Chemical Equations

Chemical Equations

shows the results of a chemical process

reactants (reagents) $\longrightarrow$ products

coefficients

the numbers in front of formulas in chemical equations

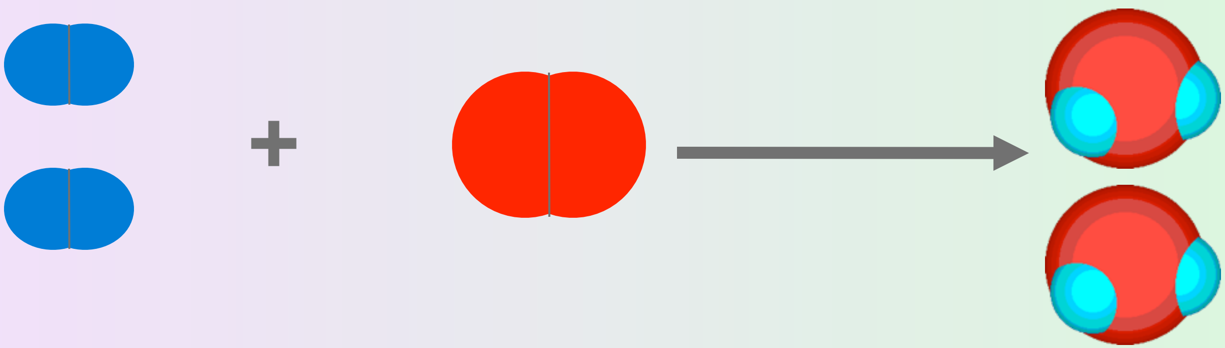
gives the relative number of molecules taking part in a reaction

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$$

Chemical bonds have been broken and new chemical bonds have been formed

Writing Chemical Equations

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O\ (g)$$



2 moles	1 mole	2 moles
4.04 g	32.00 g	36.04 g

Parentheses show physical state of substances

Example

molar interpretation

	$H_2(g)$	+ $Cl_2(g)$	$\longrightarrow$	$2HCl(g)$
Start:	**1 mol**	**1 mol**		**0**
Finish:	**0**	**0**		**2 mol**

Example

mass interpretation

	$2H_2$	+	O_2	$\longrightarrow$	$2H_2O$
	4 g		32 g		36 g
Start:	4g		32g		0
Finish:	0		0		36 g

The law of conservation of mass requires that chemical equations must balance.

What goes in $\longrightarrow$ **Must come out**

Quantitative Relationships

Stoichiometry

the study of quantitative between amounts of reactants used and amounts of products formed

how much reactant is needed to yield a certain amount of product?

Types of problems

- **Mole - mole calculations**
- **Mass - mass calculations**
- **Limiting/Excess reagent calculations**

The mole method

1. Write and balance the equation.
2. Convert the given quantities into moles.
3. Use the coefficients in the balanced equation to relate the number of moles of known substances to the desired unknown one.
4. Convert to desired units.
5. Check your answer.

The mole method

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Stoichiometry

Molar ratio of Y to X

$$\boxed{\frac{n_y}{n_x}}$$

Moles of X

Moles of Y

Mass of X

$X \longrightarrow Y$

Mass of Y

Example

How many grams of nitrogen dioxide can be formed by reaction of 1.44 g of nitrogen monoxide with oxygen?

