JOE PERFECT

Practice Test - Chapter 3 - Stoichiometry

Target 1. I can predict the products for and write balanced equations for the following types of reactions: combustion, decomposition, sysnthesis (called combination reactions), single displacement and double displacement reactions.

1. Predict the products for each equation that follows. Balance each equation. Classify it as combustion (C), decomposition (D), sysnthesis (S), single displacement (SD) or double displacement (DD). Place the letters of the classification on the blanks at left.

DD a. $Al_2(CO_3)_3$ + 6 $NaNO_3$ → 2 $Al(NO_3)_3$ + 3 Na_2CO_3 (Al^{+3}, CO_3^{-2}, Na^{+1}, NO_3^{-1})

C b. 4 C_5H_9O + 27 O_2 → 20 CO_2 + 18 H_2O

D c. 2 Al_2O_3 → 4 Al + 3 O_2

SD d. 2 LiF + Ca → CaF_2 + 2 Li

S e. 2 Sr + O_2 → 2 SrO ($Sr^{+2}O^{-2}$)

2. When sodium carbonate decomposes during heating, two products are formed. What are the two products?
a. Na and CO_3 b. Na_2O and C c. Na and CO_2 d. Na and CO (e.) Na_2O and CO_2

$Na_2CO_3 \xrightarrow{\Delta} Na_2O + CO_2$

You always get a metal oxide & CO_2 whenever a carbonate decomposes.

Target 2: I can interconvert between the number of moles and mass of a substance. I can also use Avogadro's number and molar mass to calculate the number of particles (atoms, molecules or formula units) making up a substance.

3. What is the mass in grams of 7.2×10^{22} molecules of H_2O? (No calculator!)
(a.) 2.2 g b. 0.0022 g c. 2.2×10^3 g d. 220,000 g e. 2.2×10^{45} g

4. Which of the following is the most massive?
a. 5.85 grams of NaCl
(b.) 0.500 mole of NaCl
c. 115,000 atoms of gold
d. 1.00×10^5 ng of lead
e. 250 molecules of propane (C_3H_8)

see work page!

NOTE: No calculators are allowed for the multiple choice!!

5. How many sulfur atoms are there in 25 molecules of $C_4H_4S_2$?

a. 1.5×10^{25}
b. 4.8×10^{25}
c. 3.0×10^{25}
d. 50
e. 6.02×10^{23}

Target 3: I can calculate the percentage composition of a compound by mass.

6. What is the percent by mass of hydrogen in perchloric acid?

a. 1.0 % b. 3.0 % c. 6.0 % d. 23 % e. 46%

7. Which element in sodium acetate has the greatest percentage by mass?

a. Na b. C c. H d. O e. Xe

Target 4: I can calculate the empirical formula of a compound, having been given either:
a) mass or % composition, or
b) the mass of CO_2 and H_2O produced by combustion.

8. A compound that is composed of only hydrogen and carbon contains 80.0% carbon and 20.0% hydrogen. What is the empirical formula of this compound?

a. $C_{20}H_{60}$
b. C_7H_{20}
c. CH_3
d. C_3H_6
e. $C_{20}H_7$

9. Consider the following table of molar masses for elements X, Y and Z.

Element	**X**	**Y**	**Z**
Molar mass (g/mol)	20.0	30.0	40.0

An unknown compound contained 60.0 grams of X, 45.0 grams of Y and 180 grams of Z. Calculate the empirical formula of this unknown compound.

a. $X_2Y_2Z_3$
b. XY_2Z_3
c. XYZ_2
d. X_4Y_2Z
e. X_2YZ_3

Target 5: I can calculate the molecular formula, having been given the empirical formula and the molecular weight.

10. The empirical formula of a compound is C_3H_8O. The molar mass of the compound is 180 g/mol. What is the molecular formula of the compound?
    a. C_3H_8O b. $C_3H_{16}O_2$ c. $C_6H_{16}O_2$ d. $C_9H_{24}O_3$ e. $C_9H_{16}O$

11. The empirical formula of a compound is N_2O. The molar mass of the compound is 44 g/mol. What is the molecular formula of the compound?
    a. N_2O b. N_2O_2 c. N_2O_4 d. N_3O_7 e. N_4O_8

12. A phosphorous oxide compound contains 43.7% oxygen by mass. The molecular formula of this compound could be ______ .
    a. P_7O_2 b. PO_7 c. P_2O d. P_5O_3 e. P_4O_6

Target 6: I can use stoichiometry to solve problems involving chemical reactions.

