TOPIC 5.
FORMULAS AND EQUATIONS.

Types of chemical formulas.
In earlier Topics, structural formulas were given for a number of molecules such as water, ammonia and methane. From those examples, it can be seen that a structural formula shows which atoms are bonded to which other atoms in a molecule of the compound. Structural formulas are only one of several possible types of formulas which may be written for compounds. The various types of formulas are now examined.

Empirical formulas.
The simplest possible formula for any compound is its EMPIRICAL FORMULA. An empirical formula contains the simplest integer ratio of the atoms of each element present in the compound. The empirical formula of a compound can be obtained merely by analysis for each of its constituent elements. For example, the compound ethylene which was discussed in an earlier Topic contains only carbon and hydrogen, and the ratio of C:H is found by analysis to be 1 carbon atom to 2 hydrogen atoms. Therefore the empirical formula of ethylene is CH$_2$. From your knowledge of valence, it should be obvious that CH$_2$ could not be the formula for a molecule of ethylene, as C has a valence of 4 and here it would only be 2. Empirical formulas give no information of the actual numbers of atoms of each element in the molecule of the compound.

Molecular formulas.
The molecular formula of an element or compound only applies to elements or compounds that exist as molecules and it gives the actual number of each type of atom present in the molecule. It must necessarily be either the same as the empirical formula or a multiple of it. Using water as an example, the empirical formula, H$_2$O, is the same as the molecular formula. For ethylene, the molecular formula is C$_2$H$_4$, i.e. twice the empirical formula. Another molecule dealt with earlier, acetylene, has the empirical formula CH and the molecular formula C$_2$H$_2$, also twice its empirical formula. It is essential to realise that the molecular formula does not convey any structural information. Therefore it is incorrect to write H$_2$CCH$_2$ as the molecular formula of ethylene because this indicates the structure, i.e. which atoms are bonded to which other atoms. In order to write a structural formula, further information would be required. [Sometimes there is only one structure that would be possible for a given molecular formula, e.g. water must have two H atoms bonded to one central O atom in order that the valence of H and O be satisfied. In these cases, although the structure could be inferred from the molecular formula, it is not confirmed without suitable experimental evidence.]
Structural formulas.
These contain the highest level of information about a molecule. Structural formulas have already been given for a number of molecules in earlier Topics. The following table contains examples that illustrate the empirical, molecular and structural formulas of several compounds.

<table>
<thead>
<tr>
<th>Name</th>
<th>Empirical</th>
<th>Molecular</th>
<th>Structural</th>
</tr>
</thead>
<tbody>
<tr>
<td>hydrogen peroxide</td>
<td>HO</td>
<td>H₂O₂</td>
<td>H–O–O–H</td>
</tr>
<tr>
<td>water</td>
<td>H₂O</td>
<td>H₂O</td>
<td>H–O–H</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>H–H</td>
</tr>
<tr>
<td>ethane</td>
<td>CH₃</td>
<td>C₂H₆</td>
<td>H–C–C–H</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>H–H</td>
</tr>
</tbody>
</table>

Formulas for ionic compounds.
Thus far when considering ionic compounds the formula in each case has been written as the simplest whole number of each component ion that together represents the compound - i.e. the empirical formula. *It is not possible to write a molecular formula for ionic compounds because they do not exist as molecules.* Instead, ionic compounds consist of extremely large collections of cations and anions, their relative numbers being in the ratio given by the empirical formula of the compound. The ions are packed together in a crystal lattice such that the electrostatic forces of attraction are maximised while the forces of repulsion are minimised. As an example, the ionic compound sodium chloride has the empirical formula NaCl, indicating that it contains one Na⁺ ion for each Cl⁻ ion. However, the solid crystal of sodium chloride has an arrangement of cations and anions as illustrated below, whereby each Na⁺ ion is surrounded by 6 nearest neighbour Cl⁻ ions and likewise, each Cl⁻ ion is surrounded by 6 nearest neighbour Na⁺ ions. Thus no individual sodium or chloride ions can be regarded as bonded specifically to each other. Instead, the ions are packed into a crystal lattice in arrangements which depend on the radius of the cation and the anion, so that oppositely charged ions can be as close to as many of each other as possible. Another implication of the way in which ionic species pack together into crystal lattices is that ionic bonding is non-directional. The ions are simply packed so that the smallest volume which maximises the attractive forces is used. This is in contrast to the highly directional characteristics of covalent bonds.
Check your understanding of this section.
What is the difference between an empirical and a molecular formula?
Is the formula $\text{H}_2\text{CO}_3$ an empirical, molecular or structural formula?
Why can’t ionic compounds have a molecular formula?
What criteria determine how ions in an ionic crystal are packed?