$$2NO + O_2 \longrightarrow 2NO_2$$

Stoichiometry

Molar ratio of Y to X

$$\boxed{\frac{n_y}{n_x}}$$

Moles of X

Moles of Y

1.44 g of NO

Mass of NO_2

$$2NO + O_2 \longrightarrow 2NO_2$$

Stoichiometry

Molar ratio of Y to X

$$\boxed{\frac{n_y}{n_x}}$$

0.048 Moles NO → **Moles of Y**

1.44 g of NO → **Mass of NO_2**

$$2NO + O_2 \longrightarrow 2NO_2$$

Stoichiometry

Molar ratio NO_2 to NO

$\boxed{\frac{2}{2}}$

0.048 Moles NO

Moles of Y

1.44 g of NO

Mass of NO_2

$$2NO + O_2 \longrightarrow 2NO_2$$

Stoichiometry

Molar ratio NO_2 to NO

$$\boxed{\frac{2}{2}}$$

0.048 Moles NO

0.048 Moles NO_2

1.44 g of NO

Mass of NO_2

$$2NO + O_2 \longrightarrow 2NO_2$$

Stoichiometry

Molar ratio NO_2 to NO

$$\boxed{\frac{2}{2}}$$

0.048 Moles NO → **0.048 Moles NO_2**

1.44 g of NO → **2.21 g of NO_2**

$$2NO + O_2 \longrightarrow 2NO_2$$

Example

$$2NO + O_2 \longrightarrow 2NO_2$$

$$1.44\text{g } \cancel{NO} \times \frac{1\text{mol } \cancel{NO}}{30\text{g } \cancel{NO}} \times \frac{2\text{mol } \cancel{NO_2}}{2\text{mol } \cancel{NO}}$$

$$\times \frac{46\text{g } NO_2}{1\text{mol } \cancel{NO_2}} \quad \boxed{= 2.21\text{g } NO_2}$$

pg. 359 Problem 10a

Calculate the number of moles of CS_2 produced when 1.50 mol S_8 is used.

$$2\,CH_4 + S_8 \longrightarrow 2\,CS_2 + 4\,H_2S$$

$$1.5\ \cancel{\text{mol}}\ S_8 \times \frac{2\ \text{mol}\ CS_2}{1\ \cancel{\text{mol}}\ S_8} = \boxed{3\ \text{mol}\ CS_2}$$

pg. 359 Problem 10b

Calculate the number of moles of H_2S produced when 1.50 mol S_8 is used.

$$2\,CH_4 + S_8 \longrightarrow 2\,CS_2 + 4\,H_2S$$

$$1.5\ \cancel{mol}\ S_8 \times \frac{4\ mol\ H_2S}{1\ \cancel{mol}\ S_8} \boxed{= 6\ mol\ H_2S}$$

pg. 360 Problem 11

Titanium tetrachloride ($TiCl_4$) is extracted from titanium oxide using chlorine and carbon.

$$TiO_2 + C + 2\,Cl_2 \longrightarrow TiCl_4 + CO_2$$

if you begin with 1.25 mol TiO_2, what is the mass of Cl_2 needed?

$$1.25\ \cancel{\text{mol}}\ TiO_2 \times \frac{2\ \cancel{\text{mol}}\ Cl_2}{1\ \cancel{\text{mol}}\ TiO_2} \times \frac{70.9\ \text{g}\ Cl_2}{1\ \cancel{\text{mol}}\ Cl_2}$$

$$= 177\ \text{g}\ Cl_2$$

pg. 362 Problem 13

Air bags in cars can be inflated using the decomposition of sodium azide (NaN_3).

$$2\ NaN_3 \longrightarrow 2\ Na + 3\ N_2$$

determine the mass of N_2 produced if 100.0 g of NaN_3 is decomposed?

$$100.0\ g\ \cancel{NaN_3} \times \frac{1\ \cancel{mol}\ NaN_3}{65.0\ g\ \cancel{NaN_3}} \times \frac{3\ \cancel{mol}\ N_2}{2\ \cancel{mol}\ NaN_3}$$

$$\times \frac{28\ g\ N_2}{1\ \cancel{mol}\ N_2} \boxed{= 64.64\ g\ N_2}$$

Limiting Reagents

Limiting Reagent

Reactants are not always present (or available) in "stoichiometric" quantities.

One reactant may be present in quantities such that it is completely consumed while excess amounts of other reactants remain.

- called "limiting reactant" or "limiting reagent"

The limiting reagent will limit the amount of product produced.

