Chem Factsheet

September 2000 Number 03

Moles and Equations

To succeed with this topic, you need to:
- ensure you understand the work on 'moles' from Factsheet No. 2 - Moles and Formulae
- **learn thoroughly** the valencies of the commonest cations and anions
- **learn thoroughly** those common formulae you are expected to know
- practice writing formulae and balancing equations because unless these are correct your calculations will always give the wrong answers!

After working through this Factsheet, you will understand:
- how to write chemical formulae
- putting formulae together in an equation to describe a chemical reaction
- balancing chemical equations
- calculating quantities from balanced equations
- using molar volumes of gas in equations
- writing ionic equations

Exam Hint: *Writing formulae and balanced chemical equations is central to all questions at AS level. The time spent working on these will repay you in terms of marks and grades.*

Writing chemical formulae

Chemical formulae fall into three main categories:

a) Formulae which must be **learnt**:

Water	H_2O	Sulphuric acid	H_2SO_4
Oxygen	O_2	Hydrochloric acid	HCl
Nitrogen	N_2	Nitric acid	HNO_3
Ammonia	NH_3	Phosphoric acid	H_3PO_4
Ozone	O_3	Chlorine	Cl_2
Argon	Ar	Bromine	Br_2
Neon	Ne	Iodine	I_2
Hydrogen	H_2		

N.B. Use the Periodic Table whenever you can for elements eg. Mg, Fe, Na, etc., but notice those elements in the list above are **diatomic** (two atoms in a molecule) - O_2, H_2, and the Halogens (Cl_2, Br_2, I_2).
Noble gases (Ne, Ar, etc.) are **monatomic**.

b) Formulae that can be worked out from their **names** alone:

The list of terms used is shown below

mono=1	penta=5	octa=8
di=2	hexa=6	nona=9
tri=3	septa=7	deca=10
tetra=4		

Carbon dioxide	CO_2	Carbon monoxide	CO
Sulphur trioxide	SO_3	Phosphorus pentachloride	PCl_5
Dinitrogen trichloride	N_2Cl_3	Sulphur dioxide	SO_2

NB: Hydrocarbons have a different system of naming
eg methane, CH_4 ethane, C_2H_6
This is covered in the Factsheet No. 13.

c) Formulae that can be worked out from the charge on their **ions**:

These compounds usually contain metals (the cations) and non-metals (the anions). You are expected to know these cations and anions; some have to be learnt, but you can use the Periodic Table to help you for the ions of elemental atoms - the box below reminds you how to do this.

Using the Periodic Table to help you find the charges on ions
- Group 1 form ions with charge +1
- Group 2 form ions with charge + 2 (but beryllium compounds may not be ionic)
- Group 6 form ions with charge -2
- Group 7 form ions with charge -1

The table below contains the commonest ions at AS level, but more are used as you progress through the course. You must learn these - and you will find later work much easier if you do it now, rather than waiting until the exam.

Table 1. Cations and Anions for AS-level

CATIONS		ANIONS	
Ions that can be worked out from Periodic Table rules above		**Ions that can be worked out from Periodic Table rules above**	
Name	Formula	Name	Formula
lithium	Li^+	chloride	Cl^-
sodium	Na^+	bromide	Br^-
potassium	K^+	iodide	I^-
magnesium	Mg^{2+}	oxide	O^{2-}
calcium	Ca^{2+}	sulphide	S^{2-}
strontium	Sr^{2+}		
barium	Ba^{2+}	**Other Anions**	
		Name	Formula
Other Cations		hydroxide	OH^-
Name	Formula	nitrate (V)	NO_3^-
hydrogen	H^+	nitrate (III)	NO_2^-
zinc	Zn^{2+}	cyanide	CN^-
aluminium	Al^{3+}	hydrogencarbonate	HCO_3^-
silver	Ag^+	hydrogensulphate	HSO_4^-
cobalt	Co^{2+}	carbonate	CO_3^{2-}
* copper	Cu^+/Cu^{2+}	sulphate (IV)	SO_3^{2-}
* iron	Fe^{2+}/Fe^{3+}	sulphate (VI)	SO_4^{2-}
* lead	Pb^{2+}/Pb^{4+}	phosphate	PO_4^{3-}
* manganese	Mn^{2+}/Mn^{4+}	manganate (VII)	MnO_4^-
ammonium	NH_4^+		

