Practice Test Ch 3 Stoichiometry

Name_______________Per_____

- ***Remember, this is practice - Do NOT cheat yourself of finding out what you are capable of doing. Be sure you follow the testing conditions outlined below.***
- ***DO NOT USE A CALCULATOR. AP Chem does not allow the use of a calculator for the MC part of the exam, so it is time to start practicing without one. You may use ONLY a periodic table.***
- ***While you should practice working as fast as possible, it is more important at this point in the course, that you practice without a calculator, even if it slows you down. Look for the "easy math" − common factors and rough estimation − do not do "long division" to try to get exact values. Remember it is a MC test, use the answers***
- ***Mark which questions you would like to "go over" when we get to school in September.***

1. Balance the following equation:

   $$\_\_\_NH_3 + \_\_\_O_2 \rightarrow \_\_\_NO_2 + \_\_\_H_2O$$

   The balanced equation shows that 1.00 mole of NH_3 requires ___ mole(s) of O_2

   a. 0.57
   b. 1.25
   c. 1.33
   d. 1.75
   e. 3.5

2. Write a balanced equation for the combustion of acetaldehyde, CH_3CHO.

   When properly balanced, the equation indicates that ___ mole(s) of O_2 are required for each mole of CH_3CHO.

   a. 1
   b. 2
   c. 2.5
   d. 3
   e. 5

3. What is the total mass of products formed when 16 grams of CH_4 is burned with excess oxygen?

   a. 32 g
   b. 36 g
   c. 44 g
   d. 62 g
   e. 80 g

4. Write a balanced equation for the combustion of propane, C_3H_8.

   When balanced, the equation indicates that ___ moles of O_2 are required for each mole of C_3H_8.

   a. 1.5
   b. 3
   c. 3.5
   d. 5
   e. 8

5. Balance the following equation with the SMALLEST WHOLE NUMBER COEFFICIENTS possible. Select the number that is the sum of the coefficients in the balanced equation:

   $$\_\_\_KClO_3 \rightarrow \_\_\_KCl + \_\_\_O_2$$

   a. 3
   b. 5
   c. 6
   d. 7
   e. 8

6. Calculate the mass of hydrogen formed when 27 g of aluminum reacts with excess hydrochloric acid according to the balanced equation below.

   $$2Al + 6HCl \rightarrow 2\ AlCl_3 + 3\ H_2$$

   a. 1.5 g
   b. 2.0 g
   c. 3.0 g
   d. 6.0 g
   e. 12 g

7. How many grams of nitric acid, HNO_3, can be prepared from the reaction of 138 g of NO_2 with 54.0 g H_2O according to the equation below?

   $$3NO_2 + H_2O \rightarrow 2HNO_3 + NO$$

   a. 92
   b. 108
   c. 126
   d. 189
   e. 279

8. Which of the following statements is true?

   I. The molar mass of $CaCO_3$ is 100.1 g mol^{-1}.
   II. 50 g of $CaCO_3$ contains 9×10^{23} oxygen atoms.
   III. A 200 g sample of $CaCO_3$ contains 2 moles of $CaCO_3$

   a. I only
   b. II only
   c. III only
   d. I and III only
   e. I, II, and III

$2MnO_2 + 4KOH + O_2 + Cl_2 \rightarrow 2KMnO_4 + 2KCl + 2\ H_2O$

9. For the reaction above, there is 100. g of each reactant available. Which reagent is the limiting reagent?
   [Molar Masses: MnO_2 =86.9; KOH=56.1; O_2 =32.0; Cl_2 =70.9]
   a. MnO_2
   b. O_2
   c. KOH
   d. Cl_2
   e. They all run out at the same time.