Shapes of covalent molecules.
Unlike ionic bonding, covalent bonding is highly directional due to overlapping of the electron clouds of the bonded atoms. The surrounding atoms bonded to the central atom of a molecule repel each other (and also any non-bonding electron pairs in its valence level) to result in the most stable 3-dimensional arrangement. Consequently, molecules where the atoms are bonded covalently have specific **MOLECULAR SHAPES**. To illustrate this concept, consider the simplest molecule which consists of just two atoms that are covalently bonded, for example, $\text{H}_2$. As there are only two atoms involved, this molecule (and all diatomic molecules) must have a **LINEAR** shape. Molecules consisting of three covalently bonded atoms (triatomic molecules) could be either linear or **ANGULAR**, depending on how many lone pairs are present in the valence level of the central atom. An example of a linear triatomic molecule is carbon dioxide, $\text{CO}_2$, while the water molecule, $\text{H}_2\text{O}$, is an example of an angular shape. For a central atom bonded to four other atoms such as the methane molecule, $\text{CH}_4$, the H atoms are often located in a **TETRAHEDRAL** arrangement around the C atom.

The formulas used so far have sufficed to show which atoms are covalently bonded to which other atoms, but they convey little information about the shape of the molecules. For diatomic and triatomic molecules, formulas that convey their shape can be drawn on a two dimensional page, as for $\text{H}_2$, $\text{H}_2\text{O}$ and $\text{CO}_2$. For molecules containing more than three atoms, various methods have been devised to indicate the three dimensional shapes of molecules on a two dimensional page.
These representations are excellent for showing bond lengths and angles, but they give little information about the volume requirements of the atoms in the molecule. The following representations show the closest distance another molecule could approach a $\text{H}_2$, $\text{H}_2\text{O}$ or $\text{CH}_4$ molecule without serious repulsions between the electron clouds on both molecules.

**Formulas of the elements as they occur in nature.**

In an earlier Topic it was stated that most elements do not normally occur as individual atoms (monatomic species), but usually they occur as collections of atoms bonded together in some manner. For example, all of the seventh family of elements, the halogens, occur as diatomic molecules of molecular formulas $\text{F}_2$, $\text{Cl}_2$, $\text{Br}_2$, $\text{I}_2$. Elements from other families may have more than two atoms in their molecules - for example sulfur may be in the molecular form $\text{S}_2$, $\text{S}_6$ or $\text{S}_8$, depending on the temperature. Some of the other elements exist as infinitely large covalently bonded collections of atoms, and the concept of a molecule is not really appropriate to these elements. In particular, one form in which the element carbon occurs in nature is as diamond. A diamond is a giant "molecule" consisting of carbon atoms, each bonded to four other carbon atoms in a tetrahedral arrangement around each other, (referred to as **NETWORK COVALENT BONDING**) as illustrated below.

The concept of molecules of metals is also not appropriate, as these elements consist of infinitely large collections of atoms which are bonded by outer electrons which are very mobile and constantly jump from atom to atom, leading to what is termed metallic bonding. See Topic 3 for a detailed discussion of metallic bonding.
Consequently, it can be seen that the use of a molecular formula for an element as it occurs in nature is rather limited. Instead, most elements are represented, for example in equation writing, by their empirical formula. Being elements, the empirical formula is always just the symbol for the element. Thus, normally carbon would be represented as C, sulfur as S, sodium as Na, etc in situations where a formula for a "molecule" is needed such as in a chemical equation. The following is a complete list of the few elements which would usually be represented by a molecular formula in such cases: \( \text{H}_2, \ \text{N}_2, \ \text{O}_2, \ \text{F}_2, \ \text{Cl}_2, \ \text{Br}_2, \ \text{I}_2. \) In addition, the noble gases He, Ne, Ar, Kr, Xe, Rn occur only as individual atoms.

### Check your understanding of this section.
- Why do covalent compounds have fixed shapes?
- What shapes could occur for a triatomic covalent compound?
- How do metal atoms bond in the solid?
- Why are metals good conductors of electricity?
- How are atoms in a diamond bonded to each other?