* = elements with more than one valency
For these, **Roman Numerals** are used to show which valency is being used - eg copper (II) hydroxide contains Cu^{2+}, whilst copper (I) oxide contains Cu^+
For non-metals, this refers to the **oxidation state** (see Factsheet 11) of the non-metal involved - eg sulphate (VI) contains sulphur in the +6 oxidation state

Exam Hint: *If a question uses compound names with roman numerals - eg sodium sulphate (IV) - make sure you take note of them! Many students lose marks through assuming sodium sulphate (IV) is Na_2SO_4, rather than Na_2SO_3*

Using Ions to Find Formulae
The method used is the "cross-over rule" as shown below:

eg 1: Magnesium chloride

Mg^{2+} Cl^-
1 2
$\Rightarrow MgCl_2$

eg 2: Sodium oxide

Na^+ O^{2-}
2 1
$\Rightarrow Na_2O$

eg 3: Magnesium nitrate (V)

Mg^{2+} NO_3^-
1 2
$\Rightarrow Mg\ (NO_3)_2$

> ***Tip:*** *If the formula involves a poly-atomic ion (i.e. an ion containing more than one atom), you will need to put* ***brackets*** *round it in any formula that involves more than one of that ion.*
>
> *In the above example, you could not write* $MgNO_{32}$ *- since this would not show that we had two lots of nitrogen.*
>
> *It is also incorrect to "multiply out the brackets" and write* MgN_2O_6*, since we need to show that we have two lots of the nitrate (V) ion* NO_3^-*, not just that we have 2 nitrogen atoms and 6 oxygen atoms.*

In some cases, you can "cancel down":

eg 4: Zinc Sulphate

Zn^{2+} O^{2-}
2 2
$\Rightarrow ZnO$

i.e. 2:2 'cancels down' to 1:1,

Some compounds do have formulae that do not cancel down - for example hydrogen peroxide (H_2O_2) and ethane (C_2H_6), but these are either covalent (like ethane) or contain special ions (like the peroxide ion O_2^{2-})

You need to be happy with working out formulae before going onto the rest of this Factsheet. Questions 1 and 2 at the end will give you practice at this.

Writing and Balancing Chemical Equations

(a) Balancing chemical equations
A 'balanced' chemical equation is one which has the same number of atoms of each element on both sides of the arrow, $\rightarrow$.

The rule for balancing is very simple:

> *You can change the numbers in* ***front*** *of formulae but you* ***never*** *change the formulae themselves*

When you are balancing, you should not expect to get it all balanced in one step! You may need to change numbers as you go along. The examples below show how to go from one side to the other to balance the equation.

eg 1. $Mg + O_2 \rightarrow MgO$
$1\times Mg \quad 2\times O \rightarrow 1\times Mg\ 1\times O$

We have different numbers of oxygens on each side. We balance them by having 2 lots of MgO:

$Mg + O_2 \rightarrow 2\ MgO$
$1\times Mg \quad 2\times O \rightarrow 2\times Mg\ 2\times O$

Now the magnesiums are unbalanced! So we balance them by having two lots of Mg

$2Mg + O_2 \rightarrow 2MgO$
$2\times Mg \quad 2\times O \rightarrow 2\times Mg\ 2\times O$

Now the equation is balanced, because we have the same number of each type of atom on each side

eg 2. $Fe + O_2 \rightarrow Fe_2O_3$
$1\times Fe \quad 2\times O \rightarrow 2\times Fe\ 3\times O$

We notice with the oxygens that we have 2 on one side and 3 on the other. We won't manage to balance them by multiplying just one side by something - there's no whole number we can multiply 2 by to get 3. So we have 3 lots of O_2 and 2 lots of Fe_2O_3 - so we end with 6 oxygens on each side (this is a bit like the cross-over rule)

$Fe + 3\ O_2 \rightarrow 2\ Fe_2O_3$
$1\times Fe \quad 6\times O \rightarrow 4\times Fe\ 6\times O$

The irons still are not balanced:
$4Fe + 3\ O_2 \rightarrow 2\ Fe_2O_3$
$4\times Fe \quad 6\times O \rightarrow 4\times Fe\ 6\times O$

Question 3 at the end provides further practise with balancing equations.