Consider the following combustion reaction for #13 and #14:

$$2\,C_4H_{10} + 13\,O_2 \rightarrow 8\,CO_2 + 10\,H_2O$$

13. Assume that 6.0 moles of butane (C_4H_{10}) burn in excess oxygen. How many grams of water are produced?
    a. 5.4 b. 540 c. 1.1 d. 110 e. 1.1×10^3

14. Assume 8.0 grams of oxygen react with excess butane. How many grams of CO_2 are produced?
    a. 0.067 b. 6.7 c. 27 d. 270 e. 2.7×10^7

Target 7: I can determine the limiting reactant in a reaction and determine the amount of excess reactant left over from a reaction.

Consider the following reaction for # 15 and #16:

$$N_2 + 3\,H_2 \rightarrow 2\,NH_3$$

15. Assume that 0.10 grams of H_2 react with 0.56 grams of N_2. The limiting reactant is ______ .
    a. N_2 b. H_2 c. NH_3 d. Both H_2 and N_2

16. How many grams of NH_3 can be produced if 20.0 grams of H_2 react with 168 grams of N_2?
    a. 3.98 b. 39.8 c. 398 d. 1.13 e. 113

17. Consider the following reaction:

$$Mg + 2\,H_2O \rightarrow Mg(OH)_2 + H_2$$

Assume that 48.6 grams of magnesium react with 36.0 grams of water. Which reactant is the excess reactant? How many grams are left over or in excess after the reaction is complete?

a. Mg, 12.2 (b.) Mg, 24.3 c. H_2O, 9.00 d. H_2O, 18.0

Target 8: I can calculate the theoretical and actual yields of a chemical reaction when given the appropriate data.

18. Consider the following reaction:

$$Mg_3N_2 + 3\,H_2O \rightarrow 2\,NH_3 + 3\,MgO$$

A lab was performed by students in which they mixed a specific amount of Mg_3N_2 and H_2O. They produced 15 grams of MgO in the lab. Theoretically, they should have produced 18 grams. What their percent yield?

a. 17 % b. 25 % c. 50 % (d.) 83 % e. 117 %

19. Consider the following reaction:

$$2\,S + 3\,O_2 \rightarrow 2\,SO_3$$

Billy reacted 8.0 grams of sulfur with excess oxygen and was able to collect 15 grams of SO_3. What was Billy's percent yield?

a. 5.0 % b. 25 % c. 50 % (d.) 75 % e. 125 %

Part 2: Answer each of the following questions on separate sheets of paper.

1. Predict the products for the following reactions and write a balanced equation for each:
   a) The combustion of glucose.
   b) The synthesis reaction between potassium and chlorine gas.
   c) The decomposition of magnesium carbonate.
   d) Reacting magnesium oxide and water.
   e) The decomposition of sodium chlorate.

2. Below is a chart containing data for the three naturally occurring isotopes of Mg:

Isotope	abundance (%)	mass (u)
Mg-24	78.70	23.98504
Mg-25	10.13	24.98584
Mg-26	11.17	25.98259

Calculate the atomic mass of magnesium.

3. Calculate the percentage of oxygen (by mass) in nickel (II) acetate.

4. Assume you have 5.0 liters of water. Calculate each of the following:
   a) the number of grams of water.
   b) the number of moles of water.
   c) the number of molecules of water.
   d) the number of hydrogen atoms in this sample of water.

5. Antifreeze is composed of 51.6 % oxygen, 9.70% hydrogen, and 38.7% carbon by mass. The molar mass of antifreeze is 62.1 g/mol. Calculate its empirical and molecular formulas.

6. Menthol, the substance we can smell in mentholated cough drops, is composed of C, H, and O. A 0.1005-g sample of menthol is combusted, producing 0.2829 g CO_2 and 0.1159 g of H_2O. What is the empirical formula of menthol? If the compound has a molecular mass of 156 g/mol, what is its molecular formula?

