

## TOPIC 9.

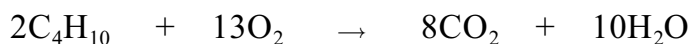
### CHEMICAL CALCULATIONS III - stoichiometry.

#### Stoichiometric calculations.

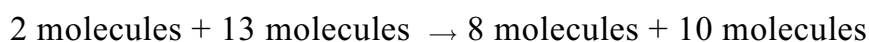
By combining a knowledge of balancing equations with the concept of the mole, it is possible to easily calculate the masses of all reactants required or products formed in any reaction. These calculations are called **STOICHIOMETRIC CALCULATIONS**. How this is done is again best illustrated by some examples.

**Example 1.** What mass of carbon dioxide and of water is produced from combustion of 10.0 g of lighter fluid, butane ( $C_4H_{10}$ ), and what mass of oxygen is consumed in the reaction?

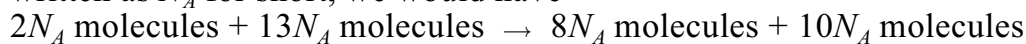
Start by writing a balanced **formula equation** for the reaction.



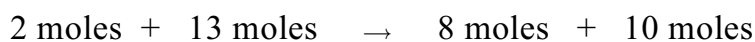
Then write down the relative numbers of molecules involved as shown in the equation (the **stoichiometric coefficients**).



If each of these numbers of molecules were multiplied by the Avogadro number, written as  $N_A$  for short, we would have



As the Avogadro number of molecules of any substance is 1 mole of the substance, then the relative numbers of moles of all species in the equation would be



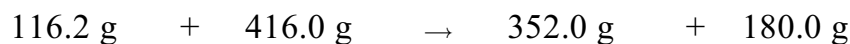
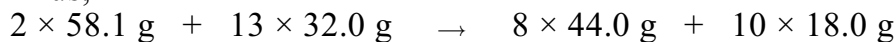
Thus the coefficients in the balanced chemical equation can be interpreted as relative numbers of moles of species as well as the relative numbers of molecules.

From Topic 7, the conversion of moles to mass for any species is achieved simply by multiplying the number of moles by the molar mass (gram formula weight) of that species, as

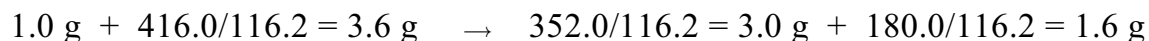
$$\text{amount in moles} = \frac{\text{mass in grams}}{\text{molar mass}}$$

or 
$$\text{mass in grams} = \text{amount in moles} \times \text{molar mass}$$

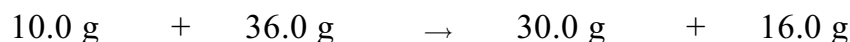
Thus,



Thence deduce amounts associated with the reaction of 1.0 g of butane:



and thus amounts associated with 10.0 g of butane:

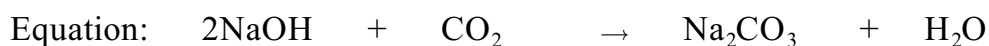


$\therefore$  10.0 g of butane uses 36.0 g of oxygen and forms 30.0 g of carbon dioxide and 16.0 g of water.

Usually the problem does not require calculation of the masses of all of the reactants and/or products in the reaction. More often only two species need to be involved in the calculation and they are referred to as the known (the reactant/product for which the masses are given) and the unknown (the reactant/product for which the masses are to be calculated).

**Example 2.** Sodium hydroxide reacts with carbon dioxide to produce sodium carbonate and water. Calculate the mass of sodium hydroxide required to prepare 53.0 g of sodium carbonate.

Using the same procedure as in Example 1,



Ratios: 2 "molecules" 1 molecule "1 molecule" 1 molecule

therefore 2 moles 1 mole 1 mole 1 mole

from which,

1 mole of the known,  $\text{Na}_2\text{CO}_3$ , requires 2 moles of the unknown, NaOH.

$$\begin{aligned} \text{Amount of Na}_2\text{CO}_3 \text{ in } 53.0 \text{ g} &= \frac{\text{mass}}{\text{molar mass}} \text{ mole} \\ &= \frac{53.0}{(2 \times 22.99 + 12.01 + 3 \times 16.00)} = 0.500 \text{ mole} \end{aligned}$$

Therefore, amount of NaOH required =  $2 \times 0.500 = 1.00$  mole

and mass of NaOH = amount of NaOH  $\times$  molar mass

$$= 1.00 \times 40.00 = 40.0 \text{ g}$$

[In the above, for convenience the term "molecules" was applied to ionic species such as NaOH which do not actually exist as molecules. Strictly a more cumbersome term such as "formula units" would be more correct.]

**Example 3.** Heating potassium chlorate,  $\text{KClO}_3$ , produces oxygen and potassium chloride. What mass of oxygen is obtained by quantitatively decomposing 3.00 g of potassium chlorate?



Mole ratios: 2 mole 2 mole 3 mole

from which 2 moles of the known,  $\text{KClO}_3$ , produces 3 moles of the unknown,  $\text{O}_2$ .

**$\therefore$  1 mole of the known ( $\text{KClO}_3$ ) produces 1.5 moles of the unknown ( $\text{O}_2$ )**

$$\text{Amount in moles of KClO}_3 \text{ in } 3.00 \text{ g} = \frac{\text{mass KClO}_3}{\text{molar mass}} = \frac{3.00}{122.6} = 0.0245 \text{ mole}$$

Therefore, amount in moles of  $\text{O}_2$  produced =  $1.5 \times 0.0245 = 0.0368$  mole

and mass of  $\text{O}_2$  produced = molar mass of  $\text{O}_2$   $\times$  moles  $\text{O}_2$

$$= 32.00 \times 0.0368 = 1.18 \text{ g}$$



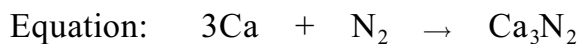
**Percentage yield calculations.**

In all the preceding examples, it has been assumed that complete conversion of reactants to products has occurred (quantitative conversion). The amount of product obtained when there is 100% conversion is called the **THEORETICAL YIELD**. For various reasons, conversion in reactions may not be complete or practical techniques used for separating and purifying materials may result in some loss of that material. Consequently, the amount of product obtained in practice, called the **ACTUAL YIELD**, is usually less than the theoretical yield. The usual way of expressing the yield obtained is as the **PERCENTAGE YIELD**. The percentage yield is given by the expression

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

Note that the units used for expressing the actual and theoretical yields are usually mass units, but they can be expressed as either mass or moles, provided that the same units are used for both.

