

Name KEY

Date \_\_\_\_\_

## AP CHEMISTRY: STOICHIOMETRY EXAM [PRACTICE]

Exam Info... The actual exam will have 10 multiple choice questions [20 points total], 3 writing reaction questions [15 points total], and 3 short answer questions [25 points total]. The entire test will be a total of 60 points.

## PART I: MULTIPLE CHOICE

How many mL of 10.0 M HCl are needed to prepare 500. mL of 2.00 M HCl?

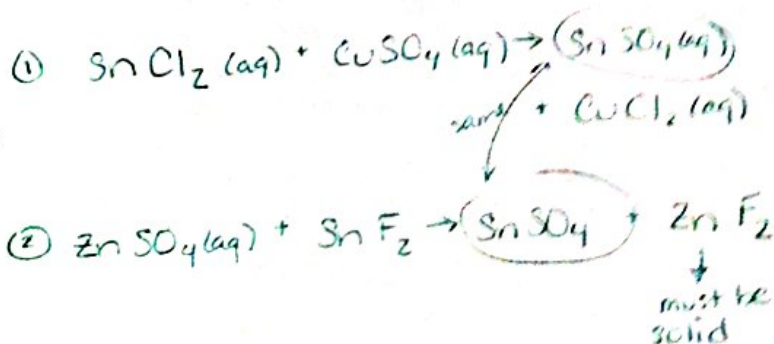
- (A) 1.00 mL  
 (B) 10.0 mL  
 (C) 20.0 mL  
 (D) 100. mL  
 (E) 200. mL

$$(10 M)(x) = (500)(2 M)$$

$$x = 100 \text{ mL}$$

A student mixes equal volumes of 1.0 M solutions of tin(II) chloride and copper(II) sulfate and observes that no precipitate forms. Then the student mixes equal volumes of 1.0 M solutions of zinc(II) sulfate and tin(II) fluoride and observes the formation of a precipitate. The formula of the precipitate must be

- (A)  $\text{SnF}_2$   
 (B)  $\text{SnSO}_4$   
 (C)  $\text{Sn}(\text{SO}_4)_2$   
 (D)  $\text{ZnF}$   
 (E)  $\text{ZnF}_2$



A 360. mg sample of aspirin,  $\text{C}_9\text{H}_8\text{O}_4$ , (molar mass 180. g), is dissolved in enough water to produce 200. mL of solution. What is the molarity of aspirin in a 50. mL sample of this solution?

- (A) 0.0800 M  
 (B) 0.0400 M  
 (C) 0.0200 M  
 (D) 0.0100 M  
 (E) 0.00250 M

$$360 \text{ mg} \cdot \frac{1 \text{ g}}{1000 \text{ mg}} = .360 \text{ g}$$

$$\frac{.360 \text{ g}}{180 \text{ g/mol}} = 2 \times 10^{-3} \text{ mol}$$

$$\frac{2 \times 10^{-3} \text{ mol}}{.2 \text{ L}} = 0.01 \text{ M}$$

$$M_1 V_1 = M_2 V_2$$

$$(0.01 \text{ M})(.2 \text{ L}) = x (.05 \text{ L})$$

$$x = 0.04 \text{ M}$$

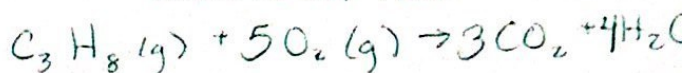
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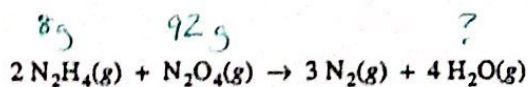
Propane gas,  $C_3H_8$ , burns in excess oxygen gas. When the equation for this reaction is correctly balanced and all coefficients are reduced to their lowest whole-number terms, the coefficient for  $O_2$  is

- (A) 4
- (B) 5
- (C) 7
- (D) 10
- (E) 22

Combustion reaction



1. look at C
2. look at H
3. look at O last



i. When 8.0 g of  $N_2H_4$  ( $32 \text{ g mol}^{-1}$ ) and 92 g of  $N_2O_4$  ( $92 \text{ g mol}^{-1}$ ) are mixed together and react according to the equation above, what is the maximum mass of  $H_2O$  that can be produced?