### Chemical reactions.
Chemical reactions occur for one of two possible reasons:
(i) the products formed by the reaction contain less energy than the reactants, and all systems tend to proceed spontaneously to the lowest energy state, or
(ii) enough energy is supplied to reactants to force them to proceed to form products, even though the products have higher energy than the reactants.
Reactions are the result of the breaking of some or all of the bonds between atoms in the reactants and the subsequent formation of new bonds to form the products.

### Chemical equations.
Regardless of why any given reaction occurs, it can be represented in a number of ways. For example, hydrogen gas combines with oxygen gas to form water which is also a gas at the temperature that the reaction reaches due to released energy. This reaction could be expressed in a word equation,

\[
\text{hydrogen} + \text{oxygen} \rightarrow \text{water}
\]

It is much less cumbersome and more informative if the equation is written using symbols rather than words. Thus,

\[
\text{H}_2 + \text{O}_2 \rightarrow \text{H}_2\text{O}
\]

tells us that hydrogen molecules react with oxygen molecules to form water molecules, and gives the formulas for all species involved.
However, note that not all the oxygen shown on the left appears on the right - i.e. the equation is not balanced. Given that water molecules must contain two H and one O atoms and that hydrogen and oxygen molecules are both diatomic, then the simplest equation which provides a proper balance so that all atoms on the left are also present on the right would be

\[ 2H_2 + O_2 \rightarrow 2H_2O \]

Writing the equation in this way states that 2 molecules of hydrogen react exactly with 1 molecule of oxygen to form 2 molecules of water. Even more information can be included in this molecular equation by indicating the physical state of each component by attaching an appropriate suffix, (g), (l) or (s) but a problem arises when species are in solution. In this case, reactants and products were all specified as being in the gaseous state.

\[ 2H_2(g) + O_2(g) \rightarrow 2H_2O(g) \]

In another example, hydrogen and chlorine react to form the compound hydrogen chloride, all species being gases. The balanced molecular equation would be

\[ H_2(g) + Cl_2(g) \rightarrow 2HCl(g) \]

The carbohydrate glucose is oxidised to carbon dioxide and water in biological systems. The overall reaction can be represented by the following molecular equation in which the physical states have been omitted.

\[ C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O \]

However, not all substances exist as molecules - as seen in earlier topics many compounds exist as ionic solids wherein no individual ion is bonded to another specific ion but instead, cations and anions are arranged in a crystal lattice in which each ion is surrounded by a number of ions of opposite charge as nearest neighbours. Consequently it would not be correct to refer to, for example, a “molecule” of NaCl. Therefore the equation for the formation of sodium chloride produced when solid sodium metal reacts with chlorine gas to form solid sodium chloride might be better referred to as a formula equation rather than a molecular equation. As this term is equally applicable to both molecules and ionic compounds, it will be used throughout the rest of these notes instead of the term “molecular equation”. For the above reaction, the equation would be

\[ 2Na(s) + Cl_2(g) \rightarrow 2NaCl(s) \]
In this example, note how the empirical formula Na is used for the metal sodium, as discussed previously. Likewise, the empirical formula C is used for carbon in the next example in which carbon solid is burnt in oxygen to form carbon dioxide.

\[ \text{C(s)} + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) \]

The use of the term “formula equation” is also more appropriate when one encounters a reaction that takes place in solution. For example, when solutions of sodium chloride and silver nitrate are mixed, solid silver chloride forms and sodium nitrate remains in solution as ions provided that its maximum solubility has not been exceeded. Sometimes chemistry texts try to cover this situation by introducing another suffix, (aq), to indicate a substance in solution, but for reasons to be given later, this is not appropriate and attempts to show the physical states in this situation are ambiguous and best not included in the equation. The formula equation for this reaction would be

\[ \text{AgNO}_3 + \text{NaCl} \rightarrow \text{AgCl} + \text{NaNO}_3 \]

For such reactions in solution, a more descriptive type of equation known as an ionic equation will be dealt with in Topic 6. Which type of equation is best to use depends upon the particular context in which it is being applied.

**Mass balance.**

In these examples, note that they all have what is termed **mass balance**, i.e. identical numbers of each type of atom appear on both sides of the equation. While chemical bonds have been broken between some atoms and new bonds formed in the reaction, **no atoms have been destroyed or created in the process**. The essential final step in writing an equation for a reaction is to **check that the same number of each type of atom on the left hand side of the equation is also present on the right hand side**.