Example

How many moles of $MgCl_2$ will be produced?

$$Mg + Cl_2 \longrightarrow MgCl_2$$

Start	1 mol	1 mol	0
Finish	0	0	1 mol

Example

How many moles of $MgCl_2$ will be produced?

$$Mg + Cl_2 \longrightarrow MgCl_2$$

	Mg	Cl_2	$MgCl_2$
Start	1 mol	2 mol	0
Finish	0	1 mol	1 mol

magnesium is the limiting reagent

1 mol of chlorine will be left unchanged

Limiting Reagent

Molar ratio Y to X

$$\boxed{\frac{n_y}{n_x}}$$

Moles of X

Moles of W

Moles of Y

Mass of X

Mass of W

$$W + X \longrightarrow Y$$

Mass of Y

Compare molar ratio W to X to their coefficients in balanced equation; identify LR

Molar ratio Y to LR

Moles of X

Moles of W

Moles of Y

Mass of X

Mass of W

$$W + X \longrightarrow Y$$

Mass of Y

Example

Determine the limiting reagent and the amount of PI_3 produced when 6.00g P_4 reacts with 25.0g of I_2.

$$P_4 + 6I_2 \longrightarrow 4PI_3$$

$$6.00\text{g } \cancel{P_4} \times \frac{1\text{mol } P_4}{124\text{g } \cancel{P_4}} = .0484\text{mol } P_4$$

$$25.0\text{g } \cancel{I_2} \times \frac{1\text{mol } I_2}{254\text{g } \cancel{I_2}} = .0984\text{mol } I_2$$

Example cont...

Determine how much I_2 would be needed to react completely with the available amount of P_4.

$$P_4 + 6I_2 \longrightarrow 4PI_3$$

if I have **I need**

$$.0484\text{mol } \cancel{P_4} \times \frac{6\text{mol } I_2}{1\text{mol } \cancel{P_4}} = \boxed{0.290\text{mol } I_2}$$

but I only have

$$\boxed{.0984\text{mol } I_2}$$

I_2 is the limiting reagent

Example cont...

…the amount of PI_3 produced from the limiting reagent...

$$P_4 + 6I_2 \longrightarrow 4PI_3$$

$$.0984\text{mol } \cancel{I_2} \quad \times \quad \frac{4\text{mol } \cancel{PI_3}}{6\text{mol } \cancel{I_2}} \quad \times \quad \frac{412\text{g } PI_3}{1\text{mol } \cancel{PI_3}} =$$

27.0g PI_3

pg. 368 problem 20a

the reaction between solid sodium and iron (III) oxide is one in a series of reactions that inflates an automobile air bag $Fe_2O_{3?}$

$$6Na + Fe_2O_3 \longrightarrow 3Na_2O + 2Fe$$

if 100.0 g Na and 100.0 g Fe_2O_3 are used in this reaction, determine: a. the limiting reactant.

$$100.0 \text{ g Na} \times \frac{1 \text{ mol Na}}{23.0 \text{ g Na}} = 4.348 \text{ mol Na}$$

$$100.0 \text{ g } Fe_2O_3 \times \frac{1 \text{ mol } Fe_2O_3}{159.9 \text{ g } Fe_2O_3} = 0.6254 \text{ mol } Fe_2O_3$$

pg. 368 problem 20a

if 100.0 g Na and 100.0 g Fe_2O_3 are used in this reaction, determine: **a. the limiting reactant.**

$$6Na + Fe_2O_3 \longrightarrow 3Na_2O + 2Fe$$

if I have

$$4.348 \text{ mol Na} \times \frac{1 \text{ mol } Fe_2O_3}{6 \text{ mol Na}} = 0.724 \text{ mol } Fe_2O_3$$

I need

0.6254 mol Fe_2O_3 is limiting

but I only have

pg. 368 problem 20a

if 100.0 g Na and 100.0 g Fe_2O_3 are used in this reaction, determine: b. the reactant in excess.

$$6Na + Fe_2O_3 \longrightarrow 3Na_2O + 2Fe$$

4.348 mol Na in excess $\times \dfrac{1 \text{ mol } Fe_2O_3}{6 \text{ mol Na}} = 0.724 \text{ mol } Fe_2O_3$

needed

0.6254 mol Fe_2O_3

available

pg. 368 problem 20a

if 100.0 g Na and 100.0 g Fe_2O_3 are used in this reaction, determine: c. the mass of iron produced.

$$6Na + Fe_2O_3 \longrightarrow 3Na_2O + 2Fe$$

$$0.6254 \cancel{\text{mol}}\ Fe_2O_3 \times \frac{2 \cancel{\text{mol}}\ Fe}{1 \cancel{\text{mol}}\ Fe_2O_3} \times \frac{55.85\ g\ Fe}{1 \cancel{\text{mol}}\ Fe}$$

$$\boxed{= 69.86\ g\ Fe}$$

produced

pg. 368 problem 20a

if 100.0 g Na and 100.0 g Fe_2O_3 are used in this reaction, determine: d. the mass of excess reactant.

