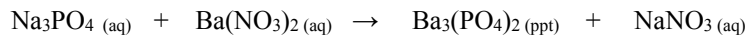


1. Consider the reaction below.



Suppose you react 4.00 g Na_3PO_4 with 4.00 g of $\text{Ba}(\text{NO}_3)_2$

- Balance the equation above.
- Which is the limiting reactant?
- Which substance is left over? What mass?

MM g/mole
sodium phosphate = 164
barium nitrate = 261
barium phosphate = 602
sodium nitrate = 87.0

2. Consider the reaction between magnesium and chlorine gas.

Given 2.0 g of magnesium, and 5.0 g of chlorine gas:

- Write a balanced equation.
- Determine which substance limits the reaction.
- Is there anything left over? Which substance? What mass?

MM g/mole
magnesium = 24.3
chlorine gas = 70.9
magnesium chloride = 95.2

3. Consider the reaction for the synthesis of H_2O from hydrogen gas and oxygen gas.

If you have been given 28.5 grams of O_2 and 26.9 grams of H_2

- Write a balanced equation
- Which gas is the limiting reactant?
- Which gas is left over? How many grams of it?

MM g/mole
hydrogen = 2.02
oxygen gas = 32.0
water = 18.0

4. Consider the reaction of phosphoric acid reacting with aluminum to produce aluminum phosphate and hydrogen gas.

Suppose you used 1.82 g of phosphoric acid and 0.659 g of aluminum.

- Write the balanced equation for the reaction described above.
- Which reactant limits the reaction?
- Which reactant is in excess?
- How many grams of the excess reactant are left over?

MM g/mole
aluminum = 27.0
phosphoric acid = 98.0
aluminum phosphate = 122
hydrogen gas = 2.02

5. Consider the reaction of magnesium nitride with water to form magnesium hydroxide and ammonia gas (aka nitrogen trihydride)

If you did this reaction with 58.1 g of magnesium nitride and 20.4 g of water,

- Write a balanced equation that describes the reaction
- What mass of each product could you make?
- How much of which reactant is left over.

MM g/mole
magnesium nitride = 100.9
water = 18.0
magnesium hydroxide = 58.3
ammonia = 17.0

6. Consider the reaction between gold(III) sulfide and hydrogen gas to produce dihydrogen sulfide gas and gold.

If 500.00 g of gold(III) sulfide is reacted with 5.67 g of hydrogen gas,

- Write a balanced equation that represents the reaction above.
- Which reactant limits the reaction?
- Which reactant is in excess? What mass of the excess is left over?
- What mass of each gold could be formed?

MM g/mole
gold(III) sulfide = 490.15
hydrogen gas = 2.02
dihydrogen sulfide = 34.0
gold = 197

7. Consider the reaction of ammonia (NH_3) with propene (C_3H_6) and oxygen gas to produce $\text{C}_3\text{H}_3\text{N}$ and water.

If you start with 22.5 g of propene and 20.6 g of ammonia and 18.1 g of oxygen gas,

- Write a balanced equation for the reaction
- How much water can be produced?

MM g/mole
C_3H_6 = 42.1
ammonia = 17.0
oxygen = 32.0
$\text{C}_3\text{H}_3\text{N}$ = 53.0
water = 18.0

8. Write the balanced equation that describes the reaction of aluminum burning in oxygen gas to produce aluminum oxide.

If you used 500.0 g of oxygen gas and 500.0 g of aluminum,

- Write the balanced equation that describes the reaction above.
- Which reactant limits the reaction?
- Which reactant is in excess?
- If you wanted to use up all of the excess reactant, how many more grams of the limiting reactant must be added.

MM g/mole
aluminum = 27.0
oxygen gas = 32.0
aluminum oxide = 102

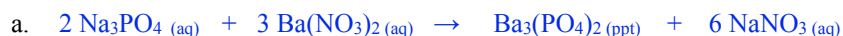
As in the P 9.4, you must recognize the fundamental difference in this problem type is that you have been given two reactant quantities, however this time it is in mass, not moles. This is distinctly different from the first three Practice Sheets in which you were given one quantity and an unlimited amount of the other reactant. As the title indicates, this problem is a limiting reactant problem. This means that one of the two quantities given will “run out” and limit the amount of product that can be made.

