Limiting reactants:

First off, you have to be able to identify a limiting reactant problem in the first place. Like I said in class, you will know that you are doing a limiting reactant problem when you are given an amount of more than one reactant.

Example:

What mass of $PbCl_2$ will be made when 15.9 grams of $CaCl_2$ is added to a solution of $Pb(NO_3)_2$?

This is NOT a limiting reactant problem because you are given an amount of only a single reactant.

Example:

What mass of $Ti(CO_3)_2$ will be made when 0.558 grams of $Ti(C_2H_3O_2)_4$ is added to a solution containing 1.13 grams of Rb_2CO_3?

This IS a limiting reactant problem because you are given an amount of more than one reactant.

If the question is NOT a limiting reactant problem, you simply convert the amount of the stuff they give you to the amount of the stuff they ask for. What do you ALWAYS have to use when going from an amount of one thing to an amount of another thing? If it IS a limiting reactant questions, it is somewhat more complicated, but since virtually all limiting reactant problems are done the same way, it isn't that bad.

In a limiting reactant problem, you have set amounts of reactants, one of which you are likely to run out of. Maybe you could think about it in terms of money. I am going to show you two examples simultaneously, one involving money, and one involving chemicals. My hope is that you will see the similarities.

Example A:

If you have $500 in your bank account and you have $436.50 worth of bills, how much of our total bills can you pay, and how much money or bills will be left over?

Example B:

Hydrogen gas and oxygen gas are mixed to make water. If 2.80 grams of hydrogen gas is reacted with 17.7 grams of oxygen gas, how much water can be formed and how much of either the hydrogen or the oxygen will be left over?

Both of these are limiting reactant problems because in each case you are given an amount of more than one reactant (money and bills in A and hydrogen and oxygen in B).

The first step, as usual, is to get balanced equations (you must do this for any of these problems), so:

$\mathbf{\$\ in\ the\ bank + \$\ of\ bills \rightarrow \$\ paid\ bill}$ (this says that for every 1$ [in the bank] + every $ of bills, you get $ of paid bills)

$$\mathbf{2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}}$$

The second step is to see how much of a product you can make from each given amount of reactant. In both cases, there is only a single product, so it is an easy choice.

$$\$500\ \textbf{in the bank} \times \frac{1\ \textbf{paid bills}}{\$1\ \textbf{in the bank}} = \$500\ \textbf{paid bills}$$

$$\$436.50\ \textbf{of bills} \times \frac{1\ \textbf{paid bills}}{\$1\ \textbf{of bills}} = \$436.50\ \textbf{paid bills}$$

So, you can use up all of your money and have $500 bills paid, or you can use up all of your bills and have $436.50 bills paid. You only have $436.50 worth of bills, so why would you give The Man any more than that? You ran out of bills before you ran out of money (lucky you), so the bills were your limiting reactant.

$$2.80\ \boldsymbol{g}\ \mathbf{H_2} \times \frac{1\ \boldsymbol{mol}\ \mathbf{H_2}}{2.01594\ \boldsymbol{g}\ \mathbf{H_2}} \times \frac{2\ \boldsymbol{mol}\ \mathbf{H_2O}}{2\ \boldsymbol{mol}\ \mathbf{H_2}} \times \frac{18.0153\ \boldsymbol{g}\ \mathbf{H_2O}}{1\ \boldsymbol{mol}\ \mathbf{H_2O}} = 25.0\ \boldsymbol{g}\ \mathbf{H_2O}$$

$$17.8\ \boldsymbol{g}\ \mathbf{O_2} \times \frac{1\ \boldsymbol{mol}\ \mathbf{O_2}}{31.9988\ \boldsymbol{g}\ \mathbf{O_2}} \times \frac{2\ \boldsymbol{mol}\ \mathbf{H_2O}}{1\ \boldsymbol{mol}\ \mathbf{O_2}} \times \frac{18.0153\ \boldsymbol{g}\ \mathbf{H_2O}}{1\ \boldsymbol{mol}\ \mathbf{H_2O}} = 20.0\ \boldsymbol{g}\ \mathbf{H_2O}$$

So, you can use up all of your H_2 and make 25.0g of H_2O or you can use up all of your O_2 and make 20.0 g of H_2O. You only have 17.8 g of O_2, so you cannot make any more H_2O than the 20.0 g shown in the second equation. You ran out of O_2 before you ran out H_2, so the O_2 was the limiting reactant.

