10 Calculations Using Chemical Equations

Introduction

In previous chapters, students have used word equations to represent chemical reactions. In this chapter, they are introduced to chemical equations. A chemical equation is a concise and universally adopted way to represent a chemical reaction. Students learn to transcribe word equations into chemical equations. Then, together with ideas from Chapters 8 and 9 relating to quantitative chemistry, students use chemical equations to do calculations involving reacting masses and volumes of gases.

Chapter Opener (page 147)

1. To open the chapter, the following questions could be discussed. Precise answers are not needed at this stage.

   What is a chemical equation and what does it show?

   Answer: A chemical equation uses chemicals symbols and formulae to show the changes that occur in a chemical reaction.

   Chemical equation should be balanced. What does this mean?

   Answer: The number and type of atoms for the reactants and the products is the same.

   For every molecule of methane gas that burns in air, one molecule of carbon dioxide gas is formed. If 10 cm³ of methane reacts, what volume of carbon dioxide gas is obtained (measured at the same temperature)? Explain.

   The volume of carbon dioxide is also 10 cm^3. From Avogardo's law, equal volumes of gas contain equal number of molecules.

2. Carry out an 'Inquiry Preview.'

Learning Outcomes

After completing this chapter, the students should be able to:

- interpret and construct chemical equations with state symbols
- calculate stoichiometric reacting masses and volumes of gases
- perform simple calculations involving the idea of limiting reactants
- calculate the percentage yield of a product in a reaction and the percentage purity of a reactant

Teaching pointers

10.1 How Do We Construct Chemical Equations? (page 148)

Stimulation

Refer to the use of chemical reactions in the chemical industry such as the extraction of iron from its ore in the Chem Mystery. Discuss the idea that chemists need to calculate the amounts of reagents needed to produce a certain amount of product of the mass of the amount of a product that can be obtained from given amounts of reagents. The teacher might, for example, show some iron ore and a piece of xxxxron and pose the question as to how mush iron could be obtained from the ore. Introduce chemical equation as necessary for such calculations. Note that the question on this extraction of iron will not be fully answered until mystery is solved.

1. Before beginning to teach this section, you may want to have a quick revision of the symbols and formulae for elements and compounds.
2. Many students find it difficult to comprehend and write chemical equations. Thus, two familiar reactions have been used in the Textbook to illustrate the writing of chemical equations (refer to pages 148 and 149 of the Textbook). If possible, use *space-filling* models to facilitate the writing and balancing of equations.
3. Emphasise the links between the macroscopic world of observable reactions and the microscopic world of particles as represented in chemical equations. The *macroscopic* world is the real world of observable matter. The *microscopic* world is the world of particles that make up that matter.
4. At this stage, students will be writing equations for many reactions they have not or will never have the chance to carry out. The practice they get now from writing equations will make the later study of the reactions easier.
5. Point out that a chemical equation contains more information about the substances involved in a reaction than a word equation because it uses formulae and states instead of names. However, you may like to point out the information not available. For example:
   - the conditions needed for a reaction to occur.
   - the speed of a reaction (whether it is fast or slow).
   - whether energy is taken in or given out during a reaction.
6. Students usually have problems balancing equations. Get them into the practice of ensuring that the numbers of each kind of atoms on both sides of the equation are equal. Pay special attention to the three notes on page 150 of the Textbook on the balancing of equations.

7. The historical method of writing and balancing a chemical equation is to carry out experiments to identify the products and to measure the relative amounts of reactants and products. This approach is seldom used now. In modern Chemistry courses, students use chemical theory to write chemical formulae and equations.
8. When carrying out Exercise 10.1 in the Theory Workbook on the writing of chemical equations, students may again use the formula cards referred to in Chapter 6 for the writing on chemical formulae.
9. Ionic equations should be omitted until students gain more experience in reactions involving ions.

