CHAPTER 5
SYMBOLS, FORMULAE, NAMES AND EQUATIONS

5.1 Chemical Symbols

As we saw in Chapter 1, each chemical element may be represented by a unique chemical symbol. This usually consists of a single capital letter or a capital letter followed by a lower case letter, as for example O for oxygen or Fe for iron. A few elements which have been discovered recently have been given three letter symbols derived from names which suggest their atomic numbers. For example, element number 110 has been given the name *ununnilium* and symbol *Uun*. Strictly speaking, only elements may be represented by symbols[1].

5.1.1 The meanings of chemical symbols

We are familiar with the idea that the meaning of a word depends very much on where it occurs. For example, consider the word *sound* in the phrases "the *sound* of music" and "his work is very *sound*". In the same way a chemical symbol may have distinct meanings depending on the context. A chemical symbol may therefore mean:

a) The element in general.
For example we may write, "The elements of group 6 include O and S." Here no particular state of combination of the elements (oxygen and sulfur) is indicated - the statement refers to the elements in general whether free or combined in a compound. This use should be contrasted with a statement such as "O_2 was collected." Here a specific form of oxygen is referred to (oxygen gas or "dioxygen") which has formula O_2.

b) A single atom of the element.
This use occurs mainly in chemical formulae and chemical equations. For example, the combustion of carbon may be written:

$$C + O_2 \rightarrow CO_2$$

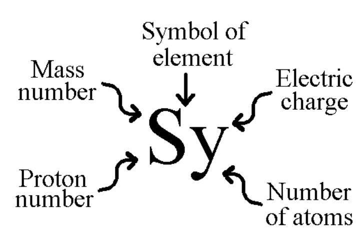
"C" identifies the element carbon, but here it does more than that. It specifically refers to a single atom of carbon. This is the reason why we should not use the symbol O instead of oxygen in the statement "O_2 was collected". Here it would imply that the oxygen consisted of individual atoms rather than molecules, and this is not true. Individual oxygen atoms have very different chemical properties from those which are present in molecules.

c) A certain quantity of the element known as a *mole*[2]. The above equation may also be read in terms of moles. Here the C may be read as "1 mole of carbon".

5.1.2 Numbers and signs accompanying a chemical symbol

A chemical symbol may be accompanied by various numbers and other symbols. (See figure 5.1.) Each one has a special meaning according to its position with respect to the symbol, as shown in figure 5.1.

Figure 5.1: *The positions of the various numbers which surround a chemical symbol.*



a) A number bottom right of a symbol refers to a certain number of atoms of the element in

---

[1] In the branch of chemistry dealing with the compounds of carbon, *organic chemistry*, it is quite common to represent groups of atoms by a symbol. For example, a carbon atom and three hydrogen atoms joined together in a group is represented by *Me*.

[2] A mole of a substance is 6×10^{23} atoms, molecules, or formula units of it, as appropriate. (See chapter 9.)

a special situation. For example, O_2 refers to two atoms of oxygen, here strongly bound together in 1 molecule of oxygen. In $NaNO_3$ the "3" means that each formula unit of the substance represented by $NaNO_3$ (sodium nitrate) contains 3 oxygen atoms. An absence of a number in this position is always read as 1. In the above formula for sodium nitrate the 3 atoms of oxygen are associated with 1 atom of sodium and 1 atom of nitrogen.

b) A number top right of a symbol is always associated with a plus sign or a minus sign. It refers to the amount of electric charge on the atom. Atoms with electric charge are called *ions*. An ion is formed when an atom gains or loses one or more electrons (see chapter 6, section 6.3). For example, S^{2-} means a sulfide ion (note the change of name from the element, sulfur). The "2-" means that it has a double negative charge, i.e.. the original atom has gained two electrons. Similarly Al^{3+} means an aluminium ion. "3+" means that it has three positive charges, i.e.. the atom has lost three electrons. As usual in chemistry, the absence of a number indicates 1. Thus Na^+ represents a sodium ion, which has one positive charge. Ordinary atoms have no charge overall and the top right hand corner of the symbol is left empty, eg. Na represents a sodium atom with no charge.

c) Two other numbers which are found adjacent to the symbols of elements will be dealt with more fully in chapter 6. A number bottom left of a symbol indicates the *atomic number* of an element, eg. ${}_{11}Na$, ${}_{8}O$ etc., whilst a number top left of a symbol indicates the *mass number* of one isotope of the element. For example, ${}^{23}Na$ is the most common isotope of sodium, sodium-23.

5.2 Chemical Formulae

A chemical formula identifies a specific substance. This may be a free element or a compound.

In its simplest form a formula may be the same as the symbol for an element. For example, "C" is both the chemical *symbol* for carbon and its chemical *formula*. On the other hand, "O_2" is the *formula* for oxygen but not the *symbol*, which is simply "O". The formula O_2 identifies oxygen as the free element in its normal state.

With compounds the situation is more clear-cut. A compound can *only* be identified by a formula. Thus "H_2O" is obviously a *formula* rather than a *symbol*.

To summarise, a formula may refer: (a) to a free element (i.e.. not combined with another element) or (b) to a compound.

5.2.1 The meaning of chemical formulae

In the same way that the meaning of a symbol may vary according to where it occurs, the same is true of a formula. Thus a formula may refer to:

a) the substance in general.
   For example "NaCl" may often be read as "sodium chloride", or "common salt" as in "the mixture contained NaCl".

b) the smallest possible amount of a substance, which may be 1 molecule or 1 formula unit of it as appropriate. (See also chapter 3, section 3.3.)
   For example, "H_2O" may represent a single molecule of water, "NaCl" a single formula unit of sodium chloride[1], and "C" a formula unit of carbon[2]. The metals, such as iron and aluminium are like carbon in this respect.

---

1 Sodium chloride does not consist of molecules, but rather of ions, so we cannot refer to a "molecule" of it. The term "formula unit" is used instead.

2 Carbon, as the free element, does not normally consist of molecules. A formula unit is the same thing as an atom in this particular case since the formula is the same as the symbol for carbon.

c) One *mole*[1] of the substance. A mole of a given substance is a certain definite quantity of it.

5.2.2 Writing the Formula for a Compound

There are some simple rules which allow the formula of a compound to be written down, correctly in most cases. The formula produced may be the empirical formula, the molecular formula, or the simplified ionic formula (see section 5.2.3) but it is not necessary to know the type of formula to apply these rules.

5.2.2.1 The valency of an element

The formula of a compound depends on the *valency* of the elements of which it is composed. The valency of an element is a measure of its *combining power*. It may be defined as *the number of hydrogen atoms with which one atom of the element combines with or replaces*. Elements often show the same valency in a wide variety of situations.

The valency of an element is determined by examining what compounds it forms. For example, carbon forms a compound with hydrogen in which one carbon atom is joined to four hydrogen atoms (CH_4, methane). This shows that the valency of carbon, in this compound at least, is 4.

Hydrogen also combines with chlorine to form the compound HCl in which, as the formula shows, one atom of hydrogen is combined with one atom of chlorine. This tells us that the valency of chlorine is 1. That is, chlorine has the same combining power as hydrogen. We therefore expect that chlorine might replace hydrogen in its compounds on a 1:1 basis. This is so, and the hydrogen in CH_4 can be replaced by chlorine to form CCl_4.

Oxygen forms compounds with hydrogen too. One of them is, of course, water[2], H_2O. The formula of this compound shows that one atom of oxygen combines with 2 atoms of hydrogen so the valency of oxygen is 2.

In view of this we might expect that an oxygen atom might replace 2 hydrogen atoms in another compound. This turns out to be the case. For example, if the 4 hydrogen atoms in CH_4 are replaced two-at-a-time by oxygen atoms, we get the formula CO_2, which most of you will recognise as carbon dioxide, a real substance.

