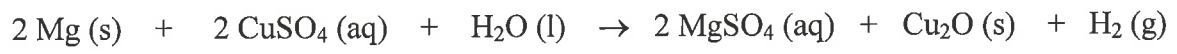


AP Test Prep: Review Packet #1 (2014)
Stoichiometry

1. Consider the following chemical equation.



(a) If 1.46 grams of Mg are added 500.0 mL of 0.200 M CuSO₄, what is the maximum volume of H₂ gas that can be produced?

$$\frac{1.46}{24.30} = .0601 \text{ moles Mg} \quad \text{LIMITING}$$

$$(.200)(.5000) = .100 \text{ moles CuSO}_4 \quad \text{EXCESS}$$

$$\frac{.0601}{2} = .0301 \text{ moles H}_2$$

$$.0301 = \frac{x}{22.4}$$

$$x = .674 \text{ Liters}$$

(b) When all of the limiting reagent has been consumed in (a), how many moles of the other reactant (not water) will remain?

$$\begin{array}{r} .100 \\ - .0601 \\ \hline .040 \end{array}$$

.040 moles excess CuSO₄ will remain unreacted

(c) What is the mass of the Cu₂O produced in (a)?

$$\frac{.0601}{2} = .0301 \text{ moles Cu}_2\text{O}$$

$$.0301 = \frac{x}{143.10}$$

$$x = 4.31 \text{ grams Cu}_2\text{O}$$

(d) What is the molarity of the Mg²⁺ in the solution at the end of this experiment? Assume the total volume of the solution remains constant.

* ALL OF THE ORIGINAL SOLID Mg⁰ REACTS TO PRODUCE SOLUBLE MgSO₄. "SPLITSVILLE", Mg²⁺ and SO₄²⁻

$$\frac{.0601 \text{ moles Mg}^{2+}}{.5000 \text{ Liters}} = .120 \text{ M Mg}^{2+}$$

2. The table below shows three common forms of copper ore.

Ore #	Empirical Formula	Percent by Weight		
		Copper	Sulfur	Iron
1	Cu ₂ S	?	?	0
2	?	34.6	34.9	30.5
3	?	55.6	28.1	16.3

(a) What is the percent by weight of copper in Cu₂S?

$$\% \text{ Cu} = \frac{\text{part of form wt from Cu}}{\text{total form wt of compd}}$$

$$\frac{(63.55)(2)}{(63.55)(2) + 32.06} = \frac{127.1}{159.16} = .799 \times 100$$

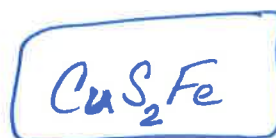
$$79.9\%$$

(b) What is the empirical formula of ore #2?

$$\text{Cu} \frac{34.6}{63.55} = .544 \div .544 = \textcircled{1}$$

$$\text{S} \frac{34.9}{32.06} = 1.09 \div .544 = \textcircled{2}$$

$$\text{Fe} \frac{30.5}{55.85} = .546 \div .544 = \textcircled{1}$$



(c) If a sample of ore #3 contains 11.0 grams of iron, how many grams of sulfur does it contain?

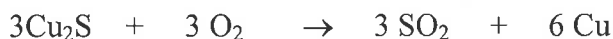
Easiest method: $\% \text{ iron}$

$$(\text{sample mass}) (.163) = 11.0$$

$$\text{sample mass} = 67.5 \text{ grams}$$

$$(67.5)(.281) = 19.0 \text{ grams S}$$

(d) Pure copper metal can be extracted from Cu₂S by the following process:



If 3.84 grams of O₂ are consumed in the process, how many grams of Cu are produced?

$$\frac{3.84}{32.00} = .120 \text{ moles O}_2 \times 2 = .240 \text{ moles Cu}$$

$$.240 = \frac{x}{63.55} \quad x = 15.3 \text{ grams Cu}$$



3. A 10.0 gram sample containing calcium carbonate and an inert material was placed in excess HCl. A reaction occurred that produced calcium chloride, water, and carbon dioxide.