**Example 6.** Calcium (50.5 g) is heated with excess nitrogen to form calcium nitride (60.2 g). Calculate the percentage yield obtained.



Ratios:        3 moles                                  1 mole

Whence 3 moles of the known, Ca, forms 1 mole of the unknown,  $\text{Ca}_3\text{N}_2$ .

$\therefore$  1 mole of the known, Ca, forms  $1/3$  mole of the unknown,  $\text{Ca}_3\text{N}_2$ .

Amount of Ca = mass / molar mass =  $50.5 / 40.08 = 1.26$  mole

For 100% conversion, amount of  $\text{Ca}_3\text{N}_2$  produced =  $1/3 \times 1.26$  mole = 0.420 mole

Mass of 0.420 mole  $\text{Ca}_3\text{N}_2$  = moles  $\times$  molar mass  
 =  $0.420 \times 148.26$  g = 62.3 g (theoretical yield)

Actual yield = 60.2 g

Therefore percentage yield =  $\frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{60.2}{62.3} \times 100 = 96.6 \%$

**Example 7.** Cobalt(II) chloride (12.99 g) is converted to cobalt(II) sulfate (14.90 g) by a series of reactions. Calculate the percentage yield obtained.

[Insufficient information has been given to write an equation for the conversion steps, but the full equation is not needed - all that is required is the formula for the two compounds, cobalt(II) chloride and cobalt(II) sulfate.]

1 mole of cobalt(II) chloride,  $\text{CoCl}_2$ , contains 1 mole of cobalt(II) ions.

1 mole of cobalt(II) sulfate,  $\text{CoSO}_4$ , contains 1 mole of cobalt(II) ions.

Therefore complete conversion of 1 mole of  $\text{CoCl}_2$  to  $\text{CoSO}_4$  would produce 1 mole of  $\text{CoSO}_4$

Moles of  $\text{CoCl}_2$  used = mass / molar mass =  $12.99 / 129.9 = 0.100$  mole

Therefore moles of  $\text{CoSO}_4$  theoretically obtainable = 0.100 mole

Theoretical yield of  $\text{CoSO}_4$  = moles  $\times$  molar mass of  $\text{CoSO}_4$  =  $0.100 \times 155.0$   
 = 15.50 g

Actual yield = 14.90 g

Therefore % yield =  $\frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{14.90}{15.50} \times 100 = 96.1 \%$

### Objectives of this Topic.

When you have completed this Topic, including the tutorial questions, you should have achieved the following goals:

1. Be able to calculate the mass of product obtained from a given mass of reactant, or vice versa.
2. Understand the concepts of theoretical yield, actual yield, % yield.
3. Be able to calculate the % yield for a reaction.
4. Understand the concept of limiting reagent in a reaction and apply it to calculations of amount of product obtained.

### SUMMARY

The coefficients in a balanced equation give the relative ratios of the formula units of each reactant and product species required for complete reaction. These ratios are also the ratios of the number of moles of each participating species, a fact that could be deduced by multiplying each coefficient by the same constant, the Avogadro number. Moles of any species can be converted to mass by multiplying by the molar mass (gram formula weight) of that species. Hence the relative masses of each reactant and product can be deduced from the balanced equation and simple mole theory. Calculations of the relative masses of reactants and products is called "stoichiometry".

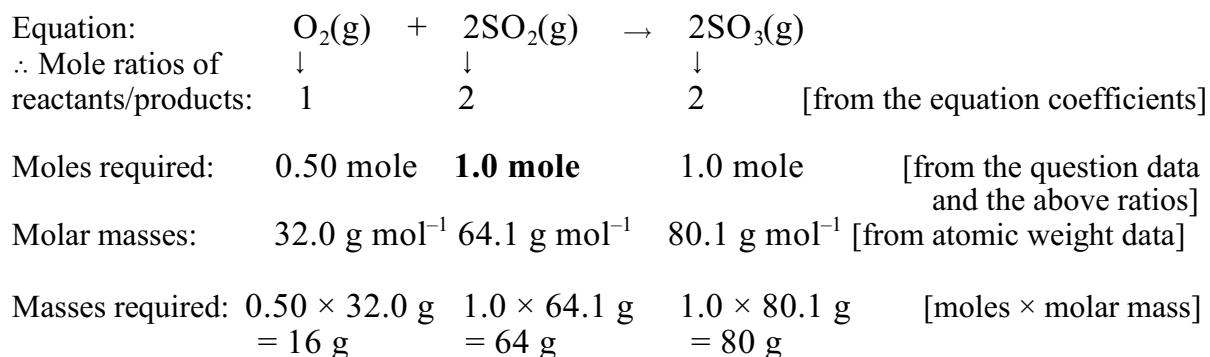
If reactants are not mixed in exactly the correct ratios required to combine completely (the stoichiometric quantities), then the reactant that is consumed totally is called the limiting reagent and the excess amounts of the other reactants that are larger than the stoichiometric amounts remain unreacted.

Stoichiometric calculations allow the theoretical yield to be calculated but in practice, the actual yield obtained may well be smaller. A quantity called the percentage yield which equals the actual yield as a percentage of the theoretical yield expresses this distinction.