7. When a mixture of 10.0 g of acetylene, C_2H_2, and 10.0 g of oxygen, O_2, is ignited, the resultant combustion produces CO_2 and H_2O.
   a) Write the balanced equation for this reaction.
   b) Which reactant is the limiting reactant?
   c) How many grams of C_2H_2, O_2, CO_2, and H_2O are present after the reaction is complete?

WORK! FOR PRACTICE TEST CHAPTER 3

#3 $x\ g\ H_2O = \frac{7.2\times10^{22}\ \text{molecules}}{} \cdot \frac{1\ \text{mole}\ H_2O}{6.02\times10^{23}\ \text{molecules}} \cdot \frac{18\ g\ H_2O}{1\ \text{mole}\ H_2O}$

$\approx \frac{(7.2\times10^{22})(18)}{6\times10^{23}} \approx (1\times10^{-1})(18) = (0.1)(18) = \boxed{1.8\ g}$ "A"

There is only one answer close to this!

#4 (a) 5.85 g

(b) $x\ g\ NaCl = \frac{0.5\ \text{mol NaCl}}{} \cdot \frac{58.5\ g\ NaCl}{1\ \text{mol}} \approx 29\ g$ By far the most massive!! ☺

(c) $x\ g\ Au = \frac{115{,}000\ \text{atoms}}{} \cdot \frac{1\ \text{mole Au}}{6.02\times10^{23}\ \text{atoms Au}} \cdot \frac{197\ g}{1\ \text{mol Au}} \approx \frac{(1.2\times10^{5})(2.0\times10^{2})}{6.02\times10^{23}}$ ← really small!

(d) $x\ g\ Pb = \frac{1\times10^{5}\ ng}{} \cdot \frac{1\ g}{1\times10^{9}\ ng} = 10^{-4}\ g\ Pb$ really small

(e) Common sense tells you that this choice isn't correct.

#5 $X\ S\ \text{atoms} = \frac{25\ \text{molecules}\ C_4H_4S_2}{} \cdot \frac{2\ \text{atoms S}}{1\ \text{molecule}\ C_4H_4S_2} = 50$ atoms of sulfur

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#6 Perchloric acid = $HClO_4$ ☺

H 1 x 1.01 = 1.01
Cl 1 x 35.5 = 35.5
O 4 x 16.0 = 64.0
100.51 g/mole

$\%H = \frac{\text{mass H}}{\text{total mass}} \times 100$

$\%H = \frac{1.01}{100.51} \times 100 \approx 1\%$

#7 $NaC_2H_3O_2$

Na $1 \times 23 = 23$
C $2 \times 12 = 24$
H $3 \times 1 = 3$
O $2 \times 16 = 32$

82 g/mole

$\frac{32}{82} \times 100$ = greatest % by mess of the 4 elements ☺

#8

$$X \text{ mol C} = \frac{80.0 \text{ g C}}{} \left| \frac{1 \text{ mole C}}{12.0 \text{ g C}} \right| = \frac{80}{12} = \frac{40}{6} = 6\tfrac{2}{3} = 6.7 \text{ mol C} \div 6.7 = \boxed{1}$$

$$X \text{ mol H} = \frac{20.0 \text{ g H}}{} \left| \frac{1 \text{ mol H}}{1.0 \text{ g H}} \right| = 20 \text{ mol H} \div 6.7 \approx \boxed{3}$$

CH_3

#9

$$X \text{ mol X} = \frac{60 \text{ g}}{} \left| \frac{1 \text{ mol}}{20 \text{ g}} \right| = 3 \text{ mol X} \div 1.5 = \boxed{2}$$

$$X \text{ mol Y} = \frac{45 \text{ g}}{} \left| \frac{1 \text{ mol}}{30 \text{ g}} \right| = 1.5 \text{ mol Y} \div 1.5 = \boxed{1}$$

$$X \text{ mol Z} = \frac{180 \text{ g Z}}{} \left| \frac{1 \text{ mol}}{40 \text{ g Z}} \right| = 4.5 \text{ mol Z} \div 1.5 = \boxed{3}$$

X_2YZ_3

10

First, find mass of empirical formula: C_3H_8O

C: $3 \times 12 = 36$
H: $8 \times 1 = 8$
O: $1 \times 16 = 16$

60 g/mole

Next, find ratio of $\frac{\text{MF mass}}{\text{EF mass}}$ ← given ☺

$\frac{180}{60} = \boxed{3}$ ... EF × 3 = MF ; $C_3H_8O \times 3 = C_9H_{24}O_3$

↑ This will always be a whole #.