From above, you have $436.50 paid bills (bills were limiting) and you can make 20.0 g of H_2O (O_2 was limiting). How much of the other reactant (the excess reactant) is left over?

In order find out how much of the excess reactant is left, you must first find out how much of the excess reactant was used in the reaction. You will do this by starting with the amount of the limiting reactant you were given in the question ($436.50 worth of bills and 17.8 g of O_2) and converting it to the excess reactant ($ in the bank and H_2)

$$\$436.50\ \textbf{of bills} \times \frac{\$1\ \textbf{in the bank}}{\$1\ \textbf{of bills}} = \$436.50\ \textbf{in the bank}$$

This is the amount of the money in the bank that you USED to pay bills!

$$17.8\ \boldsymbol{g}\ \mathbf{O_2} \times \frac{1\ \boldsymbol{mol}\ \mathbf{O_2}}{31.9988\ \boldsymbol{g}\ \mathbf{O_2}} \times \frac{2\ \boldsymbol{mol}\ \mathbf{H_2}}{1\ \boldsymbol{mol}\ \mathbf{O_2}} \times \frac{2.01594\ \boldsymbol{g}\ \mathbf{H_2}}{1\ \boldsymbol{mol}\ \mathbf{H_2}} = 2.24\ \boldsymbol{g}\ \mathbf{H_2}$$

This is the amount of the H_2 that was USED to make the H_2O!

Well, if you USED $436.50 from your bank to bay the bills, does that mean that you have $436.50 left in the bank? Of course not, although it would be nice if it worked that way. This is the amount you took out of your bank account, so you have to subtract the amount used to pay bills from the amount you started with in order to find out how much money is left in your account! **$500** (started) **- $436.50** (used) **= $63.50 LEFT OVER IN YOUR ACCOUNT!**

The same for example B, 2.24 g H_2 is not what is left over, it is what you USED. You MUST subtract this amount from what you started with to find out how much is LEFT!

2.80 g H_2 (started) **– 2.24 H_2** (used) **= 0.56 g H_2 LEFT OVER AFTER THE REACTION!**

This was fairly long winded, but hopefully helpful. The next part will be a more condensed version of “How to do limiting reactant problems”

Once you have identified a problem as a limiting reactant problem, here are the steps to go through:

Step 1: Write the balanced chemical equation
Step 2: Pick a product and find out what mass of that product can be made from **each** of the given reactants. The smallest of the calculated product masses is how much of the product can actually be made! The reactant leading to the smallest amount of product is the **limiting reactant**. Remember, all of the limiting reactant is used up, so the amount remaining is ALWAYS zero.
The following only needs to be done if the question asks for masses of the other products and reactants
Step 3: For the rest of the calculations, start with the amount of the limiting reactant given in the question. Calculate the masses of other products made from the given amount of the limiting reactant (since these are products, they are the amount MADE and therefore no subtraction is needed).
Step 4: Calculate the masses of the excess reagent(s) USED starting with the amount of the amount of the limiting reagent given. Reagents are always USED, so once you have calculated the amount used, you MUST subtract it from the amount of that reactant you were given.

I recommend making a table like I showed in class to keep track of everything!

Example 1: 4.529 g of antimony (III) sulfate and 13.98 g of lead (II) chlorate are reacted together. How much of all products and reactants are left after the reaction is complete?

This IS a limiting reactant question because amounts of the two reactants are given.

