Limiting/Excess reactant problem

What's it look like

Chemical reaction, ***exact amounts*** (in mols or grams) ***for two reactants***.
Question may include "Which reactant is limiting?", "Which reactant is in excess?", "How much excess reactant remains unreacted after the reaction is completed?", and/or they may just have a roadmap problem (eg how much of a certain product is produced).

Ex: What is the theoretical yield of iron(III) oxide when 45.3 g of iron reacts with 31.0 g of oxygen?

Concept behind it

This is basically a variation on the "roadmap" stoichiometry problem. In real life, rarely do you have the exact amounts of reactants A and B that will react with each other with nothing left over of either one. Usually you have extra (excess) of one of the reactants. Your job is to determine which one is in excess and by how much.

How to tackle it

If they didn't give you a balanced reaction, you'll need to do that first. Then compare the moles of your two reactants against what's needed in the balanced reaction.

Detailed steps

1) Write down the balanced chemical reation (figure this out if it wasn't given)
2) Write "GIVEN" next to the mole amounts of the reactants. (If the reactants were given to you as grams, convert the amounts to moles first.)
3) Arbitrarily pick one reactant to be the limiting reactant. (Sometimes by eyeballing the mole amounts and looking at the coefficients you'll be able to choose wisely here!) Use proportions or dimensional analysis to calculate how many moles of the other reactant would be needed. Write "NEEDED" next to that amount, and compare it to the amount of the other reactant you wrote "GIVEN" next to.
4) If the amount of the other reactant needed is less than the amount given of the other reactant, you picked the limiting reactant correctly and you can just subtract the "needed" from the "given" to find the excess. If the amount needed is *more* than the amount given, repeat step 3 with the other reactant being the limiting reactant.

Example (continued from above):

Balanced reaction: $4Fe + 3O_2 \rightarrow 2Fe_2O_3$

$$45.3\ g\ Fe \cdot \frac{1\ \text{mol}}{55.85\ \text{g}} = 0.811101\ mol\ Fe \quad GIVEN$$

$$31.0\ g\ O_2 \cdot \frac{1\ \text{mol}}{32.00\ \text{g}} = 0.96875\ mol\ O_2 \quad GIVEN$$

If there's 3 mol of oxygen, there needs to be 4 mol of iron. That's a third more moles iron than oxygen. That means for the 0.96875 mol of O_2 *there should be around 1.28 mol or so of iron. There isn't. So, I suspect iron is limiting.*

$$0.811101\ mol\ Fe \cdot \frac{3\ \text{mol}\ O_2}{4\ \text{mol Fe}} = 0.60832575\ mol\ O_2\ \ NEEDED \quad (Shows\ O_2\ is\ in\ excess)$$

$$0.811101\ mol\ Fe \cdot \frac{2\ \text{mol}\ Fe_2O_3}{4\ \text{mol Fe}} = 0.4055505\ mol\ Fe_2O_3 \cdot \frac{159.7\ \text{g}}{1\ \text{mol}} = 64.766414\ g\ Fe_2O_3$$

$$= 64.8\ g\ Fe_2O_3$$