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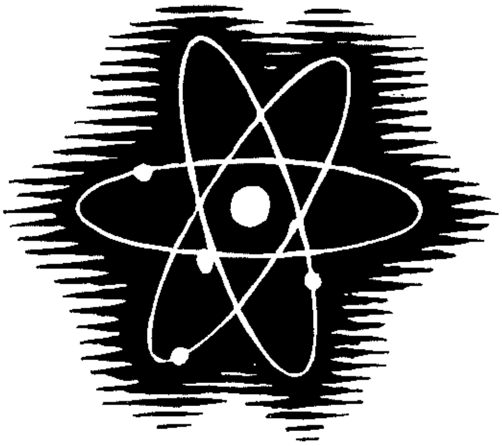
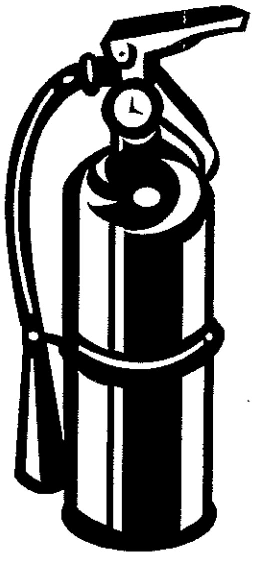
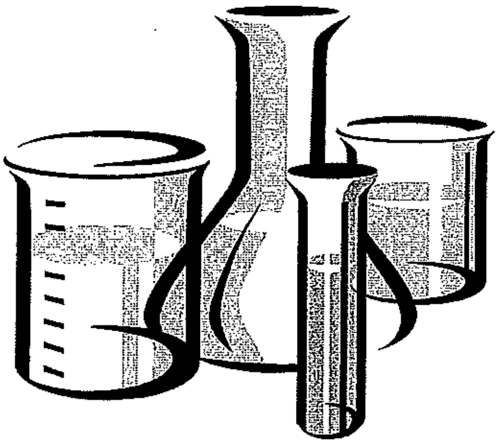




Chemical Changes and Structure Formulae/Reaction Quantities Nat 4/5



Chemical Reactions

A chemical reaction involves the formation of one or more new substances.

We can tell that a new substance has been formed by looking for one or more of the following changes :-

- Colour change
- Gas given off
- Solid appears when 2 solutions are mixed (**PRECIPITATION**)
- A smell is given off
- Heat/energy is taken in (**ENDOTHERMIC**)
- Heat/energy is given out (**EXOTHERMIC**)

The following experiments allow you to observe some of the changes mentioned above and hence determine whether a chemical reaction is taking place or not.

Teacher Demonstration

Chemical reactions happen around you everyday. Complete the table below to indicate how you know a chemical reaction has taken place.

Everyday Reaction	Change Observed
Coal burning	
A car rusting	
An egg frying	

Expt No	Breif Discription	Observation

Word Equations

A word equation can be used to describe a chemical reaction.

The equation places the substances at the start of the reaction, called the **REACTANTS**, on the left hand side and the new substance(s) formed, called the **PRODUCT(S)**, at the right hand side.

An arrow is used between them to indicate the direction and the fact that reactants have been "**changed into**" products.
NB. A = sign must never be used as by the definition of a chemical reaction a new substance has to be formed so it is **not equal** to the reactants.

Example

1. Magnesium powder burns in oxygen gas to produce a white solid called magnesium oxide.

Magnesium	+	Oxygen	→	Magnesium oxide
REACTANTS				**PRODUCT**

Word Equations-Examples

Write the word equation for the following reactions.

a) Methane reacting with oxygen to form carbon dioxide and water.

b) Nitric acid reacts with sodium hydroxide to give water and sodium nitrate.

c) Calcium reacts with water to produce calcium oxide and hydrogen.

d) Aluminium oxide and zinc are produced when aluminium and zinc oxide react together.

e) Calcium chloride is one of the products when calcium carbonate and hydrochloric acid react together, water and carbon dioxide are also produced.

Naming Compounds

A compound is formed when two or more elements **join together chemically**. ie. for a compound to be formed a chemical reaction must take place.

Most compounds can be named by applying the following rules:

1. The names of the elements are written from left to right as they appear on the periodic table. eg. sodium chloride and lithium fluoride.

2. If there are **only two elements** forming the compound then the ending of the second name is changed to '**IDE**'. For example, a compound of copper and sulphur only is called copper sulphide and a compound of sodium and oxygen only is called sodium oxide.

3. If the compound **contains oxygen and two other elements** the ending of the second name changes to '**ATE**'. For example, a compound of copper, sulphur and oxygen is called copper sulphate. (*Later you will be introduced to 'ITES' which also contain oxygen and two other elements*).

Elements present	Name of Compound
1.	magnesium oxide
2.	silver chloride
3. lead sulphur	
4. iron oxygen	
5.	copper carbonate
6.	zinc sulphate
7. iron sulphur oxygen	
8. calcium carbon oxygen	
9.	sodium nitrate

Rules Using sodium sulphide as an example

1. write down the symbol of the elements present Na S

2. write down the valencies

Na	1
S	2

3. cross over(swap) the valencies

Na	2
S	1

formula Na_2S

3. divide numbers with 'common' denominator.