$$6Na + Fe_2O_3 \longrightarrow 3Na_2O + 2Fe$$

$$0.6254 \cancel{\text{mol}}\ Fe_2O_3 \times \frac{6 \cancel{\text{mol}}\ Na}{1 \cancel{\text{mol}}\ Fe_2O_3} \times \frac{23.0\ g\ Na}{1 \cancel{\text{mol}}\ Na}$$

$= 86.30$ g Na

used

100.0 g Na - 86.30 g Na $\boxed{= 13.7\ g\ Na}$ *left over*

work sheet example

$$3CaCO_3 + 2FePO_4 \longrightarrow Ca_3(PO_4)_2 + Fe_2(CO_3)_3$$

if 100g calcium carbonate and 45 g iron(III) phosphate are used in this reaction, determine:

a. the limiting reactant.

$$100\text{ g } CaCO_3 \times \frac{1\text{ mol } CaCO_3}{100\text{ g}} = 1\text{ mol } CaCO_3$$

$$45\text{ g } FePO_4 \times \frac{1\text{ mol } FePO_4}{151\text{ g}} = 0.298\text{ mol } FePO_4$$

work sheet example

$$3CaCO_3 + 2FePO_4 \longrightarrow Ca_3(PO_4)_2 + Fe_2(CO_3)_3$$

if 100g calcium carbonate and 45 g iron(III) phosphate are used in this reaction, determine:

a. the limiting reactant.

$$1 \text{ mol } CaCO_3 \times \frac{2 \text{ mol } FePO_4}{3 \text{ mol } CaCO_3} = 0.66 \text{ mol } Fe_2O_3$$

if I have *I need*

0.298 mol $FePO_4$ is limiting

but I only have

work sheet example

$$3CaCO_3 + 2FePO_4 \longrightarrow Ca_3(PO_4)_2 + Fe_2(CO_3)_3$$

if 100g calcium carbonate and 45 g iron(III) phosphate are used in this reaction, determine:

b. grams of $Ca_3(PO_4)_2$ formed?

$$0.298\ \cancel{\text{mol}}\ FePO_4 \times \frac{1\ \cancel{\text{mol}\ Ca_3(PO_4)_2}}{2\ \cancel{\text{mol}}\ FePO_4} \times \frac{310\ \text{g}}{1\ \cancel{\text{mol}}\ Ca_3(PO_4)_2}$$

$$= 46.2\ \text{g}\ Ca_3(PO_4)_2$$

produced

work sheet example

$$3CaCO_3 + 2FePO_4 \longrightarrow Ca_3(PO_4)_2 + Fe_2(CO_3)_3$$

if 100g calcium carbonate and 45 g iron(III) phosphate are used in this reaction, determine:

b. grams of $Fe_2(CO_3)_3$ formed?

$$0.298 \cancel{\text{mol}}\ FePO_4 \times \frac{1 \cancel{\text{mol}\ Fe_2(CO_3)_3}}{2 \cancel{\text{mol}}\ FePO_4} \times \frac{291.7\ \text{g}}{1 \cancel{\text{mol}}\ Fe_2(CO_3)_3}$$

$$\boxed{= 43.5\ \text{g}\ Fe_2(CO_3)_3}$$

produced

work sheet example

$$3CaCO_3 + 2FePO_4 \longrightarrow Ca_3(PO_4)_2 + Fe_2(CO_3)_3$$

if 100g calcium carbonate and 45 g iron(III) phosphate are used in this reaction, determine:

c. grams of $CaCO_3$ unreacted?

$$0.298 \cancel{\text{mol}}\ FePO_4 \times \frac{3\ \cancel{\text{mol}}\ CaCO_3}{2\ \cancel{\text{mol}}\ FePO_4} \times \frac{100\ g}{1\ \cancel{\text{mol}}\ CaCO_3}$$