(b) Writing balanced chemical equations from word equations

The method is to replace the names with their formulae and then balance them.

eg 1. Calcium carbonate + hydrochloric acid → calcium chloride + water + carbon dioxide

Formulae:
$CaCO_3 + HCl \rightarrow CaCl_2 + H_2O + CO_2$

Balance:
$CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$

eg 2. methane + oxygen → carbon dioxide + water

Formulae:
$CH_4 + O_2 \rightarrow CO_2 + H_2O$

Balance:
$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$

(c) Using the description to find the word equation

The process is:

description of reaction → word equation → balanced chemical equation

Sometimes the description gives you all the reactants and products, and in other cases you have to apply your knowledge of chemical reactions to find the products. Both types are shown in the following examples:

eg 1. Calcium hydroxide is the only product when calcium oxide dissolves in water.

Calcium oxide	+	water	→	calcium hydroxide
CaO	+	H_2O	→	$Ca(OH)_2$

eg 2. Sulphuric acid reacts with potassium hydroxide

Acid	+	Alkali	→	salt	+	water
Sulphuric acid	+	potassium hydroxide	→	potassium sulphate	+	water
H_2SO_4	+	KOH	→	K_2SO_4	+	H_2O
H_2SO_4	+	$2KOH$	→	K_2SO_4	+	$2H_2O$

Question 4 at the end provides further practise with writing equations.

Calculation work based on equations

All calculation work must be based on **balanced chemical equations** - ones that have the same number of atoms of each type on each side of the arrow. Any calculation based on an unbalanced equation will automatically be wrong!

Calculation work also requires you to use **moles**. You need to ensure you are happy converting between masses and moles (see Factsheet 2 - Moles and Formulae) before you go any further.

e.g. $2Mg + O_2 \rightarrow 2\,MgO$

The numbers in front of the formulae tell us the **mole ratio** in the reaction - for every 2 moles of magnesium, we need 1 mole of oxygen, and we will produce 2 moles of magnesium oxide.

So if you wanted to react 4 moles of magnesium, you'd need 2 moles of oxygen and you would get 4 moles of magnesium oxide. Similarly, if you reacted 1 mole of magnesium, you'd need 0.5 moles of oxygen and you'd get 1 mole of magnesium oxide.

> ***Tip:*** *The equation can never tell you how much of a substance is actually reacting - that depends on how much of the chemicals you decide to use! It only tells you the* ***ratios***

We can use these mole ratios to find out **mass ratios** - which are what we need to work with to find masses. To find these, we need to use the equation **mass = moles × M_r**. So for the equation above, we have:

	$2Mg$	+	O_2	→	$2\,MgO$
Mole ratio	2		1		2
Multiply by M_r:	2×24		1×32		2×40
Mass ratio	48		32		80

Note that the masses balance - the total is 80 on both sides of the equation.

In all mass calculations based on equations, you must always follow these steps:

1. Write a balanced equation
2. Work out the mass ratio
3. Use the mass ratio, together with the information given in the question, to find the unknown masses.

The following examples illustrate how to do this:

eg 1. How much magnesium oxide will be made by burning 12g of magnesium?

	$2Mg$	+ O_2	→	$2MgO$
Mass ratio	48	32		80
Actual mass	12			?