10. The reaction of 7.8 g benzene, C_6H_6, with excess HNO_3 resulted in 0.90 g of H_2O. What is the percentage yield?
    Molar Mass (g/mol): C_6H_6=78 HNO_3=63 $C_6H_5NO_2$=123 H_2O=18

    $$C_6H_6 + HNO_3 \rightarrow C_6H_5NO_2 + H_2O$$

    a. 100%
    b. 90%
    c. 50%
    d. 12%
    e. 2%

11. How many grams of H_2O will be formed when 32.0 g H_2 is allowed to react with 16.0 g O_2 according to

    $$2\ H_2 + O_2 \rightarrow 2\ H_2O$$

    a. 9.00 g
    b. 16.0 g
    c. 18.0 g
    d. 32.0 g
    e. 36.0 g

12. When 2.00 g of H_2 reacts with 32.0 g of O_2 in an explosion, the final gas mixture will contain:
    a. H_2, H_2O, and O_2
    b. H_2 and H_2O only
    c. O_2 and H_2O only
    d. H_2 and O_2 only
    e. H_2O only

13. 11.2 g of metal carbonate, containing an unknown metal, M, were heated to give the metal oxide and 4.4 g CO_2.

    $$MCO_{3(s)} + \text{heat} \rightarrow MO_{(s)} + CO_{2(g)}$$

    What is the identity of the metal M?
    a. Mg
    b. Pb
    c. Ca
    d. Ba
    e. Cr

14. A given sample of some hydrocarbon is burned completely and it produces 0.44 g of CO_2 and 0.27 g of H_2O. Determine the empirical formula of the compound.
    a. CH
    b. C_2H_3
    c. CH_2
    d. C_2H_5
    e. CH_3

15. The simplest formula for a hydrocarbon that is 20.0 percent hydrogen by mass is
    a. CH
    b. CH_2
    c. CH_3
    d. C_2H_2
    e. C_2H_3

16. What mass of Al is produced when 0.500 mole of Al_2S_3 is completely reduced with excess H_2?
    a. 2.7 g
    b. 13.5 g
    c. 27.0 g
    d. 54.0 g
    e. 108 g

17. When a 16.8-gram sample of an unknown mineral was dissolved in acid, 4.4-grams of CO_2 were generated. If the rock contained no carbonate other than $MgCO_3$, what was the percent of $MgCO_3$ by mass in the limestone?
    Molar mass (g/mol): $MgCO_3$ = 84 and CO_2 = 44
    a. 33%
    b. 50%
    c. 67%
    d. 80%
    e. 100%

18. Which of the following represents the correct method for converting 11.0 g of copper metal to the equivalent number of copper atoms?
    a. $11\left(\frac{1}{63.55}\right)\left(\frac{6.02\times10^{23}}{1}\right)$
    b. $11\left(\frac{1}{63.55}\right)$
    c. $11\left(\frac{1}{63.55}\right)\left(\frac{63.55}{6.02\times10^{23}}\right)$
    d. $11\left(\frac{63.55}{1}\right)\left(\frac{6.02\times10^{23}}{1}\right)$
    e. $11\left(\frac{1}{63.55}\right)\left(\frac{1}{6.02\times10^{23}}\right)$

19. The mass of element X found in 1.00 mole of each of four different compounds is 28.0 g, 42.0 g, 56.0 g, and 70 g, respectively. The possible atomic weight of X is
   a. 8.00
   b. 14.0
   c. 28.0
   d. 38.0
   e. 42.0

$$N_{2(g)} + 2O_{2(g)} \rightarrow N_2O_{4(g)}$$

20. The above reaction takes place in a closed flask. The initial amount of $N_{2(g)}$ is 8 mole, and that of $O_{2(g)}$ is 12 mole. There is no $N_2O_{4(g)}$ initially present. The experiment is carried out at constant temperature. What is the total amount of mole of all substances in the container when the amount of $N_2O_{4(g)}$ reaches 6 mole?
   a. 0 mole
   b. 2 mole
   c. 6 mole
   d. 8 mole
   e. 20 mole