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**Check your understanding of this section.**

Write the formula equation for the combustion of methane gas in oxygen gas.

What criterion must apply to any balanced equation?

Write the formula equation for the reaction between potassium and iodine to form potassium iodide.
Objectives of this Topic.
When you have completed this Topic, including the tutorial questions, you should have achieved the following goals:

1. Understand the concepts of empirical formula; molecular formula; structural formula.
2. Understand the manner in which ions in ionic solids pack together to form crystals and that ionic bonds are non-directional.
3. Understand why the formula for ionic compounds is always an empirical formula rather than a molecular formula.
4. Understand why covalently bonded molecules have specific shapes arising from the directional nature of the bonding.
5. Be able to interpret the various ways in which covalent molecules can be represented to indicate bond angles, molecular shape and volume requirements.
6. Understand the concepts of elements occurring as network covalent bonded solids or as metallic bonded solids.
7. Know those elements which are normally represented as diatomic molecules.
8. Be able to write a formula equation for a given word equation, or to balance a skeleton formula equation.

Recommended follow up Chemcal modules: [Chemcal will not function on new browsers but will with older versions of IE or any that allow shockwave.]
Section: Properties of Molecules

Module: Molecular Shape and Structure
Topics covered: Geometric arrangements of groups in atoms; VSEPR theory.

Module: Applying VSEPR Theory
Topics covered: VSEPR theory applied to molecules with non-bonding pairs and multiple bonds.
SUMMARY

Compounds can be represented by various types of formulas. The simplest is the empirical formula which can be obtained for any pure compound by basic analytical methods and requires no other information about the compound. The empirical formula is the simplest whole number ratio of the atoms of each component element present in that compound. All pure compounds have an empirical formula.

A molecular formula applies only to elements and compounds that exist as molecules and it lists the actual number of atoms of each element present in a molecule of that substance. It is a simple multiple of the empirical formula. Molecular formulas do not convey any information about the structure of the molecule. Structural formulas for compounds indicate the arrangement of the component atoms within the molecule.

Ionic compounds exist as solid crystal lattices in which the cations and anions pack together so as to maximise their attractive forces and minimise their repulsive forces. The particular arrangement adopted in a given ionic compound depends on the relative sizes of the cation and anion. As no cation is specifically bonded to a particular anion but instead is shared by a number of neighbouring anions, then there is no molecule of an ionic compound. Consequently all ionic compounds are represented by just an empirical formula giving the simplest ratio of the component ions. Another consequence is that ionic bonding has no directional characteristics, the crystal shape being determined by the most efficient packing of the ions.

Covalent bonding is highly directional as it depends on the overlap of electron orbits. This leads to covalent molecules having specific shapes which can be represented on a two-dimensional surface by several methods to convey bond angles and also volume requirements. One molecular shape which is particularly important is the tetrahedral arrangement often found when four atoms are bonded to a central atom as for example in methane.

Only the noble gases occur in nature as individual atoms. A few elements occur as diatomic molecules (H$_2$, N$_2$, O$_2$, F$_2$, Cl$_2$, Br$_2$, I$_2$). Otherwise, all elements occur naturally as larger aggregates of atoms. Non-metals frequently consist of large numbers of atoms held together by covalent bonds - termed network covalent bonding. Metals and most non-metals do not have a molecular formula as the number of component atoms is not fixed and may be enormous. These elements are represented by their symbol alone where a formula is required.

Chemical reactions occur either because the products have less energy than the reactants or because sufficient energy is supplied externally to the reactants to force them to convert to products. Reactions can be conveniently represented by means of
chemical equations of several types.

A word equation gives the names of the reactants and products. Word equations are convenient when the species involved have large or complicated or unknown formulas.

Formula equations give the formulas of all the component reactants and products and should be balanced so that all atoms on the reactant side are also present on the products side. It is a fundamental law of chemistry that atoms are not created or destroyed in reactions. Physical states of the species (gas, liquid or solid) may also be shown as a suffix on each formula.