- (A) 9.0 g
- (B) 18 g
- (C) 36 g
- (D) 72 g
- (E) 144 g

$$\frac{8 \text{ g}}{32 \text{ g/mol}} = .25 \text{ mol } N_2H_4 \quad \leftarrow \text{limiting reagent}$$

$$\frac{92 \text{ g}}{92 \text{ g/mol}} = 1 \text{ mol } N_2O_4 \quad \leftarrow \text{excess}$$

mole ratio

$$\frac{2 N_2H_4}{1 N_2O_4} = \frac{.25 \text{ mol}}{x}$$

$x = .125 \text{ mol}$  needed to react completely with  $N_2H_4$  and we have more than enough  $N_2O_4$

- (A)  $H_2SeO_4(aq) + 2 Cl^-(aq) + 2 H^+(aq) \rightarrow H_2SeO_3(aq) + Cl_2(g) + H_2O(l)$
- (B)  $S_8(s) + 8 O_2(g) \rightarrow 8 SO_2(g)$
- (C)  $3 Br_2(aq) + 6 OH^-(aq) \rightarrow 5 Br^-(aq) + BrO_3^-(aq) + 3 H_2O(l)$
- (D)  $Ca^{2+}(aq) + SO_4^{2-}(aq) \rightarrow CaSO_4(s)$
- (E)  $PtCl_4(s) + 2 Cl^-(aq) \rightarrow PtCl_6^{2-}(aq)$

$$\frac{4 H_2O}{2 N_2H_4} = \frac{x}{.25 \text{ mol}}$$

$$x = 0.5 \text{ mol } H_2O$$

$$\frac{0.5 \text{ mol } H_2O}{18 \text{ g/mol}} = 9 \text{ g } H_2O$$

11. A precipitation reaction **D**

12. A reaction that produces a coordination complex **E**

13. A reaction in which the same reactant undergoes both oxidation and reduction **C**

14. A combustion reaction **B**

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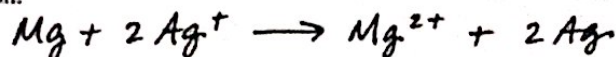
## PART II: WRITING REACTIONS

For each of the following three reactions, in part (i) write a balanced equation and in part (ii) answer the question about the reaction. In part (i), coefficients should be in terms of lowest whole numbers. Assume that solutions are aqueous unless otherwise indicated. Represent substances in solutions as ions if the substances are extensively ionized. Omit formulas for any ions or molecules that are unchanged by the reaction. You may use the empty space at the bottom of the next page for scratch work, but only equations that are written in the answer boxes provided will be scored.

**EXAMPLE:**

A strip of magnesium metal is added to a solution of silver(I) nitrate.

(i) Balanced equation:

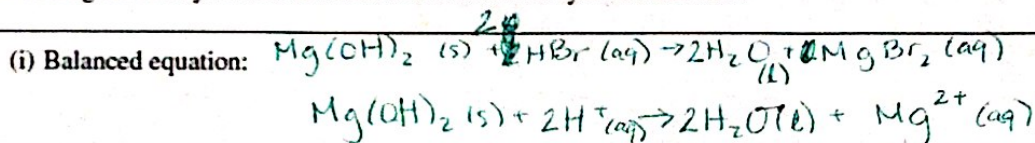


(ii) Which substance is oxidized in the reaction?

Mg is oxidized.

(a) Solid magnesium hydroxide is added to a solution of hydrobromic acid.

(i) Balanced equation:



(ii) What volume, in mL, of 2.00 M hydrobromic acid is required to react completely with 0.10 mol of solid magnesium hydroxide?

$$\frac{1 \text{ Mg}(\text{OH})_2}{2 \text{ HBr}} = \frac{0.10 \text{ mol}}{x}$$

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

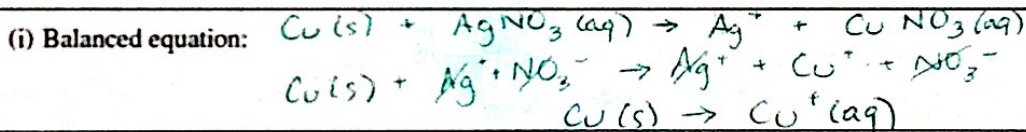
$$2.00 \text{ M} = \frac{0.02 \text{ mol}}{x}$$

$$x = 0.02 \text{ mol HBr}$$

$$x = 0.1 \text{ L} = \boxed{100 \text{ mL HBr}}$$

(c) A copper wire is dipped into a solution of silver(I) nitrate.