(page 150)
Mystery Clue

$Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(l) + 3CO_2(g)$

Skills Practice (page 151)

1. (a) $N_2(g) + \underline{3}H_2(g) \longrightarrow \underline{2}NH_3(g)$
   (b) $\underline{4}Na(s) + O_2(g) \longrightarrow \underline{2}Na_2O(s)$
   (c) $\underline{3}Fe(s) + \underline{4}H_2O(g) \longrightarrow Fe_3O_4(s) + \underline{4}H_2(g)$
2. (a) $2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$
   (b) $2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$
   (c) $Mg(s) + 2H_2O(l) \longrightarrow Mg(OH)_2(aq) + H_2(g)$
   (d) $4NH_3(g) + 3O_2(g) \longrightarrow 2N_2(g) + 6H_2O(l)$
   (e) $Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$

Notes for Teachers

Writing chemical equations

Many students have difficulties writing and comprehending chemical equations. Hence this topic must be introduced slowly. When writing equations, students must not lose sight of what an equation represents. The example on the reaction between hydrogen and oxygen on page 148 of the Textbook involves molecules. Simple molecules are easier to understand than giant structures. Giant structures are included in the second example for the reaction between magnesium (which has a giant metallic structure) and oxygen to give magnesium oxide (which has a giant ionic structure). It is important for students to understand that when we write a giant structure such as MgO, we are *not* implying that it is a molecule. Refer to Figures 10.1.and 10.2 on pages 148 and 149 respectively, so that students can visualise what an equation represents.

Students can use formula cards and models to explore the use of symbols, formulae and equations. The advantage of models is that they give a pictorial view of the change represented in an equation. Another strategy is to use coloured, transparent two-dimensional plastic models of atoms and molecules on an overhead projector to simulate the changes represented by an equation.

Teaching pointers

10.2 Calculations from Equations (page 151)

1. To do calculations involving reacting masses and volumes of gases, students need a good understanding of the *mole*, *molar mass* and *molar volume*. Revise these terms.
2. As suggested in earlier chapters, when doing calculations, get students to describe the steps *qualitatively* first before doing the numerical calculations. (If the examples in the Textbook are used, ensure that students close their textbooks so that they are not just looking at the numerical steps.)

(page 152)
Mystery Clue

Mass of iron produced = 28 g.

3. Limiting reactions are important when speeds of reactions are investigated. For example, in the reaction between magnesium and hydrochloric acid, it is necessary to know which reactant is used up first.
4. A worksheet on the use of spreadsheet for chemical calculations involving the extraction of iron is provided at the end of this chapter. You may photocopy and distribute the worksheet to the class.

Note: An easy way to check students' spreadsheets is to ask students to submit a printout of their spreadsheets. Prepare a printout of the correct spreadsheet. The numbers should match if the students' work is correct.

Skills Practice (page 153)

1. (a) $2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$
   Number of moles of magnesium = 5.2/24 = 0.217 mol
   Number of moles of magnesium oxide = 0.217 mol
   Mass of magnesium oxide = 0.217 × (24 + 16) = 8.7 g
   (b) Number of moles of magnesium = 0.84/24 = 0.035 mol
   Number of moles of magnesium oxide = 0.035 mol
   Mass of magnesium oxide = 0.035 × (24 + 16) = 1.4 g
2. Number of moles of zinc = 4.16/65 = 0.064
   Number of moles of zinc oxide = 0.064
   Mass of zinc oxide = 0.064 × (65 + 16) = 5.18 g
3. Number of moles of aluminium oxide = 3.4/102 = 0.0333
   Number of moles of aluminium = 2 × 0.0333 = 0.0666
   Mass of aluminium = 0.0666 × 27 = 1.8 g
4. Number of moles of copper(II) oxide = 8/80 = 0.1
   Number of moles of aluminium oxide = $\frac{1}{3}$ × 0.1 = 0.0333
   Mass of aluminium oxide = 0.0333 × 102 = 3.4 g

   Note: If the relative atomic mass of copper is taken as 63.5 instead of 64, the answer becomes 3.42 g.

Skills Practice (page 155)

1. Number of moles of sulfur = 24/32 = 0.75 mol
   Number of moles of sulfur dioxide = 0.75 mol
   Volume of sulfur dioxide = 0.75 × 24 = 18 dm^3 (18 000 cm^3)
2. Number of moles of hydrogen peroxide = 17/34 = 0.5 mol
   Number of moles of oxygen = 0.5 × 0.5 = 0.25 mol
   Volume of oxygen = 0.25 × 24 = 6 dm^3 (6 000 cm^3)
3.

