Sequences to solve a Chemical Reaction Problem [Draft]

Example Problem:

15.4 g of Aluminum metal is reacted with 12.2 g of Sulfuric Acid to yield 32 mg of product.

1. What is the % Yield?
2. How much excess is there of the non-limiting reagent?

To solve this problem, let's break it down into it's components.

1. Convert the English to Chemistry for each of the Reactants

A. Get the right Anions, Cations, Polyatomics, and charges on each

B. Balance the Anions and Cations – the number of plus charges equals the number of minus charges on the Reactants

$Al + H_2SO_4 \rightarrow$

2. Decide what the Products are. Do an Ion Swap. Balance the Anions and Cations – the number of plus charges must equal the number of minus charges on the Products

$Al + H_2SO_4 \rightarrow Al\ SO_4 + H_2$	Decide what the Products are. Do an Ion Swap.
$Al + H_2SO_4 \rightarrow Al_2(SO_4)_3 + H_2$	Balance the Anions and Cations in the Products
$Al + H_2SO_4 \rightarrow Al_2(SO_4)_3 + H_2 \uparrow$	Label Gases with ↑ and precipitates with ppt or ↓

3. Identify what kind of a reaction this is, what is the driving force, if it is any at all:

A. REDOX	**Is there a pure element – Metal, Gas, Halogen?**
B. Is there a ppt	**Look up each reactant and product to see if it is soluble or not**
C. Is there a gas	**The common gases are H_2, O_2, N_2 and CO_2**
D. Is H-OH formed	**An acid and a base reacting will produce H-OH**
E. If none of the above	**There Probably is NO REACTION!**

4. This is a REDOX reaction, there is a pure element – Al and a gas H_2 formed. REDOX reactions need to be balanced by the number of electrons lost and gained:

$Al^o \rightarrow Al^{+3} + 3\ e^-$	This ½ reaction generates 3 electrons
$2\ H^+ + 2\ e^- \rightarrow H_2 \uparrow$	This ½ reaction uses 2 electrons.

We need to balance the electrons lost and gained and give a summary equation

$2 * [\ Al^o \rightarrow Al^{+3} + 3\ e^-\ ]$	This ½ reaction generates 2 * 3 electrons
$3 * [\ 2\ H^+ + 2\ e^- \rightarrow H_2 \uparrow\ ]$	This ½ reaction uses 3 * 2 electrons.
$2\ Al^o + 6\ H^+ \rightarrow 2\ Al^{+3} + 3\ H_2 \uparrow$	The overall net REDOX reaction

5. Ok, now go plug the Coefficients back into the original equation:

$2\ Al + 3\ H_2SO_4 \rightarrow Al_2(SO_4)_3 + 3\ H_2 \uparrow$

6. Now determine the Molecular weights of each reactant and the desired product. Remember, the Desired Product is either the ppt, gas or water! Please try to be neat doing this and label them.

Al	2 * 26.98			=	**53.96 g / mole 2 Al**
3 H_2SO_4	6 H	6 * 1.008	6.048		
	3 S	3 * 32.08	96.24		
	12 O	12 * 16.00	192.00		
			294.288	=	**294.29 g / mole 3 H_2SO_4**
3 H_2	3 * 1.008			=	**3.024 g / mole 3 H_2**

7. Now line up the Givens and what is to be found?

Given	15.4g		12.2 g			32 mg
To Find	Xcs					% Yield
Formula	2 Al	+	3 H_2SO_4	$\rightarrow$	$Al_2(SO_4)_3$ +	3 $H_2 \uparrow$
Mw	53.96 g / mole		294.29 g / mole			3.024 g / mole

8. Ok, now with the amounts of the two reactants given, we know this is a LIMITING REACTION. We need to find out which of the two reactants is the limiting reactant. Which of the two reactants will be all used up when this reaction is complete? We do this by calculating the number of Moles of each:

Moles of 2 Al = $\dfrac{15.4\ g}{53.96\ g\ /\ mole}$ = 0.2853965 = 0.285 moles

Moles of 3 H_2SO_4 = $\dfrac{12.2\ g}{294.29\ g\ /\ mole}$ = **0.0414557 = 0.0415 moles Smallest Number**
H_2SO_4 is the Limiting Reactant

9. Now we can determine the amount of product, the amount of Hydrogen Gas that is produced:
Do the calculations based on the limiting reactant - Al

$\dfrac{12.2\ g}{294.29\ g\ /\ mole\ \mathbf{H_2SO_4}} = \dfrac{X}{3.024\ g\ /\ mole\ 3\ H_2}$ **X** = 0.1253662 = **0.125 g H_2**

10. We only were able to get 0.032 g of H_2, but the Theoretical Yield or 100% yield is 0.125 g. So, our % yield is: Note: 32 mg = 0.032 g

$\dfrac{0.032\ g}{0.125\ g}$ * 100% = 25.6 % = **26 % yield.**

11. And, now to determine the amount of excess starting reactant. Remember, this is a limiting reactant, so the limiting reactant is all used up, and the other, the non-limiting, is only partially used up. The Al is all used up and the H_2SO_4 is only partially used up. We first determine how much of the H_2SO_4 is actually used up:

$$\frac{12.2\text{ g}}{294.29\text{ g / mole } \mathbf{H_2SO_4}} = \frac{X}{53.96\text{ g / mole Al}}$$

X = 2.2369499 = **2.24 g Al is used**

12. Now subtract the amount of Al that was consumed in the reaction from the amount that we started with:

15.4 g
2.24 g
13.16 g = **13.2 g of Al in excess.**

<u>Now you should be able to solve the following:</u>

10.2 g of Magnesium Chloride is reacted with 4.4 g of Silver Nitrate and we were able to isolate 2.00 g of product. What was the % Yeild? How much excess of a reactant is there?

25.0 g of Lithium Carbonate is reacted with 5.0 g of Nitric Acid. How much of the product is produced? How much excess of a reactant is there?

10.0 g of Lithium Hydroxide is reacted with 10.0 g of Phosphoric Acid. How much of the product is produced? How much excess of a reactant is there?

10.0 g of Barium Chloride is reacted with 10.0 g of Ammonium Iodide. How much of the product is produced? How much excess of a reactant is there?