## TUTORIAL QUESTIONS - TOPIC 9.

1. For each of the reactions represented by the following equations, supply the number of moles and masses for all of the reactants and products, assuming complete reaction in which 1.0 mole of the specified reactant is used or 1.0 mole of the specified product is produced. Quote your answers to the nearest gram of reactants and products. Part (a) is an example:

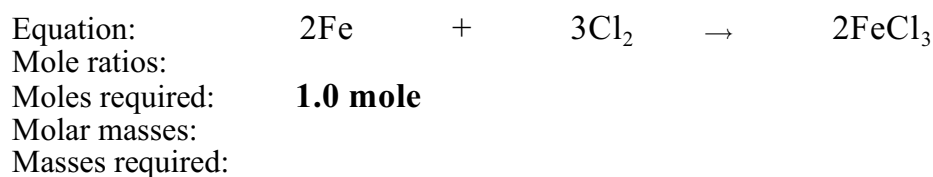
(a) For 1.0 mole of SO<sub>2</sub> reacting,



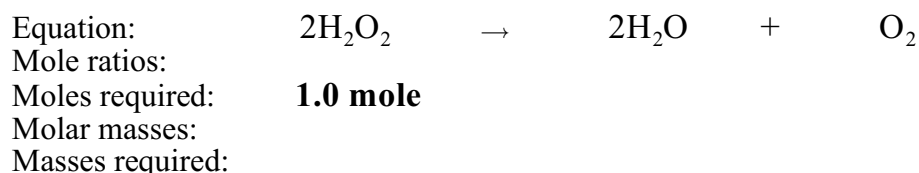
(b) For 1.0 mole of CaCO<sub>3</sub> reacting,



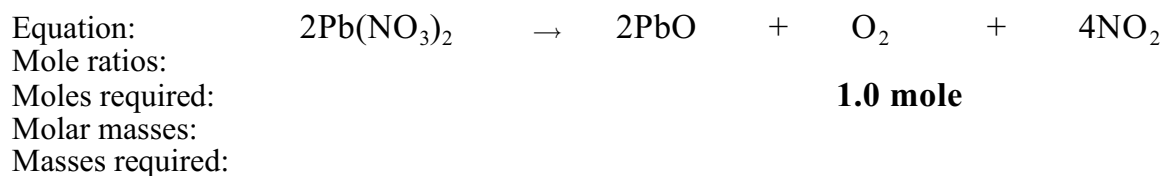
(c) For 1.0 mole of Fe reacting



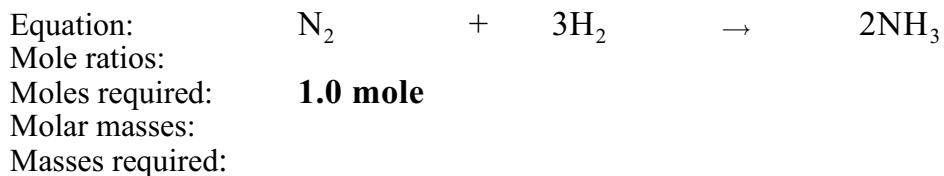
(d) For 1.0 mole of H<sub>2</sub>O<sub>2</sub> reacting,



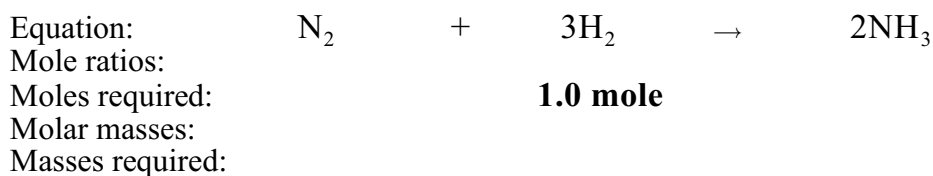
(e) For 1.0 mole of O<sub>2</sub> produced,



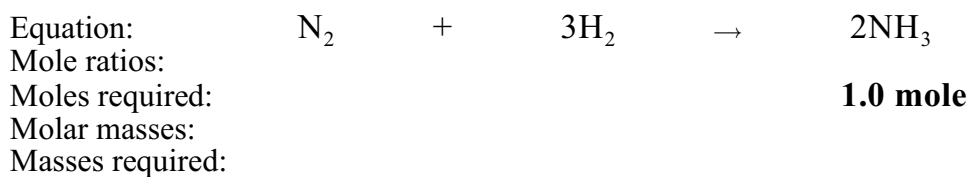
(f) For 1.0 mole of  $\text{N}_2$  reacting,



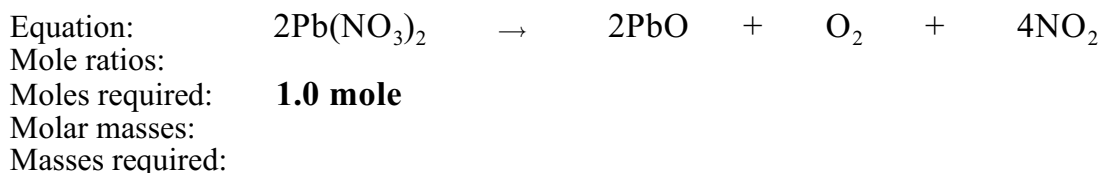
(g) For 1.0 mole of  $\text{H}_2$  reacting,



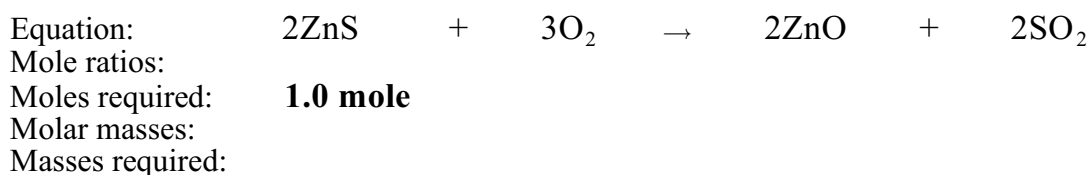
(h) For 1.0 mole of  $\text{NH}_3$  produced,



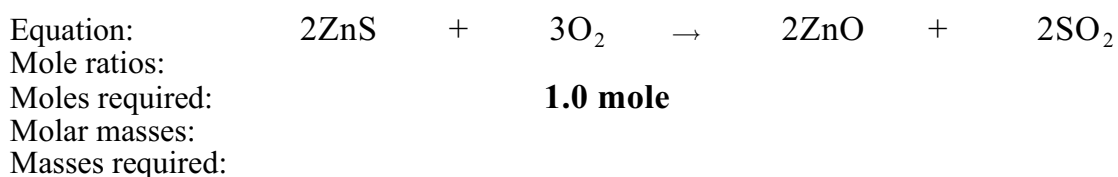
(i) For 1.0 mole of  $\text{Pb}(\text{NO}_3)_2$  reacting