Eg calcium carbide

Ca	2	4	2
C	4	2	1

formula Ca_2C

PRACTICE

Write the formula for the following:-

a) potassium oxide
b) strontium iodide
c) magnesium chloride
d) lithium nitride
e) aluminium fluoride
f) calcium oxide
g) lithium sulphide
h) calcium bromide
i) sodium chloride
j) aluminium oxide
k) magnesium nitride
l) caesium iodide
m) strontium carbide
n) aluminium nitride

EXCEPTIONS

A. There are some elements (usually the **TRANSITION METALS**) that can have more than one valency. The valency of these elements is given by a **ROMAN NUMERAL** after the element name.

I	one	II	two
III	three	IV	four

How to write a chemical formula

1. Elements

The formula of an element is simply given by it's symbol on the periodic table.

Eg.	Potassium	K
	Sodium	Na
	Sulphur	S
	Magnesium	Mg

EXCEPTION The formula of the elements that exist as DIATOMIC MOLECULES must show that 2 atoms are in the molecule.

Eg.	Hydrogen	H_2
	Nitrogen	N_2
	Oxygen	O_2
	Fluorine	F_2
	Chlorine	Cl_2
	Bromine	Br_2
	Iodine	I_2

2. Compounds

The formula of a compound can be worked out using the **VALENCY** of each atom present in the compound. (remember –IDE ending tells you that **2 elements are present**, -ATE or -ITE tells you that OXYGEN is present and at least two other elements.)

The VALENCY of an element is related to the number of half filled clouds it has or the number of bonds it is able to form when combining with another element. Valency is therefore related to the position in the periodic table.

GROUP	1	2	3	4	5	6	7	8
VALENCY	1	2	3	4	3	2	1	0

Apart from being given the valency by the Roman Numeral these formula are worked out using RULES 1-4 as before.

PRACTICE

Write the formula for the following:-

a)	nickel(II)fluoride	b)	silver(I)chloride
c)	copper(I)sulphide	d)	cobalt(II)oxide
e)	silver(I)oxide	f)	copper(II)nitride
g)	lead(II)chloride	h)	manganese(IV)oxide
i)	iron(III)sulphide	j)	scandium(III)nitride
k)	tin(II)nitride	l)	vanadium(V)oxide

B. There are some groups whose valency is given by their **CHARGE**, which is shown on the top right hand side of the formula given in the data booklet as "**ions containing more than one kind of atom**" (page 4)

Apart from being given the valency by the **charge** these formula are worked out using RULES 1-4 as before.

PRACTICE

Write the formula for the following:-

a)	lithium sulphate	b)	sodium phosphate
c)	magnesium nitrate	d)	calcium carbonate
e)	potassium carbonate	f)	calcium phosphate
g)	magnesium hydroxide	h)	aluminium sulphate
i)	sodium nitrate	j)	potassium sulphate
k)	ammonium sulphate	l)	ammonium phosphate

C. Their are some molecules whose names are not worked out using the valency rules. Their names contain **PREFIXES** that will indicate the number of each atom present in a molecule.
NO CROSS OVER OR CANCELING DOWN NEEDS TO BE DONE.

Prefixes to be learned

Mono	= 1
Di	= 2
Tri	= 3
Tetra	= 4
Penta	= 5
Hexa	= 6

Examples: Phosphorus tribromide = PBr_3
Dinitrogen tetroxide = N_2O_4

PRACTICE

Write the formula for the following:-

a) carbon monoxide
b) carbon tetrabromide
c) nitrogen dioxide
d) xenon hexafluoride
e) phosphorus pentafluoride
f) dinitrogen trioxide

2. When naming a compound from a formula when they contain a transition metal with a variable valency you must be sure to check the valency of the other atom/ion present and work the rules back.

Examples: $NiCl_2$ - nickel(II)chloride
$CuSO_4$ - copper(II)sulphate
Cu_2SO_4 - copper(I)sulphate

PRACTICE

Write the name for the following:-

a) Ag_2S b) ScF_3 c) $Fe(NO_3)_3$

d) $CuCl_2$ e) $CoBr_2$ f) $Mn_3(PO_4)_4$

3. More Practice

Write the formula for the following:-

a) potassium sulphide
b) sodium sulphate
c) aluminium iodide
d) copper(I)carbonate
e) lithium carbonate
f) sulphur dioxide
g) tin(II)oxide
h) sodium nitride
i) phosphorus triiodide
j) iron(II)nitrate
k) potassium phosphate
l) magnesium oxide
m) sulphur trioxide
n) aluminium sulphide
o) nickel(II)bromide
p) dinitrogen pentoxide
q) magnesium carbide
r) silver(I)fluoride
s) sulphur hexafluoride
t) ammonium hydroxide
u) calcium phosphide
v) barium iodide
w) carbon tetrachloride
x) magnesium phosphate
y) strontium hydroxide
z) nitrogen monoxide
A) rubidium oxide
B) aluminium carbonate
C) calcium chloride
D) lead(IV)oxide
E) ammonium carbonate
F) iron(III)hydroxide
G) lead(IV)nitrate
H) vanadium(IV)sulphate

The significance of the formula

Since discrete molecules consist of definite numbers of atoms held together by covalent bonds, the formula of a covalent compound shows the actual number of atoms in each molecule.

e.g. H_2O shows **two** hydrogen atoms joined with **one** oxygen atom.

$C_6H_{12}O_2$ shows **six** carbon atoms , **twelve** hydrogen atoms and **six** oxygen atoms in a molecule.

On the other hand, because a covalent network structure consists of a giant lattice of covalently bonded atoms, the formula of a covalent network compound gives the simplest ratio of atoms of each element.

e.g. SiO_2 shows **twice** the number of Si atoms as O atoms

In a similar way, because a large number of positive and negative ions are held together by ionic bonds to form a giant lattice, the formula of an ionic compound only indicates the relative numbers of ions present.

e.g. NaCl shows **equal** numbers of Na^+ and Cl^- ions

K_2S shows **twice** as many K^+ ions as S^{2-} ions.