When starting one of these problems, you first must determine which reactant is the “limiting reactant.” There is more than one way to determine this, but one of the easiest is to convert your quantities to moles (“When in doubt, change to moles.”) and divide each reactant’s mole amount by its own coefficient in the balanced equation. The resulting smallest number will indicate the limiting reactant. Then do not use these numbers for any further calculations. Start with the original mass or mole amount of the limiting reactant to answer any further questions.

Consistent with the rest of the Practices in this unit, the **ratio** that is from the coefficients in the balanced equation highlighted in **red**. The answers are in **blue**. Do ALL your significant figure rounding at the END.

The answers are shown in **blue**. It is important to put your work on paper and label your numbers with **units** (moles, grams, etc), **identifiers** (O₂, H₂O, etc,) and **descriptors** (needed, produced, etc).

1. Suppose you react 4.00 g Na₃PO₄ with 4.00 g of Ba(NO₃)₂



- b. First determine which is the limiting reactant:

$$4.00 \text{ g Na}_3\text{PO}_4 \left(\frac{1 \text{ mol Na}_3\text{PO}_4}{164 \text{ g Na}_3\text{PO}_4} \right) = \frac{0.0244 \text{ mol Na}_3\text{PO}_4}{2} = 0.0122$$

$$4.00 \text{ g Ba}(\text{NO}_3)_2 \left(\frac{1 \text{ mol Ba}(\text{NO}_3)_2}{262 \text{ g Ba}(\text{NO}_3)_2} \right) = \frac{0.0153 \text{ mol Ba}(\text{NO}_3)_2}{3} = 0.00509 \text{ this number is smaller, thus Ba}(\text{NO}_3)_2 \text{ limits}$$

- c. Use the limiting reactant to determine the mass of the excess reactant that must be used (needed) to go with the limiting reactant.

$$4.00 \text{ g Ba}(\text{NO}_3)_2 \left(\frac{1 \text{ mol Ba}(\text{NO}_3)_2}{262 \text{ g Ba}(\text{NO}_3)_2} \right) \left(\frac{2 \text{ Na}_3\text{PO}_4}{3 \text{ Ba}(\text{NO}_3)_2} \right) \left(\frac{164 \text{ g Na}_3\text{PO}_4}{1 \text{ mol Na}_3\text{PO}_4} \right) = 1.67 \text{ g of Na}_3\text{PO}_4 \text{ needed}$$

Thus, 4.00 g Na₃PO₄ started with – 1.67 g of Na₃PO₄ needed = **2.33 g sodium phosphate left over**

2. Given 2.0 g of magnesium, and 5.0 g of chlorine gas:



- b. First determine which is the limiting reactant:

$$2.0 \text{ g Mg} \left(\frac{1 \text{ mol Mg}}{24.3 \text{ g Mg}} \right) = \frac{0.0823 \text{ mol Mg}}{1} = 0.0823$$

$$5.0 \text{ g Cl}_2 \left(\frac{1 \text{ mol Cl}_2}{71 \text{ g Cl}_2} \right) = \frac{0.0704 \text{ mol Cl}_2}{1} = 0.0704 \text{ this number is smaller, thus Cl}_2 \text{ limits}$$

- c. Use the limiting reactant to determine the mass of the excess reactant that must be used (needed) to go with the limiting reactant.

$$5.0 \text{ g Cl}_2 \left(\frac{1 \text{ mol Cl}_2}{71 \text{ g Cl}_2} \right) \left(\frac{1 \text{ Mg}}{1 \text{ Cl}_2} \right) \left(\frac{24.3 \text{ g Mg}}{1 \text{ mol Mg}} \right) = 1.71 \text{ g of Mg needed}$$

Thus, 2.0 g Mg started with - 1.71 g of Mg needed = **0.29 g Mg left over**

3. If you have been given 20.0 grams of O₂ and 40.0 grams of H₂



- b. First determine which is the limiting reactant:

$$26.9 \text{ g H}_2 \left(\frac{1 \text{ mol H}_2}{2.02 \text{ g H}_2} \right) = \frac{13.3 \text{ mol H}_2}{2} = 6.7$$

$$28.5 \text{ g O}_2 \left(\frac{1 \text{ mol O}_2}{32 \text{ g O}_2} \right) = \frac{0.891 \text{ mol O}_2}{1} = 0.891 \text{ this number is smaller, thus oxygen limits}$$