(a) $2CuO(s) + C(s) \longrightarrow 2Cu(s) + CO_2(g)$ Number of moles of CuO = 80.5/80 = 1 mol Number of moles of copper = 1 mol Mass of copper = 1 × 64 = 64 g (b) Number of moles of carbon dioxide = 0.5 × 1 = 0.5 mol Volume of carbon dioxide = 0.5 × 24 = 12 dm^3 (12 000 cm^3)
(a) $2ZnS(s) + 3O_2(g) \longrightarrow 2ZnO(s) + 2SO_2(g)$ Number of moles of SO_2 = 12/24 = 0.5 mol Number of moles of zinc sulfide = 0.5 mol Mass of zinc sulfide = 0.5 × 97 = 48.5 g (b) Number of moles of zinc oxide = 0.5 mol Mass of zinc oxide = 0.5 × 81 = 40.5 g
$2Pb(NO_3)_2(s) \longrightarrow 2PbO(s) + 4NO_2(g) + O_2(g)$ Number of moles of NO_2 = 6 ÷ 24 = 0.25 mol Number of moles of lead(II) oxide = 0.5 × 0.25 = 0.125 mol Mass of lead(II) oxide = 0.125 × (207 + 16) = 27.9 g

Teaching pointers

10.3 How Do We Calculate Volumes of Reacting Gases from Equations? (page 156)

1. This topic uses knowledge of:
   (a) Relative numbers of particles/moles of particles in the chemical equation, and
   (b) Avogadro's Law.
2. In Figure 10.3 on page 156 of the Textbook, we are only concerned with the relative numbers of molecules. Hence, the term volume is used rather than any specific volume. Any number of molecules can be drawn for 1 volume; in this case, four molecules were drawn. It might be an interesting exercise to get the class to calculate the actual numbers of molecules involved if 1 cm^3 of nitrogen was used.

Skills Practice (page 156)

1.

Nitrogen, N_2	Hydrogen, H_2	Ammonia, NH_3
4 dm^3	12 dm^3	8 dm^3
24 dm^3	72 dm^3	48 dm^3

Table 9.2

2. (a) (i) 200 cm^3 of H_2 react with 100 cm^3 of O_2
   (ii) 200 cm^3 $H_2O(g)$ is produced.
   (b) (i) 200 cm^3 SO_2 react with 100 cm^3 of O_2
   (ii) 200 cm^3 SO_3 is produced.
   (c) (i) 50 cm^3 CH_4 react with 100 cm^3 of O_2
   (ii) 50 cm^3 CO_2 is produced.
   (d) (i) 80 cm^3 NH_3 react with 100 cm^3 of O_2
   (ii) 80 cm^3 NO_2 is produced.
3. Volume of SO_2 = 0.5 × 40 = 20 cm^3
4. (a) (i) (ii)

oxygen, O_2 sulfur dioxide, SO_2

   (b) Volume of sulfur dioxide produced = 50 cm^3

Teaching pointers

10.4 What are Limiting Reactants? (page 157)

Limiting reactions will also be important later when speeds of reactions are investigated. For example, in the reaction of magnesium with hydrochloric acid, it is necessary to know which reactant is used up first.

Skills Practice (page 157)

(a) $Fe(s) + S(s) \longrightarrow FeS(s)$
(b) The molar masses of Fe and S are 56 g and 32 g respectively. For a given mass of sulfur, a larger mass of iron is needed to react with all the sulfur. In the question, 8 g each of iron and sulfur are used. When all the Fe has reacted, some S remains. Therefore, Fe is the limiting reactant.
(c) Fe is the limiting reactant, thus all 8 g of the iron filings reacted.
Number of moles of sulfur = number of moles of iron
= 8/56 = 0.143 mol
Mass of sulfur reacted = 0.143 × 32 = 4.57 g

Number of moles of FeS formed = 0.143 mol
Mass of FeS formed = 0.143 × 88 = 12.57 g.

(page 157)

Mystery Clue

The limiting reactant is iron(III) oxide which allows all the oxide to be converted to iron, which is the desired aim. Thus the carbon monoxide needs to be in excess for this to occur.