Using the Periodic Table (i)

The number of covalent bonds which an atom forms is equal to the number of "extra" electrons which an atom requires to reach the same electron arrangement as a noble gas. This is the same as the number of half-filled electron clouds

Complete the table below to show the number of bonds formed by atoms of the groups of elements.

	C	N	O	F	Ne
	Si	P	S	Cl	Ar
		As	Se	Br	Kr
Number of outer electrons					8
Number of extra electrons					0
Number of bonds formed					0

NOTE:

(i) The atom of hydrogen (one half-filled cloud) needs 1 electron to reach the same electron arrangement as an atom of helium. Hydrogen forms 1 bond.

(ii) It is easier for atoms of elements in Group 1 to 3 to lose electrons to reach a noble gas electron arrangement. This explains why atoms of metal elements do not form covalent bonds.

The Periodic Table can be used to find the formulae for covalent compounds from bonding diagrams.

Atoms of elements with 4 electrons in the outer energy level (carbon, silicon, etc.) form 4 bonds.
Atoms of elements with 5 electrons in the outer energy level (nitrogen, phosphorus, etc.) form 3 bonds.
Atoms of elements with 6 electrons in the outer energy level (oxygen, sulphur, etc.) form 2 bonds.
Atoms of elements with 7 electrons in the outer energy level (fluorine, chlorine, etc.) form 1 bond.

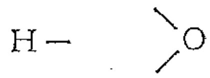
Using the Periodic Table (ii)
(continued)

Hydrogen oxide

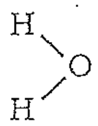
1. Use the Periodic Table to write symbols for the elements.

   H　　　O

2. Use the Periodic Table to put in the number of bonds which will be formed by each atom.



3. Find the number of atoms of each element required to complete all the bonds.



4. Write the formula based on the number of atoms of each element which are used:
   2 hydrogen atoms and 1 oxygen atom

   formula　　H_2O

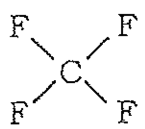
Carbon fluoride

1. Use the Periodic Table to write symbols for the elements.

   C　　　F

2. Use the Periodic Table to put in the number of bonds which will be formed by each atom.

3. Complete the bonding picture.



4. Write the formula based on the number of atoms of each element which are used:
   4 fluorine atoms and 1 carbon atom

   formula　　CF_4

Using the PeriodicTable (ii)

The charge on many ions can be worked out from the electron arrangements shown in the Periodic Table on page ..... of the data booklet.

Metal atoms lose electrons from the outer energy level to reach the stable electron arrangement of the 'nearest' noble gas. Since electrons have a negative charge, the loss of electrons gives metal ions a positive charge.
Non-metal atoms gain electrons in the outer energy level to reach the stable electron arrangement of the 'nearest' noble gas. Since electrons have a negative charge, the gain of electrons gives non-metal ions a negative charge.

	Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7
Number of outer electrons	1	2	3	4	5	6	7
Charge on ions	1+	2+	3+		3-	2-	1-

In an ionic compound, the charge on all positive ions must balance the charge on all negative ions.

Since the overall charge is neutral, the formula for an ionic compound can be worked out by finding the relative number of each ion required to make the overall charge zero.

Sodium chloride positive ion negative ion

Na^+ Cl^-

formula Na^+Cl^- or NaCl

Potassium oxide positive ion negative ion

K^+ O^{2-}

formula $(K^+)_2O^{2-}$ or K_2O

NOTE: if charges are included in the formula, the symbol and charge must be put in brackets when the balancing number is greater than 1.
$(K^+)_2$ means **two** K^+ ions; K_2^+ would represent **one** $(K_2)^+$ ion.

Using combining powers (valency)

The chemical formula for a compound can always be worked out by considering the bonding. There is, however, a shorter method which uses the combining powers (valency). This method works for both covalent and ionic compounds.

The combining power can be found from the Periodic Table.

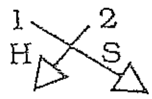
Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7
1	2	3	4	3	2	1

For metals which show variable charge the combining power corresponds to the charge on the ion,
e.g. in iron(II) oxide the combining power of the iron ion is 2;
in copper(I) oxide the combining power of the copper ion is 1.

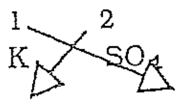
For group ions, the combining power corresponds to the charge on the ion,
e.g. in SO_4^{2-} the combining power of the ion is 2;
in NO_3^- the combining power of the ion is 1.

This method will always give the correct answer, but it does not show you why it is correct; use in emergency when all else fails!

Example 1:
hydrogen sulphide

Step 1	Write atoms and combining powers in this form	1 2 H S
Step 2	Exchange the combining powers	
Step 3	Ignore the number 1 to give the correct chemical formula	H_2S

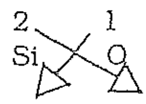
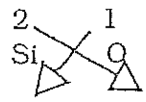
Example 2:
potassium sulphate

Step 1	As before	1 2 K SO_4
Step 2	As before	
Step 3	As before	K_2SO_4

Using combining powers (valency) (continued)

An extra step is sometimes necessary.

Example 3:
silicon oxide

Step 1	As before	4 2 Si O
Step 2	Cancel the numbers 2 and 4 to give 1 and 2	
Step 3	As step 2 before	
Step 4	As step 3 before	SiO_2

Using Roman numerals

The transition metals are a large group of elements in the Periodic Table which start after calcium (atomic number 20). The elements lie between the second and third columns.

