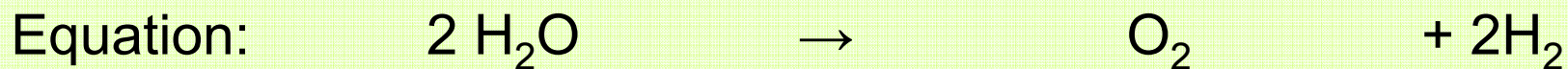

Stoichiometry Worksheet 2:

Percent Yield

For each of the problems:

- a. Write the balanced chemical equation
 - b. Identify the given (with units) and what you want to find (with units)
 - c. Show set up with units. Check sig figs, give final answer with units and label.
-

- 1. Using the Hoffman apparatus for electrolysis, a chemist decomposes 36 g of water into its gaseous elements. How many grams of hydrogen gas should she get (theoretical yield)?



Before:

Change:

After:

Change grams to moles!

Only moles go in the BCA table!

- $36 \text{ g} \times \frac{1 \text{ mole H}_2\text{O}}{18.02 \text{ g}} = 2.0 \text{ mole}$

- 1. Using the Hoffman apparatus for electrolysis, a chemist decomposes 36 g of water into its gaseous elements. How many grams of hydrogen gas should she get (theoretical yield)?

Equation: 2 H₂O → O₂ + 2H₂

Before: 2.0 0 0

Change:

After:

Calculate the change!

$$2.0 \text{ mol } H_2O \times \frac{2 \text{ mol } H_2}{2 \text{ mol } H_2O} = 2 \text{ mol } H_2$$

$$2.0 \text{ mol } H_2O \times \frac{1 \text{ mol } O_2}{2 \text{ mol } H_2O} = 1 \text{ mol } O_2$$

Equation: 2 H₂O → O₂ + 2H₂

Before: 2.0 0 0

Change:

After:

- 1. Using the Hoffman apparatus for electrolysis, a chemist decomposes 36 g of water into its gaseous elements. How many grams of hydrogen gas should she get (theoretical yield)?

Equation:	$2 \text{H}_2\text{O}$	\rightarrow	O_2	$+ 2\text{H}_2$
Before:	2.0		0	0
Change:	-2.0		1.0	2.0
After:	0		1.0	2.0

Change moles to grams!

- $2.0 \text{ mole} \times \frac{2.02 \text{ g}}{1 \text{ mole H}_2} = 4.0 \text{ g H}_2$

- 2. Recall that liquid sodium reacts with chlorine gas to produce sodium chloride. You want to produce **581 g of sodium chloride**. How many **grams of sodium** are needed?



Before:

Change:

After:

Change grams to moles!

■ 581 g of NaCl x $\frac{1 \text{ mole H}_2\text{O}}{58.44 \text{ g}}$ = 9.94 mole

Na : 22.99

Cl: +35.45

NaCl: 58.44g

- 2. Recall that liquid sodium reacts with chlorine gas to produce sodium chloride. You want to produce **581 g of sodium chloride**. How many **grams of sodium** are needed?



Before: ? XS 0

Change: 9.94

After: 9.94

$$9.94 \text{ mol NaCl} \times \frac{1 \text{ mol Cl}_2}{2 \text{ mol NaCl}} = 4.97 \text{ mol Cl}_2$$

$$9.94 \text{ mol NaCl} \times \frac{2 \text{ mol Na}}{2 \text{ mol NaCl}} = 9.94 \text{ mol Na}$$

Equation: 2 Na + Cl₂ → 2 NaCl

Before: 9.94 XS 0

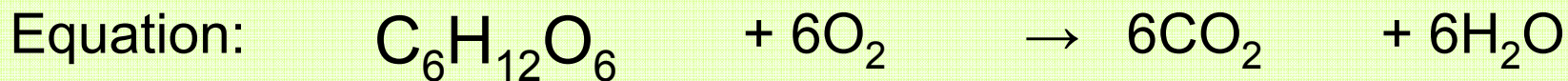
Change: -9.94 -4.97 9.94

After: 0 XS 9.94

Change moles to grams!

- $9.94 \text{ mole Na} \times \frac{22.99 \text{ g}}{1 \text{ mole Na}} = 228 \text{ g Na}$

- 3. You eat 180.0 g of glucose (90 M&Ms). If glucose, $C_6H_{12}O_6$, reacts with oxygen gas to produce carbon dioxide and water, how many grams of oxygen will you have to breathe in to burn the glucose?



Before:

Change:

After:

Change grams to moles!