21. Given that there are two naturally occurring isotopes of gallium, ^{69}Ga and ^{71}Ga, the natural abundance of the ^{71}Ga isotope must be approximately
   a. 25 %
   b. 40 %
   c. 50 %
   d. 71 %
   e. 90 %

$$2Ca_3(PO_4)_2 + 10C + 6SiO_2 \rightarrow P_4 + 6CaSiO_3 + 10CO$$

22. Elemental phosphorus can be produced by the reduction of phosphate minerals in an electric furnace. What mass of carbon would be required to produce 0.2 mol of P_4 in the presence of 1 mol of calcium phosphate and 3 mol of silicon dioxide?
   a. 2 g
   b. 12 g
   c. 24 g
   d. 60 g
   e. 120 g

23. In which of the following compounds is the mass ratio of element X to oxygen closest to 2.5 to 1? (The molar mass of X is 40.0 g/mol.)
   a. X_5O_2
   b. X_3O_2
   c. X_2O
   d. XO_2
   e. XO

$$6H^+ + 5H_2O_2 + 2MnO_4^- \rightarrow 5O_2 + 2Mn^{2+} + 8H_2O$$

24. According to the balanced equation above, how many moles of the permanganate ion are required to react completely with 25.0 ml of 0.100 M hydrogen peroxide?
   a. 0.000500 mol
   b. 0.00100 mol
   c. 0.00500 mol
   d. 0.00625 mol
   e. 0.0100 mol

For the Free Response you may use a calculator and Periodic Table.

25. A 10.0 g sample containing calcium carbonate and an inert material was placed in excess hydrochloric acid. A reaction occurred producing calcium chloride, water, and carbon dioxide.
   (a) Write a balanced equation for the reaction.

   (b) When the reaction was complete, 1.55 g of carbon dioxide gas was collected. How many moles of calcium carbonate were consumed in the reaction?

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

   (d) If the inert material was only silicon dioxide, what was the mole fraction of silicon dioxide in the mixture?

$$\text{mole fraction} = N_1 = \frac{n_1}{n_{total}}$$

   (e) In fact perhaps there had been some other material present in the original sample that was not so inert and generated a gas during the reaction. Would this have caused the calculated percentage of calcium carbonate in the sample to be higher, lower or have no effect? Justify your response.