Some of the transition metals can have ions with more than one charge. In these cases the charge is shown in Roman numerals after the name of the metal element.

e.g. in iron(II) oxide the charge of the iron is 2-positive (Fe^{2+})
in copper(I) oxide the charge of the copper is 1-positive (Cu^{+})

The formula for an ionic compound with a Roman numeral in the name can be found in the same way as for two-element ionic compounds.

Copper(I) oxide

	positive ion	negative ion
	Cu^{+}	O^{2-}
formula	$(Cu^{+})_2O^{2-}$	or Cu_2O

Iron(II) oxide

	positive ion	negative ion
	Fe^{2+}	O^{2-}
formula	$Fe^{2+}O^{2-}$	or FeO

Group ions

A number of ions consist of a group of atoms which tend to stay together during reactions. These are called **group ions.**
The charge is on the whole group and not on any particular atom.

e.g. the sulphate ion

the charge of the ion is 2-negative

$$SO_4^{2-}$$

formula for the ion

The formula and charge of some group ions can be found on page ..... of the data booklet.

Complete the following table by writing in the formula and charge of each of the group ions.

Group ion	Formula	Charge
carbonate	CO_3	2-negative
nitrate		
sulphate		
phosphate		
hydroxide		
sulphite		
ammonium		

The presence of a group ion can usually be recognised from the -ate or -ite name ending which indicates the presence of oxygen. The exceptions are the ammonium ion and the hydroxide ion. Apart from the ammonium ion, which has a positive charge like the metal ions, all the group ions have a negative charge.

The formula for an ionic compound with a group ion can be found in the same was as for two-element ionic compounds.

Sodium nitrate

	positive ion		negative ion		
	Na^+		(NO_3^-)		
formula	$Na^+(NO_3^-)$	or	$Na^+NO_3^-$	or	$NaNO_3$

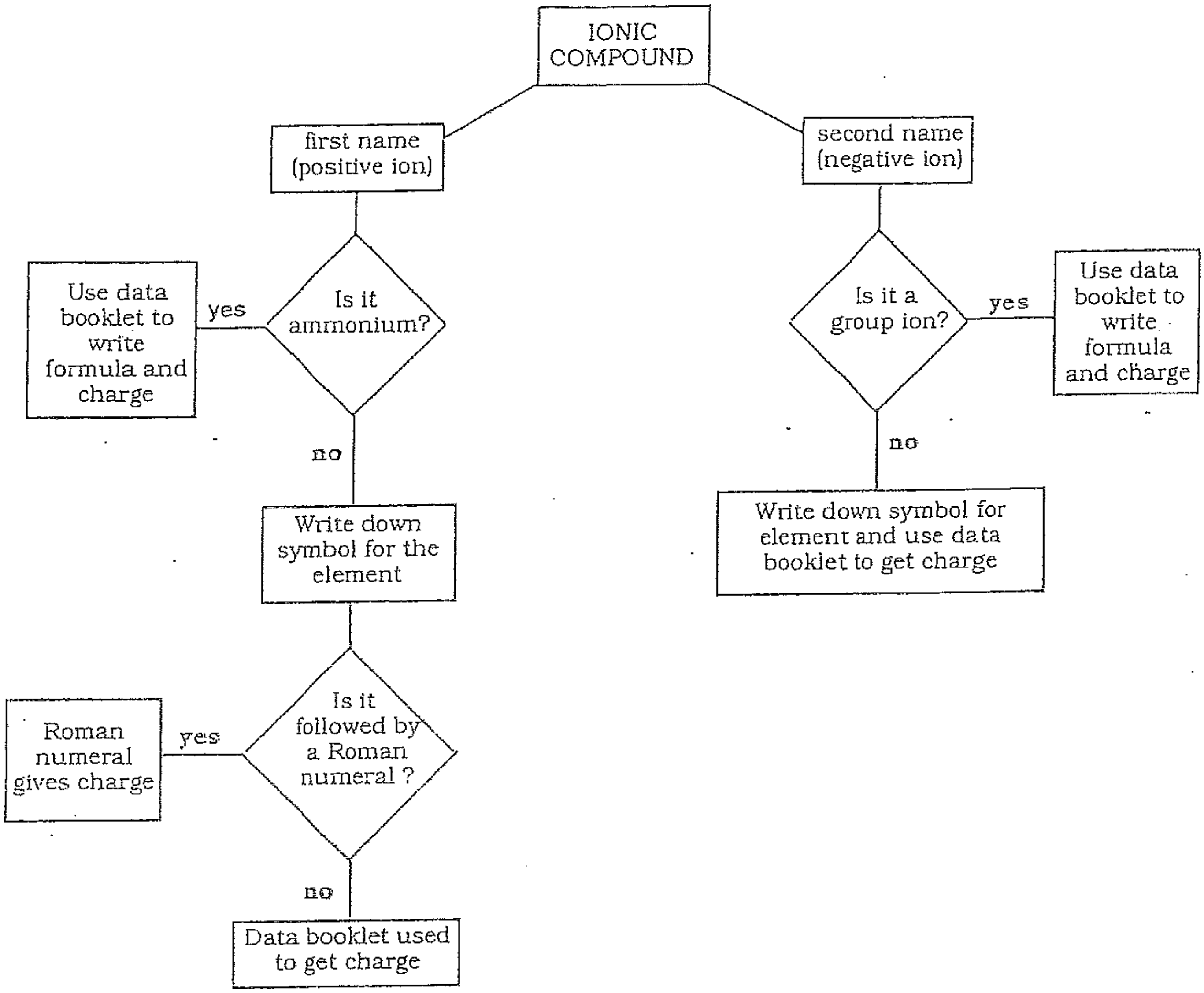
NOTE:
Always put the formula of the group ion in brackets. When the subscript numeral for the group is 1, as above, the brackets can be removed. **When the subscript numeral for the group is greater than 1, brackets are essential.**