■ $180.0 \text{ g of } \text{C}_6\text{H}_{12}\text{O}_6 \times \frac{1 \text{ mole } \text{H}_2\text{O}}{180.18 \text{ g}} = 0.9990 \text{ mole}$

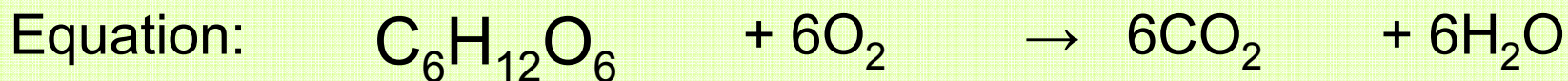
6C : 6(12.01)

12H: 12(1.01)

6O: +6(16.00)

180.18g

- 3. You eat **180.0 g of glucose** (90 M&Ms). If glucose, $C_6H_{12}O_6$, reacts with oxygen gas to produce carbon dioxide and water, how many **grams of oxygen** will you have to breathe in to burn the glucose?



Before: 0.9990 XS 0 0

Change:

After:

$$.9990\text{mol} \times \frac{6\text{molO}_2}{1\text{mol}} = 5.994\text{molO}_2$$

$$.9990\text{mol} \times \frac{6\text{molCO}_2}{1\text{mol}} = 5.994\text{molCO}_2$$

$$.9990\text{mol} \times \frac{6\text{molH}_2\text{O}}{1\text{mol}} = 5.994\text{molH}_2\text{O}$$

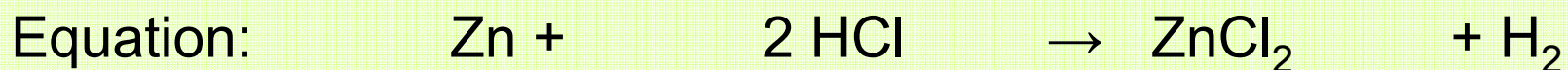
Equation:	$\text{C}_6\text{H}_{12}\text{O}_6$	+ 6 O_2	→ 6 CO_2	+ 6 H_2O
Before:	0.9990	XS	0	0
Change:	-0.9990	-5.994	5.994	5.994
After:	0	XS	5.994	5.994

Change moles to grams!

- $5.994 \text{ mole O}_2 \times \frac{32.00 \text{ g}}{1 \text{ mole O}_2} = 191.8 \text{ g O}_2$

4. Suppose 4.61 g of zinc was allowed to react with hydrochloric acid to produce zinc chloride and hydrogen gas. How much zinc chloride should you get?

Suppose that you actually recovered 8.56 g of zinc chloride. What is your percent yield?



Before:

Change:

After:

Change grams to moles!

- 4.61 g of Zn x $\frac{1 \text{ mole Zn}}{65.38 \text{ g}}$ = 0.0705 mole

4. Suppose 4.61 g of zinc was allowed to react with hydrochloric acid to produce zinc chloride and hydrogen gas. How much zinc chloride should you get?

Suppose that you actually recovered 8.56 g of zinc chloride. What is your percent yield?

Equation:	Zn +	2 HCl	→	ZnCl ₂	+ H ₂
Before:	0.0705	XS		0	0
Change:	- 0.0705	-0.141		0.0705	0.0705
After:	0	XS		0.0705	0.0705

Change moles to grams!

- $0.0705 \text{ mole ZnCl}_2 \times \frac{136.28 \text{ g}}{1 \text{ mole ZnCl}_2} = 9.61 \text{ g ZnCl}_2$

4. Suppose 4.61 g of zinc was allowed to react with hydrochloric acid to produce zinc chloride and hydrogen gas. How much zinc chloride should you get?

Suppose that you actually recovered 8.56 g of zinc chloride. What is your percent yield?

Equation:	Zn +	2 HCl	→	ZnCl ₂	+ H ₂
Before:	0.0705	XS		0	0
Change:	- 0.0705	-0.141		0.0705	0.0705
After:	0	XS		0.0705	0.0705

Find %Yield!

- 0.0705 mole ZnCl_2 x $\frac{136.28 \text{ g}}{1 \text{ mole ZnCl}_2}$ = 9.61 g ZnCl_2

$$\% \text{Yield} = \frac{\text{ACTUAL}}{\text{THEORETICAL}} \times 100$$

$$= \frac{8.56 \text{ g}}{9.61 \text{ g}} \times 100 = 89.1\% \text{ yield}$$

- 5. Determine the mass of carbon dioxide that should be produced in the reaction between 3.74 g of carbon and excess O₂. What is the % yield if 11.34 g of CO₂ is recovered?