$= 44.7\ g\ CaCO_3$ *used*

100 g - 44.7 g $= 55.3\ g\ CaCO_3$ *unused*

Example

From the reaction between of 10.0g of Hg and 9.0g of Br_2 . What mass of which reagent is left unreacted?

$$Hg + Br_2 \longrightarrow HgBr_2$$

Hg is limiting

$$10.0g\ Hg \times \frac{1\ molHg}{200.6gHg} = 4.99 \times 10^{-2}\ molHg$$

$$9.0g\ Br_2 \times \frac{1\ molBr_2}{159.8gBr_2} = 5.63 \times 10^{-2}\ molBr_2$$

$$Hg + Br_2 \longrightarrow HgBr_2$$

$$4.99 \times 10^{-2}\ \text{molHg} \times \frac{1\ \text{molBr}_2}{1\ \text{molHg}} = 4.99 \times 10^{-2}\ \text{molBr}_2$$

moles of Br_2 needed to use up Hg available

$$4.99 \times 10^{-2}\ \text{molBr}_2 \times \frac{159.8\text{gBr}_2}{1\ \text{molBr}_2} = 7.97\ \text{g Br}_2$$

grams of Br_2 used

$$9.0\text{g Br}_2 - 7.97\ \text{g Br}_2 = \boxed{1.03\ \text{g Br}_2\ \text{excess}}$$

Reaction Yield

Theoretical yield

the amount of product that would result if all the limiting reagent reacted

Actual yield

the amount of product actually obtained from the reaction

Almost always less than the theoretical yield

Percent Yield

$$\%\text{yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\%$$

Determines how efficient a reaction is

Example

In a certain industrial operation 3.54×10^7g of $TiCl_4$ is reacted with 1.13×10^7g of Mg. (a) Calculate the theoretical yield of Ti in grams. (b) Calculate the percent yield if 7.91×10^6g are actually obtained.

$$TiCl_4 + 2Mg \longrightarrow Ti + 2MgCl_2$$

Calculate theoretical yield

$$3.54 \times 10^7 \cancel{g}\ TiCl_4 \times \frac{1\,mol\ TiCl_4}{187.7\,\cancel{g}\ TiCl_4} = 1.87 \times 10^5\,mol\ TiCl_4$$

I have

$$25.0\,\cancel{g}\ Mg \times \frac{1\,mol\ Mg}{24.31\,\cancel{g}\ Mg} = 4.65 \times 10^5\ mol\ Mg$$

if I have

$$1.87 \times 10^5\,\cancel{mol}\ TiCl_4 \times \frac{2\,mol\ Mg}{1\,\cancel{mol}\ TiCl_4} = 3.74 \times 10^5\ mol\ Mg$$

I need

there is more than enough Mg

$TiCl_4$ is limiting

$$3.54 \times 10^7 \cancel{\text{g TiCl}_4} \times \frac{1\cancel{\text{mol TiCl}_4}}{187.\cancel{7\text{g TiCl}_4}} \times \frac{1\cancel{\text{mol Ti}}}{1\cancel{\text{mol TiCl}_4}}$$

$$\times \frac{47.88\text{g Ti}}{1\cancel{\text{mol Ti}}} = 8.93 \times 10^6 \text{g Ti}$$ *theoretical*