**TUTORIAL QUESTIONS - TOPIC 5.**

1. Balance the following formula equations.

   (i) \( \text{K} + \text{Cl}_2 \rightarrow \text{KCl} \)
   
   (ii) \( \text{Ba} + \text{H}_2\text{O} \rightarrow \text{Ba(OH)}_2 + \text{H}_2 \)
   
   (iii) \( \text{HCl} + \text{O}_2 \rightarrow \text{H}_2\text{O} + \text{Cl}_2 \)
   
   (iv) \( \text{Li} + \text{N}_2 \rightarrow \text{Li}_3\text{N} \)
   
   (v) \( \text{Al} + \text{O}_2 \rightarrow \text{Al}_2\text{O}_3 \)
   
   (vi) \( \text{H}_2\text{S} + \text{O}_2 \rightarrow \text{H}_2\text{O} + \text{SO}_2 \)
   
   (vii) \( \text{C}_2\text{H}_6 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} \)
   
   (viii) \( \text{Fe}_2\text{O}_3 + \text{C} \rightarrow \text{Fe} + \text{CO}_2 \)
   
   (ix) \( \text{C}_6\text{H}_{14} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} \)

2. Balance each of the following formula equations:

   (i) \( \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l}) \)
   
   (ii) \( \text{N}_2(\text{g}) + \text{H}_2(\text{g}) \rightarrow \text{NH}_3(\text{g}) \)
   
   (iii) \( \text{Na}(\text{s}) + \text{Br}_2(\text{l}) \rightarrow \text{NaBr}(\text{s}) \)
   
   (iv) \( \text{Fe}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{Fe}_2\text{O}_3(\text{s}) \)
   
   (v) \( \text{Hl}(\text{g}) \rightarrow \text{H}_2(\text{g}) + \text{I}_2(\text{g}) \)
3. Write balanced formula equations for each of the following reactions. Physical states of reactants and products need not be shown.

(i) The combustion of ethane gas, C\(_2\)H\(_6\), in excess air to form carbon dioxide and water.

(ii) The decomposition of solid calcium carbonate by heat to form calcium oxide and carbon dioxide.

(iii) The decomposition of hydrogen peroxide to form water and oxygen

(iv) The reaction of iron with chlorine to form iron(III) chloride.

(v) The formation of ammonia from nitrogen and hydrogen.

(vi) The action of heat on potassium nitrate to form potassium nitrite and oxygen gas.

(vii) The action of heat on lead(II) nitrate to form lead(II) oxide and the gases, oxygen and nitrogen dioxide (nitrogen(IV) oxide).

(viii) The action of oxygen on calcium metal to form calcium oxide.
(ix) The reaction of glucose, $C_6H_{12}O_6$, to form ethanol, $C_2H_5OH$, and carbon dioxide gas.

(x) The reaction of copper(II) oxide with hydrogen gas to form copper metal and water.

(xi) The reaction of zinc sulfide with oxygen gas to form zinc oxide and sulfur dioxide gas.

(xii) The action of heat on sodium hydrogen carbonate to form sodium carbonate, carbon dioxide gas and water.

(xiii) The burning of magnesium metal in oxygen to form magnesium oxide.

(xiv) The action of heat on copper(II) hydroxide to form copper(II) oxide and water.

(xv) The reaction of calcium hydroxide with carbon dioxide to form calcium carbonate and water.

4. Give the empirical formulas for each of the following molecules:

   (i) $N_2O_4$  (ii) $C_4H_{10}$  (iii) $C_2H_6O$
   (iv) $N_2H_4$  (v) $P_4O_{10}$  (vi) $C_4H_8O_4$
   (vii) $NSF_3O_3$  (viii) $C_2H_4$  (ix) $C_3H_6$
   (x) $C_6H_{12}$  (xi) $CS_2$  (xii) $C_6H_{12}O_6$

5. Explain why an ionic compound is always represented by its empirical formula.
6. Define each of the following terms:
Empirical formula, molecular formula, structural formula, network covalent bonding, metallic bonding.

7. Chemical Crossword No. 4. Cryptic clues to any element or compound.
CHEMICAL CROSSWORD No. 4
CRYPTIC CLUES TO ANY ELEMENT OR COMPOUND

RULES:
1. The symbol for each element is written in a single square, even if the symbol consists of two letters.
2. Where a compound's formula contains brackets, each is written in its own square.
3. For metals in compounds, the same oxidation state must apply to both the across and down formula where that metal is common to both.