In a similar way we may predict the formula of many compounds, once we know either the formula of one compound of each of the elements involved, or we know the valency of the elements.

GROUP	VALENCY
Group I	1
Group II	2
Group III	3
Group IV	4
Group V	3 or 5
Group VI	2 or 6
Group VII	1 or 7
Group 0	0

Table 5.1: *The normal valencies of the main group elements. The higher valencies of the elements of groups V, VI and VII are only shown in certain compounds, usually when the element combines with oxygen. Certain elements, especially those lower in the periodic table often show other valencies.*

In many cases the valency of an element is easy to determine from its position in the periodic table (see inside back cover). For the elements of groups I through IV it is the same as the group number. For elements of groups V, VI, and VII it is usually equal to (8 - group no.), though in compounds with oxygen and chlorine it may sometimes be equal to the group number again.

For the noble gases (in group 0) the valency is zero since these elements form no compounds. The valencies of these *main group* elements is summarised in table 5.1.

---

[1] One mole of a substance is 6×10^{23} atoms, molecules, or formula units of it, as appropriate.

[2] Others are hydrogen peroxide (H_2O_2) and dihydrogen trioxide (recently discovered, H_2O_3) which do not obey these simple rules of valency.

Exercise
Determine the normal valencies of the following elements:
i) sodium
ii) potassium
iii) sulfur
iv) silicon
v) oxygen
vi) calcium
vii) aluminium
viii) phosphorus
Answers are given below[1].

5.2.2.2 *Rules for writing the formula of binary compounds of main group elements.*

a) Binary compounds are those which contain only two elements. Write down the symbols for the elements in the order that they appear in the name. Their names generally consist of the names of the separate elements though the last one takes the termination *-ide*. (See section 5.3.) Some have special names, such as water, ammonia and silica. For these it may not be immediately obvious what elements the compound is made of. Others contain the prefixes *di-*, *tri-* etc. which may be ignored at this stage.
For example, suppose you wish to write the formulae of potassium sulfide and silicon dioxide. You will start by writing:
K S for potassium sulfide,
and Si O for silicon dioxide
b) Determine the valencies of the elements by finding them in the periodic table.
For example, K (potassium) is in group I, so its valency is 1, S (sulfur) is in group VI so its valency is 2, Si (silicon) is in group IV, so its valency is 4, and O (oxygen) is in group VI, so its valency is 2.
c) If the valencies have a common factor, divide them by it.
For example the valencies of silicon and oxygen are 4 and 2, so they have the common factor 2. Divide the numbers by this to get 2 and 1. On the other hand the valencies of potassium and sulfur are 1 and 2 with no common factor, so these are left unchanged.
d) Write the numbers bottom right of the *other* symbol in the formula.
For example, for sulfur we have determined 2, and so this is written bottom right of the K. For silicon we have also determined 2 (after dividing by the common factor) and we write this number bottom right of the O. The number 1 is not usually written in a formula. When no number is written, 1 is understood.
This gives us
K_2S for potassium sulfide
and SiO_2 for silicon dioxide.
These are the correct formulae of the compounds.

Exercise
Write chemical formulae for the following:
i) sodium fluoride
ii) calcium sulfide
iii) aluminium oxide
iv) sulfur dichloride
v) phosphorus trichloride
vi) dinitrogen trioxide
Answers are given below[2].

5.2.2.3 *Formulae of binary compounds[3] of variable valency and non-main group elements.*

There are many elements which do not lie in the main groups. i.e.. they are not in groups I through VII, nor group 0. Such elements are known as *transition elements* (see section 6.2.8.3.2). They occupy the central block in your periodic table and typically show variable valency. In other words they may show different

[1] 1,1,2,4,2,2,3,3
[2] NaF, CaS, Al_2O_3, SCl_2, PCl_3, N_2O_3
[3] "Binary compounds are those which contain only two elements.

valencies in different compounds. For example copper forms two oxides, copper(I) oxide where the valency of the copper is 1, and copper(II) oxide where the valency of the copper is 2. Some of them only show one valency, however. For example, silver only shows a valency of 1.

There are also some main group elements which show variable valency, especially those lower down the periodic table. Lead (group IV) is one such. It shows valencies of 4 and 2.

Some of the more common of these elements and their valencies are shown in table 5.2. The valencies of non-main group elements which only show one valency, such as zinc and silver, must be learned, but those of the variable valency elements are always indicated in the name, as with copper(I) oxide and copper(II) oxide above.

ELEMENT	VALENCIES
Copper	1 or 2
Iron	2 or 3
Lead	2 or 4
Silver	1
Tin	2 or 4
Zinc	2

Table 5.2: *The normal valencies of some variable-valency and non-main group elements.*

For example, to write the formulae of silver oxide and tin(IV) oxide we proceed by exactly the same stages as previously.

a) Writing symbols:
    Ag O for silver oxide
    and Sn O for tin(IV oxide

b) Finding valencies:
    silver: 1 (must be learned)
    oxygen: 2 (from group number)
    tin: 4 (from number in name)

c) Finding common factors where there are any gives:
    1 and 2 for silver oxide (unaltered)
    and 2 and 1 for tin(IV) oxide (divided by 2).

d) Writing these numbers by the *opposite* symbols gives:
    Ag_2O and SnO_2
    which are the correct formulae for silver oxide and tin(IV) oxide respectively.

Exercise

Write formulae for the following:

i) silver sulfide
ii) copper(I) oxide
iii) iron(III) sulfide
iv) iron(II) chloride
v) lead(II) bromide
vi) tin(II) carbide

Answers are given below[1].

5.2.2.4 Formulae of more complex compounds

There are many compounds which contain more than two elements. Where these are compounds of non-metals, writing a formula is not usually straightforward and will not be dealt with in this book. The few examples which are encountered are dealt with individually.

Many common compounds of more than two elements are compounds of metals, however, and they can be dealt with under the same scheme as the binary compounds of metals. The only difference is that instead of containing a nonmetal they contain a nonmetallic group, i.e.. a group of atoms that takes the place of a nonmetal. This is better referred to as a *polyatomic ion* (see *section 5.2.3.2.1*.

For example, sulfate is a nonmetallic group, or polyatomic ion, with formula SO_4^{2-}. The "2-" is the charge on the ion and indicates the valency. The valency of the sulfate ion is just "2" (without the sign[2]) just the same as the number of charges. Once this is recognised it is easy to write the formula of, for example, sodium sulfate. We just follow the same rules as before. The charge is not included.

a) Na SO_4 - Writing symbols & formulae.

[1] Ag_2S, Cu_2O, Fe_2S_3, $FeCl_2$, $PbBr_2$, Sn_2C

[2] Some people use signed valencies, positive for metals and negative for non-metals.

b) 1 2 - Valencies: sodium is in group I, sulfate must be learned.
c) 1 2 - No common factor so no change.
d) Na_2SO_4 - Position of numbers is swapped.

NAME OF ION	FORMULA	VALENCY
carbonate	CO_3^{2-}	2
hydrogencarbonate	HCO_3^-	1
sulfate	SO_4^{2-}	2
hydrogensulfate	HSO_4^-	1
sulfite	SO_3^{2-}	2
hydrogensulfite	HSO_3^-	2
nitrate	NO_3^-	1
phosphate	PO_4^{3-}	3
hydrogenphosphate	HPO_4^{2-}	2
dihydrogenphosphate	$H_2PO_4^-$	1
hydroxide	OH^-	1
ammonium	NH_4^+	1
hydronium	H_3O^+	1

Table 5.3: *Names, formulae and valencies of some common polyatomic ions (radical ions). Note that the valency is the same as the number of charges.*

Sometimes we need to use brackets. Consider, for example magnesium nitrate. This contains the nitrate ion, NO_3^-, valency 1. Using the same rules as before, we have:

a) Mg NO_3
b) 2 1
c) 2 1
d) $Mg(NO_3)_2$

The brackets are necessary to show that everything within them is doubled. That is, there are two nitrate ions for every magnesium atom.
There are only a few common polyatomic ions. They are given in table 5.3 together with their valencies. Unfortunately there is no simple way to predict the valencies or formulae of such polyatomic ions from the elements which make them up. They have to be learned.