Practice Test Ch 3 Stoichiometry *ANSWERS*

1. d It might be easiest to balance the equation with mostly whole numbers: $2\,NH_3 + \frac{7}{2}O_2 \rightarrow 2NO_2 + 3H_2O$. The question asks about the amount of oxygen reacting with ONE mole of ammonia, thus cut the $\frac{7}{2}$ (3.5) of oxygen in half to 1.75
2. c Balance: $CH_3CHO + \frac{5}{2}O_2 \rightarrow 2CO_2 + 2H_2O$ *Note: If you are having trouble balancing equations, you MUST act fast on the first day of school and get in for some extra help.*
3. e Balance: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ Then do some stoichiometry using "easy math" 16 g of methane (MM = 16) is 1 mole and 1 mole of methane will produce 1 mole of CO_2 = 44 g, and 2 moles of H_2O which is 36 g for a total of 80 g
4. d Balance: $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$
5. d Balance: $2KClO_3 \rightarrow 2KCl + 3O_2$
6. c In multiple choice questions without a calculator, you must look for the "easy math" − You will be most successful at this if you put all the numbers in the dimensional analysis on the page and look for common factors you can cancel out.
$$27gAl\left(\frac{1mol}{27g}\right)\left(\frac{3H_2}{2Al}\right)\left(\frac{2g}{1mol}\right) = 3\text{ g }H_2$$
7. c First you must realize this is a limiting reactant problem. You can tell this since you are given quantities for both reactants. Convert both values to moles: $138gNO_2\left(\frac{1mol}{46g}\right) = 3molNO_2$ and $54gH_2O\left(\frac{1mol}{18g}\right) = 3molH_2O$ Clearly the NO_2 limits since the balanced equation tells us that 3 moles are required for every one mole of water, thus use the limiting reactant to determine the amount of acid that can be produced. $3molNO_2\left(\frac{2HNO_3}{3NO_2}\right)\left(\frac{63g}{1mol}\right) = 126gHNO_3$
8. e Add the molar mass of $CaCO_3$ ($40 + 12 + 3\times16 = 100$), thus I is correct. Remember that molar mass is the mass of one mole. The units are often shown as g/mol, but g mol^{-1} is equivalent since raised to the −1 power means reciprocal, or in the denominator. Since I is true, III is equally true, because 2 moles would weigh twice as much. II is also true because 50 g is 0.5 mole of the compound, and since there are 3 moles of oxygen atoms per mole of compound,
$$50gCaCO_3\left(\frac{1molCaCO_3}{100gCaCO_3}\right)\left(\frac{3molO}{1molCaCO_3}\right)\left(\frac{6.02\times10^{23}atoms}{1mol}\right) = 9\times10^{23}Oatoms$$
9. c The quickest way to determine limiting reactant is to convert to moles and divide each mole value by the coefficient in the balanced equation. Whichever substance turns up as the smallest number will be the limiting reactant. Again remember to use approximations since you would not have access to a calculator to do any of these MC questions. Thus,
$$100.gMnO_2\left(\frac{1mol}{86.9g}\right) \div 2 = just > 0.5molMnO_2$$
$$100.gKOH\left(\frac{1mol}{56.1g}\right) \div 4 = just < 0.5molKOH \quad \text{thus the KOH limits the reaction}$$
$$100.gO_2\left(\frac{1mol}{32g}\right) \div 1 = \ \sim 3molO_2$$
$$100.gCl_2\left(\frac{1mol}{71g}\right) \div 1 = \ \sim 1.5molCl_2$$
10. c Look for easy approximations: $7.8gC_6H_6\left(\frac{1mol}{78g}\right) = 0.1molC_6H_6$ and continue $0.1molC_2H_6\left(\frac{1H_2O}{1C_6H_6}\right)\left(\frac{18g}{1mol}\right) = 1.8gH_2O$ then percent yield also will be easy values: $\left(\frac{0.9gH_2O}{1.8gH_2O}\right)\times100 = 50\%H_2O$
11. c Again, this is a limiting reactant problem, 32 g of H_2 is 16 moles and is far in excess of the 0.5 moles of O_2 available. Thus the O_2 limits the reaction. Set up the dimensional analysis and look for easy factors.
$$16gO_2\left(\frac{1mol}{32g}\right)\left(\frac{2H_2O}{1O_2}\right)\left(\frac{18g}{1mol}\right) = 18gH_2O$$
12. c First you should recognize that you need to write a balanced equation: $2H_2 + O_2 \rightarrow 2H_2O$ Next, realize that this is a limiting reactant problem 2 g of H_2 is 1 mole and 32 g of O_2 is 1 mole. The 1 mole of H_2 will limit the reaction, thus we know that after the reaction stops, O_2 will remain in the container with the H_2O produced.

13. e This problem requires a bit more clever thought and understanding. Work the problem "backwards" starting with the information that the carbon dioxide product tells you. Since $4.4 g CO_2 \left(\frac{1 mol}{44 g}\right) = 0.1 mol CO_2$, the balanced equation tells us that 0.1 mol of CO_2 must come from 0.1 mole of MCO_3. Further, we know that 0.1 mole of MCO_3 contains 0.1 mole of M and also 0.1 mole of CO_3. Since CO_3 has a mass of 60 g/mol, we can calculate that the 0.1 mole of the CO_3 has a mass of 6 g and the remainder of the compound (whose mass we were told was 11.2 g), the M part, must weigh 5.2 g. But remember there is 0.1 mole of M in this compound, thus $\left(\frac{5.2 g}{0.1 mol}\right) = \left(\frac{52 g}{1 mol}\right)$ and so you are looking for an element that has a molar mass of 52, which is Cr.