Example:

<table>
<thead>
<tr>
<th></th>
<th>A</th>
<th>I</th>
<th></th>
<th>S</th>
<th>O</th>
<th>4</th>
<th></th>
</tr>
</thead>
</table>

ACROSS
1. in epsom salts
3. copper(I) cyanide
6. gas used in San Quentin
9. backward water
11. original breathalyser ingredient
13. lithium bromide
14. caesium carbonate
15. nitrogen dioxide
16. aluminium nitride
17. carbon tetrachloride
18. potassium hydrogensulfide
19. blonde's secret
20. protects from UV light
21. element; main component of air
23. antacid remedy
25. Bond's enemy; a pollutant from cars
26. used in cloud seeding
27. element; caused Hindenberg disaster
28. rubidium hydride
29. rust converter; a weak acid
30. another cobalt(II) compound
31. element; red liquid
33. phosphorus pentoxide

DOWN
2. component of acid rain
4. toxic gas in car exhaust
5. ammonium bromide
7. main acidic component of vinegar
8. rubidium hydroxide
10. selenium trioxide
12. chromium(III) phosphate
13. lithium hydrogensulfite
14. caesium amide
15. nitrogen trichloride
17. main greenhouse gas
18. potassium carbonate
21. nitrogen triiodide
22. lead(IV) oxide
23. in iodised table salt
24. copper(I) hydrogensulfide
25. hydride of nitrogen
26. silver dihydrogenphosphate
28. a halide of rubidium
29. acid with one less oxygen atom than nitric acid
30. a cobalt(II) compound
32. silver salt in photography
35. hello; a Group VII hydride
37. rotten egg gas
34. calcium phosphide 39. sulfur analogue of 10 down
35. hydride of bromine 40. manganese dioxide
36. lead(II) oxide 41. sodium compound; in oven
37. aluminium sulphate 42. greenhouse gas from flatulent
cleaners and draino
cows
38. kidney stones 43. sodium compound; in oven
44. component of teeth and bones 44. phosphine
45. lead(II) oxide 43. greenhouse gas from flatulent
cows
47. strontium fluoride 48. chromium(III) chloride
46. a hypochlorite in bleaches 49. a fertilizer
47. strontium phosphide 50. element; green gas
49. potassium iodide 51. on every dinner table
50. element; green gas 52. sodium hydride
51. sodium cyanide 53. a chill in the air
52. hypo in photography 54. potassium bromide
53. nickel(ll) chloride 55. hydrogen peroxide,
54. potassium bromide 55. structural formula
55. water, structural formula 56. sodium permanganate
56. sodium phosphate 57. compound in alumina
57. occasional acid 58. nitric acid
58. occasional acid 59. element; purple solid
59. sodium hydride 60. laughing gas
60. laughing gas 61. calcium selenide
61. calcium selenide 62. element; the most reactive
62. element; the most reactive
63. hydride of selenium 64. aluminium sulfite
64. aluminium sulfite
64. aluminium sulfite
66. laughing gas
ANSWERS TO TUTORIAL TOPIC 5

1. (i)  \[ 2K + Cl_2 \rightarrow 2KCl \]

As chlorine occurs in the form of diatomic molecules, it must be shown as \( Cl_2 \) in the equation. Thus two Cl atoms must appear on the products side, so two KCl is required there. Now there are two K atoms on the right hand side but only one on the left, so it is necessary to show two K atoms as reactants.
(ii) \[ \text{Ba} + 2\text{H}_2\text{O} \rightarrow \text{Ba(OH)}_2 + \text{H}_2 \]

Because the formula of barium hydroxide on the right hand side requires two H atom and two O atoms, the minimum number of water molecules on the left hand side that can supply this would be two \( \text{H}_2\text{O} \) molecules. This would leave an excess of two H atoms on the right hand side and, as hydrogen is a diatomic gas, they would be shown as a hydrogen molecule, \( \text{H}_2 \).