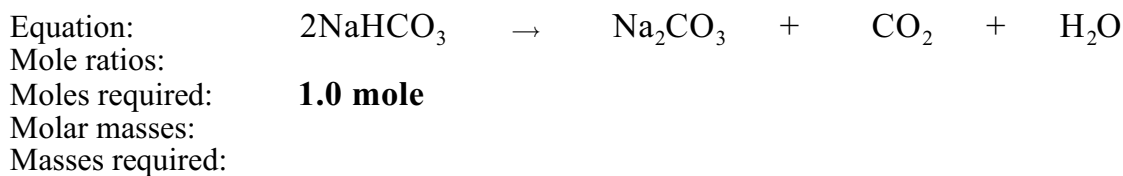
(j) For 1.0 mole of  $\text{ZnS}$  reacting,



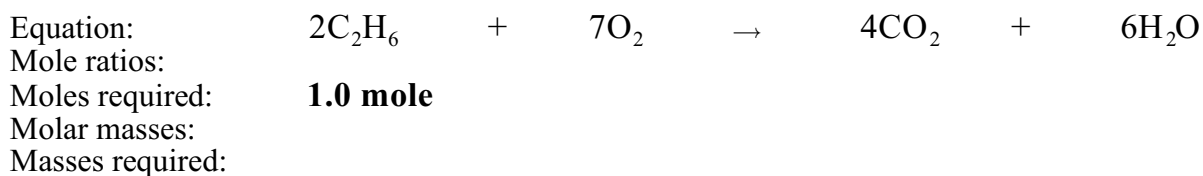
(k) For 1.0 mole of  $\text{O}_2$  reacting,



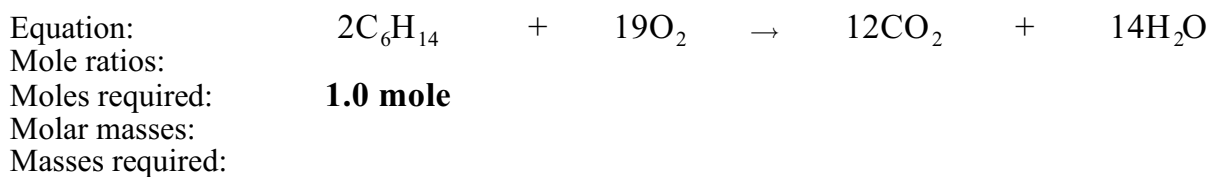
(l) For 1.0 mole of  $\text{NaHCO}_3$  reacting,



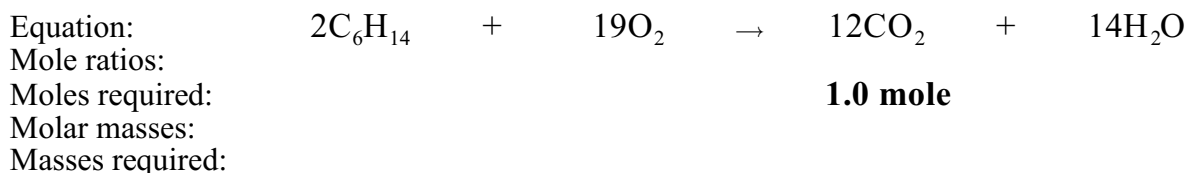
(m) For 1.0 mole of  $\text{C}_2\text{H}_6$  reacting,



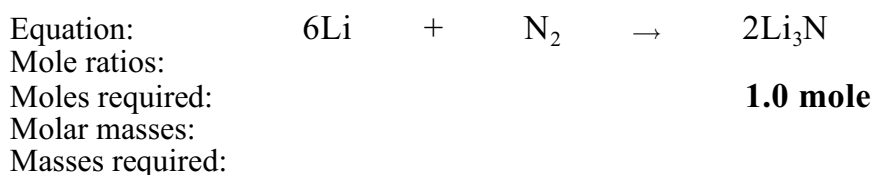
(n) For 1.0 mole of  $\text{C}_6\text{H}_{14}$  reacting,



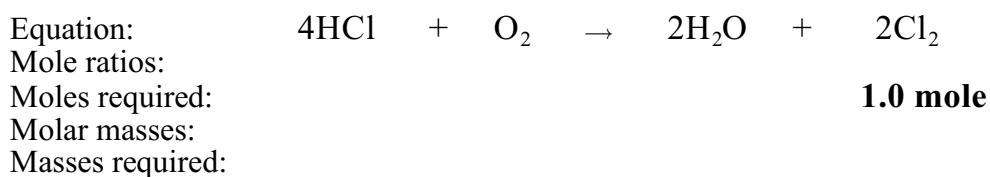
(o) For 1.0 mole of  $\text{CO}_2$  produced,



(p) For 1.0 mole of  $\text{Li}_3\text{N}$  produced,



(q) For 1.0 mole of  $\text{Cl}_2$  produced,



2. Hydrogen gas combines with chlorine gas to form hydrogen chloride. If 2.00 mole of hydrogen reacts with excess chlorine, what mass of hydrogen chloride can be obtained?

3. What mass of carbon dioxide is produced when 18 g of carbon is burnt in excess oxygen?

4. Hydrogen and chlorine gases combine to form hydrogen chloride gas, HCl. Calculate the minimum mass of hydrogen gas and of chlorine gas needed to produce 10.0 g of hydrogen chloride.

5. Calculate the mass of chlorine needed to produce 5.85 g of sodium chloride when excess sodium is heated with chlorine gas.

6. Iron burns in air to form the compound  $\text{Fe}_3\text{O}_4$ .

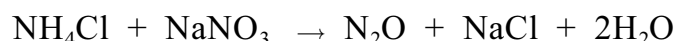
(a) Write the equation for the reaction.

(b) How many moles of oxygen gas are needed to burn 1.0 mole of iron?

(c) What is the mass of this amount of oxygen gas?

(d) Would 0.050 mole of oxygen be enough to cause 5.6 g of iron to react completely to form  $\text{Fe}_3\text{O}_4$ ?

7. The gas  $\text{N}_2\text{O}$  can be prepared by heating a mixture of ammonium chloride and sodium nitrate according to the following equation:



(i) What mass of  $\text{N}_2\text{O}$  can be prepared from 100.0 g of ammonium chloride?