Exercise
Write the formula for:
i) Calcium dihydrogen phosphate
ii) Sodium carbonate
iii) Aluminium nitrate
iv) Iron(III) phosphate
v) Sodium hydroxide
vi) Zinc sulfate
Answers are given below[1].

5.2.2.5 Formulae of compounds containing metallic groups

These are polyatomic ions which have a positive charge (cations). They take the place of a metal in many simple compounds. The only common one encountered in compounds is the ammonium[2] ion, NH_4^+, though the hydronium ion, H_3O^+, is found in aqueous solutions. They are shown in table 5.3. They both have valency 1.
The formula of an ammonium compound is simply written according to the same rules as before. For example, suppose we want to write the formula of ammonium sulfate. We simply apply the rules:

a) NH_4 SO_4 - appropriate formulae with charges omitted.
b) 1 2 - valencies
c) 1 2 - no common factor
d) $(NH_4)_2SO_4$ - correct formula

5.2.2.6 Compounds which do not obey the simple rules of valency

There are a large number of compounds whose formulae do not obey the normal rules of valency as given above and are therefore unpredictable. Some of the more common, together with their names are given in table 5.4.

[1] $Ca(H_2PO_4)_2$, Na_2CO_3, $Al(NO_3)_3$, $FePO_4$, NaOH, $ZnSO_4$
[2] It is most important to distinguish the *ammonium ion* from *ammonia*. Ammonia is a substance in its own right, whereas the ammonium ion is only found as part of a substance and cannot exist on its own under normal circumstances.

5.2.3 *Types of chemical formula*

The chemical formulae mentioned in the preceding sections were of several different types and no attempt was made to distinguish between them. The particular formula used depends both on the type of compound involved and the type of information which needs to be conveyed.

NAME OF COMPOUND	FORMULA
sulfur dioxide	SO_2
nitrogen dioxide	NO_2
dinitrogen oxide	N_2O
mercury(I) chloride	Hg_2Cl_2
carbon monoxide	CO
hydrogen peroxide	H_2O_2

Table 5.4: *Some compounds which do not obey the simple rules of valency.*

5.2.3.1 Molecular formulae

A molecular formula shows both the number and types of atoms in one molecule of a free element or of a compound. The term "molecular formula" cannot apply to ionic compounds, which include most of the common compounds of metals.

All liquid and gaseous elements[1] and compounds consist of molecules and can therefore be identified by molecular formulae. Solid compounds of non-metals also, are often molecular[2].

The gaseous elements, hydrogen (H_2), fluorine (F_2), chlorine (Cl_2), nitrogen (N_2), and oxygen (O_2) consist of diatomic molecules. A few other elements which are not gases also consist of diatomic molecules, for example: bromine (Br_2) and iodine (I_2)[3]. (The terms dioxygen, dihydrogen, di-iodine etc. may be used to indicate these elements in their normal molecular form).

Some elements occur (or can be made) in more than one molecular form. One example is oxygen, which also occurs as ozone (O_3). Ozone may be referred to as *trioxygen*. It has very different properties from ordinary oxygen, being extremely poisonous[4].

Some elements occur naturally in the form of free atoms. These are the noble gases: helium (He), neon (Ne) etc. This means, according to the definition (see chapter 3, section 3.3, that, in this particular case, a molecule is the same as an atom and so the symbol is also the molecular formula.

Compounds which are molecular include water, a liquid, H_2O; carbon dioxide, a gas, CO_2 and sucrose, a solid, $C_{12}H_{22}O_{11}$.

The molecular formula may not be sufficient to identify a substance uniquely. Various arrangements of the atoms may be possible giving rise to different substances. For example, lactose has the same molecular formula as sucrose but they are different substances.

5.2.3.2 Ionic formulae

Many compounds, but no free elements, are ionic, or electrovalent. That is, they are made up of a lattice of ions of opposite charges. (See chapter 3, figure 3.4.) These are almost always compounds of a metal and one or more non-metals (i.e.. oxides, hydroxides or salts of metals), though ammonium compounds form an important exception.

A molecular formula cannot be written for ionic compounds, since they do not contain any molecules. An ionic formula is appropriate here.

The "ordinary" formulae for compounds of metals that we met in a previous section can be termed simplified ionic formulae. For example, the simplified ionic formula for sodium chloride is NaCl and for aluminium oxide, Al_2O_3. The full ionic formula shows the charges on the ions as

---

1 Liquid or gaseous at room temperature, that is.

2 *Molecular* compounds should not be confused with *giant molecular* compounds. The latter do not contain molecules in the true sense.

3 Sulfur and white phosphorus, which are solids, are also molecular (S_8 and P_4 respectively), but traditionally these elements are referred to simply by S and P.

4 In a way it is not so different from ordinary oxygen as you might think. Pure oxygen is actually toxic to living things. Although necessary for life, oxygen also wreaks havoc on the proteins and DNA that make up living organisms and they have to struggle continuously to repair this damage. According to some theories this damage is at least partly responsible for ageing. Ozone is just a very active form of oxygen and overwhelms the body's repair mechanisms.

well. For sodium chloride it is Na^+Cl^- and for aluminium oxide $(Al^{3+})_2(O^{2-})_3$.

The ionic formula is like the empirical formula (see section 5.2.3.3) in one way: it shows only relative numbers of ions, not actual numbers.

Note that the number of positive charges and the number of negative charges in an ionic formula are always equal. The formula of sodium chloride shows one positive charge (on the sodium ion) and one negative charge (on the chloride ion). In the formula of aluminium oxide there are six positive charges (2 aluminium ions with 3 positive charges each) and six negatives (3 oxide ions with 2 negative charges each). This shows that an ionic compound is electrically neutral, like any other substance. This is a very useful guide for writing the formula of an ionic compound (see section 5.2.2).

Alternatively the same principle can be used to find the charges on the ions from the simplified ionic formula, once it is realised that ions with more than 4 charges are never encountered.

On the other hand the charge on the ions may be found from the valency, given that metals form positive ions and non-metals form negative ions (when they form ions at all). Thus calcium is in group II and so forms an ion with a double charge. All group II elements are metals so the charge is positive and the formula of the ion is Ca^{2+}. In the same way, nitrogen (if it forms an ion at all) will form one of formula N^{3-}, since it is a nonmetal in group V, and group V elements are non-metals with a normal valency of 3.

Exercise

Write formulae for the following ions:

i) sodium ion
ii) aluminium ion
iii) lead(II) ion
iv) oxide ion
v) sulfide ion
vi) iodide ion

Answers are given below[1].

Once the charges on the ions have been decided, it is a simple matter to write the full ionic formula for a compound. One may start either with the "ordinary" formula, determined by the rules of valency, and simply put the charges in. Extra brackets will often be necessary to make things clear.

Alternatively one may take the ions as given and put them together in such numbers that the total charge is zero.