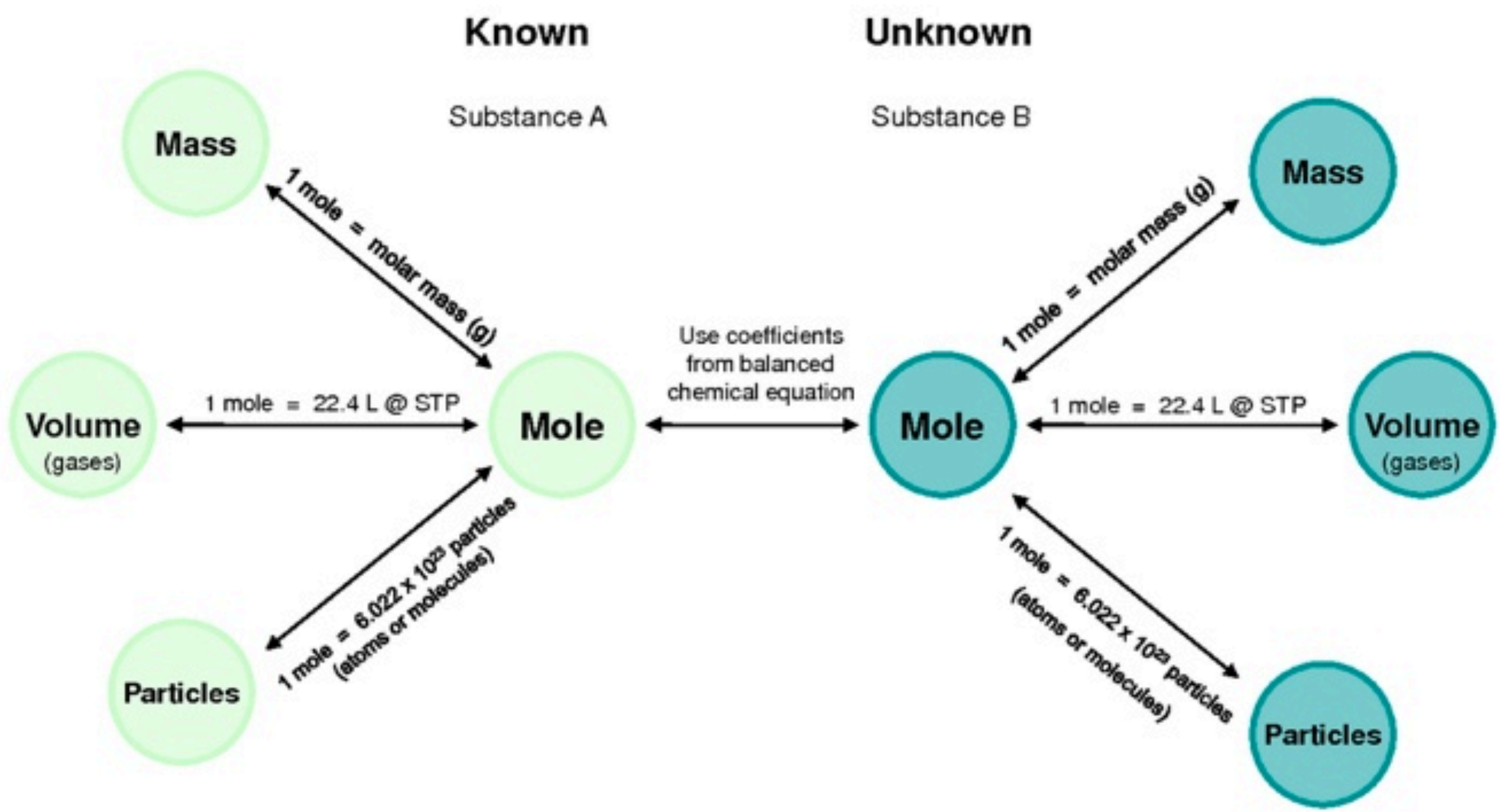
$$\%\text{yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\%$$

$$100\% \times \frac{7.91 \times 10^6 \text{g Ti}}{8.93 \times 10^6 \text{g Ti}} =$$

$$= 88.6\%$$

Stoichiometry Mole Island Diagram

When in doubt…convert to moles!



another method

Determine the limiting reagent and the amount of PI_3 produced when 6.00g P_4 reacts with 25.0g of I_2.

$$P_4 + 6I_2 \longrightarrow 4PI_3$$

$$6.00\text{g } \cancel{P_4} \times \frac{1\text{mol } P_4}{124\text{g } \cancel{P_4}} = .0484\text{mol } P_4$$

$$25.0\text{g } \cancel{I_2} \times \frac{1\text{mol } I_2}{254\text{g } \cancel{I_2}} = .0984\text{mol } I_2$$

another method

$$P_4 + 6I_2 \longrightarrow 4PI_3$$

$$\frac{.0984\text{mol } I_2}{.0484\text{mol } P_4} = 2.033$$

The actual I_2/P_4 ratio is less than the stiochiometic ratio

$$\frac{6 \text{ mol } I_2}{1 \text{ mol } P_4} = 6$$

So there is not enough I_2 to react with all the P_4

I_2 is the limiting reagent