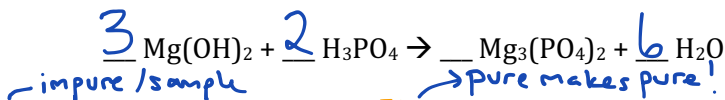


Chemistry 11

Percent Purity

Name: Key
 Date:
 Block:

1. Consider the reaction of magnesium hydroxide with phosphoric acid:

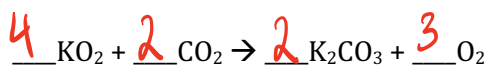


Calculate the mass of Mg(OH)_2 needed to make 127 g of $\text{Mg}_3(\text{PO}_4)_2$. Assume the Mg(OH)_2 is 88.5% pure.

$$127 \text{ g Mg}_3(\text{PO}_4)_2 \times \frac{1 \text{ mol Mg}_3(\text{PO}_4)_2}{262.87 \text{ g Mg}_3(\text{PO}_4)_2} \times \frac{3 \text{ mol Mg(OH)}_2}{1 \text{ mol Mg}_3(\text{PO}_4)_2} \times \frac{58.33 \text{ g Mg(OH)}_2}{1 \text{ mol Mg(OH)}_2} = 84.5 \text{ g pure Mg(OH)}_2$$

$$\% \text{ Purity} = \frac{\text{Pure}}{\text{Sample}} \times 100\% \rightarrow 88.5\% = \frac{84.5 \text{ g}}{x} \times 100\% \rightarrow x = \boxed{95.5 \text{ g sample Mg(OH)}_2}$$

2. Consider the reaction:



a. A 30.0 g sample of KO_2 is 59.3% pure. What mass of K_2CO_3 can the sample produce?

$$\% \text{ Purity} = \frac{\text{Pure}}{\text{Sample}} \times 100\% \rightarrow 59.3\% = \frac{x}{30.0 \text{ g KO}_2} \times 100\% \rightarrow x = 17.8 \text{ g pure KO}_2$$

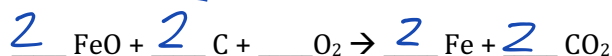
$$17.8 \text{ g pure KO}_2 \times \frac{1 \text{ mol KO}_2}{71.10 \text{ g KO}_2} \times \frac{2 \text{ mol K}_2\text{CO}_3}{4 \text{ mol KO}_2} \times \frac{138.21 \text{ g K}_2\text{CO}_3}{1 \text{ mol K}_2\text{CO}_3} = \boxed{17.3 \text{ g pure K}_2\text{CO}_3}$$

b. Another sample of KO_2 with a mass of 150.0 g is reacted so as to produce 89.7 g of K_2CO_3 . What is the percentage purity of KO_2 ?

$$89.7 \text{ g pure K}_2\text{CO}_3 \times \frac{1 \text{ mol K}_2\text{CO}_3}{138.12 \text{ g K}_2\text{CO}_3} \times \frac{4 \text{ mol KO}_2}{2 \text{ mol K}_2\text{CO}_3} \times \frac{71.10 \text{ g KO}_2}{1 \text{ mol KO}_2} = 92.3 \text{ g pure KO}_2$$

$$\% \text{ Purity} = \frac{\text{Pure}}{\text{Sample}} \times 100\% \rightarrow \frac{92.3 \text{ g}}{150.0 \text{ g}} \times 100\% \rightarrow \boxed{61.5\% \text{ KO}_2}$$

3. If 72.1 g of FeO produces 12.9 g of pure Fe according to the reaction:



What is the percentage purity of the FeO used?

$$12.9 \text{ g pure Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} \times \frac{2 \text{ mol FeO}}{2 \text{ mol Fe}} \times \frac{71.85 \text{ g FeO}}{1 \text{ mol FeO}} = 16.6 \text{ g pure FeO}$$

$$\% \text{ purity} = \frac{\text{Pure}}{\text{Sample}} \times 100\% \rightarrow \frac{16.6 \text{ g FeO}}{72.1 \text{ g FeO}} \times 100\% = \boxed{23.0\% \text{ FeO}}$$

Chemistry 11

Percent Yield

Name: Key
 Date:
 Block:

4. Potassium chlorate decomposes to form potassium chloride and oxygen gas.

a. What is the balanced reaction?



b. When 5.95 g of potassium chlorate decomposes, 1.45 g of oxygen gas is given off. Calculate the percentage yield of oxygen.