Calcium nitrate

	positive ion		negative ion
	Ca^{2+}		(NO_3^-)
formula	$Ca^{2+}(NO_3^-)_2$	or	$Ca(NO_3)_2$

NOTE:
The formula for calcium nitrate is $Ca^{2+}(NO_3^-)_2$ and **not** $Ca^{2+}NO_3^-{}_2$. The formula has one calcium ion for every two nitrate ions. This gives a total of one calcium, two nitrogens and six oxygens.

Flow diagram summary for writing the formula for an ionic compound



The formula can be worked out by finding the relative number of each ion required to make the overall charge zero.

Word equations

In a chemical reaction, substances present at the start change to make new substances, e.g.

STARTING SUBSTANCE A and STARTING SUBSTANCE B
change to make
NEW SUBSTANCE C and NEW SUBSTANCE D

or

STARTING SUBSTANCE K	**breaks up to give**	NEW SUBSTANCE L + NEW SUBSTANCE M

or

STARTING SUBSTANCE X and STARTING SUBSTANCE Y
join up to make
NEW SUBSTANCE Z

The starting substances in chemical reactions are called the **reactants**; the new substances which are produced are called the **products**.

The chemical reaction can be written in a short-hand form, called a **word equation** e.g.

STARTING SUBSTANCE A + STARTING SUBSTANCE B
↓
NEW SUBSTANCE C + NEW SUBSTANCE D

or

STARTING SUBSTANCE K	→	NEW SUBSTANCE L + NEW SUBSTANCE M

or

STARTING SUBSTANCE X + STARTING SUBSTANCE Y
↓
NEW SUBSTANCE Z

In a word equation:
(i) the '+' sign means '**and**'
(ii) the ' → ' sign means '**changed into**'
(iii) the **reactants** come **in front** of the arrow; the **products** come **after** the arrow

Here is a real example of a chemical reaction.

The food in our bodies joins up with oxygen taken from the air to produce water vapour and carbon dioxide which we breathe out.

The word equation is:

food + oxygen → water vapour + carbon dioxide

the reactants — **the products**

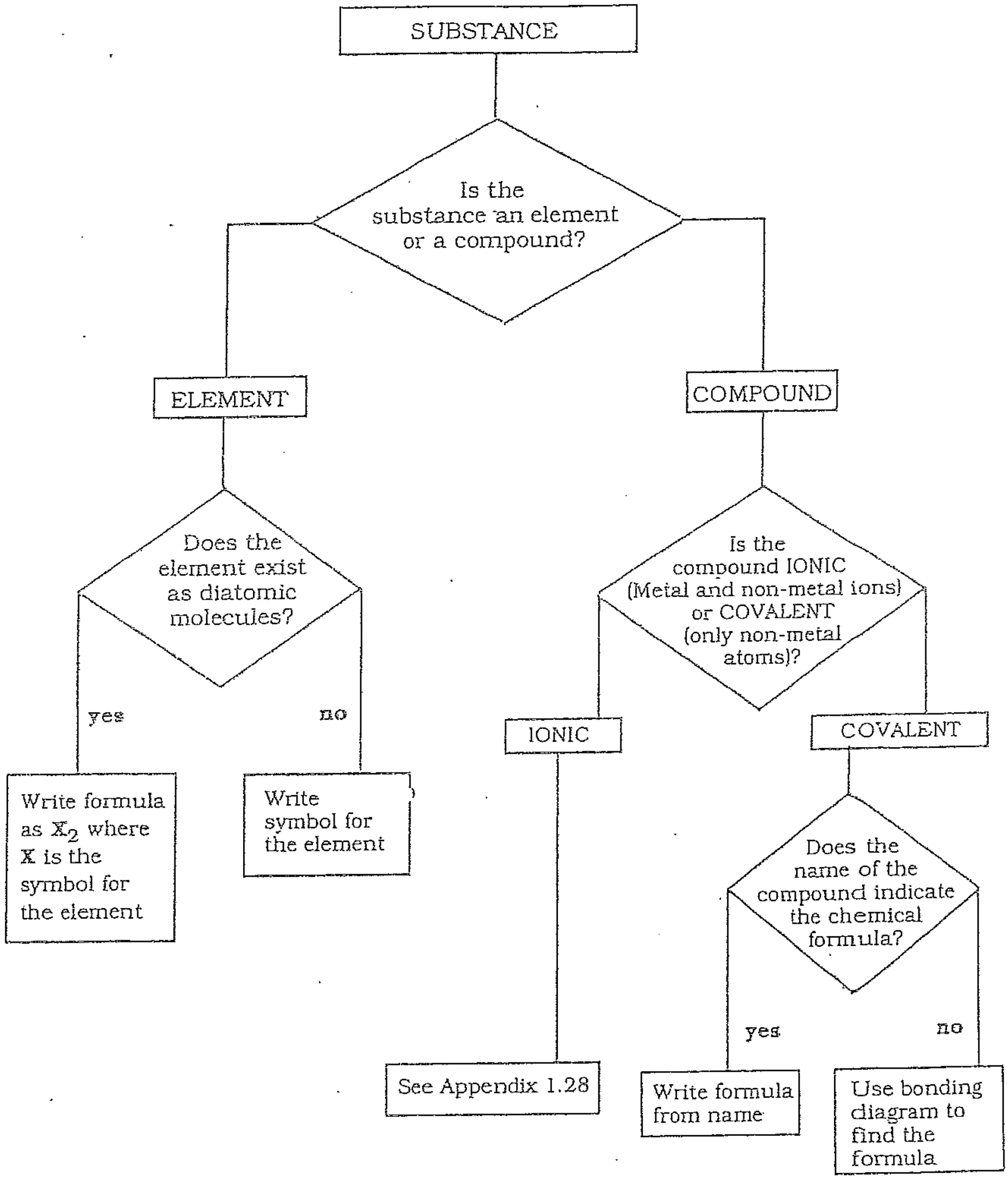
Equations using symbols and formulae

Equations are mainly written using symbols and formulae.

e.g. carbon reacting with oxygen to form carbon dioxide

$$C + O_2 \rightarrow CO_2$$

For elements and compounds, the following flow diagram may be of value when writing formulae. It should be followed for each substance in the reaction in turn.



The mole

One mole of any substance (abbreviate to mol) is defined as the relative formula mass in grams, i.e. the gram formula mass.

The relative formula mass of any substance is first calculated from the formula.
It his has no units.
To calculate the mass of one mole of the substance, simply express the relative formula mass in grams.

Example 1:
1 mol of sodium

formula	Na
relative formula mass	23
gram formula mass i.e. 1 mol	23 g

Example 2:
1 mol of chlorine

formula	Cl_2
relative formula mass	71
gram formula mass i.e. 1 mole	71 g

Example 3:
1 mol of carbon dioxide

formula	CO_2	
elements	C	O
number of each element	1	2
	12	16
relative atomic mass of each element	12	32
total relative mass of each element		
formula mass	44	
gram formula mass i.e. 1 mol	44 g	

Example 4:

1 mol of calcium carbonate

formula		$CaCO_3$	
elements	Ca	C	O
number of each element	1	1	3
relative atomic mass of each element	40	12	16
total relative mass of each element	40	12	48
formula mass		100	
gram formula mass i.e. 1 mol		**100 g**	

The mole - further examples

Example 1:
Calculate the mass of 2 mol of sodium chloride.

formula		NaCl	
elements		Na	Cl
number of each element		1	1
relative atomic mass of each element		23	35.5
relative formula mass		58.5	
1 mol	<---------------->	58.5 g	
1 mol	<---------------->	58.5 g	
2 mol		**117 g**	

The <----------------> symbol is used to show a simple proportion.

Example 2:
How many moles of water are there in a 36 g of water?

formula		H_2O	
elements		H	O
number of each element		2	1
relative atomic mass of each element		1	16
relative formula mass		18	
1 mol		18 g	
18 g	<---------------->	1 mol	
36 g	<---------------->	**2 mol**	

1. For a given Mass, **1 mole = formula mass in grams**

$$\text{No. moles} = \frac{\text{mass}}{\text{gram formula mass}} = \frac{m}{gfm}$$

So mass = No. moles x gfm

1a. Changing Mass to Moles

Example : Calculate the mass of 4 moles of carbon tetrachloride, CCl_4

C	1 x 12	=	12
Cl	4 x 35.5	=	142
So,	1 mole	=	154g

Therefore, 4 moles = 4 x 154 = 616g

1b. Changing Moles to Mass

Example : How many moles are contained in 150g of calcium carbonate $CaCO_3$?

Ca	1 x 40	=	40
C	1 x 12	=	12
O	4 x 16	=	64
So, 1 mole		=	100g

$$\text{No. moles} = \frac{\text{mass}}{gfm} = \frac{150}{100} = 1.5 \text{ mol}$$

The Mole

Substances can be 'equated' in chemistry if we express them as MOLES. A mole of any substance contains the same number of particles. This is important to know in a chemical reaction because it is particles that are reacting.

We can only quantify this by defining **a mole** as the **formula mass expressed in grams**.

For example:	1 mole of Cu	= 64g
	1 mole of H_2	= 2g
	1 mole of NaCl	= 23 + 35.5 = 58.5g
	1 mole of CO_2	= 12 + (2x16) = 44g

Worked Examples

A. Calculate the mass of 2.5 mol of copper.

The relative atomic mass of copper = 64

1 mol ---------→ 64 g

2.5 ----------→ $\frac{64 \times 2.5}{1}$

= **160g**

[note that since the answer has to be in grams we place that on the RHS of the calculation]

B. How many moles of sodium chloride are present in 11.7g of the compound?

Formula of sodium chloride is NaCl
The formula mass of NaCl = 23 + 35.5 = 58.5
So 1 mole of NaCl = 58.5g

Since we want our answer to come out as a number of moles we must place that on the RHS of the calculation.