(a) Write the balanced chemical equation for this reaction.



(b) When the reaction was complete, 900 mL of CO_2 gas was collected at 740 Torr and 30°C . How many moles of calcium carbonate were consumed in this reaction?

↑ ↑
non STP conditions

$$PV = nRT$$

$$P = \frac{740}{760} = .974 \text{ atm}$$

$$(.974)(.900) = n(.0821)(303)$$

$$n = .0352 \text{ moles CO}_2$$

$$V = .900 \text{ Liters}$$

$$n = \text{moles CO}_2$$

$$R = .0821$$

$$T = 303 \text{ K}$$

$$\text{moles CO}_2 = \text{moles CaCO}_3$$

$$.0352 \text{ moles CaCO}_3$$

(c) If all of the calcium carbonate initially present in the sample was consumed in the reaction, what percentage of the original by mass of the sample was due to calcium carbonate?

$$.0352 = \frac{x}{100.09}$$

$$x = 3.52 \text{ grams CaCO}_3$$

$$\% \text{ CaCO}_3 = \frac{3.52}{10.0} \times 100 = 35.2\%$$

↑
ORIGINAL MASS
OF SAMPLE
(SEE BEGINNING
OF PROBLEM)

(d) If the inert material was silicon dioxide (SiO_2), what was the mole ratio of the calcium carbonate to silicon dioxide (SiO_2) in the original sample?

$$\begin{array}{r} \text{ORIGINAL SAMPLE} = 10.0 \text{ grams} \\ - 3.52 \\ \hline 6.48 \text{ grams SiO}_2 \end{array}$$

$$\frac{6.48}{60.09} = .108 \text{ moles SiO}_2$$

from above,
.0352 moles
 CaCO_3

$$.0352 : .108 \\ \text{moles CaCO}_3 : \text{moles SiO}_2$$

$$1 : 3$$

contains only C + H

4. A gaseous hydrocarbon sample is completely burned in air, producing 1.80 Liters of carbon dioxide at STP conditions and 2.16 grams of water.

(a) What is the empirical formula for this hydrocarbon?



$$\frac{1.80}{22.4} = .0804 \text{ moles CO}_2$$

$$\frac{2.16}{18.016} = .120 \text{ moles H}_2\text{O}$$

$$.0804 \text{ moles CO}_2 = .0804 \text{ moles C} \div .0804 = 1$$

$$.120 \text{ moles H}_2\text{O} = .240 \text{ moles H} \div .0804 = 2.99 \approx 3$$

EMPIRICAL FORMULA



(b) What was the mass of the hydrocarbon that was consumed?

$$\text{mass of hydrocarbon} = \text{mass C} + \text{mass H}$$

get mass of each element from # of moles

$$.0804 = \frac{x}{12.01}$$

$$x = .966 \text{ grams C}$$

$$.240 = \frac{x}{1.008}$$

$$x = .242 \text{ grams H}$$

$$.966 + .242 = 1.208 \text{ g}$$

(c) The hydrocarbon was initially contained in a closed 1.00 Liter vessel at a temperature of 32°C and a pressure of 760 Torr. What is the molecular formula of the hydrocarbon?

Ideal Gas Law Needed $PV = nRT$

$$P = \frac{760}{760} = 1.00 \text{ atm}$$

$$V = 1.00 \text{ liter}$$

$$n =$$

$$R = .0821$$

$$T = 273 + 32 = 305 \text{ K}$$

$$(1.00)(1.00) = n(.0821)(305)$$

$$n = .0399 \text{ moles gas hydrocarbon}$$

Answer from (b) gives me total mass of hydrocarbon

$$.0399 = \frac{1.208}{\text{fwt}}$$

$$\text{fwt} = 30.3$$

(d) Write the balanced equation for the combustion of the hydrocarbon.

each emp formula = CH₃

15 = fwt

2 bricks needed