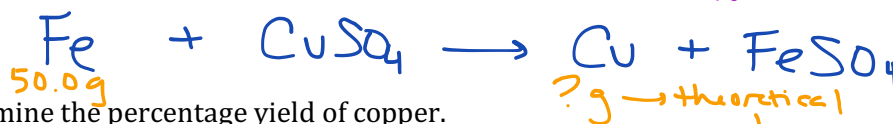
$$5.95 \text{g KClO}_3 \times \frac{1 \text{mol KClO}_3}{122.55 \text{g KClO}_3} \times \frac{3 \text{mol O}_2}{2 \text{mol KClO}_3} \times \frac{32.00 \text{g O}_2}{1 \text{mol O}_2} = 2.33 \text{g O}_2$$

← theoretical

$$\% \text{ yield} = \frac{\text{actual}}{\text{theoretical}} \times 100\% = \frac{1.45 \text{g O}_2}{2.33 \text{g O}_2} \times 100\% = \boxed{62.2\% \text{ O}_2}$$

5. When 50.0 g of iron metal is reacted with copper (II) sulfate, 43.0 g of copper metal is recovered.

a. What is the balanced reaction?

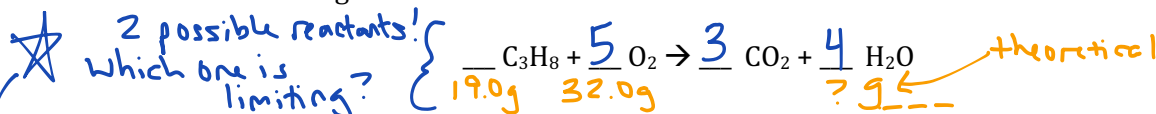


b. Determine the percentage yield of copper.

$$50.0 \text{g Fe} \times \frac{1 \text{mol Fe}}{55.85 \text{g Fe}} \times \frac{1 \text{mol Cu}}{1 \text{mol Fe}} \times \frac{63.55 \text{g Cu}}{1 \text{mol Cu}} = 56.9 \text{g Cu}$$

$$\% \text{ yield} = \frac{\text{actual}}{\text{theoretical}} \times 100\% = \frac{43.0 \text{g Cu}}{56.9 \text{g Cu}} \times 100\% = \boxed{75.6\% \text{ Cu}}$$

6. Consider the following reaction:



32.0 g of oxygen reacts with 19.0 g of C₃H₈. The experiment gives 2.00 g H₂O. What is the % yield?

$$19.0 \text{g C}_3\text{H}_8 \times \frac{1 \text{mol C}_3\text{H}_8}{44.11 \text{g C}_3\text{H}_8} \times \frac{4 \text{mol H}_2\text{O}}{1 \text{mol C}_3\text{H}_8} \times \frac{18.02 \text{g H}_2\text{O}}{1 \text{mol H}_2\text{O}} = 31.0 \text{g H}_2\text{O}$$

← actual
excess

$$32.0 \text{g O}_2 \times \frac{1 \text{mol O}_2}{32.00 \text{g O}_2} \times \frac{4 \text{mol H}_2\text{O}}{5 \text{mol O}_2} \times \frac{18.02 \text{g H}_2\text{O}}{1 \text{mol H}_2\text{O}} = 14.4 \text{g H}_2\text{O}$$

← theoretical
limiting!

$$\% \text{ yield} = \frac{\text{actual}}{\text{theoretical}} \times 100\% = \frac{2.00 \text{g H}_2\text{O}}{14.4 \text{g H}_2\text{O}} \times 100\% = \boxed{13.9\% \text{ H}_2\text{O}}$$

We have to try both reactants...
 ☺