Teaching pointers

10.5 How are Percentage Yield and Percentage Purity Calculated? (page 158)

1. Percentage yield is particularly important in *organic* reactions where side reactions can reduce the yield of the main product quite significantly.
2. In calculations involving percentage purity, we usually assume that the yield of product from the reaction is 100%. Without this assumption, or without knowing the actual yield, it would not be possible to do the calculations.

Skills Practice (page 157)

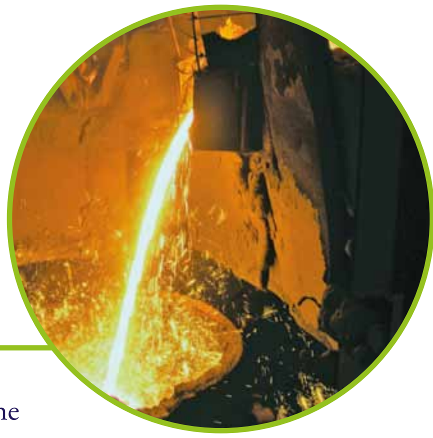
1. $Zn(s) + S(s) \longrightarrow ZnS(s)$
   Number of moles of zinc = 6.5/65 = 0.1 mol
   Number of moles of ZnS = 0.1 mol
   Mass of ZnS expected to be obtained = 0.1 × 97 = 9.7 g
   $$\text{Percentage yield} = \frac{9.0}{9.7} \times 100$$
   $$= 92.8\%$$
2. $CaCl_2(aq) + Na_2CO_3(aq) \longrightarrow CaCO_3(s) + 2NaCl(aq)$
   Number of moles of calcium carbonate = 0.2 mol
   Mass of calcium carbonate expected to be obtained = 0.2 × 100 = 20 g
   $$\text{Percentage yield} = \frac{18}{20} \times 100$$
   $$= 90\%$$

(page 159)

Mystery Clue

100 g of haematite ore provides 85 g of Fe_2O_3 to give 59.5 g of iron. 100 g of magnetite ore provides 40 g of Fe_3O_4 to give 29.0 g of iron. Hence, based on these data only, haematite would be the better choice as the overall yield is higher.

Solving the Mystery (page 160)



Producing iron — how does the chemist produce 30 000 tonnes of iron?

This exercise brings together the key ideas introduced in Chapters 8 to 10 that are needed to do this calculation, namely: knowledge of chemical formulae, chemical equations and how to balance then, knowledge of molar masses, how to calculate numbers of numbers of moles of substances using chemical equations and knowledge of limiting reaction and percentage purity.
To provide further practice with the equation for the production of iron using a spreadsheet, Additional Exercise 3 is provided at the end of the chapter. The exercise involves calculating masses of iron that can be obtained from different quantities of iron(III) oxide and is similar to the kind of calculations that an industrial chemist might carry out. Teachers can again check students work with the spreadsheet using an overhead projector.

Infer
Mass of iron from 30 000 tonnes of iron(III) oxide is 21 000 tonnes.

Connect
An alloy is a mixture of a mixture containing two or more metals or metals and non-metals fused together. Brass is an alloy of two metals, zinc and copper. Steel is an alloy of iron, carbon and usually other elements.

Further Thought
Examples can be found in Chapter 14, Table 14.2 on page 212.

10 Chapter Review

Concept Link

Self-Management

Misconception Analysis (page 161)

1. **True** For example, $2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$
For every 2 moles of hydrogen, one mole of oxygen is needed and 2 moles of water are formed. The numbers 2, 1 and 2 give the relative numbers of moles of hydrogen, oxygen and water respectively.

2. **False** Formulae are fixed. Thus, the formula of water is always H_2O. To balance an equation, the numbers of particles in the equation are changed.

3. **True** There are often 'side reactions' which produce unwanted products.

4. **False** Equal volumes of gases contain equal numbers of *molecules*.

5. **False** One mole of any gas has the same volume (at the same temperature and pressure) but *not* the same mass. For example, 1 mole of H_2 has a mass of 2 g whereas 1 mole of O_2 has a mass of 32 g.