Note: An alternate method would be to find the molar mass of the 5 choices. Only one of them will be 180 g/mole. This might be faster. ☺

11 N_2O : N 2 x 14 = 28
O 1 x 16 = 16
44 g/mole ⇒ This is the unique case in which the EMPIRICAL FORMULA is the same as the MOLECULAR FORMULA.

EF = MF

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12 First, find the EMPIRICAL FORMULA! 43.7 % oxygen & 56.3 % phosphorous

$$X \text{ mol P} = \frac{56.3 \text{ g P}}{} \Big| \frac{1 \text{ mole}}{31 \text{ g P}} \approx 2$$

$$X \text{ mol O} = \frac{43.7 \text{ g O}}{} \Big| \frac{1 \text{ mol}}{16. \text{ g O}} \approx 3$$

P_2O_3 E.F. — only reasonable answer is P_4O_6

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13 $$X \text{ g } H_2O = \frac{6 \text{ mol } C_4H_{10}}{} \Big| \frac{10 \text{ mol } H_2O}{2 \text{ mol } C_4H_{10}} \Big| \frac{18 \text{ g } H_2O}{1 \text{ mol } H_2O} = (6)(10)(9) = 60 \times 9$$

= 540 grams

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14 $$X \text{ g } CO_2 = \frac{8 \text{ g } O_2}{} \Big| \frac{1 \text{ mol } O_2}{32 \text{ g } O_2} \Big| \frac{8 \text{ mol } CO_2}{13 \text{ mol } O_2} \Big| \frac{44 \text{ g } CO_2}{1 \text{ mol } CO_2} = \frac{(8)(8)(44)}{(32)(13)} = \frac{(8)(44)}{(4)(13)} = \frac{(2)(44)}{13}$$

$$= \frac{88}{13} \approx$$ {6 or 7 ish} choice "B" is the only logical choice!

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15 Calculate moles of each reactant... then use the balanced equation!

$$X \text{ mol } H_2 = \frac{0.1 \text{ g } H_2}{} \Big| \frac{1 \text{ mol } H_2}{2.0 \text{ g}} = 0.05 \text{ mol } H_2$$

$$X \text{ mol } N_2 = \frac{0.56 \text{ g } N_2}{} \Big| \frac{1 \text{ mol } N_2}{28 \text{ g } N_2} = 0.02 \text{ mol } N_2$$

⇒ H_2 is limiting reactant. The balanced equation says you need 3X more H_2 than N_2... and 0.05 is NOT 3X bigger than 0.02.

[16] $x \text{ mol } H_2 = \frac{20.0\text{ g } H_2 \mid 1\text{ mol } H_2}{\mid 2.0\text{ g } H_2} = 10 \text{ mol } H_2$

$x \text{ mol } N_2 = \frac{168\text{ g } N_2 \mid 1\text{ mol } N_2}{\mid 28\text{ g } N_2} = 6 \text{ mol } N_2$

Use the bal. equation to determine the limiting reactant... which is H_2! Use H_2 to solve.

$x \text{ g } NH_3 = \frac{10\text{ mol } H_2 \mid 2\text{ mol } NH_3 \mid 17\text{ g } NH_3}{\mid 3\text{ mol } H_2 \mid 1\text{ mol } NH_3} = \frac{340}{3} \approx$ {113 grams}

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[17] Convert each reactant to moles. Then, use the balanced equation to determine the excess reactant.

$x \text{ mol } Mg = \frac{48.6\text{ g } Mg \mid 1\text{ mol } Mg}{\mid 24.3\text{ g } Mg} = 2 \text{ mol } Mg$ (excess)

$x \text{ mol } H_2O = \frac{36.0\text{ g } H_2O \mid 1\text{ mol } H_2O}{\mid 18\text{ g } H_2O} = 2 \text{ mol } H_2O$ (L.R.)

Using the bal. equation, you can see that the Mg is in excess. Only 1 mol of Mg reacts with the 2 mol H_2O. You will have one mole of Mg left = 24.3 grams!