14. e For this problem you must realize that the combustion of any hydrocarbon (CH compounds) produces only water and carbon dioxide. Use the amount of CO_2 to inform how much C must be in the original hydrocarbon, then use the amount of water produced to determine the amount of H that must be in the original hydrocarbon:

For carbon: $0.44 g CO_2 \left(\frac{1 mol}{44 g}\right) = 0.01 mol CO_2 \left(\frac{1C}{1CO_2}\right) = 0.01 mol C$ and

For hydrogen: $0.27 g H_2O \left(\frac{1 mol}{18 g}\right) = 0.015 mol H_2O \left(\frac{2H}{1H_2O}\right) = 0.03 mol H$ a 0.01/0.03 ratio is 1:3, thus CH_3

15. c Again, it is important to know that a hydrocarbon is any compound made of just hydrogen and carbon. Probably the easiest way to work this problem would be to use the answers and calculate the mass ratio for the mass of H to the total mass of compound: CH is 1/13, CH_2 is 3/14, CH_3 is 3/15, C_2H_2 is 2/26, and C_2H_3 is 3/27.

The CH_3 compound: 3/15 should pop out as ⅕ which is of course 20%

16. c You could consider the chemical reaction, ($Al_2S_3 \rightarrow 2 Al + 3S$) or you could simply realize that since the Al would be written as a monatomic atom in the balanced equation and since there are 2 Al's in the compound, there must be a 2:1 ratio between the Al_2S_3 and Al : $0.5 mol Al_2S_3 \left(\frac{2Al}{1Al_2S_3}\right) = 1 mole Al \left(\frac{27 g}{1 mol}\right) = 27 g Al$

17. b As in problem 13, it is an important concept to realize that metal carbonates decompose or react with acid to produce the same amount of moles of carbon dioxide as the moles of the original compound. All metal carbonates will react with acid to produce carbon dioxide and a metal oxide. This reaction is specifically: $MgCO_3 + H^+ \rightarrow MgO + CO_2$. The original 16.8 g sample contains both $MgCO_3$ and some other inert substances that do not react with acid to produce any gas.

$4.4 g MgCO_3 \left(\frac{1 mol}{44 g}\right) = 0.1 mol MgCO_3 \left(\frac{1 mol MgCO_3}{1 mol CO_2}\right)\left(\frac{84 g}{1 mol}\right) = 8.4 g MgCO_3$ thus $\left(\frac{8.4 g MgCO_3}{16.8 g sample}\right) \times 100 = 50\%$

18. a Putting in the correct units, allows you to see the dimensional analysis will work.

$11 g \left(\frac{1 mol}{63.55 g}\right)\left(\frac{6.02 \times 10^{23} atoms}{1 mol}\right) = Cu\ atoms$, the actual value would be approximately 1×10^{23}, but the value is unimportant for this question.

19. b In any compound that contains some element X, the number of atoms of X will always be whole numbers 1, 2, 3, etc (since you can't have half an atom). Thus you must look for a factor that is common to each of the masses of X provided.

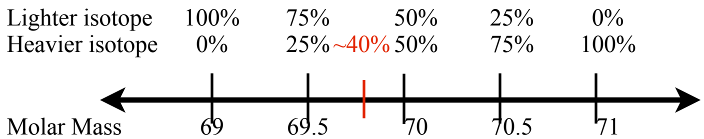
20. d This problem is a simple stoichiometry problem that you can certainly do without a calculator. When the reaction reaches a quantity of 6 mole of product (remember the problem states that there was no product, N_2O_4 to start with), use the coefficients in the balanced equation to determine that 12 mole of O_2 and 6 mole of N_2 must have reacted to produce the 6 mole of product.

$6 mol N_2O_4 \left(\frac{2 mol O_2}{1 mol N_2O_4}\right) = 12 mol O_2$ and since the reaction was started with only 12 mole of O_2, there will be none left

$6 mol N_2O_4 \left(\frac{1 mol N_2}{1 mol N_2O_4}\right) = 6 mol N_2$ and since the reaction was started with 8 mole of N_2, there will be 2 mole of N_2 left.

Thus 2 mole N_2 left over, no O_2 left over, and 6 mole of N_2O_4 produced, will mean a total of 8 mole of substances left in the flask.