Practice

Structured Questions (pages 162)

1. (a) $4Fe(s) + 3O_2(g) \longrightarrow 2Fe_2O_3(s)$
(b) (i) Number of moles of O_2 that react = 6 ÷ 24 = 0.25 mol
(ii) Mass of Fe_2O_3 produced
= 0.25 ÷ 3 x 2 x (56 x 2 + 16 x 3) = 26.7g

2. (a) Number of moles of carbon that reacts = 0.5 × 3 = 1.5 mol
Mass of carbon that reacts = 1.5 × 12 = 18 g
(b) Number of moles of CO produced = 1.5 mol
Volume of CO produced at r.t.p. = 1.5 × 24 = 36 dm^3

3. (a) Relative molecular mass of P_4O_{10} = 4(31) + 10(16) = 284
(b) Number of moles of 48 dm^3 of phosphine = 48 ÷ 24 = 2 mol
Mass of phosphine = 2 × 34 = 68 g
(c) Volume of oxygen required at r.t.p. = 2 × 72 = 144 dm^3

4. (a) Number of moles of sodium carbonate used = 4 × 2 = 8 mol
(b) Number of moles of sulfuric acid used = 8 mol
(c) Moles of water produced = 8 mol
Mass of water produced = 8 × 18 = 144 g

5. (a) The element is bismuth.
(b) (i) Number of moles of carbon needed = 12 × 3 = 36 mol
(ii) Number of moles of carbon monoxide = 36 mol
Volume of carbon monoxide produced = 36 × 24 = 864 dm^3
(c) (i) Relative molecular mass of Bi_2O_3 = 2(209) + 3(16) = 466
(ii) Number of moles of 1864 kg of Bi_2O_3 = 1864 000 ÷ 466
= 4000 mol
Number of moles of Bi = 2 × 4000 = 8000 mol
Maximum mass of Bi that can be extracted = 8000 × 209
= 1672 kg

6. (a) (i) Number of moles of zinc = 6.5 ÷ 65 = 0.1 mol
(ii) Number of moles of dilute hydrochloric acid = 7.3 ÷ 36.5
= 0.2 mol
(b) The reaction stops because all the reactants have been used up. 0.1 mole of zinc chloride and 0.1 mole of hydrogen gas are present in the beaker at the end.

7. (a) Number of moles of oxygen produced = 2 mol
(b) Number of moles of silver nitrate that were decomposed
= 4 mol
(c) Mass of silver nitrate heated = 4 × 170 = 680 g

8. (a) $2AmF_3(s) + 3Ba(s) \longrightarrow 2Am(s) + 3BaF_2(s)$
(b) Barium fluoride
(c) 1.5 moles
(d) Mass of americium = 324 g

Free Response Question (page 162)

$MgCO_3(s) \longrightarrow MgO(s) + CO_2(g)$
Relative mass of magnesium carbonate = 24 + 12 + 3(16) = 84
Number of moles of magnesium carbonate = 21/84 = 0.25 mol
Number of moles of magnesium oxide formed = 0.25 mol
Mass of magnesium oxide formed = 0.25 x 24 = 10 g

Extension (page 123)

Using volumes, molar volume of gases and number of moles in Section 9.3, the spreadsheet formulae for the cells are:

Cell B3: '=A3/24'
Cell C3: '=B3*3'
Cell D3: '=C3*24'
Cell E3: '=B3*2'
Cell F3: '=E3*24'

Spreadsheet printout:

	A	B	C	D	E	F
1	volume N_2	moles N_2	moles H_2	volume H_2	moles NH_3	volume NH_3
2	dm^3	mol	mol	dm^3	mol	dm^3
3	24	1	3	72	2	48
4	40	1.6666667	5	120	3.3333333	80
5	60	2.5	7.5	180	5	120
6	125	5.2083333	15.625	375	10.416667	250
7						

Note: The calculations of the volumes of hydrogen and ammonia can be carried out directly using the balanced equation and Avogadro's Law. Thus, the volume of $H_2 = 3 \times$ volume of N_2 and volume of $NH_3 = 2 \times$ volume N_2.

Additional Teaching Material

Additional Exercise 1: Obtaining Equations from Experiments

1. 10 cm^3 of hydrogen fluoride gas (HF) react with 5 cm^3 of dinitrogen difluoride gas (N_2F_2) to form 10 cm^3 of a single gas. All volumes are measured at r.t.p.