Exercise

Write ionic formulae for the following compounds.

i) sodium oxide
ii) aluminium sulfide
iii) lead(II) iodide
iv) sodium iodide
v) sodium sulfide
vi) lead(II) oxide

Answers are given below[2]

A rule that must always be followed in writing formulae is that *if any charges are shown, all must be shown*. Thus NaCl and Na^+Cl^- are both correct, but any mixture of the two is incorrect. If isolated ions are represented, such as the chloride ion ($C1^-$) or the magnesium ion (Mg^{2+}), the charges *must always be shown*.

5.2.3.2.1 Polyatomic ions

Many ionic compounds contain more complex ions, sometimes referred to as *radicals* or (better) *radical ions* or *polyatomic ions*. The most common ones include ammonium[3] (NH_4^+), sulfate (SO_4^{2-}), carbonate (CO_3^{2-}), and nitrate (NO_3^-). A more complete list is given in table 5.3.

Polyatomic ions occur in ionic formulae in just the same way as simple ions. Again, these must be written so that the negative and positive charges are present in equal numbers. Thus sodium carbonate is $(Na^+)_2(CO_3^{2-})$ and ammonium sulfate is $(NH_4^+)_2(SO_4^{2-})$, whilst magnesium nitrate is $(Mg^{2+})(NO_3^-)_2$.

Exercise

---

1 Na^+, Al^{3+}, Pb^{2+}, O^{2-}, S^{2-}, I^-

2 $(Na^+)_2O^{2-}$ $(Al^{3+})_2(S^{2-})_3$, $Pb^{2+}(I^-)_2$, Na^+I^-, $(Na^+)_2S^{2-}$, $Pb^{2+}O^{2-}$

3 Ammonia is a substance, NH_3 and can be stored in a container. *Ammonium* is an ion, and can only form part of a substance. It cannot be prepared in a pure form.

Write ionic formulae for the following:

i) sodium sulfate
ii) ammonium sulfite
iii) calcium phosphate
iv) calcium hydrogenphosphate
v) aluminium sulfate
vi) magnesium nitrate

Answers are given below[1].

5.2.3.3 Empirical formulae

An empirical formula shows only the *ratio* of the numbers of the various types of atoms present in a substance in its lowest terms. An empirical formula can be written for any compound. Ethanoic acid, for example, has molecular formula $C_2H_4O_2$ (often written CH_3COOH). The molecule contains 2 atoms of carbon, 4 atoms of hydrogen and 2 atoms of oxygen. Thus any amount of ethanoic acid contains these atoms in the ratio 2:4:2. This ratio can be expressed in its lowest terms as 1:2:1. Remembering that the figure "1" is always omitted from formulae, we get the empirical formula CH_2O.

Similarly hydrogen peroxide, molecular formula H_2O_2, has empirical formula HO.

In many cases the empirical formula and the molecular formulae are the same. This is the case with H_2O and CO_2, for example. Here the ratio of numbers of atoms is already in its lowest terms.

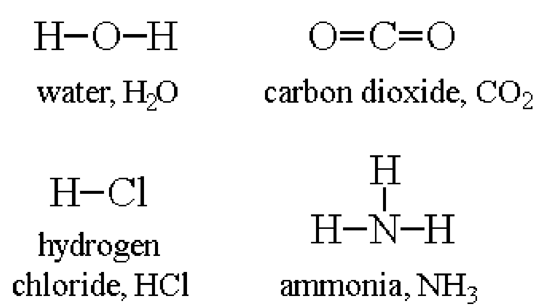
Many elements and compounds do not consist of either molecules or ions, and therefore neither a molecular formula nor an ionic formula can be written. Metals, for example, consist of a lattice of atoms strongly bound together by metallic bonds (see chapter 7, section 7.4 and figure 3.3). Silicon dioxide and carbon consist of giant covalent lattices (see chapter 7, section 7.3.12. Compounds of this type may only be represented by an empirical formula. The formula for silicon dioxide, SiO_2, shows only the types of atoms present and their numerical ratio in its lowest terms.

Elements of this type may only be represented by their symbols. For example carbon is represented by C and iron by Fe. Symbols used in this way may also be classified as empirical formulae.

Ionic formulae, especially when charges are omitted, are in some ways similar to ionic formulae. They show a ratio of numbers of *ions* in its lowest terms. Often, however, the ionic and the empirical formulae are different. For example, sodium peroxodisulfate has ionic formula $(Na^+)_2(S_2O_8^{2-})$ (or just $Na_2S_2O_8$). The empirical formula, however, is $NaSO_4$[2].

5.2.3.4 Structural formulae

A structural formula shows not only the types and numbers of atoms in a molecule (or polyatomic ion), but also how they are arranged, i.e.. which atom is connected to which. The structural



formulae for some simple molecules is shown in figure 5.2

Figure 5.2: *Structural formulae of some simple molecules. The precise meanings of the double and single lines joining the atoms is dealt with in chapter 7.*

In many simple compounds, but by no means all, the molecular or ionic formula serves to identify the compound uniquely. For example, there is only one compound of formula CO_2. On the other hand there are several compounds with the molecular formula $C_2H_4O_2$. This is because there is more than one way of arranging the atoms in the molecule[3]. In such cases a structural formula

---

1 $(Na^+)_2SO_4^{2-}$, $(NH_4^+)_2SO_3^{2-}$, $(Ca^{2+})_3(PO_4^{3-})_2$, $Ca^{2+}(HPO_4^{2-})_2$, $(Al^{3+})_2(SO_4^{2-})_3$, $Mg^{2+}(NO_3^-)_2$

2 Be careful! The common compound sodium sulfate has formula Na_2SO_4 which should not be confused with this rarely-encountered empirical formula.

3 Compounds which differ only in the arrangement of atoms in the molecule are known as *isomers*.

is useful to distinguish between the various substances. For example, both ethanoic acid and methyl methanoate have the same molecular formulae. Their structural formulae are shown in figure 5.3. The structural formulae show that the same atoms are connected in a different way and so help to explain why the two substances are different.

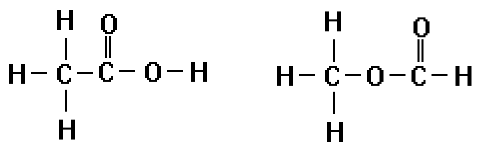


Figure 5.3: *The structural formulae of two compounds with the same molecular formulae.*

A *condensed structural formula* is often more convenient. For example, ethanoic acid may be represented as CH_3COOH and methyl methanoate as CH_3OCHO. It shows most of the information of a full structural formula, but on one line.

5.3 Naming simple compounds

The name of a compound is simply derived from the formula in most cases, though a few compounds have traditional names which do not follow any such rules. In addition the more complex compounds of carbon have their own system of naming.
Amongst the simple compounds there is a broad division between those containing a metal (or metallic group) and those which do not.

5.3.1 Naming binary compounds of metals

Essentially the name is based on the names of the elements the compound contains, except that the last one takes the termination *-ide*.
Thus NaF is sodium fluoride, K_2O is potassium oxide etc.
Some of the more common elements and their modified names are given in table 5.5. Note that they are all non-metals. Normally, metals do not occupy the second position in a name or formula. The modified names are also the names of the anions formed by the elements. Metals do not form anions.

Exercise
Write chemical names for the following:
i) MgS
ii) KF
iii) ZnO
iv) AlI_3
v) Li_2S
vi) AgBr
Answers are given below[1].

NAME OF ELEMENT	NAME OF ANION
bromine	bromide
chlorine	chloride
fluorine	fluoride
iodine	iodide
oxygen	oxide
sulfur	sulfide

Table 5.5: *Some of the more common elements and their modified names when found in binary compounds as the last element in the formula. These names are the same as the names of the corresponding anions - see chapter 6, section 6.3.*

5.3.1.1 Names of binary compounds of variable valency metals

The names of compounds of variable valency metals show the valency of the metal in roman numerals placed in brackets immediately after the name of the metal. Thus $PbCl_4$ is lead(IV) chloride, Fe_2O_3 is iron(III) oxide and so on.