- c. Use the limiting reactant to determine the mass of the excess reactant that must be used (needed) to go with the limit reactant.

$$28.5 \text{ g O}_2 \left(\frac{1 \text{ mol O}_2}{32 \text{ g O}_2} \right) \left(\frac{2 \text{ H}_2}{1 \text{ O}_2} \right) \left(\frac{2.02 \text{ g H}_2}{1 \text{ mol H}_2} \right) = 3.60 \text{ g H}_2 \text{ needed}$$

Thus 26.5 g H₂ given – 3.60 g H₂ needed(used) = 22.9 g hydrogen(H₂) gas left over

4. Suppose you used 1.82 g of phosphoric acid and 0.659 g of aluminum



- b. First determine which is the limiting reactant:

$$\bullet \quad 1.82 \text{ g H}_3\text{PO}_4 \left(\frac{1 \text{ mol H}_3\text{PO}_4}{98 \text{ g H}_3\text{PO}_4} \right) = \frac{0.0186 \text{ mol H}_2}{3} = 0.00619 \text{ this number is smaller, thus H}_3\text{PO}_4 \text{ limits the reaction}$$

$$\bullet \quad 0.659 \text{ g Al} \left(\frac{1 \text{ mol Al}}{27.0 \text{ g Al}} \right) = \frac{0.0244 \text{ mol Al}}{2} = 0.0122$$

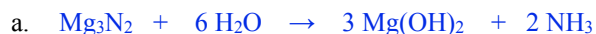
- c. Aluminum is in excess.

- d. Use the limiting reactant to determine the mass of the excess reactant that must be used (needed) to go with the limit reactant.

$$1.82 \text{ g H}_3\text{PO}_4 \left(\frac{1 \text{ mol H}_3\text{PO}_4}{98 \text{ g H}_3\text{PO}_4} \right) \left(\frac{2 \text{ Al}}{2 \text{ H}_3\text{PO}_4} \right) \left(\frac{27.0 \text{ g Al}}{1 \text{ mol Al}} \right) = 0.501 \text{ g Al needed}$$

Thus 0.659 g Al given – 0.501 g Al needed(used) = 0.158 g of aluminum is left over.

5. If you did this reaction with 58.1 g of magnesium nitride and 20.4 g of water,



- First determine which is the limiting reactant:

$$\bullet \quad 58.1 \text{ g Mg}_3\text{N}_2 \left(\frac{1 \text{ mol Mg}_3\text{N}_2}{101 \text{ g Mg}_3\text{N}_2} \right) = \frac{0.575 \text{ mol Mg}_3\text{N}_2}{1} = 0.575$$

$$\bullet \quad 20.4 \text{ g H}_2\text{O} \left(\frac{1 \text{ mol H}_2\text{O}}{18.0 \text{ g H}_2\text{O}} \right) = \frac{1.13 \text{ mol H}_2\text{O}}{6} = 0.189 \text{ this number is smaller, thus water limits the reaction}$$

b. $20.4 \text{ g H}_2\text{O} \left(\frac{1 \text{ mol H}_2\text{O}}{18.0 \text{ g H}_2\text{O}} \right) \left(\frac{3 \text{ Mg(OH)}_2}{6 \text{ H}_2\text{O}} \right) \left(\frac{58.3 \text{ g Mg(OH)}_2}{1 \text{ mol Mg(OH)}_2} \right) = 33.0 \text{ g Mg(OH)}_2 \text{ formed}$

$$20.4 \text{ g H}_2\text{O} \left(\frac{1 \text{ mol H}_2\text{O}}{18.0 \text{ g H}_2\text{O}} \right) \left(\frac{2 \text{ NH}_3}{6 \text{ H}_2\text{O}} \right) \left(\frac{17.0 \text{ g NH}_3}{1 \text{ mol NH}_3} \right) = 6.42 \text{ g NH}_3 \text{ formed}$$

c. $20.4 \text{ g H}_2\text{O} \left(\frac{1 \text{ mol H}_2\text{O}}{18.0 \text{ g H}_2\text{O}} \right) \left(\frac{1 \text{ Mg}_3\text{N}_2}{6 \text{ H}_2\text{O}} \right) \left(\frac{101 \text{ g Mg}_3\text{N}_2}{1 \text{ mol Mg}_3\text{N}_2} \right) = 19.1 \text{ g Mg}_3\text{N}_2 \text{ needed}$