(a) What is the whole number ratio of the numbers of moles of reactants and products?

(b) From your answer in (a), deduce the formula of the product.

(c) Write the balanced equation for the reaction.

2. When hydrogen gas is passed over heated copper(II) oxide, the oxide changes into copper. In an experiment, the mass of copper(II) oxide was 12.0 g and the mass of copper formed was 9.6 g.

(a) Write the chemical formula for copper(II) oxide.

(b) Calculate the number of moles of copper(II) oxide reacting and the number of moles of copper produced.

(c) Use the figures in (b) to write the balanced equation for this reaction.

3. 20.0 cm^3 of chlorine gas reacted with exactly with 10.0 cm^3 of oxygen to give 20.0 cm^3 of a gaseous oxide of chlorine. All the volumes are measured at r.t.p.

(a) From the data, work out the formula of the oxide of chlorine (let its formula be Cl_xO_y).

(b) Use the formula in **(a)** to write the balanced equation for this reaction.

4. A solid hydrocarbon **X** has a relative molecular mass 128. Exactly 64 g of **X** burnt completely in oxygen to form 220 g of carbon dioxide and 36 g of water. An equation for the reaction can be represented as:

$$C_xH_y + \text{oxygen} \longrightarrow 10CO_2 + 4H_2O$$

(a) Calculate the number of moles of **X**, carbon dioxide and water.

(b) What is the whole number ratio of moles of C_xH_y, CO_2 and H_2O in the reaction?

____

(c) What are the values for x and y in C_xH_y and hence, what is the chemical formula for the hydrocarbon?

____

(d) Write the balanced equation for the reaction.

____

Additional Teaching Material

Additional Exercise 2: Calculations Using a Spreadsheet

Objective

- To calculate amounts of reactants and products from a chemical reaction

Strategy

- Using a spreadsheet

Magnesium metal reacts with dilute hydrochloric acid to produce magnesium chloride and hydrogen. The equation for the reaction is shown below.

$$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$

Starting with a known mass of Mg in grams, a spreadsheet is now constructed to calculate:

- the mass of HCl that reacted,
- the mass of $MgCl_2$ produced, and
- the mass of H_2 produced.

The procedure for doing this is given below.

The equation shows that 1 mole of Mg reacts with 2 moles of HCl to produce 1 mole of $MgCl_2$ and 1 mole of H_2.

		Procedure ↓
(i)	Mass of Mg = 24 g	Enter [24] in cell A3.
	Number of moles of Mg = Mass of Mg/24	Enter [=A3/24] in cell B3.
(ii)	Number of moles of HCl = 2 × Number of moles of Mg	
	Mass of HCl = Number of moles × 36.5	Enter [=2*B3*36.5] in cell C3.
(iii)	Number of moles of $MgCl_2$ = Number of moles of Mg	
	Mass of $MgCl_2$ = Number of moles × 95	Enter [=B3*95] in cell D3.
(iv)	Number of moles of H_2 = Number of moles of Mg	
	Volume of H_2 = Number of moles × 24	Enter [=B3*24] in cell E3.

(Relative atomic mass of Mg = 24; relative molecular mass of HCl = 36.5; relative molecular mass of $MgCl_2$ = 95)

Now enter the data into the spreadsheet as shown on the next page. Note that you only enter what is inside the square brackets into row 3.

Enter **48**, **100**, **300** and **5** grams into column A so that you can work out what amounts of HCl would be used and what amounts of $MgCl_2$ and H_2 would be produced if these masses of Mg reacted.

	A	B	C	D	E
1	Mass of Mg	Moles of Mg	Mass of HC*l*	Mass of MgCl_2	Volume of H_2
2	grams	moles	grams	grams	dm^3
3	24	= A3/24	= 2*B3*36.5	= B3*95	= B3*24
4	48				
5	100				
6	300				
7	5				

If you have everything correctly placed, row 3 should read as follows.

3	24	1	73	95	24

Now hold down the **left** mouse key at the plus sign of cell B3 (at the bottom right corner) and then drag down to cell B7. Repeat this step for cells C3, D3 and E3.

Questions

Use the spreadsheet you created to answer the following questions.