(iii) \[ 4\text{HCl} + \text{O}_2 \rightarrow 2\text{H}_2\text{O} + 2\text{Cl}_2 \]

(iv) \[ 6\text{Li} + \text{N}_2 \rightarrow 2\text{Li}_3\text{N} \]

(v) \[ 4\text{Al} + 3\text{O}_2 \rightarrow 2\text{Al}_2\text{O}_3 \]

(vi) \[ 2\text{H}_2\text{S} + 3\text{O}_2 \rightarrow 2\text{H}_2\text{O} + 2\text{SO}_2 \]

In the unbalanced equation, there are 2 H atoms in \( \text{H}_2\text{S} \) and they require 1 O atom to form a molecule of \( \text{H}_2\text{O} \) on the right. The single S atom of \( \text{H}_2\text{S} \) on the left requires 2 O atoms to form a molecule of \( \text{SO}_2 \) on the right, so the total number of O atoms required on the left is 3. This would mean using \( 3/2 \) molecules of \( \text{O}_2 \) on the left in the equation \( \text{H}_2\text{S} + 3/2\text{O}_2 \rightarrow \text{H}_2\text{O} + \text{SO}_2 \). To avoid a fractional number of \( \text{O}_2 \) molecules appearing, multiply throughout by 2.

(vii) \[ 2\text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O} \]

(viii) \[ 2\text{Fe}_2\text{O}_3 + 3\text{C} \rightarrow 4\text{Fe} + 3\text{CO}_2 \]

Each iron(III) oxide on the left will result in two Fe atoms on the right hand side. Then there are three O atoms on the left to be removed as carbon dioxide, so \( 3/2 \) C atoms are needed on the left and \( 3/2 \) \( \text{CO}_2 \) molecules would result on the right. In order that all the coefficients are integers, multiply throughout by two.

(ix) \[ 2\text{C}_6\text{H}_{14} + 19\text{O}_2 \rightarrow 12\text{CO}_2 + 14\text{H}_2\text{O} \]

Starting with one hexane molecule on the left would result in six carbon dioxide molecules plus seven water molecules appearing on the right. The total number of O atoms on the right would then be 19 and they would be required to be shown on the left hand side. But oxygen occurs as a diatomic gas so it must be shown as \( \text{O}_2 \) on the left. One could write the equation with \( 19/2 \) \( \text{O}_2 \) molecules on the left or multiply throughout by two in order to avoid non-integer coefficients, as in the answer given above.
2. (i) $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O(1)}$

(ii) $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$

(iii) $2\text{Na}(\text{s}) + \text{Br}_2(\text{l}) \rightarrow 2\text{NaBr}(\text{s})$

(iv) $4\text{Fe}(\text{s}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{Fe}_2\text{O}_3(\text{s})$

(v) $2\text{HI}(\text{g}) \rightarrow \text{H}_2(\text{g}) + \text{I}_2(\text{g})$

(vi) $3\text{Mg}(\text{s}) + \text{N}_2(\text{g}) \rightarrow \text{Mg}_3\text{N}_2(\text{s})$

(vii) $2\text{P}(\text{s}) + 5\text{F}_2(\text{g}) \rightarrow 2\text{PF}_5(\text{s})$

(viii) $\text{O}_2(\text{g}) + 2\text{SO}_2(\text{g}) \rightarrow 2\text{SO}_3(\text{g})$

(ix) $4\text{HCl}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{Cl}_2(\text{g}) + 2\text{H}_2\text{O(g)}$

(x) $\text{Fe}_3\text{O}_4(\text{s}) + 4\text{H}_2(\text{g}) \rightarrow 3\text{Fe}(\text{s}) + 4\text{H}_2\text{O(g)}$

(xi) $2\text{Pb(NO}_3)_2(\text{s}) \rightarrow 2\text{PbO}(\text{s}) + 4\text{NO}_2(\text{g}) + \text{O}_2(\text{g})$

One lead(II) nitrate on the left would result in one lead(II) oxide plus two nitrogen dioxide molecules on the right hand side. The total number of O atoms on the left would be six while only five are so far accounted for on the right, leaving a single O atom to appear there as oxygen gas. Because oxygen is a diatomic element, it needs to be shown as $\text{O}_2$ on the right hand side and this would require the equation to be multiplied by two throughout.

(xii) $\text{C}_6\text{H}_{12}\text{O}_6(\text{s}) + 6\text{O}_2(\text{g}) \rightarrow 6\text{CO}_2(\text{g}) + 6\text{H}_2\text{O(1)}$

Starting with one glucose molecule on the left would result in the formation of six carbon dioxide molecules plus six water molecules on the right hand side. The total number of O atoms on the right would then be 18. However, glucose already contains six O atoms so only 12 additional O atoms are required on the left, shown as six $\text{O}_2$ molecules.