To work out the required mass of magnesium oxide, we use the fact that 48:80 and 12:? are in the same ratio. One easy way of dealing with ratios is using the "cross method", shown in the box below

> **Working with ratios using the "cross method"**
> This method involves 3 easy steps:
> 1. Write down the two ratios underneath each other, putting in a ? for the number you don't know.
> 2. Draw a cross
> 3. Multiply the two joined numbers and divide by the other one.
>
> Using the example above, our two ratios are:
>
> 48 : 80
> 12 : ?
>
> So we do $12 \times 80 \div 48 =$ **20**
> If you are not happy working with ratios, see Factsheet 14: Maths for Chemists 1

So mass of MgO = 20g

eg 2. What mass of calcium carbonate is needed to make 0.12g of calcium oxide? (M_r values: Ca = 40, C = 12, O = 16)

	$CaCO_3$	→	CaO	+	CO_2
	1 mole	→	1 mole	+	1 mole
Mass ratios	100	→	56	+	44
Actual mass	?		0.12		

So our ratios are 100 : 56 and ? : 0.12
So ? = $100 \times 0.12 \div 48$ = **0.25g**

Gas Molar Volumes

Up to this point **state symbols** have not been used in any of the equations:

(s)	=	solid
(l)	=	liquid
(g)	=	gas
(aq)	=	solution (i.e. dissolved in water ≡ 'aqueous')

They are important and in effect complete any balanced chemical equation. They are used in examination questions, and are useful because you may need to specify that something is precipitated (so it will have an (s) not an (aq)), and, most importantly in this section, because some formulae - and hence some methods - only work for gases! The key fact is:

> *1 mole of any gas has a volume of 24 dm^3 (24000cm^3) at room temperature and pressure (rtp) - which is 1 atmosphere and 25°C*
>
> *At standard temperature and pressure (stp) - which is 1 atmosphere and 0°C - the volume of 1 mole of any gas is 22.4dm^3 (22400cm^3)*

There are two types of questions that use this fact:

1. Equations involving **only gas volumes**

The method here relies on the fact that **mole ratio = gas volume ratio**
So the steps are:

1 Write a balanced equation
2 Write down the mole ratio
3 Use the mole ratio, together with the gas volume information given in the question, to find the unknown volume.

eg 1. What volume of SO_3 would be made from 200cm^3 O_2 reacting with SO_2?

	$2SO_2$ (g)	+	O_2 (g)	➔	$2SO_3$ (g)
mole ratio	2		1		2
volumes			200		?

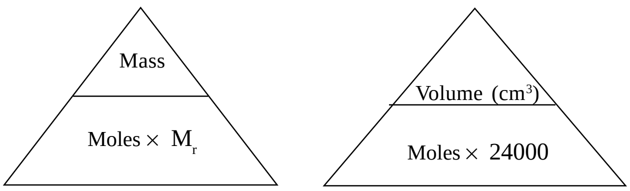
So 1: 2 and 200 : ? are in the same ratio.
So ? = 200 × 2 ÷ 1 = **400cm³**

2. Equations involving **gas volumes** and **masses**.
Here, we cannot rely on simple ratio methods. We need to convert between masses/volumes and moles, then work with moles and the mole ratio.
NB: This approach will also work with all the previous types of calculation, so if you'd rather remember just one method, use this one!
The procedure is:

1 Write a balanced equation
2 Write down the mole ratio
3 For the substance you have information about, work out how many moles of it there are
4 Use the mole ratio to find out how many moles there are of the substance you're asked about
5 Find out the mass or volume you're asked for using the correct moles equation

You may find the "triangles" below helpful:

Triangles for moles formulae



Cover up the thing in the triangle you want to find. Then, what you can see tells you the calculation to do. For example, if you want to find moles, cover it up and you are left with mass/M_r

NB: Convert litres (= dm³) to cm³ by multiplying by 1000.
Convert cm³ to litres by dividing by 1000

	$CaCO_3$(s)	➔	CaO(s)	+	CO_2(g)
Mole ratio	1	➔	1	+	1

We know we have 20g of $CaCO_3$. So we work out how many moles this is:
$M_r(CaCO_3)$ = 40 + 12 + 48 = 100
moles of $CaCO_3$ = mass÷M_r
= 20÷ 100 = 0.2 moles
We're asked about CO_2. The mole ratio $CaCO_3$: CO_2 is 1:1.