(a) What is the mass of HC*l* that reacts with 48 g of Mg? Mass = ____________

(b) What is the mass of MgCl_2 obtained from 100 g of Mg? Mass = ____________

(c) What is the volume of H_2 obtained from 100 g of Mg? Volume = ____________

(d) What is the mass of MgCl_2 obtained from 300 g of Mg? Mass = ____________

(e) What is the volume of H_2 obtained from 5 g of Mg? Volume = ____________

Additional Teaching Material

Additional Exercise 3: Calculations with a Spreadsheet for the Extraction of Iron

A spreadsheet can also be used to calculate the amounts of substances from a chemical reaction. Consider the equation for the extraction of iron from iron(III) oxide:

$$Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(l) + 3CO_2(g)$$

Starting with different masses of iron(III) oxide, the masses of the iron produced can be calculated using a spreadsheet.

1. Enter the headings as shown in rows 1 and 2 of the spreadsheet below. Then enter the masses of iron(III) oxide (16 g, 40 g, 85 g, 196 g and 2600 g) into Column A of your spreadsheet.

2. Calculate the number of moles of 16 g of iron(III) oxide. To do this, enter the spreadsheet formula '=A3/160' in cell B3. (160 is the relative formula mass of Fe_2O_3.)

3. From the chemical equation, the number of moles of Fe = 2 × number of moles of Fe_2O_3. Enter the spreadsheet formula '=2*B3' in cell C3.

4. Calculate the mass of Fe produced using the formula: mass = no. of moles × relative atomic mass (A_r of Fe = 56). To do this, enter the spreadsheet formula '=C3*56' in cell D3.

5. Hold down the left mouse key to highlight the grey area and then click 'Edit' and go to 'Fill'. Click on 'Down'.

Note: Steps 3 and 4 could be combined to calculate the mass of iron directly. To do this, place 'mass Fe' in column C and enter the spreadsheet formula '=2*B3*56' in cell C3.

	A	B	C	D
1	mass Fe_2O_3	moles Fe_2O_3	moles Fe	mass Fe
2	grams	mol	mol	grams
3	16	=A3/160	=2*B3	=C3*56
4	40			
5	85			
6	196			
7	2600			

Answers

Additional Exercise 1:

1. (a) $HF : N_2F_2$: product = 2 : 1 : 1

   (b) NHF_2

   (c) $2HF(g) + N_2F_2(g) \longrightarrow 2NHF_2(g)$

2. (a) CuO

   (b) Moles of CuO = number of moles of Cu = 0.15 mol

   (c) $CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(l)$

3. (a) From the data, 1 mole of Cl_2 + 0.5 mole of $O_2 \longrightarrow$ 1 mole of Cl_xO_y
   Therefore, the formula is Cl_2O

   (b) $2Cl_2(g) + O_2(g) \longrightarrow 2Cl_2O(g)$

4. (a) 64 g of **X** = 64/128 = 0.5 mol
   220 g of carbon dioxide = 220/44 = 5 mol
   36 g of water = 36/18 = 2 mol

   (b) 1 : 10 : 4

   (c) $x = 10$, $y = 8$. $C_{10}H_8$

   (d) $C_{10}H_8(s) + 12O_2(g) \longrightarrow 10CO_2(g) + 4H_2O(l)$

Additional Exercise 2:

	A	B	C	D	E
1	Mass of Mg	Moles of Mg	Mass of HC*l*	Mass of $MgCl_2$	Volume of H_2
2	grams	moles	grams	grams	dm^3
3	24	1	73	95	24
4	48	2	146	190	48
5	100	4.1666667	304.16667	395.833333	100
6	300	12.5	912.5	1187.5	300
7	5	0.208333	15.208333	19.7916667	5

Questions

(a) 146 g (b) 395.8 g (c) 100 dm^3 (d) 1187.5 g (e) 5.0 dm^3

Additional Exercise 3:

	A	B	C	D
1	mass Fe_2O_3	moles Fe_2O_3	moles Fe	mass Fe
2	grams	mol	mol	grams
3	16	0.1	0.2	11.2
4	40	0.25	0.5	28
5	85	0.53125	1.0625	59.5
6	196	1.225	2.45	137.2
7	2600	16.25	32.5	1820