5.3.2 More complex compounds of metals

5.3.2.1 Metallic groups

These are polyatomic cations which take the place of a metal in a compound, as has been

[1] magnesium sulfide, potassium fluoride, zinc oxide, aluminium iodide, lithium sulfide, silver bromide.

mentioned above. They may also be referred to as *radical cations*. There is only one metallic group which is commonly encountered in simple compounds, the *ammonium ion*, NH_4^+).
The names of compounds involving the ammonium ion mirror the names of compounds of metals. Thus we have NH_4Cl, ammonium chloride etc.

5.3.2.2 Non-metallic groups

These are polyatomic anions and may also be referred to as radical anions. They behave as a nonmetal in compounds with metals. Their names and formulae (see table 5.3) cannot be predicted from simple rules, though some patterns are evident. For example, with the exception of hydroxide, OH^-, their names all end in *-ate* or *-ite* and all the common ones contain oxygen. Those with the termination *-ite* contain less oxygen that those with the termination *-ate*. Those containing hydrogen are named according to the amount present eg. $H_2PO_4^-$ is dihydrogenphosphate.
Naming of compounds containing nonmetallic groups follows the same rules as other simple compounds of metals. Thus Na_2SO_4 is sodium sulfate, $Mg(NO_3)_2$ is magnesium nitrate, $(NH_4)_2HPO_4$ is ammonium hydrogenphosphate, and so on.

Exercise
Name the following compounds:
i) Na_2SO_3
ii) $Al_2(SO_4)_3$
iii) NH_4Br
iv) $NaHCO_3$
v) $Fe_2(SO_4)_3$
vi) Cu_3PO_4
Answers are given below[1].

5.3.3 Naming compounds where hydrogen replaces a metal - acids

Many, but by no means all, compounds containing hydrogen are acids. The names and formulae of the common acids are given in table 5.6. Most of these compounds contain oxygen. In these compounds the terminations *-ic* and *-ous* have the same significance as *-ate* and *-ite* in the polyatomic anions. Thus we have sulfuric acid, H_2SO_4 and sulfurous acid, H_2SO_3.
Where the compound does not contain oxygen the name takes the prefix *hydro-* and the termination *-ic*. Where these are binary compounds, such as HCl or H_2S, an alternative name where hydrogen is treated as metal may often be used. Thus HCl may be referred to as hydrogen chloride. This name is usually reserved for the pure gaseous compound whereas the solution in water is referred to as hydrochloric acid. H_2S is almost always referred to as hydrogen sulfide rather than "hydrosulfuric acid".

NAME OF ACID	FORMULA
Carbonic acid	H_2CO_3
Hydrochloric acid	HCl
Nitric acid	HNO_3
Nitrous acid	HNO_2
Phosphoric acid	H_3PO_4
Sulfuric acid	H_2SO_4
Sulfurous acid	H_2SO_3

Table 5.6: *The names and formulae of some common acids. (See also chapter 8.).*

5.3.4 Naming compounds of non-metals

Only the *binary* compounds of non-metals will be dealt with here. All but the simplest compounds of carbon have their own naming system which is beyond the scope of this book.
The most common system is very like that for compounds of metals except that the number of each type of atom in the formula is indicated by a prefix.
Thus SiO_2 is silicon *di*oxide, N_2O_4 is *di*nitrogen *tetra*oxide, SO_3 is sulfur *tri*oxide and N_2O is *di*nitrogen oxide. The common prefixes are given in table 5.7.

Exercise

[1] sodium sulfite, aluminium sulfate, ammonium bromide, sodium hydrogencarbonate, iron(III) sulfate, copper(I) phosphate.

Write names for the following:
i) $SiCl_4$
ii) S_2Cl_2
iii) N_2O_5
iv) P_4O_{10}
v) C_3N_4
vi) N_2S_3
Answers are given below[1].

PREFIX	NUMBER
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
deca-	10

Table 5.7: *The most common prefixes used to indicate the numbers of atoms in the formula when naming the compound. The prefix mono- is only used in a few names such as carbon monoxide. Normally the absence of a prefix indicates one atom.*

5.3.5 Alternative names for compounds

In many cases more than one name for a compound is in common use. Some examples regarding acids have already been mentioned. In addition, compounds of metals are sometimes named according to the rules for non-metals. For example $PbCl_4$ is sometimes known as *lead tetra-chloride* rather than *lead(IV) chloride*.
In other cases, and this is more common, compounds of non-metals are named according to the rules for compounds of metals. For example, N_2O_5 could be called *nitrogen(V) oxide*.

In other cases a completely non-systematic name exists alongside a systematic one. For example, SiO_2 is known both as *silicon dioxide* and *silica*; NaOH is known both as *sodium hydroxide* and *caustic soda*.

5.4 Chemical Equations

A chemical equation shows how one or more substance turns into one or more different substances as a result of a chemical change. The starting materials are known as the *reactants* and the new substances produced are known as the *products*. The change is indicated by an arrow (→) and each substance is shown by its chemical formula. Reactants lie to the left of the arrow and products to the right.
Chemical equations also show the *relative amounts* of substances used up and produced during the change. To make this possible a chemical equation must be balanced, that is, all the atoms amongst the reactants must turn up amongst the products, both in terms of number and of kind. This is a necessary consequence of Dalton's atomic theory, according to which atoms can neither be created nor destroyed.

5.4.1 Word equations

As a first step on the way to writing a chemical equation a word equation is often helpful. A word equation identifies the reactants and products by their names. All must be shown.
In your practical class you will have seen how sodium reacts with water. You will recall that the sodium forms a ball which skates around the surface of the water, hissing and getting smaller and smaller. Sometimes there is an explosion. Bubbles of hydrogen gas (which may be lit) are also produced, and the water at the end contains a substance which turns red litmus blue. This is in fact sodium hydroxide. Once *all* the reactants and products have been identified, the word equation my be written:

Sodium + water → sodium hydroxide + hydrogen

[1] silicon tetrachloride, disulfur dichloride, dinitrogen pentaoxide (or pentoxide), tetraphosphorus decaoxide (or decoxide), tricarbon tetranitride, dinitrogen trisulfide.

The reactants are sodium and water and are placed to the left of the arrow. The products are sodium hydroxide and hydrogen. They are placed to the right of the arrow. The arrow may be read “change into”, “give” or “yield”. The different substances amongst the reactants and products are separated by plus signs.

Note that none of the violence of the reaction is conveyed by this equation. This is always true of chemical equations. A lot of imagination and experience is required to fill in these missing details.

In a similar way the burning of magnesium ribbon in air, which produces a white ash of magnesium oxide may be shown by:

Magnesium + oxygen → magnesium oxide

Oxygen, not air, is written as one of the reactants. This is because it is only the oxygen in the air which is used up. Note that there is only one product in this case.

5.4.2 *Writing symbolic chemical equations*

There are essentially two steps in producing a full chemical equation once a word equation has been written. First the chemical formulae are written in place of the names of the reactants and products. For the reaction between sodium and water this gives:

$$Na + H_2O \rightarrow NaOH + H_2$$

It is absolutely essential that the correct formulae are used. Unless this is done the equation will be completely incorrect.

Secondly the equation must be *balanced* as mentioned in section 5.4. Only the *amounts* of substances may be altered: formulae cannot be changed. In practice this means that the only changes that can be made in balancing an equation are putting big numbers in front of formulae. The numbers *inside* formulae cannot be changed.