3. (i) $2\text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O}$

(ii) $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$

(iii) $2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2$

(iv) $2\text{Fe} + 3\text{Cl}_2 \rightarrow 2\text{FeCl}_3$

(v) $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$

(vi) $2\text{KNO}_3 \rightarrow 2\text{KNO}_2 + \text{O}_2$

(vii) $2\text{Pb(NO}_3)_2 \rightarrow 2\text{PbO} + \text{O}_2 + 4\text{NO}_2$
(viii) \[ 2\text{Ca} + \text{O}_2 \rightarrow 2\text{CaO} \]
(ix) \[ \text{C}_6\text{H}_{12}\text{O}_6 \rightarrow 2\text{C}_2\text{H}_5\text{OH} + 2\text{CO}_2 \]
(x) \[ \text{CuO} + \text{H}_2 \rightarrow \text{Cu} + \text{H}_2\text{O} \]
(xi) \[ 2\text{ZnS} + 3\text{O}_2 \rightarrow 2\text{ZnO} + 2\text{SO}_2 \]
(xii) \[ 2\text{NaHCO}_3 \rightarrow \text{Na}_2\text{CO}_3 + \text{CO}_2 + \text{H}_2\text{O} \]
(xiii) \[ 2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO} \]
(xiv) \[ \text{Cu(OH)}_2 \rightarrow \text{CuO} + \text{H}_2\text{O} \]
(xv) \[ \text{Ca(OH)}_2 + \text{CO}_2 \rightarrow \text{CaCO}_3 + \text{H}_2\text{O} \]

4. (i) \[ \text{NO}_2 \]
(ii) \[ \text{C}_2\text{H}_5 \]
(iii) \[ \text{C}_2\text{H}_6\text{O} \]
(iv) \[ \text{NH}_2 \]
(v) \[ \text{P}_2\text{O}_5 \]
(vi) \[ \text{CH}_2\text{O} \]
(vii) \[ \text{NSF}_3\text{O}_3 \]
(viii) \[ \text{CH}_2 \]
(ix) \[ \text{CH}_2 \]
(x) \[ \text{CH}_2 \]
(xi) \[ \text{CS}_2 \]
(xii) \[ \text{CH}_2\text{O} \]

5. Ionic compounds do not exist as molecules but instead consist of ions packed into a crystal lattice in such a manner that attractions between oppositely charged ions are maximised and repulsions between ions having charges with the same sign are minimised. Thus each cation has up to 6 nearest neighbour anions and vice versa, so no individual cation and anion can be regarded as bonded to each other in the way that atoms are in covalent molecules. Consequently, in ionic compounds the concept of a molecular formula is not relevant and only an empirical formula which gives the simplest whole number ratio of the cations and anions is possible.

6. Empirical formula: The simplest whole number ratio of atoms or ions in a compound.

Molecular formula: Applicable only to compounds that contain covalently bonded molecules, and is the actual number of each atom in the molecule. The molecular formula is a simple multiple of the empirical formula and conveys no structural information.

Structural formula: Based on a molecular formula, the structural formula shows which atoms are bonded to which other atoms.

Network covalent bonding: Present in elements or compounds where an extremely large number of atoms are covalently bonded to form the solid.
Metallic bonding: Atoms of metals in the solid state are bonded by mobile outer electrons that readily move between atoms. Metallic bonding explains the unique properties of metals such as their electrical conduction and malleability.

CHEMICAL CROSSWORD No. 4
CRYPTIC CLUES TO ANY ELEMENT OR COMPOUND

Mg S O 4 Cu C N H C N
O Rb O H 2 H Se
K 2 Cr 2 O 7 4 3 O
P H Li Br Cs 2 C O 3
N O 2 H Al N O
C Cl 4 K H S H 2 O 2
O 3 2 O N 2 H Pb
2 Na H C O 3 I Cu N O
Ag I O 3 H H 2
Rb H H 3 P O 4 Co S O 3
Br 2 N H Ag
P 2 O 5 Cu 3 P 2 H Br
Pb O 2 H O I
4 Al 2 ( S O 4 ) 3 Mn
Na S O O
Ca C 2 O 4 Ca 3 ( P O 4 ) 2
H H H H
4 Na Cl O Sr 3 P 2
Cr 2 3 F K I
Cl 2 Na C N Na 2 S 2 O 3
3 Ni Cl 2 K H P
3 O Br H O H
Na 3 P O 4 Al H I O 4
Mn 2 2 N 2 O Ca
O F O O H 2 Se
4 Al 2 ( S O 3 ) 3