}_3\text{PO}_4 \times \frac{3\text{ mol Mg}(\text{OH})_2}{2\text{ mol H}_3\text{PO}_4} \times \frac{58.33\text{g}}{1\text{ mol Mg}(\text{OH})_2} = 2.353\text{g of pure Mg}(\text{OH})_2$$

$$\frac{n}{CV} \text{ (moles)} \quad \frac{n}{CV} \text{ (volume)}$$

(concentration = molarity)

$$M = \frac{\text{moles}}{\text{L}}$$

**Percentage Yield:**

Some reactions complete themselves only partially (not all of the limiting reactant has been converted into product). To compare how much has actually been obtained to what was expected we use:

$$\text{Percentage yield} = \frac{\text{amount of product obtained}}{\text{amount of product expected}} \times 100\%$$

Amount of product expected is commonly referred to as the theoretical yield

**Common Questions**

1.  $\text{GeF}_3\text{H}$  is synthesized in the reaction:  $\text{GeH}_4 + 3\text{GeF}_4 \rightarrow 4\text{GeF}_3\text{H}$ . If the reaction yield is 91.5%, how many moles of  $\text{GeH}_4$  are needed to produce 12.00 mol of  $\text{GeF}_3\text{H}$ ?

$$\%y = \frac{\text{pdt. ob}}{\text{pdt. ex}} \times 100\%$$

$$13.11 \text{ mol} \times \frac{1 \text{ mol GeH}_4}{4 \text{ mol GeF}_3\text{H}} = 3.278 \text{ mol of GeH}_4$$

$$\text{pdt ex} = \frac{\text{pdt ob} \times 100\%}{\%y} = \frac{12.00 \text{ mol} \times 100\%}{91.5\%} = 13.11 \text{ mol of GeF}_3\text{H expected}$$

2. What mass of silver could be formed if a large zinc wire is placed in a beaker containing 145.0 mL of 0.095 mol/L silver nitrate and allowed to react overnight? Assume the reaction has a 97% yield



$$\text{mol} = 0.095 \text{ mol/L} \times \left( \frac{145.0 \text{ mL} \times 1 \text{ L}}{1000 \text{ mL}} \right)$$

$$= 0.013775 \text{ mol of AgNO}_3 \times \frac{2 \text{ mol Ag}}{2 \text{ mol AgNO}_3} \times \frac{107.87 \text{ g}}{1 \text{ mol Ag}} = 1.486 \text{ g of Ag obtained}$$

$$\%y = \frac{\text{pdt ob}}{\text{pdt ex}} \times 100 \rightarrow \text{pdt ex} = \frac{\text{pdt ob} \times 100\%}{\%y} = \frac{1.486 \text{ g} \times 100\%}{97\%} = 1.532 \text{ g of Ag expected}$$

3. Copper (II) oxide reacts with hydrogen gas to form water and copper metal. From this reaction, 3.6 g of copper metal was obtained with a yield of 32.5%. What mass of copper(II) oxide was reacted with the excess hydrogen gas? Begin with a balanced equation.



$$\text{pdt. ex} = \frac{\text{pdt. ob} \times 100\%}{\%y} = \frac{3.6 \text{ g} \times 100\%}{32.5\%} = 11 \text{ g of Cu} \times \frac{1 \text{ mol Cu}}{63.55 \text{ g}} \times \frac{1 \text{ mol CuO}}{1 \text{ mol Cu}} \times \frac{79.55 \text{ g}}{1 \text{ mol CuO}} = 14 \text{ g of CuO}$$

4. 32g of  $\text{O}_2$  react with 11g of  $\text{C}_3\text{H}_8$ :

a. Balanced equation?

b. Limiting reagent?

c. Theoretical yield of water?



c). start with limiting reagent: 32g  $\text{O}_2$

$$32 \text{ g} \times \frac{1 \text{ mol}}{32.00 \text{ g}} = 1.0 \text{ mol of O}_2 \times \frac{1 \text{ mol C}_3\text{H}_8}{5 \text{ mol O}_2} = 0.20 \text{ mol of C}_3\text{H}_8 \text{ required}$$

$$32 \text{ g} \times \frac{\text{mol O}_2}{32.00 \text{ g}} \times \frac{4 \text{ mol H}_2\text{O}}{5 \text{ mol O}_2} \times \frac{18.02 \text{ g}}{1 \text{ mol H}_2\text{O}} = 14.92 \text{ g of H}_2\text{O}$$

$$11 \text{ g} \times \frac{1 \text{ mol}}{44.10 \text{ g}} = 0.25 \text{ mol of C}_3\text{H}_8 \times \frac{5 \text{ mol O}_2}{1 \text{ mol C}_3\text{H}_8} = 1.25 \text{ mol of O}_2 \text{ required}$$

b)  $\text{O}_2$  (oxygen) is our limiting reagent.  $\text{C}_3\text{H}_8$  is in xs as the amount required < given

$$\text{c) } = 14.9 \text{ g of H}_2\text{O}$$