It is extremely important to be systematic in balancing equations. Work steadily from left to right, dealing with each element in turn. Looking first at the Na on the left we note that there is a single Na on the right which balances it. Since the Na is balanced, we move to the next element, H. There are 2 H's on the left and a total of 3 on the right. Thus H needs to be balanced. Since there are more H's on the right than on the left we need to add more on the left. The only thing that can be done is to double the amount of H_2O by writing $2H_2O$:

$$Na + 2H_2O \rightarrow NaOH + H_2$$

This means that there are four H's on the left (and also 2 O's instead of 1) since the big "2" doubles everything in the formula that follows it.

At first this seems no better, since there are still 3 H's on the right, but now it is possible to double the NaOH by writing 2NaOH:

$$Na + 2H_2O \rightarrow 2NaOH + H_2$$

The H_2O is not doubled since that would make things even worse. We always try the easiest things first.

Now it will be seen that the H is balanced, there being 4 H's both left and right of the arrow. However this action has also increased the Na on the right. There are now 2 atoms of Na on the right and only 1 on the left still. The remedy is to double the Na on the left as well. This does not affect any other element. We now have:

$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$

We next move on to the O in H_2O on the left. There are a total of 2 O's on the left and 2 on the right. That means that O is already balanced and nothing needs to be done. Finally we should go back and check each element in turn once again. This is a simple check that should be carried out whenever an equation is written. It is guaranteed to save you marks on exams!

Looking now at the reaction of magnesium with oxygen we first write:

$$Mg + O_2 \rightarrow MgO$$

Checking each element in turn we first come to the Mg. Looking to the right we see that Mg is balanced so we move on to the next element, O. There are 2 O's on the left and only 1 on the right. The remedy is to double the MgO on the right, giving:

$$Mg + O_2 \rightarrow 2MgO$$

This means that Mg is no longer balanced and we must check it again. The remedy is to double the Mg on the left:

$$2Mg + O_2 \rightarrow 2MgO$$

A careful check reveals that every element is balanced and so the equation is complete.

Exercise

Write balanced equations in the following cases:

i) sodium chloride + silver nitrate → sodium nitrate + silver chloride

ii) magnesium + hydrochloric acid → magnesium chloride + hydrogen

iii) copper(II) nitrate → copper(II) oxide + nitrogen dioxide + oxygen

Answers are given below[1].

5.4.3 State symbols

The states of substances involved in chemical reactions should also be indicated in the equation. This is done by putting one of the following abbreviations after each formula in the equation:

(s) - solid
(l) - liquid
(g) - gas
(aq) - aqueous, i.e.. dissolved in water.

The pure liquid state (l) is distinguished from the similar state where the substance is dissolved in water (aq), although strictly the state is liquid in both cases. (aq) may indicate a substance which is normally a gas or a liquid or a solid when it is dissolved in water. (aq) is very common because many reactions are carried out in solution. Indeed many reactions will not proceed except in solution - especially those between solids.

Including state symbols in the equations for the reaction of sodium with water and of magnesium with oxygen gives:

$$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g) \text{ and}$$

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

Particularly in the case of sodium, it is evident that the sodium melts as soon as it touches the water. For this reason Na(l) would also be acceptable. Generally it is best to use the states that apply to the pure substances, or their solutions, at room temperature.

Exercise

Write the state symbols in the equations you wrote for the previous exercise. You will need to consider whether substances are soluble (see chapter 8, section 8.12.3, table 8.6) and in what form they are usually used. In some cases alternative states may be acceptable.

Answers are given below[2].

5.4.4 Ionic equations

Just as ionic substances may be represented by ionic formulae, reactions involving ions may be represented by ionic equations. These are often much simpler than so-called "molecular" equations, since many of the ions in the reactants are also present in the products and so can be omitted from the equation.

For example, when sodium chloride and silver nitrate are mixed, a dense white precipitate of silver chloride is produced whilst sodium nitrate is left behind in solution. The "molecular" equation is:

$$NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

but in fact all these substances are ionic, and in solution the ions move around independently. Sodium chloride may be represented as $Na^+(aq) + Cl^-(aq)$[3], silver nitrate as $Ag^+(aq) + NO_3^-(aq)$, and sodium nitrate as $Na^+(aq) + NO_3^-(aq)$. Silver chloride is represented as $Ag^+Cl^-(s)$ or simply as AgCl(s), since the ions in a solid are not independent. The equation may now be rewritten as:

$$Na^+(aq) + Cl^-(aq) + Ag^+(aq) + NO_3^-(aq) \rightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq)$$

We can now see that the sodium and nitrate ions are completely unaffected by the reaction, remaining in the solution throughout. They are often referred to as *spectator* ions, since they

[1] $NaCl + AgNO_3 \rightarrow NaNO_3 + AgCl$; $Mg + 2HCl \rightarrow MgCl_2 + H_2$; $2Cu(NO_3)_2 \rightarrow 2CuO + 4NO_2 + O_2$

[2] $NaCl(aq) + AgNO_3(aq) \rightarrow NaNO_3(aq) + AgCl(s)$; $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$ (or $Mg(s) + 2HCl(g) \rightarrow MgCl_2(s) + H_2(g)$)
$2Cu(NO_3)_2(s) \rightarrow 2CuO(s) + 4NO_2(g) + O_2(g)$

[3] The "+" sign between the formulae for the ions indicates that they are independent of one another, as well as that they are coming together in the reaction.

remain apart from the action, just looking on from the sidelines, so to speak.

Since the spectator ions play no part in the reaction in any real sense, they are omitted. This leaves us with:

$$Cl^-(aq) + Ag^+(aq) \rightarrow AgCl(s) \quad ... (1)$$

This equation is obviously much simpler than the "molecular" equation, and as well as being closer to the truth, it is much more general. It shows that chloride ions simply bump into silver ions and stick together to form insoluble silver chloride, which then precipitates. It is general because it shows us that any solution containing chloride ions and silver ions will behave in the same way. We would write exactly the same equation if we mixed hydrochloric acid and silver fluoride solutions.

Exercise

Write an ionic equation to represent the reaction of solutions of sodium sulfate and barium chloride ($BaCl_2$) to give a precipitate of barium sulfate and a solution of sodium chloride.

The answer is given below.

$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$When writing ionic equations it is most important that only ionic substances be represented in terms of ions (see section 5.2.3.2). Generally speaking, liquids and gases are not ionic and so water must always be represented as H_2O, with no charges. Free elements are never ionic.

For example the reaction of sodium with water:

$$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$$

is written, in ionic terms, as:

$$2Na(s) + 2H_2O(l) \rightarrow 2Na^+(aq) + 2OH^-(aq) + H_2(g) \quad ... (2)$$

In this case little simplification results since the only ionic compound present is the sodium hydroxide.

5.4.4.1 Balancing ionic equations

Ionic equations, like any equation, must be balanced. However, as well as balancing the number and type of atom, the charge must also be balanced. This means that the total charge on the left (i.e.. to the left of the arrow) must be the same as the total charge on the right.

For example, in equation (1) above, we see Ag^+ on the left and Ag on the right. These are the same atom, only differing in their charge, and so the equation is balanced with respect to silver. The same applies to chlorine.

As for charge, shown by the signs top right of the symbol, the total charge on the left is (-1) + (+1) = 0. There is no charge shown on the right, and so the total charge there is also 0. This satisfies the requirement that the total charge on the left is equal to the total charge on the right.

In equation (2) balancing the number and type of atom is a little more complicated. For example, there are two Na's on the left and two on the right, differing only in their charge. Thus sodium is balanced. The same applies to the H's and the O's. In this case there is no charge shown on the left but on the right there are two sodium ions, each with a positive charge (making +2 in total) and two hydroxide ions, each with a negative charge (making -2 in total). So overall the total charge on the right is (+2) + (-2) = 0. Once again the total charge is the same left and right.