21. b For this problem, you must have the good sense to go to the periodic table and look up the average molar mass of gallium and see that it has a value of 69.7, and then go back to view the answers to realize that you do NOT need to do an exact calculation. If there were 50% of each of the two isotopes, the average molar mass would have to be right in the middle at 70, and since the molar mass is lower than 70, you now have narrowed it down to either a) or b). If there were only 25 % of the heavier isotope and 75% of the lighter isotope, the molar mass would have to be 69.5, and since the molar mass is greater than 69.5 is not, the remaining choice, between 50% and 25% leaves the option of b) 40%. Perhaps the number line shown below will help.



22. c In this problem you want to produce 0.2 mol of P_4, thus you must figure the minimum amount of C to produce that. The quantities given for calcium phosphate and silicon dioxide are just distractors. Although it is important that you confirm that you have enough of these two reactants to produce the 0.2 mol of P_4 that the problem asks for.

$0.2 molP_4 \left(\frac{10C}{1P_4}\right) = 2molC$ and finish the mass calculation: $2molC\left(\frac{12g}{1mol}\right) = 24gC$

It is worth confirming that you have enough of each of the other reactants as given in the problem.

So calculate: $0.2molP_4\left(\frac{2Ca_3(PO_4)_2}{1P_4}\right) = 0.4molCa_3(PO_4)_2 required$ is enough as 1 mol was given

and $0.2molP_4\left(\frac{6SiO_2}{1P_4}\right) = 1.2molSiO_2 required$ is also enough as 3 mol was given.

23. e In this problem, if the mass ratio is to be 2.5 X to 1 Oxygen, since the mass of oxygen is 16 calculate the mass for X that is 2.5 times greater than 16. This would be 40, which is the molar mass of X, thus the formula must be XO. Remember, the MC questions without your calculator will be easy math, and I will try to show you the easiest ways to "see" the "easy math."

24. b Remember the definition of molarity $M = \left(\frac{molesSolute}{Vol(inL)ofSolution}\right)$ Thus you can use the volume and molarity to calculate the moles of hydrogen peroxide given and then use the stoichiometry to convert to moles of permanganate required

$(0.025L)(0.1\frac{mol}{L})\left(\frac{2MnO_4^-}{5H_2O_2}\right) = 0.001mol\ MnO_4^- required$ watch for your "easy math" factors.

25. **Free Response** − Remember you will be able to use a calculator for FR questions.

a. $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$ (This should "appear" to be a double replacement reaction: ($CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2CO_3$) This is correct, however, whenever carbonic acid, H_2CO_3, is written as a product, it is best (expected by AP) that you will rewrite it as decomposed into H_2O and CO_2, thus the first reaction written is the correct version.

b. Convert the mass of carbon dioxide to moles, and use stoichiometry back to the moles of calcium carbonate:

$1.55gCO_2 \times \frac{1mol}{44g} = 0.0352molCO_2$ then use the stoichiometry $0.0352molCO_2\left(\frac{1CaCO_3}{1CO_2}\right) = 0.0352molCaCO_3$

c. change moles of b) to grams: $0.0352molCaCO_3\left(\frac{100g}{1mol}\right) = 3.52gCaCO_3$ and then calculate the part out of whole

$\left(\frac{3.52gCaCO_3}{10gSample}\right) \times 100\% = 35.2\%$

d. subtract for the remaining mass of the sample $10gSample - 3.52gCaCO_3 = 6.48gSiO_2 \times\left(\frac{1mol}{60.1g}\right) = 0.108molSiO_2$

then use the moles of calcium carbonate previously determined to calculate the total moles of substances in the mixture $0.1078molSiO_2 + 0.0352molCaCO_3 = 0.143molTotal$

and lastly calculate silicon dioxide moles out of the total moles to get mole fraction: $\frac{0.1078molSiO_2}{0.143molTotal} = 0.754$

e. if an other material (other than silicon dioxide) were present, more gas would have formed that you would have thought was from carbonate, resulting in an inaccurate *higher percentage of calcium carbonate.*