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[18] $\% \text{ Yield} = \frac{AY}{TY} \times 100$; $\% \text{ Yield} = \frac{15\text{ g}}{18\text{ g}} \times 100 = \frac{5}{6} \times 100 =$ {83 %} only reasonable answer

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[19] First, calculate Billy's theoretical yield... that is, the amount that he should have gotten in the lab.

$x \text{ g } SO_3 = \frac{8.0\text{ g } S \mid 1\text{ mol } S \mid 2\text{ mol } SO_3 \mid 80\text{ g } SO_3}{\mid 32\text{ g } S \mid 2\text{ mol } S \mid 1\text{ mol } SO_3} = \frac{(8)(2)(80)}{(32)(2)} =$ **20 grams SO_3** → THEO. YIELD!

$\% \text{ Yield} = \frac{AY}{TY} \times 100$; $\frac{15\text{ g}}{20\text{ g}} \times 100 =$ **75 % Yield** {25 % error!}

PART 2 ANSWER KEY

① (a) $C_6H_{12}O_6 + \underline{6}\, O_2 \rightarrow \underline{6}\, CO_2 + \underline{6}\, H_2O$

(b) $2K + Cl_2 \rightarrow 2KCl$

(c) $MgCO_3 \rightarrow MgO + CO_2$ * Carbonates decompose forming CO_2.

(d) $MgO + H_2O \rightarrow Mg(OH)_2$ * Metal oxides added to water produces metal hydroxides!

(e) $2\,NaClO_3 \rightarrow 2\,NaCl + 3\,O_2$ * Chlorates decompose forming O_2.

② ATOMIC MASS $= \dfrac{(78.70)(23.98504) + (10.13)(24.98584) + (11.17)(25.98259)}{100} = $ **24.33 amu**

③ $Ni(C_2H_3O_2)_2$

Ni $1 \times 58.69 = 58.69$
C $4 \times 12.01 = 48.04$
H $6 \times 1.008 = 6.048$
O $4 \times 16.00 = 64.00$
176.78 g/mol

% oxygen $= \dfrac{\text{mass O}}{\text{total mass}} \times 100$

% O $= \dfrac{64.00}{176.78} \times 100$

% O = 36.20 %

④ (a) Since given no other information, we will assume the density of water is 1.0 g/mL.

$$x\ g\ H_2O = \frac{5.0\ L}{} \left| \frac{1000\ mL}{1\ L} \right| \frac{1\ g}{1\ mL} = 5{,}000\ g = \boxed{5.0 \times 10^3\ g}$$

(2 sig. fig's)

(b)
$$x\ mol\ H_2O = \frac{5.0 \times 10^3\ g}{} \left| \frac{1\ mol\ H_2O}{18.0\ g\ H_2O} \right| = 277.8\ mol\ H_2O = \boxed{280\ moles\ H_2O}$$

(2 sig. fig's)

(c)
$$x\ molecules = \frac{277.8\ mol}{} \left| \frac{6.02 \times 10^{23}\ molecules}{1\ mole} \right| = 1.672 \times 10^{26}\ molecules$$
$$= \boxed{1.7 \times 10^{26}\ molecules\ H_2O}$$

(d)
$$x\ H\ atoms = \frac{1.672 \times 10^{26}\ molecules\ H_2O}{} \left| \frac{2\ H\ atoms}{1\ molecule\ H_2O} \right| = 3.345 \times 10^{26}\ atoms$$
$$= \boxed{3.3 \times 10^{26}\ atoms\ H}$$

(5) Assume you have a 100-g sample.

$$X \text{ mol } O = \frac{51.6 \text{ g } O}{} \cdot \frac{1 \text{ mol } O}{16.00 \text{ g } O} = 3.225 \text{ mol } O \div 3.222 \approx 1$$

$$X \text{ mol } H = \frac{9.70 \text{ g } H}{} \cdot \frac{1 \text{ mol } H}{1.008 \text{ g } H} = 9.623 \text{ mol } H \div 3.222 \approx 3$$

$$X \text{ mol } C = \frac{38.7 \text{ g } C}{} \cdot \frac{1 \text{ mol } C}{12.01 \text{ g } C} = 3.222 \text{ mol } C \div 3.222 = 1$$