Consider now the reaction between chlorine in solution ($Cl_2(aq)$) and iron(II) chloride to give iron(III) chloride. The "molecular" equation is:

$$2FeCl_2(aq) + Cl_2(aq) \rightarrow 2FeCl_3(aq) \quad ... (3)$$

In ionic terms this becomes, before balancing:

$$Fe^{2+}(aq) + Cl_2(aq) \rightarrow Fe^{3+}(aq) + Cl^-$$

Clearly, as far as atoms go, the Cl^- on the right must be doubled to give:

$$Fe^{2+}(aq) + Cl_2(aq) \rightarrow Fe^{3+}(aq) + 2Cl^-$$

But this is not enough, since the charges do not balance yet. The total charge on the left is +2, and the total charge on the right is

$$(+3) + 2\times(-1) = +1.$$

However, if both the Fe^{2+} and the Fe^{3+} are doubled, the atoms are still balanced, and the charges balance also:

$$2Fe^{2+}(aq) + Cl_2(aq) \rightarrow 2Fe^{3+}(aq) + 2Cl^-$$

The total charge on the left is now $2\times(+2) = +4$ and the total charge on the right is $2\times(+3) + 2\times(-1) = +4$.

5.5 Endothermic and exothermic reactions

An endothermic reaction is one which is accompanied by a drop in temperature. An exothermic reaction is one which is accompanied by a rise in temperature. This may be represented symbolically by the value of "ΔH" (pronounced "delta aitch")[1]. The endothermic reaction is represented by "ΔH positive" and the exothermic by "ΔH negative".

For example the reaction between sodium and water to form sodium hydroxide and hydrogen is exothermic. This may be expressed at the end of the equation as follows:

$$2Na(s) + 2H_2O(l) \rightarrow 2Na^+(aq) + 2OH^-(aq) + H_2(g) \text{ ... } \Delta H \text{ -ve...(2)}$$

If the reaction is carried out in a vessel equipped with a thermometer *the reading on the thermometer will rise* and the vessel will get *hot*. We say that *heat is being given out* by the reaction.

Exothermic reactions are very common and very useful as sources of heat. All combustion reactions are exothermic. You make use of them in such devices as the gas stove and the car engine.

The reaction between ammonia and iodine, which leads to the slow formation of a dark brown precipitate of nitrogen triiodide when the substances are mixed, is endothermic. The equation is written:

$$2NH_3(aq) + 6I_2(aq) \rightarrow 2NI_3(s) + 6HI(aq), \quad \Delta H \text{ +ve}$$

We say in this case that *heat is taken in* by the reaction. A thermometer in the reaction vessel would show a *fall* in its reading, and the vessel would feel *cool*.

Endothermic reactions are common too, but usually they need an external source of heat to make them proceed. For example copper(II) nitrate has to be heated over a Bunsen to make it decompose to copper(II) oxide. This is typical of an endothermic reaction, but here the heat supplied by the Bunsen completely swamps any cooling effect due to the reaction itself.

Heat is a form of energy (see chapter 4, section 4.3). Whilst it may be converted into other forms, it can neither be created nor destroyed. In an endothermic chemical reaction heat is converted into "chemical energy", which arises from the strong forces (bonding: see chapter 7) operating between atoms in compounds. The heat is taken from the materials which are undergoing reaction and so their temperature falls. This explains why we say "heat is taken in" in such cases. In an exothermic reaction chemical energy is converted into heat, and we say "heat is given out".

5.6 Types of chemical reaction

There are a number of different types of reaction. The types are not always mutually exclusive. In other words a reaction may belong to more than one type. They include decomposition, combination, reversible, dissociation, partner exchange, neutralisation, proton exchange, reduction, oxidation, and redox reactions.

5.6.1 Decomposition

This is a reaction in which one substance changes into two or more simpler substances. One example is the decomposition of lead(II) nitrate when it is heated. It turns into lead(II) oxide, nitrogen dioxide and oxygen:

$$2Pb(NO_3)_2(s) \rightarrow 2PbO(s) + 4NO_2(g) + O_2(g)$$

This may also be described as *thermal decomposition* since it is brought about by heat.

5.6.2 Combination

This is the opposite of decomposition. Here two or more relatively simple substances (elements or compounds) combine together to form a single more complex substance. One example is the combination of sulfur and iron to form iron(II) sulfide:

$$Fe(s) + S(s) \rightarrow FeS(s)$$

[1] "H" may be taken to stand for "heat", but more strictly refers to a quantity called *enthalpy*, which may be taken to be the same as heat at this stage. "Δ" is a mathematical symbol meaning "a change in", so ΔH is a change in heat, or enthalpy.

5.6.3 Reversible reactions

Generally chemical reactions are, or at least appear to be, difficult to reverse. Reactants change into products which are difficult to change back into reactants. The decomposition of lead(II) nitrate above is a good example of this. In some cases, however, a slight change in conditions (such as temperature, acidity or pressure) is sufficient to reverse the reaction.

The colour change of litmus is a good example of this. Acid will turn blue litmus red, and exposing the red litmus to base or alkali (see chapter 8, section 8.6) will change it back to blue again.

Instead of a normal arrow ($\rightarrow$) a reversible sign ($\rightleftharpoons$) is used. In the case of litmus we could write:

Blue litmus $\rightleftharpoons$ Red litmus

(The formulae of red and blue litmus are too complex to write here.)

5.6.4 Dissociation

This is very similar to decomposition, except that it is reversible. In other words decomposition is followed by recombination as the conditions change.

A good example is the dissociation of ammonium chloride. When it is heated it decomposes to form a mixture of ammonia gas and hydrogen chloride gas, but when this mixture is cooled the gases recombine to form solid ammonium chloride. It may look as if the ammonium chloride simply sublimes, but really it is a chemical change.

The equation is written with the reversible sign:

$$NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$$

5.6.5 Partner exchange (double decomposition[1]).

This type of reaction is commonly encountered in precipitation reactions involving ionic compounds. The precipitation of silver chloride when sodium chloride and silver nitrate solutions are mixed is a good example:

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

As you can see the partner of the silver ion (the nitrate ion) is exchanged with the partner of the sodium ion (the chloride ion). (See also section 5.4.4 on ionic equations.)

5.6.6 Neutralisation

This is the reaction between an acid and a basic oxide or hydroxide to form a salt plus water, for example the reaction between sulfuric acid and sodium hydroxide:

$$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$

For further information see chapter 8, section 8.11.

5.6.7 Proton exchange

These are reactions which involve the exchange of a *proton*, i.e.. a hydrogen ion, H^+. They are particularly important in acid/base reactions. They are described in chapter 8.

5.6.8 Reduction

This term was originally used to describe reactions in which a metal was extracted from its ore. In such reactions oxygen, or a similar element such as sulfur or chlorine, is removed. Thus copper(II) oxide can be *reduced* to copper by heating it in a stream of hydrogen:

$$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(g)$$

It is now realised that the underlying process is the addition of electrons to an atom, molecule, or ion. In the above example the essential process is the conversion of copper(II) ions (Cu^{2+}) present in copper(II) oxide (CuO), into copper atoms (Cu) by adding electrons to them. We write:

$$Cu^{2+} + 2e^- \rightarrow Cu$$

where e^- represents an electron.

5.6.9 Oxidation

This term describes a reaction which is opposite to reduction. Originally it was used to describe reactions in which an element combined with oxygen or a similar element such as sulfur or

[1] Not at good term, since the reaction is neither double, nor decomposition.

chlorine. For example, when copper is heated in air or oxygen, copper(II) oxide is formed:

$$2Cu(s) + O_2(g) \rightarrow 2CuO(s)$$

As with reduction it is a shift of electrons that is really important. In this case a *removal* of electrons from an atom, ion, or molecule. In the oxidation of copper to copper(II) oxide, electrons are removed from copper atoms to change them into copper(II) ions:

$$Cu(s) \rightarrow Cu^{2+} + 2e^{-}$$

All oxidations may be described in these terms.