So we have 0.2 moles of CO_2.

Now we must find the volume of CO_2.

Volume CO_2 = moles × 24000cm³ = **4800cm³**

eg 2. Iron reacts with oxygen to make iron (III) oxide. Calculate the mass of iron and the volume of oxygen required to produce 3 grammes of iron (III) oxide. (A_r values are Fe: 56 O:16)

	4Fe +	$3O_2$	➔ $2Fe_2O_3$
Mole ratio	4	3	2

We're told about iron (III) oxide, so find moles of it:
$M_r(Fe_2O_3)$ = 112 + 48 = 160
moles of Fe_2O_3 = 3 ÷ 160 = 0.01875
NB: We do **not** round at this stage - it leads to loss of accuracy.

We need to find the moles of iron and the moles of oxygen:
Moles of iron = 0.01875 × 4 ÷ 2 = 0.0375
Moles of oxygen = 0.01875 × 3 ÷ 2 = 0.028125
Now we need to find the mass of iron and volume of oxygen:
Mass Fe = moles × 56 = **2.1g**
Volume O_2 = moles × 24000cm³ = **675cm³**

Tip: If you are at all unsure what to multiply by and what to divide by when you are using ratios:
- *Use the "cross method"*
- *Double check that the substances with the larger number in the mole ratio has the higher the number of moles.*

Writing ionic equations for precipitation reactions

In precipitation reactions, the reactants are solutions, but one of the products is a solid. So state symbols are very important when writing equations for these reactions.

The key idea used here is that when an ionic substance is in solution, the ions seperate - so we can consider sodium chloride solution, for example, to consist of Na^+ (aq) and Cl^- (aq).

The worked example below shows how it works:

eg. Write the following equation in its ionic form.

$FeSO_4$ (aq) + 2NaOH(aq) ➔ Fe $(OH)_2$ (s) + Na_2SO_4 (aq)

Three of the substances are in solution (we know this from the (aq)), so we split them into ions:

Fe^{2+}(aq) + SO_4^{2-}(aq) + $2Na^+$(aq) +$2OH^-$(aq)➔ $Fe(OH)_2$(s) + $2Na^+$ (aq) + SO_4^{2-} (aq)

Note that some ions appear on both sides of the equation - they started off in solution and they stay in solution. These are called **spectator ions**. We must "cancel them out" to give the final **ionic equation:**

Fe^{2+}(aq) + ~~SO_4^{2-}~~(aq) + 2~~Na^+~~(aq) +$2OH^-$(aq)➔ $Fe(OH)_2$(s) + 2~~Na^+~~ (aq) + ~~SO_4^{2-}~~ (aq)

Fe^2 (aq) + 2 OH^- (aq) ➔ Fe $(OH)_2$ (s)

Practice Questions

1. Write the formulae for the following compounds:

a) Sodium iodide	b) Iron (III) nitrate (V)	c) Lithium sulphide
d) Copper (II) nitrate(V)	e) Magnesium carbonate	f)Dihydrogen monoxide
g) Silver (I) oxide	h) Calcium chloride	i) Barium sulphide
j)Potassium sulphate(VI)	k) Carbon disulphide	l) Zinc carbonate
m) Magnesium oxide	n) Sulphur dioxide	o) Sulphur dichloride
p) Aluminium hydroxide	q) Hydrochloric acid	r) Lead (IV) oxide
s) Hydrogen chloride	t) Hydrofluoric acid	u)Trioxygen (Ozone)
v) Hydrogen gas	w)Ammonium chloride	x) Argon gas
y) Dinitrogen monoxide	z)Aluminium fluoride	aa) Magnesium iodide
ab)Disulphur dichloride	ac) Copper (I) carbonate	ad)Carbon tetrafluoride

