Mr. Bracken
AP Chemistry

Name BRACKEN KEY
Period ______

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

1. Consider the following chemical equation.

$$2\,Mg\,(s) + 2\,CuSO_4\,(aq) + H_2O\,(l) \rightarrow 2\,MgSO_4\,(aq) + Cu_2O\,(s) + H_2\,(g)$$

(a) If 1.46 grams of Mg are added 500.0 mL of 0.200 M $CuSO_4$, what is the maximum volume of H_2 gas that can be produced?

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

$(.200)(.5000) = \boxed{.100 \text{ moles } CuSO_4}$

EXCESS

$\frac{.0601}{2} = \boxed{.0301 \text{ moles } H_2}$

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

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

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

$.100 - .0601$

.040 moles excess $CuSO_4$ will remain unreacted

(c) What is the mass of the Cu_2O produced in (a)?

$\frac{.0601}{2} = .0301 \text{ moles } Cu_2O$

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

$\boxed{x = 4.31 \text{ grams } Cu_2O}$

(d) What is the molarity of the Mg^{2+} in the solution at the end of this experiment? Assume the total volume of the solution remains constant.

★ ALL OF THE ORIGINAL SOLID Mg° REACTS TO PRODUCE SOLUBLE $MgSO_4$. "SPLITSVILLE" Mg^{2+} and SO_4^{2-}

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

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

Ore #	Empirical Formula	Percent by Weight		
		Copper	Sulfur	Iron
1	Cu_2S	?	?	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 Cu_2S?

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

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

x100

$\boxed{79.9\%}$

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

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

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

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

$\boxed{CuS_2Fe}$

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

Easiest method: % iron

(sample mass)(.163) = 11.0

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

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

(d) Pure copper metal can be extracted from Cu_2S by the following process:

$$3Cu_2S + 3\,O_2 \rightarrow 3\,SO_2 + 6\,Cu$$

If 3.84 grams of O_2 are consumed in the process, how many grams of Cu are produced?

$\frac{3.84}{32.00} = \boxed{.120 \text{ moles } O_2} \times 2 = \boxed{.240 \text{ moles Cu}}$

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

$CaCO_3$

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.

$$CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$$

(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

$PV = nRT$

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

$V = .900$ Liters

$n =$ moles CO_2

$R = .0821$

$T = 303$ K

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

$\boxed{n = .0352 \text{ moles } CO_2}$

moles CO_2 = moles $CaCO_3$

$\boxed{.0352 \text{ moles } CaCO_3}$

(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?

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

$\boxed{x = 3.52 \text{ grams } CaCO_3}$

$\% CaCO_3 = \frac{3.52}{10.0} \times 100 = \boxed{35.2\%}$

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?

ORIGINAL SAMPLE = 10.0 grams

$- 3.52$

6.48 grams SiO_2

$\frac{6.48}{60.09} = \boxed{.108 \text{ moles } SiO_2}$

from above, .0352 moles $CaCO_3$

.0352 moles $CaCO_3$: .108 moles SiO_2

$\boxed{1 : 3}$

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?

$CH + O_2 \longrightarrow CO_2 + H_2O$

$\frac{1.80}{22.4} = \boxed{.0804 \text{ moles } CO_2}$ $\frac{2.16}{18.016} = \boxed{.120 \text{ moles } H_2O}$

.0804 moles CO_2 = .0804 moles C ÷ .0804 = 1

.120 moles H_2O = .240 moles H ÷ .0804 = 2.99 ≈ 3

EMPIRICAL FORMULA

CH_3

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

mass of hydrocarbon = mass C + mass H

get mass of each element from # of moles

$.0804 = \frac{x}{12.01}$ $\boxed{x = .966 \text{ grams C}}$

$.240 = \frac{x}{1.008}$ $\boxed{x = .242 \text{ grams H}}$

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

(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 $PV = nRT$

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

$V = 1.00$ Liter

$n =$

$R = .0821$

$T = 273 + 32 = 305$ K

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

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

Answer from (b) gives me total mass of hydrocarbon

$.0399 = \frac{1.208}{fwt}$ $\boxed{fwt = 30.3}$

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

each emp formula = CH_3

15 = fwt

2 bricks needed

CH_3 | CH_3

C_2H_6

$$2\,C_2H_6 + 7\,O_2 \longrightarrow 4\,CO_2 + 6\,H_2O$$