5.7.10 Redox

In more recent times it has been realised that reduction and oxidation are processes which always occur together. When one substance is reduced, another is oxidised.

For example, when copper reacts with oxygen to form copper(II) oxide as above, the copper atoms lose electrons (i.e.. they are oxidised) which are accepted by oxygen atoms (present in oxygen molecules). Since the oxygen atoms gain electrons they are reduced. They form oxide ions (O^{2-}):

$$O_2(g) + 4e^{-} \rightarrow 2O^{2-}$$

5.7.10.1 Oxidants and Reductants

Substances which bring about the oxidation of other substances are termed *oxidants* or *oxidising agents*. Common laboratory oxidising agents include hydrogen peroxide, nitric acid acidified potassium permanganate solution, and chlorine as well as oxygen itself. Substances which bring about the reduction of other substances are termed *reductants* or *reducing agents*. Common laboratory reducing agents include hydrogen (often produced on the spot with tin and hydrochloric acid or zinc and hydrochloric acid), sodium thiosulfate, sulfur dioxide and iron(II) chloride solution.

PROBLEMS FOR CHAPTER 5

1) Give the correct chemical symbol or formula (as appropriate) for each of the following substances.
   a) sodium metal
   b) helium gas
   c) silicon
   d) zinc metal
   e) chlorine gas
   f) fluorine gas
   g) copper(II) oxide
   h) sulphur trioxide*
   i) nitrogen dioxide*
   j) diphosphorus hexoxide
   k) copper(I) chloride
   l) carbon monoxide*
   m) lithium hydroxide
   n) nitric acid
   o) iron(III) sulphate
   p) potassium sulphate
   q) potassium nitrate
   r) magnesium chloride
   s) iron(II) oxide
   t) boron trichloride
   u) ammonium sulphate
   v) hydrochloric acid
   w) nitrogen monoxide*
   x) sulphuric acid
   y) calcium hydrogencarbonate

   * Formulae marked in this way do not obey the simple rules of valency.

2) Write the empirical formula for each of the following compounds.
   a) C_6H_6
   b) C_2H_6O
   c) $C_2H_4O_2$
   d) N_2O_4
   e) H_2O_2
   f) H_2O

3) W, X, Y and Z are used here to represent elements from groups I, II, IV and VI (respectively) in the periodic table. Give the most probable formula (in terms of W, X, Y and Z) and state the likely bonding

type of each of the following compounds. List the likely physical properties of each.

a) A compound between X and Z.
b) A compound between Y and Z.
c) A compound between W and Z

4) Give the chemical names of each of the following compounds.

a) Na_2CO_3
b) Al_2S_3
c) $Ca(HSO_4)_2$
d) $CuCl_2$
e) H_2SO_4
f) PbO_2
g) $CuSO_4$
h) Na_2SO_4
i) $NaNO_3$
j) $Al(NO_3)_3$
k) $CuCl$
l) $MgCl_2$
m) MgO
n) $Zn(NO_3)_2$
o) $NaHCO_3$

****

5) Imagine that X and Y are the symbols of elements. Given the valencies of X and Y suggested below, write the most probable formula of the compound they would form in each case. Two answers are given as examples.

VALENCY OF X		1	2	3	4	5	6
VALENCY OF Y	1	XY					
	2				XY_2		
	3						
	4						
	5						
	6						
	7						

****

6) a) What is an ion? How are ions formed? What special names are given to ions with (i) a negative charge, and (ii) a positive charge?

b) Write formulae (including charges) for each of the following ions:

i) sodium
ii) chloride
ii) magnesium
iii) aluminium
iv) fluoride
vi) iodide
vii) oxide
viii) sulfide
ix) nitrate
x) ammonium
xi) calcium
xii) sulfate
xiii) carbonate
xiv) hydroxide
xv) sulfite
xvi) hydrogensulfate
xvii) iron(II)
xviii) copper(II)
xix) hydrogencarbonate
xx) iron(III)
xxi) lead(II)

c) Ionic compounds normally contain two types of ions: one positive and the other negative. For example potassium sulfate contains potassium ions and sulfate ions. Its ionic formula is $(K^+)_2SO_4^{2-}$. The relative numbers of each of these types of ion is such that the total charge on the formula is zero. Using this principal write the formula of all the compounds that might be formed between pairs of the ions listed in part (b) of this question. There are 108 of them and most can be prepared as pure solids.

7) Various chemical reactions are outlined below. Rewrite them in balanced form.

a) $H_2 + Cl_2 \rightarrow HCl$

b) $Mg + O_2 \rightarrow MgO$

c) $H_2O + F_2 \rightarrow O_2 + HF$

d) $Fe + HCl \rightarrow FeCl_2 + H_2$

e) $AgNO_3 + FeCl_3 \rightarrow AgCl + Fe(NO_3)_3$

****

8) In each of the following cases the reactants and products of a chemical reaction are given. Write a balanced chemical equation in each case. It is essential to decide on the correct formula of each reactant and each product before attempting to balance the equation.

	REACTANTS	**PRODUCTS**
a)	lead(II) oxide and hydrogen gas	water and lead metal
b)	sodium metal and oxygen gas	sodium oxide
c)	silicon and oxygen gas	silicon dioxide
d)	zinc metal and dilute sulphuric acid	Zinc sulphate and hydrogen gas
e)	carbon and steam	carbon dioxide and hydrogen gas
f)	phosphorus and chlorine gas	phosphorus trichloride
g)	aluminium metal and iron(II) oxide	aluminium oxide and iron metal
h)	calcium hydroxide and hydrochloric acid	water and calcium chloride
i)	iron(III) nitrate	iron(III) oxide, oxygen gas and nitrogen dioxide
j)	lithium metal and water	lithium hydroxide and hydrogen gas

****

9) Write a balanced chemical equation to describe each of the following reactions. As before you must decide on the formulae of the reactants and the products before attempting to balance the equation, but in this question not all the reactants and products will be mentioned. If you get stuck look for them in your practical schedules or your textbook.

a) The combustion of carbon.

b) The reaction between magnesium metal and dilute hydrochloric acid.

c) The reaction between potassium metal and water.

d) The reaction occurring when copper(II) carbonate is heated.

e) The reaction between copper(II) sulphate solution and sodium hydroxide solution.

10) a) Chemical reactions may be classified as ***endothermic*** or ***exothermic***. Explain the meaning of these terms, making it clear what would happen to the temperature of a solution in which these types of reaction was occurring. How does the symbol ΔH apply to these types of reaction.

b) The reaction between nitrogen and hydrogen to form ammonia is exothermic. Write a balanced equation for the reaction and indicate that the reaction is exothermic by writing ΔH in the appropriate place.

c) The formation of nitrogen triiodide from ammonia and iodine in solution is an endothermic reaction. What is the sign of ΔH? What would happen to the reading on a thermometer dipped into a solution in which ammonia and iodine were reacting.

d) The temperature in a solution where sulfuric acid was being reacted with sodium hydroxide was observed to rise. Was the reaction exothermic or endothermic. What is the sign of ΔH?

11) Rewrite the following equations in ionic form:

a) $AgF(aq) + NaCl(a) \rightarrow AgCl(s) + NaF(aq)$

b) $Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$

c) $Ba(NO_3)_2(aq) + Na_2SO_3(aq) + BaSO_3(s) + 2NaNO_3(aq)$