2. Write the names of the following compounds, as fully as possible:

a) SO_3	b) $Ca(OH)_2$	c) $FeCl_2$	d) ZnS	e) FeS
f) $ZnCO_3$	g) I_2	h) $Sr(HSO_4)_2$	i) Li_2S	j) H_2O_2
k) NO	l) $Fe(OH)_3$	m) $(NH_4)_2SO_4$	n) CuF	o) $NaHSO_4$
p) NaOH	q) P_2O_3	r) BeO	s) $Mg(CO_3)_2$	t) $CuSO_4$

3. Balance each of the following equations:

a)	___Na + O_2	➔	___Na_2O
b)	K_2CO_3 + ___HCl	➔	___KCl + CO_2 + H_2O
c)	Ba + ___H_2O	➔	$Ba(OH)_2$ + H_2
d)	C_4H_8 + ___O_2	➔	___CO_2 + ___H_2O
e)	$Al_2(SO_4)_3$ + ___NaOH	➔	___Al $(OH)_3$ + ___ Na_2SO_4
f)	$SrCO_3$ + ___ HNO_3	➔	Sr $(NO_3)_2$ + H_2O + CO_2
g)	___ Fe + ___ O_2	➔	Fe_3O_4
h)	___HNO_3 + ___Cu	➔	___$Cu(NO_3)_2$ + ___NO + ___H_2O
i)	___NaOH + H_3PO_4	➔	Na_3PO_4 + ___H_2O
j)	Na_2CO_3 + $CuCl_2$	➔	$CuCO_3$ + ___NaCl
k)	___$NaNO_3$	➔	___$NaNO_2$ + O_2
l)	___Fe + ___H_2O	➔	Fe_3O_4 + ___H_2
m)	___SO_2 + O_2	➔	___SO_3
n)	N_2 + ___H_2	➔	___NH_3

4. Write the balanced chemical equation for each of the following:
a) Ethane (C_2H_6) burns in oxygen to produce carbon dioxide and water vapour.
b) Silicon tetrachloride reacts in water to make silicon dioxide and hydrogen chloride.
c) Heating strontium carbonate to produce the metal oxide and carbon dioxide.
d) Hydrogen reacting with oxygen to produce water.
e) Magnesium burning in are to make the oxide.

5. Write the balanced chemical equation for each of the following reactions:
a) Nitric acid with copper carbonate.
b) Burning potassium in air
c) Adding sodium hydroxide solution to zinc chloride solution
d) Potassium oxide with sulphuric acid
e) Burning octane (C_8H_{18}) in air

6. a) What mass of iron (III) oxide would be made by reacting 50g of iron with oxygen?
b) What mass of sulphur needs to be burnt in oxygen to produce 5g of sulphur dioxide?
c) What mass of calcium oxide and carbon dioxide would be made by heating 2g of calcium carbonate?
d) What mass of hydrogen would be produced by adding 10g of calcium to water?
e) What mass of oxygen would need to be added to 0.5g of carbon to turn it all into carbon dioxide?

7. a) In the reaction N_2 (g) + 3 H_2 (g) ➔ 2 NH_3 (g)
i) What volume of hydrogen would react with 50cm^3 of nitrogen
ii) What volume of ammonia would be made?
b) 2 SO_2 (g) + O_2 (g) ➔ 2 SO_3 (g)
What volumes of SO_2 and O_2 would be needed to produce 25 cm$_3$ of SO_3?
c) CO_2 (g) + C(s) ➔ 2 CO (g)
What volume of carbon monoxide would be made from 100cm^3 of carbon dioxide?

8. a) The equation below shows the fermentation process:
$C_6H_{12}O_6$ (aq) ➔ 2 C_2H_5OH(l) + 2 CO_2 (g)
What volume of CO_2 would be made from 5g of $C_6H_{12}O_6$?
b) 2 $NaNO_3$ (s) ➔ 2 $NaNO_2$ (s) + O_2 (g)
What volume of hydrogen would be made by adding 5.8g magnesium to excess hydrochloric acid?