58.5g -------→ 1 mol

11.7g -------→ $\frac{1 \times 11.7}{58.5}$

= **0.2 mol**

Worked Examples

A. What is the concentration of a solution containing 2.5 mol of substance dissolved in 5 litres of solution?

$$\text{Concentration} = \frac{\text{number of moles}}{\text{Volume}}$$

$$= \frac{2.5}{5}$$

$$= \mathbf{0.5\ mol/l}$$

B. How many moles of substance are present in $25cm^3$ of a 0.2 mol/l solution?

Moles = concentration x volume(in litres)

= 0.2 x 0.025

= **0.005 mol**

C. 2g of sodium hydroxide, NaOH, are dissolved in water to make a 0.4 mol/l solution. What volume is the solution?

In order to calculate the volume we need to know the concentration and the number of moles, but we are only told the concentration: first we must work out the number of moles of sodium hydroxide from it's mass:

Formula of sodium hydroxide is NaOH
The formula mass of NaOH = 23 + 16 + 1 = 40
So 1 mole of NaOH = 40g

40g -------------→ 1 mol
2g -------------→ $\frac{1 \times 2}{40}$

= **0.05 mol of NaOH**

[Now fit this value into the appropriate equation]

$$\text{Volume (in litres)} = \frac{\text{number of moles}}{\text{Concentration}}$$

$$= \frac{0.05}{0.4}$$

$$= \mathbf{0.125\ litres}\ (125cm^3)$$

D. What mass of sodium carbonate, Na_2CO_3, must be dissolved to make 0.25 litres of a 0.2 mol/l solution?

We have the information to calculate the number of moles of sodium carbonate required. This can be fitted into the appropriate equation directly:

$$\text{Moles} = \text{concentration} \times \text{volume (in litres)}$$

$$= 0.2 \times 0.25$$

$$= 0.05\ \text{mol}$$

However the question asks for the mass of sodium carbonate which this represents.

$$\text{Formula mass of } Na_2CO_3 = (2 \times 23) + 12 + (3 \times 16)$$

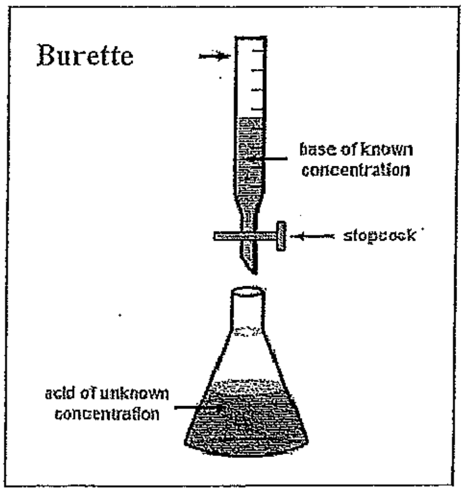
$$= 106g$$

1 mol ------------→ 106g

0.05 mol --------→ $\frac{106 \times 0.05}{1}$

$$= \mathbf{5.3g}$$

4. <u>Acid/ Alkali Titrations</u>



Example :

Sodium hydroxide and nitric acid react as follows :

$NaOH + HNO_3 \longrightarrow NaNO_3 + H_2O$

In a titration, 30 cm^3 of nitric acid neutralised 20 cm^3 of 0.2 mol l-1 sodium hydroxide.

Calculate the concentration of the nitric acid.

<u>Steps</u>

1. Write equation $NaOH + HNO_3 \longrightarrow NaNO_3 + H_2O$

2. Mole equation 1 mol 1 mol

0.004mol

3. Look at info given, the reactant for which there is both concentration and volume data then that is the place to start by calculating the number of moles present.

For NaOH, $n = CV$ $n = 0.2 \times 0.02$ $n = 0.004$ mol

4. From the mole equation above, one mole of alkali reacts with one mole of acid so the number of moles of HNO_3 which react exactly with the alkali is 0.004 mol.

5. From the number of moles calculated above and the volume of acid given in the question, the required concentration can be calculated.

So. For the acid $n = CV$, therefore $C = \frac{n}{V} = \frac{0.004}{0.03}$

$= 0.133 \text{ mol } l^{-1}$

Balanced chemical equations (i)

The word equation for the burning of natural gas is:

natural gas + oxygen → carbon dioxide + water

This equation can be written using the formula for each reactant and product:

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

In this reaction, carbon, hydrogen and oxygen, which form the reactants, also form the products. In a reaction, the elements which take part (either as an element or as part of a compound) also make up what is formed.

An equation using symbols and formulae gives more information than a word equation - it shows the elements involved and the way in which they are joined up in the reactant(s) and product(s).

1. *Make a model of a molecule of natural gas, CH_4 and oxygen O_2.*
2. *Break the bonds to form carbon, hydrogen and oxygen atoms.*
3. *Now form new bonds to try to make carbon dioxide, CO_2 and water H_2O.*

There are not enough oxygen atoms to form both CO_2 and H_2O. So another oxygen molecule, O_2 is needed.

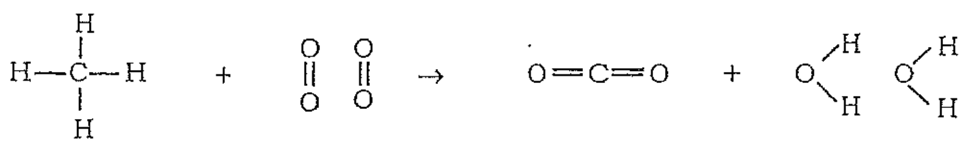
$$CH_4 + 2O_2 \rightarrow CO_2 + H_2O$$

4. *Using a molecule of natural gas, CH_4 and 2 molecules of oxygen, O_2, break the bonds to form carbon dioxide, CO_2 and water, H_2O.*

Now 2 molecules of water, H_2O can be formed as well as the 1 molecule of carbon dioxide, CO_2.

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

This is called a balanced chemical equation.



The number of atoms on the reactant side is equal to the number of atoms on the product side.