(i) Balanced equation:



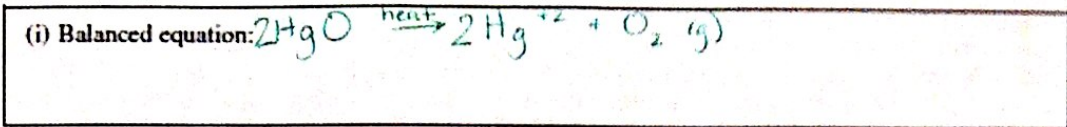
(ii) Describe what is observed as the reaction proceeds.

Copper wire dissolves into solution and creates copper ions

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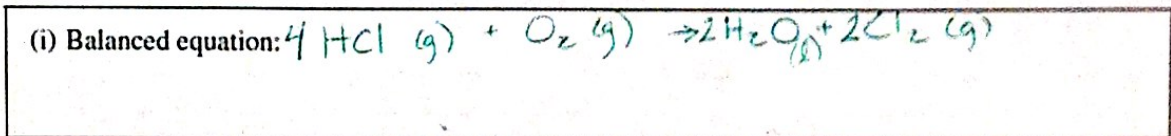
Solid mercury(II) oxide decomposes as it is heated in an open test tube in a fume hood.



(ii) After the reaction is complete, is the mass of the material in the test tube greater than, less than, or equal to the mass of the original sample? Explain.

Less than, because oxygen gas is formed and released from system since the test tube is open.

Hydrogen chloride gas is oxidized by oxygen gas.



(ii) If three moles of hydrogen chloride gas and three moles of oxygen gas react as completely as possible, which reactant, if any, is present in excess? Justify your answer.

$\frac{3 \text{ mol HCl}}{x} = \frac{4 \text{ HCl}}{1 \text{ O}_2}$   
 $x = .75 \text{ mol O}_2 \text{ needed, and there is 3 mol so O}_2 \text{ is in excess.}$   
 to react with 3 mol HCl

(ii)  
 $\text{C: } \frac{.0433 \text{ mol}}{.024} = 3.5 \times 2 = 7$

PART III: SHORT ANSWER

Answer the following questions about a pure compound that contains only carbon, hydrogen, and oxygen.

(a) A 0.7549 g sample of the compound burns in  $\text{O}_2(\text{g})$  to produce 1.9061 g of  $\text{CO}_2(\text{g})$  and 0.3370 g of  $\text{H}_2\text{O}(\text{g})$ .

(i) Calculate the individual masses of C, H, and O in the 0.7549 g sample.

(ii) Determine the empirical formula for the compound.  $\text{C}_7\text{H}_{10}\text{O}_2$

$\text{H: } \frac{.0374 \text{ mol}}{.024} = 3.0 \times 2 = 6$   
 $\text{O: } \frac{.0124 \text{ mol}}{.024} = 1 \times 2 = 2$   
 $\frac{1.9061 \text{ g CO}_2}{12 + 2(16) \text{ g/mol}} = \frac{1.9061 \text{ g}}{44 \text{ g/mol}} = 0.0433 \text{ mol CO}_2 = 0.0433 \text{ mol C} \times 12 \text{ g/mol} = .52 \text{ g C}$   
 (1:1 ratio)

$\frac{.3370 \text{ g H}_2\text{O}}{18 \text{ g/mol H}_2\text{O}} = .0187 \text{ mol H}_2\text{O}$   
 $\frac{.0187 \text{ mol H}_2\text{O}}{x} = \frac{1 \text{ H}_2\text{O}}{2 \text{ H}}$   
 $x = .0374 \text{ mol H} \times 1 \text{ g/mol} = .0374 \text{ g H}$

mass reactants = mass products  
 $1.9061 \text{ g} + 0.3370 \text{ g} = 2.2431 \text{ g}$   
 $0.7549 \text{ g} + x = 2.2431 \text{ g}$   
 $x = 1.4882 \text{ g O}_2 = 0.0465 \text{ mol O}_2$   
 $\frac{1.9061 \text{ g}}{44 \text{ g/mol}} = 0.0433 \text{ mol C}$   
 $\frac{0.3370 \text{ g}}{18 \text{ g/mol}} = 0.0187 \text{ mol H}_2\text{O} \rightarrow 0.0374 \text{ mol H}$   
 mass of O in compound =  $0.7549 \text{ g} - (.52 + .0374) = .1975 \text{ g O}$   
 $\frac{.1975 \text{ g O}}{16 \text{ g/mol}} = .0124 \text{ mol O}$   
 $\frac{.0433 \text{ mol C}}{.0124 \text{ mol O}} = 3.5 \times 2 = 7$   
 $\frac{.0374 \text{ mol H}}{.0124 \text{ mol O}} = 3.0 \times 2 = 6$   
 $\text{C}_7\text{H}_6\text{O}_2$

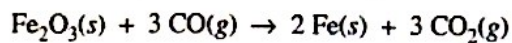
~~0.7549 g sample~~

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Answer the following questions that relate to chemical reactions.