(ii) The mass of  $\text{N}_2\text{O}$  actually obtained from this experiment was 75.4 g. Calculate the percentage yield.

8. The oxide of iron,  $\text{Fe}_3\text{O}_4$ , can be converted to iron metal by heating with excess hydrogen gas. What mass of  $\text{Fe}_3\text{O}_4$  is needed to produce 112 g of iron?

9. Sodium nitrate decomposes on heating to form sodium nitrite and oxygen gas.

(i) What mass of sodium nitrate would be required to produce 64.0 g of oxygen gas if 100 % conversion occurred?

(ii) If the actual yield from this reaction was only 80.0 %, what mass of sodium nitrate would have to be used in order to produce 64.0 g of oxygen?

10. Carbon (1.20 g) is burned in a closed container with 8.00 g of oxygen.

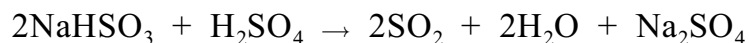
(i) What mass of carbon dioxide would be formed?

(ii) Which reagent is in excess and what mass of it remains after the reaction?

11. Aluminium sulfide,  $\text{Al}_2\text{S}_3$ , can be prepared by reacting aluminium with sulfur.

- (i) What mass of aluminium sulfide can be prepared from 10.0 g of aluminium mixed with 15.0 g of sulfur?
- (ii) Which is the non-limiting reagent and by what mass is it in excess?

12. Sulfur dioxide gas can be obtained by heating solid sodium hydrogensulfite with sulfuric acid as shown in the following equation:



- (i) Industrially, instead of using moles as gram formula weights for the molar masses of reactants and products, often units of tonne formula weights are preferred. How many tonnes of sulfur dioxide gas can be obtained from 1.00 tonne of sodium hydrogensulfite?
- (ii) The standard glob was a unit of mass used by ancient Martians before they succumbed to the loss of atmosphere through the greenhouse effect. How many globs of  $\text{SO}_2$  can be prepared from 5.0 globs of  $\text{NaHSO}_3$  in the reaction represented by the above equation?

13. Calculate the mass of sodium hydroxide required to react exactly with 0.50 mole of nitric acid,  $\text{HNO}_3$ .

14. Nickel(II) nitrate-6-water has the formula  $\text{Ni}(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O}$ . If 29.1 g of this salt were dissolved completely in water, calculate the following for the resulting solution:

- (a) The number of moles of  $\text{Ni}^{2+}$  ions present.
- (b) The number of moles of  $\text{NO}_3^-$  ions present.
- (c) The number of nickel ions present.
- (d) The number of nitrate ions present.

15. A piece of pure zinc weighing 15.8 g is dissolved completely in hydrochloric acid. Calculate:

- (a) the mass of hydrogen gas formed
- (b) the mass of zinc chloride that could be formed by evaporating the water from the resultant solution.

**ANSWERS TO TUTORIAL TOPIC 9**1. (b) For 1.0 mole of  $\text{CaCO}_3$  reacting,

Equation:	$\text{CaCO}_3$	$\rightarrow$	$\text{CaO}$	$+$	$\text{CO}_2$
Mole ratios:	1		1		1
Moles required:	<b>1.0</b>		1.0		1.0
Molar masses:	$100.1 \text{ g mol}^{-1}$		$56.1 \text{ g mol}^{-1}$		$44.0 \text{ g mol}^{-1}$
Masses required:	$1.0 \times 100.1$		$1.0 \times 56.1$		$1.0 \times 44.0$
	= 100 g		= 56 g		= 44 g

(c) For 1.0 mole of Fe reacting

Equation:	$2\text{Fe}$	$+$	$3\text{Cl}_2$	$\rightarrow$	$2\text{FeCl}_3$
Mole ratios:	2		3		2
Moles required:	<b>1.0</b>		1.5		1.0
Molar masses:	$55.8 \text{ g mol}^{-1}$		$70.9 \text{ g mol}^{-1}$		$162.2 \text{ g mol}^{-1}$
Masses required:	$1.0 \times 55.8$		$1.5 \times 70.9$		$1.0 \times 162.2$
	= 56 g		= 106 g		= 162 g

(d) For 1.0 mole of  $\text{H}_2\text{O}_2$  reacting,

Equation:	$2\text{H}_2\text{O}_2$	$\rightarrow$	$2\text{H}_2\text{O}$	$+$	$\text{O}_2$
Mole ratios:	2		2		1
Moles required:	<b>1.0</b>		1.0		0.50
Molar masses:	$34.0 \text{ g mol}^{-1}$		$18.0 \text{ g mol}^{-1}$		$32.0 \text{ g mol}^{-1}$
Masses required:	$1.0 \times 34.0$		$1.0 \times 18.0$		$0.50 \times 32.0$
	= 34 g		= 18 g		= 16 g

(e) For 1.0 mole of  $\text{O}_2$  produced,

Equation:	$2\text{Pb}(\text{NO}_3)_2$	$\rightarrow$	$2\text{PbO}$	$+$	$\text{O}_2$	$+$	$4\text{NO}_2$
Mole ratios:	2		2		1		4
Moles required:	2.0		2.0		<b>1.0</b>		4.0
Molar masses:	$331.2 \text{ g mol}^{-1}$		$223.2 \text{ g mol}^{-1}$		$32.0 \text{ g mol}^{-1}$		$46.0 \text{ g mol}^{-1}$
Masses required:	$2.0 \times 331.2$		$2.0 \times 223.2$		$1.0 \times 32.0$		$4.0 \times 46.0$
	= 662 g		= 446 g		= 32 g		= 184 g

(f) For 1.0 mole of  $\text{N}_2$  reacting,

Equation:	$\text{N}_2$	+	$3\text{H}_2$	$\rightarrow$	$2\text{NH}_3$
Mole ratios:	1		3		2
Moles required:	<b>1.0</b>		3.0		2.0
Molar masses:	$28.0 \text{ g mol}^{-1}$		$2.0 \text{ g mol}^{-1}$		$17.0 \text{ g mol}^{-1}$
Masses required:	$1.0 \times 28.0$		$3.0 \times 2.0$		$2.0 \times 17.0$
	= 28 g		= 6 g		= 34 g