CH_3O

EMPIRICAL FORMULA

C 1 × 12.0 = 12.0
H 3 × 1.01 = 3.03
O 1 × 16.0 = 16.0

31.0 g/mol

$$\frac{MF \text{ (mass)}}{EF \text{ (mass)}} = \frac{62.1 \text{ g/mol}}{31.0 \text{ g/mol}} = 2 \qquad MF = 2 \times EF$$

$MF = C_2H_6O_2$

(6) Use the mass of CO_2 & mass of H_2O to calculate the mass of C & H in the menthol! (calculate moles first)

$$X \text{ mol } C = \frac{0.2829 \text{ g } CO_2}{} \cdot \frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2} \cdot \frac{1 \text{ mol } C}{1 \text{ mol } CO_2} = 0.006428 \text{ mol } C \Rightarrow 0.07720 \text{ grams } C$$

$$X \text{ mol } H = \frac{0.1159 \text{ g } H_2O}{} \cdot \frac{1 \text{ mol } H_2O}{18.016 \text{ g } H_2O} \cdot \frac{2 \text{ mol } H}{1 \text{ mol } H_2O} = 0.01287 \text{ mol } H \Rightarrow 0.01297 \text{ grams } H$$

By subtraction, determine mass of oxygen in the menthol.

Total mass of sample = Mass C + mass H + mass O

$0.1005 \text{ g} = 0.07720 \text{ g } C + 0.01297 \text{ g } H + \text{mass } O$

mass oxygen = $0.01033 \text{ g } O \Rightarrow 0.0006456$ moles of oxygen

$$C_{\frac{0.006428}{0.0006456}} H_{\frac{0.01287}{0.0006456}} O_{\frac{0.0006456}{0.0006456}} \Rightarrow C_{10}H_{20}O_1 \Rightarrow \text{molar mass is } 156.3 \text{ g/mole}$$

Empirical Formula: $C_{10}H_{20}O_1$

$$\frac{MF \text{ (mass)}}{EF \text{ (mass)}} = \frac{156 \text{ g/mole}}{156.3 \text{ g/mole}} \approx 1$$

The molecular formula is also $C_{10}H_{20}O$. ☺

(7) (a)

$$\underset{(10.0g)}{\underline{\underline{2}}\, C_2H_2} + \underset{(10.0g)}{\underline{\underline{5}}\, O_2} \rightarrow \underline{\underline{4}}\, CO_2 + \underline{\underline{2}}\, H_2O$$

(b) $x \text{ mol } C_2H_2 = \frac{10.0\text{ g } C_2H_2 \mid 1 \text{ mol } C_2H_2}{26.04\text{ g } C_2H_2} = 0.3840 \text{ mol } C_2H_2$ (EXCESS REACTANT)

$x \text{ mol } O_2 = \frac{10.0\text{ g } O_2 \mid 1 \text{ mol } O_2}{32.00\text{ g } O_2} = 0.3125 \text{ mol } O_2$ (LIMITING REACTANT)

(c) $x \text{ g } CO_2 = \frac{0.3125 \text{ mol } O_2 \mid 4 \text{ mol } CO_2 \mid 44.01\text{ g } CO_2}{5 \text{ mol } O_2 \mid 1 \text{ mol } CO_2} =$ **11.0 grams CO_2**

$x \text{ g } H_2O = \frac{0.3125 \text{ mol } O_2 \mid 2 \text{ mol } H_2O \mid 18.02\text{ g } H_2O}{5 \text{ mol } O_2 \mid 1 \text{ mol } H_2O} =$ **2.25 g H_2O**

$x \text{ g } C_2H_2$ "used" $= \frac{0.3125 \text{ mol } O_2 \mid 2 \text{ mol } C_2H_2 \mid 26.04\text{ g } C_2H_2}{5 \text{ mol } O_2 \mid 1 \text{ mol } C_2H_2} = 3.26\text{ g } C_2H_2$ used

  10.0 g present initially
− 3.26 g reacts

6.74 g C_2H_2 excess

6.7 g C_2H_2 excess

no oxygen is left at the end of the reaction as it is the limiting reactant!

To double check that the Law of Conservation of Mass is obeyed ... (INITIAL MASS = FINAL MASS)

INITIAL MASS OF REACTANTS: 10.0 g + 10.0 g = **20.0 grams**

FINAL MASS OF PRODUCTS: 11.0 g CO_2 + 2.25 g H_2O + 6.7 g C_2H_2

= 20.0 grams