9.Write the ionic equation for each of the following:

a) NaOH (aq) + HCl (aq)	➔	NaCl (aq) + H_2O (l)
b) Mg (s) + H_2SO_4 (aq)	➔	$MgSO_4$ (aq) + H_2 (g)
c) $Al_2(SO_4)_3$ (aq) + 6 NaOH (aq)	➔	2Al $(OH)_3$ (s) + 3 Na_2SO_4 (aq)
d) Na_2CO3 (aq) + 2 HNO_3 (aq)	➔	2 $NaNO_3$ (aq) + H_2O (l) + CO_2 (g)
e) 2 $AgNO_3$ (aq) + $CuCl_2$ (aq)	➔	$Cu(NO_3)_2$ (aq) + 2 AgCl (s)

Answers

1.a) NaI	b) $Fe(NO_3)_2$	c)Li_2S	d) $Cu(NO_3)_2$	e) $MgCO_3$	f) H_2O
g) Ag_2O	h) $CaCl_2$	i) BaS	j)K_2SO_4	k) CS_2	l) $ZnCO_3$
m)MgO	n) SO_2	o) SCl_2	p) $Al(OH)_3$	q) HCl	r) PbO_2
s) HCl	t) HF	u) O_3	v) H_2	w) NH_4 Cl	x) Ar
y) N_2O	z)AlF_3	aa)MgI_2	ab) S_2Cl_2	ac) Cu_2CO_3	ad)CF_4

2.a)sulphur trioxide b) calcium hydroxide c) iron (II) chloride
d) zinc sulphide e) iron (II) sulphide f) zinc carbonate
g)iodine h) strontium hydrogen sulphate (VI)
i) lithium sulphide j) dihydrogen dioxide k) nitrogen monoxide
l) iron(III) hydroxide m)ammonium sulphate (VI)
n) copper (I) fluoride o) sodium hydrogen sulphate (VI)
p) sodium hydroxide q) diphosphorus trioxide
r) beryllium oxide s) magnesium carbonate t)copper (II) sulphate

3. a).4, 2	b) 2, 2	c) 2	d) 6, 4, 4	e) 6, 2, 3
f) 2	g) 3, 2	h) 8, 3, 3, 2, 4	i) 3, 3	j) 2
k) 2, 2	l) 3, 4, 4	m) 2, 2	n) 3, 2	

4.a)	2 C_2H_6 + 7 O_2	➔	4 CO_2 + 6 H_2O
b)	$SiCl_4$ + 2 H_2O	➔	SiO_2 + 4 HCl
c)	$SrCO_3$	➔	SrO + CO_2
d)	2 H_2 + O_2	➔	2 H_2O
e)	2 Mg + O_2	➔	2 MgO

5.a)	2 HNO_3 + $CuCO_3$	➔	Cu $(NO_3)_2$ + H_2O + CO_2
b)	4 K + O_2	➔	2 K_2O
c)	2 NaOH + $ZnCl_2$	➔	$Zn(OH)_2$ + 2 NaCl
d)	K_2O + H_2SO_4	➔	K_2SO_4 + H_2O
e)	2 C_8H_{18} + 25 O_2	➔	16 CO_2 + 18 H_2O

6. a) 71.52g b)2.496g c)1.12g, 0.88g d) 0.5g e) 0.73g

7. a) i) 150cm^3 ii)100cm^3 b)25cm^3, 12.5cm^3 c) 200cm^3

8. a) 1344cm^3 b) 5760cm^3

9. a)	H^+ + OH^-	➔	H_2O
b)	Mg + $2H^+$	➔	Mg^{2+} + H_2
c)	Al^{3+} + 3 OH^-	➔	Al $(OH)_3$
d)	CO_3^{2-} + 2 H^+	➔	H_2O + CO_2
e)	Ag^+ + Cl^-	➔	AgCl

Acknowledgements:
This Factsheet was researched and written by Sam Goodman & Kieron Heath
Curriculum Press, Unit 305B, The Big Peg, 120 Vyse Street, Birmingham, B18 6NF
 ISSN 1351-5136