(g) For 1.0 mole of  $\text{H}_2$  reacting,

Equation:	$\text{N}_2$	+	$3\text{H}_2$	$\rightarrow$	$2\text{NH}_3$
Mole ratios:	1		3		2
Moles required:	0.333		<b>1.0</b>		0.667
Molar masses:	$28.0 \text{ g mol}^{-1}$		$2.0 \text{ g mol}^{-1}$		$17.0 \text{ g mol}^{-1}$
Masses required:	$0.333 \times 28$		$1.0 \times 2.0$		$0.667 \times 17.0$
	= 9.3 g		= 2.0 g		= 11.3 g
	= 9 g		= 2 g		= 11 g

(h) For 1.0 mole of  $\text{NH}_3$  produced,

Equation:	$\text{N}_2$	+	$3\text{H}_2$	$\rightarrow$	$2\text{NH}_3$
Mole ratios:	1		3		2
Moles required:	0.50		1.5		<b>1.0</b>
Molar masses:	$28.0 \text{ g mol}^{-1}$		$2.0 \text{ g mol}^{-1}$		$17.0 \text{ g mol}^{-1}$
Masses required:	$0.50 \times 28.0$		$1.5 \times 2.0$		$1.0 \times 17.0$
	= 14 g		= 3 g		= 17 g

(i) For 1.0 mole of  $\text{Pb}(\text{NO}_3)_2$  reacting

Equation:	$2\text{Pb}(\text{NO}_3)_2$	$\rightarrow$	$2\text{PbO}$	+	$\text{O}_2$	+	$4\text{NO}_2$
Mole ratios:	2		2		1		4
Moles required:	<b>1.0</b>		1.0		0.50		2.0
Molar masses:	$331 \text{ g mol}^{-1}$		$223 \text{ g mol}^{-1}$		$32 \text{ g mol}^{-1}$		$46 \text{ g mol}^{-1}$
Masses required:	$1.0 \times 331$		$1.0 \times 223$		$0.50 \times 32$		$2.0 \times 46$
	= 331 g		= 223 g		= 16 g		= 92 g

(j) For 1.0 mole of ZnS reacting,

Equation:	$2\text{ZnS}$	+	$3\text{O}_2$	$\rightarrow$	$2\text{ZnO}$	+	$2\text{SO}_2$
Mole ratios:	2		3		2		2
Moles required:	<b>1.0</b>		1.5		1.0		1.0
Molar masses:	$97.5 \text{ g mol}^{-1}$		$32.0 \text{ g mol}^{-1}$		$81.4 \text{ g mol}^{-1}$		$64.1 \text{ g mol}^{-1}$
Masses required:	$1.0 \times 97.5$		$1.5 \times 32.0$		$1.0 \times 81.4$		$1.0 \times 64.1$
	= 97.5 g		= 48.0 g		= 81.4 g		= 64.1 g
	= 97 g		= 48 g		= 81 g		= 64 g

(k) For 1.0 mole of O<sub>2</sub> reacting,

Equation:	$2\text{ZnS}$	+	$3\text{O}_2$	$\rightarrow$	$2\text{ZnO}$	+	$2\text{SO}_2$
Mole ratios:	2		3		2		2
Moles required:	$2/3 = 0.667$		<b>1.0</b>		$2/3 = 0.667$		$2/3 = 0.667$
Molar masses:	$97.5 \text{ g mol}^{-1}$		$32.0 \text{ g mol}^{-1}$		$81.4 \text{ g mol}^{-1}$		$64.1 \text{ g mol}^{-1}$
Masses required:	$0.667 \times 97.5$		$1.0 \times 32.0$		$0.667 \times 81.4$		$0.667 \times 64.1$
	= 65.0 g		= 32.0 g		= 54.3 g		= 42.7 g
	= 65 g		= 32 g		= 54 g		= 43 g

(l) For 1.0 mole of NaHCO<sub>3</sub> reacting,

Equation:	$2\text{NaHCO}_3$	$\rightarrow$	$\text{Na}_2\text{CO}_3$	+	$\text{CO}_2$	+	$\text{H}_2\text{O}$
Mole ratios:	2		1		1		1
Moles required:	<b>1.0</b>		0.50		0.50		0.50
Molar masses:	$84.0 \text{ g mol}^{-1}$		$106.0 \text{ g mol}^{-1}$		$44.0 \text{ g mol}^{-1}$		$18 \text{ g mol}^{-1}$
Masses required:	$1.0 \times 84.0$		$0.50 \times 106.0$		$0.50 \times 44.0$		$0.50 \times 18.0$
	= 84 g		= 53 g		= 22 g		= 9 g

(m) For 1.0 mole of C<sub>2</sub>H<sub>6</sub> reacting,

Equation:	$2\text{C}_2\text{H}_6$	+	$7\text{O}_2$	$\rightarrow$	$4\text{CO}_2$	+	$6\text{H}_2\text{O}$
Mole ratios:	2		7		4		6
Moles required:	<b>1.0</b>		3.5		2.0		3.0
Molar masses:	$30.1 \text{ g mol}^{-1}$		$32.0 \text{ g mol}^{-1}$		$44.0 \text{ g mol}^{-1}$		$18.0 \text{ g mol}^{-1}$
Masses required:	$1.0 \times 30.1$		$3.5 \times 32.0$		$2.0 \times 44.0$		$3.0 \times 18.0$
	= 30 g		= 112 g		= 88 g		= 54 g

(n) For 1.0 mole of  $C_6H_{14}$  reacting,

Equation:	$2C_6H_{14}$	+	$19O_2$	$\rightarrow$	$12CO_2$	+	$14H_2O$
Mole ratios:	2		19		12		14
Moles required:	<b>1.0</b>		9.5		6.0		7.0
Molar masses:	$86.2 \text{ g mol}^{-1}$		$32.0 \text{ g mol}^{-1}$		$44.0 \text{ g mol}^{-1}$		$18.0 \text{ g mol}^{-1}$
Masses required:	$1.0 \times 86.2$ = 86 g		$9.5 \times 32.0$ = 304 g		$6.0 \times 44.0$ = 264 g		$7.0 \times 18.0$ = 126 g