Thus 58.1 g Mg₃N₂ given – 19.1 g Mg₃N₂ needed(used) = 39.0 g Mg₃N₂ left over

6. If 500.00 g of gold(III) sulfide is reacted with 5.67 g of hydrogen gas,



b. First determine which is the limiting reactant:

$$\bullet \quad 500.0 \text{ g Au}_2\text{S}_3 \left(\frac{1 \text{ mol Au}_2\text{S}_3}{490.15 \text{ g Au}_2\text{S}_3} \right) = \frac{1.02 \text{ mol Au}_2\text{S}_3}{1} = 1.02$$

$$\bullet \quad 5.67 \text{ g H}_2 \left(\frac{1 \text{ mol H}_2}{2.02 \text{ g H}_2} \right) = \frac{2.807 \text{ mol H}_2}{3} = 0.936 \text{ this number is smaller, thus hydrogen gas limits}$$

c. $5.67 \text{ g H}_2 \left(\frac{1 \text{ mol H}_2}{2.02 \text{ g H}_2} \right) \left(\frac{1 \text{ Au}_2\text{S}_3}{3 \text{ H}_2} \right) \left(\frac{490.15 \text{ g Au}_2\text{S}_3}{1 \text{ mol Au}_2\text{S}_3} \right) = 458.6 \text{ g Au}_2\text{S}_3 \text{ needed}$

Thus 500.0 g Au₂S₃ given – 458.6 g Au₂S₃ needed(used) = 41.4 g gold(III) sulfide is left over

d. $5.67 \text{ g H}_2 \left(\frac{1 \text{ mol H}_2}{2.02 \text{ g H}_2} \right) \left(\frac{2 \text{ Au}}{3 \text{ H}_2} \right) \left(\frac{197 \text{ g Au}}{1 \text{ mol Au}} \right) = 369 \text{ g of gold can be produced}$

7. If you start with 22.5 g of propene and 20.6 g of ammonia and 18.1 g of oxygen gas.



Even though the question does not ask, you must first determine the limiting reactant.

$$\bullet \quad 22.5 \text{ g C}_3\text{H}_6 \left(\frac{1 \text{ mol C}_3\text{H}_6}{42.1 \text{ g C}_3\text{H}_6} \right) = \frac{0.534 \text{ mol C}_3\text{H}_6}{2} = 0.267$$

$$\bullet \quad 20.6 \text{ g NH}_3 \left(\frac{1 \text{ mol NH}_3}{17.0 \text{ g NH}_3} \right) = \frac{1.21 \text{ mol NH}_3}{2} = 0.606$$

$$\bullet \quad 18.1 \text{ g O}_2 \left(\frac{1 \text{ mol O}_2}{32 \text{ g O}_2} \right) = \frac{0.566 \text{ mol O}_2}{3} = 0.189 \text{ this number is smallest, thus oxygen gas limits}$$

b. $18.1 \text{ g O}_2 \left(\frac{1 \text{ mol O}_2}{32 \text{ g O}_2} \right) \left(\frac{6 \text{ H}_2\text{O}}{3 \text{ O}_2} \right) \left(\frac{18.0 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \right) = 20.4 \text{ g H}_2\text{O formed}$

8. If you used 500.0 g of oxygen gas and 500.0 g of aluminum,



b. First determine which reactant limits

$$\bullet \quad 500.0 \text{ g Al} \left(\frac{1 \text{ mol Al}}{27.0 \text{ g Al}} \right) = \frac{18.5 \text{ mol Al}}{4} = 4.63 \text{ this number is smallest, thus Aluminum limits the reaction.}$$

$$\bullet \quad 500 \text{ g O}_2 \left(\frac{1 \text{ mol O}_2}{32 \text{ g O}_2} \right) = \frac{15.6 \text{ mol O}_2}{3} = 5.21$$

c. Oxygen is in excess.

d. $500 \text{ g O}_2 \left(\frac{1 \text{ mol O}_2}{32 \text{ g O}_2} \right) \left(\frac{4 \text{ Al}}{3 \text{ O}_2} \right) \left(\frac{27.0 \text{ g Al}}{1 \text{ mol Al}} \right) = 562.5 \text{ g Al needed, thus 62.5 g more of aluminum needed to use up all of the oxygen}$