Step 1: balance equation:

$$Sb_2(SO_4)_{3(aq)} + 3\ Pb(ClO_3)_{2(aq)} \rightarrow 2\ Sb(ClO_3)_{3(aq)} + 3\ Pb(SO_4)_{2(s)}$$

Step 2: I am going to pick the lead (II) sulfate as my product, just because I can.

$$4.529\ \boldsymbol{g\ Sb_2(SO_4)_3} \times \frac{1\ \boldsymbol{mol\ Sb_2(SO_4)_3}}{531.68\ \boldsymbol{g\ Sb_2(SO_4)_3}} \times \frac{3\ \boldsymbol{mol\ PbSO_4}}{1\ \boldsymbol{mol\ Sb_2(SO_4)_3}} \times \frac{303.25\ \boldsymbol{g\ PbSO_4}}{1\ \boldsymbol{mol\ PbSO_4}} = 7.750\ \boldsymbol{g\ PbSO_4}$$

$$13.98\ \boldsymbol{g\ Pb(ClO_3)_2} \times \frac{1\ \boldsymbol{mol\ Pb(ClO_3)_2}}{374.09\ \boldsymbol{g\ Pb(ClO_3)_2}} \times \frac{3\ \boldsymbol{mol\ PbSO_4}}{3\ \boldsymbol{mol\ Pb(ClO_3)_2}} \times \frac{303.25\ \boldsymbol{g\ PbSO_4}}{1\ \boldsymbol{mol\ PbSO_4}} = 11.33\ \boldsymbol{g\ PbSO_4}$$

7.750 g $PbSO_4$ is the smallest amount of product calculated, so that is how much can actually be made. This also shows that $Sb_2(SO_4)_3$ is the limiting reactant.

$Sb_2(SO_4)_3$ = 0 g	$Sb(ClO_3)_3$
$Pb(ClO_3)_2$	$Pb(SO_4)_2$ = 7.750 g

Step 3:

$$4.529\ \boldsymbol{g\ Sb_2(SO_4)_3} \times \frac{1\ \boldsymbol{mol\ Sb_2(SO_4)_3}}{531.68\ \boldsymbol{g\ Sb_2(SO_4)_3}} \times \frac{2\ \boldsymbol{mol\ Sb(ClO_3)_3}}{1\ \boldsymbol{mol\ Sb_2(SO_4)_3}} \times \frac{372.10\ \boldsymbol{g\ Sb(ClO_3)_3}}{1\ \boldsymbol{mol\ Sb(ClO_3)_3}} = 6.339\ \boldsymbol{g\ Sb(ClO_3)_3}$$

This is the amount of the second product that can actually be made!

$Sb_2(SO_4)_3$ = 0 g	$Sb(ClO_3)_3$ = 6.339 g
$Pb(ClO_3)_2$	$Pb(SO_4)_2$ = 7.750 g

Step 4:

$$4.529\ g\ \mathbf{Sb_2(SO_4)_3} \times \frac{1\ mol\ \mathbf{Sb_2(SO_4)_3}}{531.68\ g\ \mathbf{Sb_2(SO_4)_3}} \times \frac{3\ mol\ \mathbf{Pb(ClO_3)_2}}{1\ mol\ \mathbf{Sb_2(SO_4)_3}} \times \frac{374.09\ g\ \mathbf{Pb(ClO_3)_2}}{1\ mol\ \mathbf{Pb(ClO_3)_2}} = 9.560\ g\ \mathbf{Pb(ClO_3)_2}$$

This is the amount of the excess reactant that was USED!!

11.33 g $Pb(ClO_3)_2$ (started) **– 9.560 g $Pb(ClO_3)_2$** (used) **= 1.77 g $Pb(ClO_3)_2$ LEFT!!**

This last step is the step the most people forget, so DON'T forget!