(o) For 1.0 mole of  $CO_2$  produced,

Equation:	$2C_6H_{14}$	+	$19O_2$	$\rightarrow$	$12CO_2$	+	$14H_2O$
Mole ratios:	2		19		12		14
Moles required:	$2/12 = 0.167$		$19/12 = 1.58$		<b>1.0</b>		$14/12 = 1.17$
Molar masses:	$86.2 \text{ g mol}^{-1}$		$32.0 \text{ g mol}^{-1}$		$44.0 \text{ g mol}^{-1}$		$18.0 \text{ g mol}^{-1}$
Masses required:	$0.167 \times 86.2$ = 14.4 g = 14 g		$1.58 \times 32.0$ = 50.6 g = 51 g		$1.0 \times 44.0$ = 44.0 g = 44 g		$1.17 \times 18.0$ = 21.0 g = 21 g

(p) For 1.0 mole of  $Li_3N$  produced,

Equation:	$6Li$	+	$N_2$	$\rightarrow$	$2Li_3N$
Mole ratios:	6		1		2
Moles required:	3.0		0.50		<b>1.0</b>
Molar masses:	$6.9 \text{ g mol}^{-1}$		$28.0 \text{ g mol}^{-1}$		$34.8 \text{ g mol}^{-1}$
Masses required:	$3.0 \times 6.9$ = 21 g		$0.50 \times 28.0$ = 14 g		$1.0 \times 34.8$ = 35 g

(q) For 1.0 mole of  $Cl_2$  produced,

Equation:	$4HCl$	+	$O_2$	$\rightarrow$	$2H_2O$	+	$2Cl_2$
Mole ratios:	4		1		2		2
Moles required:	2.0		0.50		1.0		<b>1.0</b>
Molar masses:	$36.5 \text{ g mol}^{-1}$		$32.0 \text{ g mol}^{-1}$		$18.0 \text{ g mol}^{-1}$		$70.9 \text{ g mol}^{-1}$
Masses required:	$2.0 \times 36.5$ = 73 g		$0.5 \times 32.0$ = 16 g		$1.0 \times 18.0$ = 18 g		$1.0 \times 70.9$ = 71 g

**Worked solutions to questions 2 - 15 are provided on pages 15 - 20.**

2. 146 g
3. 66 g
4. 0.277 g H<sub>2</sub> and 9.71 g Cl<sub>2</sub>
5. 3.55 g
6. (a)  $3\text{Fe} + 2\text{O}_2 \rightarrow \text{Fe}_3\text{O}_4$   
 (b) 0.67 mol  
 (c) 21 g  
 (d) No. 5.6 g Fe (0.10 mol) requires 0.067 mol O<sub>2</sub>.
7. (i) 82.30 g                      (ii) 91.6 %
8. 154 g
9. (i) 340 g                      (ii) 425 g
10. (i) 4.40 g                      (ii) O<sub>2</sub> by 4.80 g
11. (i) 23.4 g                      (ii) Al by 1.59 g
12. (i) 0.615 ton                      (ii) 3.1 globs
13. 20 g
14. (a) 0.100 mol              (b) 0.200 mol              (c)  $6.02 \times 10^{22}$               (d)  $12.0 \times 10^{22}$
15. (a) 0.488 g                      (b) 33.0 g

**WORKED SOLUTIONS FOR QUESTIONS 2 - 15.**

2. Balanced equation:  $\text{H}_2 + \text{Cl}_2 \rightarrow 2\text{HCl}$

Mole ratios:                      1 mol                      2 mol

Moles of H<sub>2</sub> available = 2.00 mol

∴ moles of HCl produced =  $2 \times 2.00 \text{ mol} = 4.00 \text{ mol}$

Mass of HCl = moles of HCl  $\times$  molar mass of HCl  
 $= 4.00 \times (1.008 + 35.45) = 146 \text{ g}$



Mole ratios:  $\quad\quad\quad 1 \text{ mol} \quad\quad\quad 1 \text{ mol}$

Mass of carbon available = 18 g

$\therefore$  moles of C available =  $18 / 12.01 = 1.50 \text{ mol}$

From the mole ratios above, moles of  $\text{CO}_2$  formed = 1.50 mol

Mass of carbon dioxide = moles  $\times$  molar mass of  $\text{CO}_2$   
 $= 1.50 \times (12.01 + 2 \times 16.00) = 66 \text{ g}$



Mole ratios:  $\quad\quad\quad 1 \text{ mol} \quad 1 \text{ mol} \quad 2 \text{ mol}$

Mass of hydrogen chloride = 10.0 g

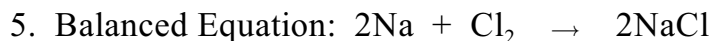
$\therefore$  moles of HCl = mass / molar mass of HCl =  $10.0 / (1.01 + 35.45) = 0.274 \text{ mol}$

1 mole of HCl requires  $\frac{1}{2}$  mole of  $\text{H}_2$  and  $\frac{1}{2}$  mole of  $\text{Cl}_2$

$\therefore$  moles of  $\text{H}_2$  = moles of  $\text{Cl}_2$  =  $0.500 \times 0.274 \text{ mol} = 0.137 \text{ mol}$

Mass of hydrogen = moles  $\times$  molar mass of  $\text{H}_2$  =  $0.137 \times 2.02 \text{ g} = 0.277 \text{ g}$

Mass of chlorine = moles  $\times$  molar mass of  $\text{Cl}_2$  =  $0.137 \times 70.90 \text{ g} = 9.71 \text{ g}$



Mole ratios:  $\quad\quad\quad 1 \text{ mol} \quad 2 \text{ mol}$

Mass of sodium chloride = 5.85 g

$\therefore$  moles of NaCl =  $5.85 / \text{molar mass of NaCl} = 5.85 / 58.5 = 0.100 \text{ mol}$

Moles of  $\text{Cl}_2$  required =  $\frac{1}{2} \times$  moles of NaCl =  $0.500 \times 0.100 = 0.0500 \text{ mol}$

Mass of  $\text{Cl}_2$  = moles  $\times$  molar mass of  $\text{Cl}_2$  =  $0.0500 \times 70.90 \text{ g} = 3.55 \text{ g}$



(b) From the stoichiometric coefficients in the equation, 3 moles of Fe require 2 moles of  $\text{O}_2$  so 1.0 mole of Fe requires  $2/3$  moles of  $\text{O}_2$  = 0.667 mol

(c) Molar mass of  $\text{O}_2$  =  $2 \times 16.00 \text{ g} = 32.00 \text{ g mol}^{-1}$

$\therefore$  0.667 mole of  $\text{O}_2$  has a mass =  $0.667 \times 32.00 = 21 \text{ g}$

(d) From the equation, 1 mole of  $\text{O}_2$  requires 1.5 moles of Fe for complete reaction.  
 $\therefore$  0.050 mole of  $\text{O}_2$  would require  $0.050 \times 1.5 = 0.075$  mole of Fe.