Equation:	C	+ O ₂	→	CO ₂
Before:	0.311	XS		0
Change:	- 0.311	- 0.311		0.311
After:	0	XS		0.311

Calculations

$$3.74\text{gC} \times \frac{1\text{molC}}{12.01\text{g}} = 0.311\text{molC}$$

$$0.311\text{molC} \times \frac{1\text{molCO}_2}{1\text{molC}} = 0.311\text{molCO}_2$$

$$0.311\text{molCO}_2 \times \frac{44.01\text{gCO}_2}{1\text{molCO}_2} = 13.7\text{gCO}_2$$

$$\frac{11.34\text{ g CO}_2}{13.7\text{ g CO}_2} \times 100\% = 82.8\% \text{ yield}$$

- 6. In the reaction between excess K(s) and 4.28 g of O₂(g), potassium oxide is formed . What mass would you expect to form (theoretical yield)? If 17.36 g of K₂O is actually produced, what is the percent yield?



Before:

Change:

After:

- 6. In the reaction between excess K(s) and 4.28 g of O₂(g), potassium oxide is formed . What mass would you expect to form (theoretical yield)? If 17.36 g of K₂O is actually produced, what is the percent yield?

Equation:	4K +	O ₂	→	2K ₂ O
Before:	XS	0.134		0
Change:	-0.536	-0.134		0.266
<hr/>				
After:	XS	0		0.266

Calculations

$$4.28 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} = 0.134 \text{ mol } O_2$$

$$0.134 \text{ mol } O_2 \times \frac{2 \text{ mol } K_2O}{1 \text{ mol } O_2} = 0.268 \text{ mol } K_2O$$

$$0.268 \text{ mol } K_2O \times \frac{94.20 \text{ g } K_2O}{1 \text{ mol } K_2O} = 25.2 \text{ g } K_2O$$

$$\frac{17.36 \text{ g } K_2O}{25.2 \text{ g } K_2O} \times 100\% = 68.9 \% \text{ yield}$$

- 7. Determine the mass of carbon dioxide one could expect to form (and the percent yield) for the reaction between excess CH_4 and 11.6 g of O_2 if 5.38 g of carbon dioxide gas is produced along with some water vapor.

Equation:	$\text{CH}_4 +$	2O_2	\rightarrow	CO_2	$+ 2\text{H}_2\text{O}$
Before:	XS	0.363		0	0
Change:					
After:					0.363

- 7. Determine the mass of carbon dioxide one could expect to form (and the percent yield) for the reaction between excess CH_4 and 11.6 g of O_2 if 5.38 g of carbon dioxide gas is produced along with some water vapor.

Equation:	CH_4	+	2O_2	\rightarrow	CO_2	+	$2\text{H}_2\text{O}$
Before:	XS		0.363		0		0
Change:	- 0.181		- 0.363		0.181		0.363
After:	XS		0		0.181		0.363

Calculations

$$11.6\text{gO}_2 \times \frac{1\text{molO}_2}{32.00\text{gO}_2} = 0.363\text{molO}_2$$

$$0.363\text{molO}_2 \times \frac{1\text{molCO}_2}{2\text{molO}_2} = 0.181\text{molCO}_2$$

$$0.181\text{molCO}_2 \times \frac{44.01\text{gCO}_2}{1\text{molCO}_2} = 7.98\text{gCO}_2$$

$$\frac{5.38 \text{ g CO}_2}{7.98 \text{ g CO}_2} \times 100\% = 67.4 \% \text{ yield}$$

- 8. Determine the mass of water vapor you would expect to form (and the percent yield) in the reaction between 15.8 g of NH_3 and excess oxygen to produce water and nitric oxide (NO). The mass of water actually formed is 21.8 g.



Before: 0.928 XS 0 0

Change:

After:

- 8. Determine the mass of water vapor you would expect to form (and the percent yield) in the reaction between 15.8 g of NH_3 and excess oxygen to produce water and nitric oxide (NO). The mass of water actually formed is 21.8 g.

Equation:	$4\text{NH}_3 +$	5O_2	\rightarrow	$6\text{H}_2\text{O}$	$+ 4\text{NO}$
Before:	0.928	XS		0	0
Change:	- 0.928	-1.16		1.39	0.928
After:	0	XS		1.39	0.928

Calculations

$$15.8\text{gNH}_3 \times \frac{1\text{molNH}_3}{17.04\text{gNH}_3} = 0.928\text{molNH}_3$$

$$0.928\text{molNH}_3 \times \frac{6\text{molH}_2\text{O}}{4\text{molNH}_3} = 1.39\text{molH}_2\text{O}$$

$$1.39\text{molH}_2\text{O} \times \frac{18.02\text{gH}_2\text{O}}{1\text{molH}_2\text{O}} = 25.1\text{gH}_2\text{O}$$

$$\frac{21.8\text{ g H}_2\text{O}}{25.1\text{ g H}_2\text{O}} \times 100\% = 86.9\% \text{ yield}$$