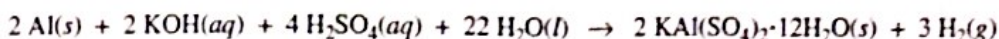
(a) Iron(III) oxide can be reduced with carbon monoxide according to the following equation.



A 16.2 L sample of  $\text{CO}(g)$  at 1.50 atm and  $200.^\circ\text{C}$  is combined with 15.39 g of  $\text{Fe}_2\text{O}_3(s)$ .

- (i) How many moles of  $\text{CO}(g)$  are available for the reaction?  $\emptyset$
- (ii) What is the limiting reactant for the reaction? Justify your answer with calculations.
- (iii) How many moles of  $\text{Fe}(s)$  are formed in the reaction?

OMIT  
Sorry. Haven't  
learned this yet.



In an experiment, a student synthesizes alum,  $\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}(s)$ , by reacting aluminum metal with potassium hydroxide and sulfuric acid, as represented in the balanced equation above.

- (a) In order to synthesize alum, the student must prepare a 5.0 M solution of sulfuric acid. Describe the procedure for preparing 50.0 mL of 5.0 M  $\text{H}_2\text{SO}_4$  using any of the chemicals and equipment listed below. Indicate specific amounts and equipment where appropriate.

10.0 M $\text{H}_2\text{SO}_4$	50.0 mL volumetric flask
Distilled water	50.0 mL buret
100 mL graduated cylinder	25.0 mL pipet
100 mL beaker	50 mL beaker

- (b) Calculate the minimum volume of 5.0 M  $\text{H}_2\text{SO}_4$  that the student must use to react completely with 2.7 g of aluminum metal.

- (c) As the reaction solution cools, alum crystals precipitate. The student filters the mixture and dries the crystals, then measures their mass.

- (i) If the student weighs the crystals before they are completely dry, would the calculated percent yield be greater than, less than, or equal to the actual percent yield? Explain.
- (ii) Cooling the reaction solution in an ice bath improves the percent yield obtained. Explain.

- (d) The student heats crystals of pure alum,  $\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}(s)$ , in an open crucible to a constant mass. The mass of the sample after heating is less than the mass before heating. Explain.

① ~~Use 100 mL graduated cylinder~~  
~~Put 25.0 mL of 10 M  $\text{H}_2\text{SO}_4$  in 50.0 mL volumetric flask. Put distilled water into 50 mL beaker and fill volumetric flask with water until just before the line. Use a pipette to fill rest of flask w/ water drop by drop.~~

Stock solution of 10.0 M  $\text{H}_2\text{SO}_4$   
 $M_1 V_1 = M_2 V_2$   
 $(10 \text{ M})(x) = (5 \text{ M})(50 \text{ mL})$   
 $x = 25 \text{ mL}$  stock solution needed

Use 100 mL graduated cylinder  
 Put 25.0 mL of 10 M  $\text{H}_2\text{SO}_4$  in 50.0 mL volumetric flask. Put distilled water into 50 mL beaker and fill volumetric flask with water until just before the line. Use a pipette to fill rest of flask w/ water drop by drop.

②  $\frac{2.7 \text{ g}}{27 \text{ g/mol}} = 0.1 \text{ mol Al}$   
 $\frac{2 \text{ Al}}{4 \text{ H}_2\text{SO}_4} = \frac{0.1 \text{ mol Al}}{x}$   
 $x = 0.2 \text{ mol H}_2\text{SO}_4$   
 $5.0 \text{ M} = \frac{0.2 \text{ mol}}{x}$   
 $x = 0.04 \text{ L} = \boxed{40. \text{ mL}}$

- ③ (i) Greater than because of the residual water adding more mass to the precipitate.

- (ii) ~~It~~ Cooling in an ice bath ensures the formation of more crystals ~~and~~ by reducing the amount of ions floating in solution.

- ④ The alum is a hydrate, so heating the alum in an open crucible will cause the loosely bound  $\text{H}_2\text{O}$  molecules to evaporate and reduce the mass of the alum.