Mass of 0.075 mole of Fe =  $0.075 \times \text{molar mass of Fe} = 0.075 \times 55.85 \text{ g} = 4.2 \text{ g}$ .

Thus there would be an excess of iron =  $5.6 - 4.2 = 1.4 \text{ g}$





Mole ratios: 1 mol 1 mol 1 mol

[Note that this is a limiting reagent type of problem where both reactants have a specified mass and one of them is in excess. The first task is to find out which is the limiting reagent and this will determine how much carbon dioxide can form.]

Moles of C = mass / molar mass of C =  $1.20 / 12.01 = 0.100$  mol

Moles of  $O_2$  = mass / molar mass of  $O_2$  =  $8.00 / 32.00 = 0.250$  mol

From the equation, 1 mole of C uses 1 mole of  $O_2$  to produce 1 mole of  $CO_2$ .

If all the carbon (0.100 mole) were used it would react with 0.100 mole of  $O_2$ , leaving  $(0.250 - 0.100)$  mole of  $O_2$  in excess unreacted and form 0.100 mole of  $CO_2$ , so the carbon is the limiting reagent.

Mass of carbon dioxide formed = moles  $\times$  molar mass of  $CO_2$   
 $= 0.100 \times 44.0 = 4.40$  g

(ii) From (i), the oxygen is in excess by 0.150 mole.

Mass of oxygen in excess = moles  $\times$  molar mass of  $O_2$  =  $0.150 \times 32.0$  g = 4.80 g.



Mole ratios: 2 mol 3 mol 1 mol

[This is another limiting reagent type of problem.]

Moles of Al = mass / molar mass of Al =  $10.0 / 26.98 = 0.371$  mol

Moles of S = mass / molar mass of S =  $15.0 / 32.07 = 0.468$  mol

From the equation, 1 mole of S requires  $2/3$  mole of Al and produces  $1/3$  mole of  $Al_2S_3$ .

$\therefore$  if all the sulfur reacted, it would require  $(2/3) \times 0.468$  mole of Al  
 $= 0.312$  mole and produce  $(1/3) \times 0.468$  mole of  $Al_2S_3$ .

As 0.371 mole of Al are available, then the sulfur is the limiting reagent and the aluminium is in excess by  $(0.371 - 0.312)$  mole.

Mass of  $Al_2S_3$  produced = moles  $\times$  molar mass of  $Al_2S_3$   
 $= (1/3) \times 0.468 \times 150.2 = 23.4$  g.

(ii) From (i), the aluminium is in excess by  $(0.371 - 0.312) = 0.059$  mol

Mass of aluminium in excess = moles  $\times$  molar mass of Al  
 $= 0.059 \times 26.98$  g = 1.59 g



14. (a) Molar mass of  $\text{Ni}(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O} = 290.8 \text{ g mol}^{-1}$ .

From its formula, each mole of  $\text{Ni}(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O}$  contains one mole of  $\text{Ni}^{2+}$  ions.

$$\begin{aligned} \therefore \text{moles of } \text{Ni}^{2+} &= \text{moles of } \text{Ni}(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O} = \text{mass} / \text{molar mass} \\ &= 29.1 / 290.8 \text{ mol} = 0.100 \text{ mol} \end{aligned}$$

(b) From the formula for nickel(II) nitrate-6- water,

$$\text{moles of } \text{NO}_3^- = 2 \times \text{moles of } \text{Ni}(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O}.$$

$$\text{Moles of } \text{Ni}(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O} = 0.100 \text{ mol}$$

$$\therefore \text{moles of } \text{NO}_3^- = 2 \times 0.100 = 0.200 \text{ mol}$$

$$\begin{aligned} \text{(c) Number of } \text{Ni}^{2+} \text{ ions} &= N_A \times \text{moles of } \text{Ni}^{2+} = 6.022 \times 10^{23} \times 0.100 \\ &= 6.02 \times 10^{22} \text{ Ni}^{2+} \text{ ions.} \end{aligned}$$

$$\begin{aligned} \text{(d) Number of } \text{NO}_3^- \text{ ions} &= N_A \times \text{moles of } \text{NO}_3^- = 6.022 \times 10^{23} \times 0.200 \\ &= 12.0 \times 10^{22} \text{ NO}_3^- \text{ ions.} \end{aligned}$$

15. (a) Equation:  $\text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2$

Mole ratios:  $1 \text{ mol} \quad 2 \text{ mol} \quad 1 \text{ mol} \quad 1 \text{ mol}$

Moles of  $\text{H}_2$  produced = moles of Zn reacting.

$$\text{Moles of Zn in } 15.8 \text{ g} = \text{mass} / \text{molar mass of Zn} = 15.8 / 65.39 = 0.242 \text{ mol}$$

$$\therefore \text{moles of } \text{H}_2 = 0.242 \text{ mol}$$

$$\begin{aligned} \text{Mass of } \text{H}_2 &= \text{moles of } \text{H}_2 \times \text{molar mass of } \text{H}_2 \\ &= 0.242 \times 2.016 \text{ g} = 0.488 \text{ g.} \end{aligned}$$

(b) Moles of  $\text{ZnCl}_2$  formed = moles of Zn reacting = 0.242 mol

$$\therefore \text{mass of } \text{ZnCl}_2 = \text{moles} \times \text{molar mass of } \text{ZnCl}_2 = 0.242 \times 136.3 \text{ g} = 33.0 \text{ g.}$$