Balanced chemical equations (i)
(continued)

This example shows the three kinds of chemical equation.

a) **a word equation**

natural gas + oxygen → carbon dioxide + water

b) **an equation using formulae**

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

c) **a balanced chemical equation**

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

Remember -

Never change a formula to make an equation balance. The formula of carbon dioxide is always CO_2 . CO is carbon monoxide, a quite different gas. The formula of water is always H_2O; HO does not exist.

Equations can only be balanced by putting a number in front of a formula. e.g. $2H_2O$ or $2O_2$.

Balanced chemical equations (ii)

The following examples show a method of balancing equations.

Example 1:

Hydrogen reacts with fluorine to form hydrogen fluoride

$$H_2 + F_2 \rightarrow HF$$

H_2 represents two hydrogen atoms joined together.
F_2 represents two fluorine atoms joined together.

The equation can be written as :

H H	+	F F	→	HF

There are two hydrogens on the left side but only one on the right side (the product side).

Write another HF on the right side to increase the number of hydrogens (and at the same time the number of fluorines).

H H	+	F F	→	HF HF

There are now the same number of each type of atom on both sides.
The equation is balanced.

Rewrite the equation $H_2 + F_2 \rightarrow 2HF$

Example 2:

$$C + Br_2 \rightarrow CBr_4$$

C	+	Br Br	→	C Br Br Br Br

There are two bromine atoms on the left side (the reactant side) but four are required on the product side.

C	+	Br Br Br Br	→	C Br Br Br Br

$$C + 2Br_2 \rightarrow CBr_4$$

Using balanced equations

A balanced equation is taken to give the **relative number of moles** of each reactant and product. Since the mass of one mole of any substance is expressed in grams, the masses involved can then be calculated as shown.

Example 1:
Calculate the mass of water produced on burning 1 g of methane.

balanced equation	CH_4	+	$2O_2$	$\rightarrow$	CO_2	+	$2H_2O$
relative number of moles	1 mol						2 mol

note: It is not necessary to calculate the masses of carbon dioxide and oxygen these substances are not included in the question.

	CH_4		$2H_2O$
relative atomic masses	12 + (4 x 1)		2[(2 x 1) + 16)]
mass in grams	16 g	<---------------->	36 g
	1 g	<---------------->	$\frac{36 \times 1}{16}$
			= **2.25 g**

note: The last part of the calculation is a simple proportion and hence the use of the symbol <---------------->.

Example 2:
Calculate the mass of lead(II) carbonate required to produce 2.2 g carbon dioxide on heating.

balanced equation	$PbCO_3$	$\rightarrow$	PbO	+	CO_2
number of moles	1 mol				1 mol
relative atomic masses	207 + 12 + (3 x 16)				12 + (2 x 16)
mass in grams	267 g	<---------------->			44 g
	$\frac{267 \times 2.2}{44}$	<---------------->			2.2 g
	= **13.35 g**				

Example 3:
An industrial plant produces ammonia by the reaction of nitrogen with hydrogen. An output of 7.5×10^3 kg of ammonia is required each day.
Calculate the mass of nitrogen used each day assuming that the factory is working at 80% efficiency.

balanced equation	N_2	+	$3H_2$	$\rightarrow$	$2NH_3$
relative number of moles	1 mol				2 mol
relative atomic masses	(2 x 14)				2 {14 + (3 x 1)}
mass in grams	28 g			<------>	34 g
	$\frac{28 \times 7.5 \times 10^3}{34}$			<------>	7.5×10^3 kg
	$= 6.18 \times 10^3$ kg				

This assumes 100% efficiency
Actual efficiency is 80% and hence a greater amount of nitrogen is required to produce the required amount of ammonia.

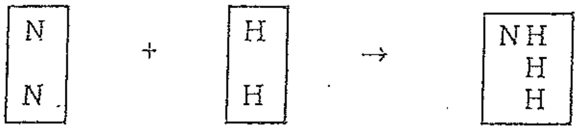
Mass of nitrogen used each day $= \frac{100 \times 6.18 \times 10^3}{80}$ kg

$= 7.725 \times 10^3$ kg

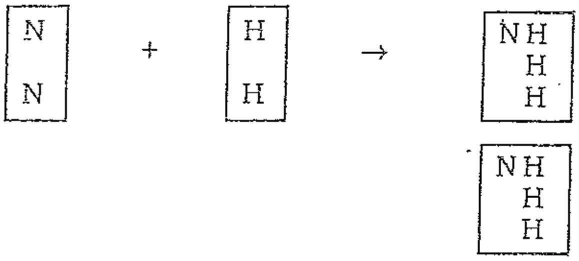
Balanced chemical equations (ii)
(continued)

Example 3:

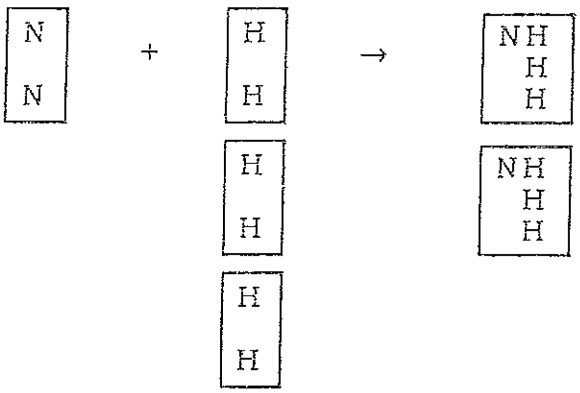
$$N_2 + H_2 \rightarrow NH_3$$



Two nitrogens are required on the product side.



Six hydrogens are required on the reactant side.



$$N_2 + 3H_2 \rightarrow 2NH_3$$