Final answer:

$Sb_2(SO_4)_3$ = 0 g	$Sb(ClO_3)_3$ = 6.339 g
$Pb(ClO_3)_2$ = 1.77 g	$Pb(SO_4)_2$ = 7.750 g

Example 2: What mass of all products and reactants are present after the reaction of 1.14g of potassium carbonate with 100. mL of a 0.0341 M solution of tin (IV) nitrate?

$2\ K_2CO_{3(aq)} + Sn(NO_3)_{4(aq)} \rightarrow Sn(CO_3)_{2(s)} + 4\ KNO_{3(aq)}$

$$1.14\ g\ \mathbf{K_2CO_3} \times \frac{1\ mol\ \mathbf{K_2CO_3}}{138.213\ g\ \mathbf{K_2CO_3}} \times \frac{1\ mol\ \mathbf{Sn(CO_3)_2}}{2\ mol\ \mathbf{K_2CO_3}} \times \frac{238.71\ g\ \mathbf{Sn(CO_3)_2}}{1\ mol\ \mathbf{Sn(CO_3)_2}} = 0.984\ g\ \mathbf{Sn(CO_3)_2}$$

$$0.100\ L\ \mathbf{solution} \times \frac{0.0341\ mol\ \mathbf{Sn(NO_3)_4}}{1\ L\ \mathbf{solution}} \times \frac{1\ mol\ \mathbf{Sn(CO_3)_2}}{1\ mol\ \mathbf{Sn(NO_3)_4}} \times \frac{238.71\ g\ \mathbf{Sn(CO_3)_2}}{1\ mol\ \mathbf{Sn(CO_3)_2}} = \boxed{0.814\ g\ \mathbf{Sn(CO_3)_2}}$$

0.814 g $Sn(CO_3)_2$ can be made and $Sn(NO_3)_4$ is the L.R.

K_2CO_3	$Sn(CO_3)_2$ = 0.814 g
$Sn(NO_3)_4$ = 0 g	KNO_3

$$0.100\ L\ \mathbf{solution} \times \frac{0.0341\ mol\ \mathbf{Sn(NO_3)_4}}{1\ L\ \mathbf{solution}} \times \frac{4\ mol\ \mathbf{KNO_3}}{1\ mol\ \mathbf{Sn(NO_3)_4}} \times \frac{101.107\ g\ \mathbf{KNO_3}}{1\ mol\ \mathbf{KNO_3}} = 1.38\ g\ \mathbf{KNO_3}$$

K_2CO_3	$Sn(CO_3)_2$ = 0.814 g
$Sn(NO_3)_4$ = 0 g	KNO_3 = 1.38 g

$$0.100\ L\ \mathbf{solution} \times \frac{0.0341\ mol\ \mathbf{Sn(NO_3)_4}}{1\ L\ \mathbf{solution}} \times \frac{2\ mol\ \mathbf{K_2CO_3}}{1\ mol\ \mathbf{Sn(NO_3)_4}} \times \frac{138.213\ g\ \mathbf{K_2CO_3}}{1\ mol\ \mathbf{K_2CO_3}} = 0.943\ g\ \mathbf{K_2CO_3}$$

1.14 g K_2CO_3 (started) **– 0.943 g K_2CO_3** (used) = **0.20 g K_2CO_3 LEFT!**

K_2CO_3 = 0.20 g	$Sn(CO_3)_2$ = 0.814 g
$Sn(NO_3)_4$ = 0 g	KNO_3 = 1.38 g

You try:

1) 25.00 g of nickel (II) chloride is reacted with 25.00 g of silver sulfate
2) 18.9 grams of lithium phosphate reacts with gold (I) acetate
3) 0.500 grams of copper (II) nitrate is reacted with 1.25 grams of sodium oxalate
4) 525 mL of 1.199 M sulfuric acid is added to 23.66 grams of strontium bromide
5) 3.8 grams of octane (C_8H_{18}) is burned in 22.1 grams of oxygen gas.
6) 10.0 mL of 1.50 M solution of hydroiodic acid is poured onto 0.95 g of solid calcium carbonate.