TOPIC 4.

COVALENT COMPOUNDS: bonding, naming, polyatomic ions.

Covalent bonding.
In Topic 3, one type of chemical bond, the ionic bond, was discussed. Ionic bonding is essentially electrostatic attraction between ions. The ions examined in that Topic were shown to form because the energetically favoured outer electron arrangement of the noble gases is attained in the process. However, there is another means by which an atom can attain at least a share of the number of electrons required for it to become isoelectronic with a noble gas without having the electrons transferred totally to or from that atom. In the type of chemical bond called the COVALENT BOND, two atoms share electrons which are spread over both of the bonded atoms.

Consider the simplest possible molecule, the \( \text{H}_2 \) molecule. In Topic 2, it was shown that each hydrogen atom has a nucleus consisting of a single proton around which one electron orbits. If two hydrogen atoms come sufficiently close together, there is a repulsion between the two nuclei, as they both have the same positive electrical charge and like charges repel. Similarly, there is a repulsion between the two electrons. At the same time however, there is attraction between each nucleus and the two electrons, and this attraction allows the nuclei to remain close together in spite of the repulsions present. Calculations show that the electrons become more concentrated in the region between the nuclei than elsewhere, and this accounts for the stability of the bond. Recall that the hydrogen atom's electron is in an orbit that can contain a maximum of two electrons. Formation of the \( \text{H}_2 \) molecule results in each \( \text{H} \) atom having a share of 2 electrons and therefore achieving the stable structure of the inert gas, helium. This is the basis for all covalent bonds, the resultant molecule having a lower energy than the individual atoms.

Thus \( \text{H}_2 \) is the stable hydrogen molecule rather than \( \text{H}_3^+, \text{H}_4^- \), etc. Note that unlike ionic bonds, there has not been a complete transfer of charge from one atom to another, so no ions form. A covalent bond is often represented as a line between the two bonded atoms. The line indicates two shared electrons (a pair) between the atoms which it joins. Thus the hydrogen molecule can be shown as \( \text{H}–\text{H} \). Just as the charges on the component ions in ionic compounds are not normally shown, so too the bonds between covalently bonded atoms are often understood in the formulas of such molecules and so the molecule of hydrogen would usually be written as \( \text{H}_2 \) rather than \( \text{H}–\text{H} \).
Covalent bonding in the water molecule.

Now consider the molecule of water. Earlier the formula H₂O was given for this molecule. Consider the electron arrangement around the O atom. As its atomic number is 8, then there are 8 electrons surrounding the O nucleus. Of these electrons, there are 2 in the first energy level and 6 in the second energy level. As discussed in Topic 2, a further 2 electrons are required in the second level to bring it up to 8 electrons and thereby attain the same electron arrangement as in the neon atom. This arrangement is present in the O²⁻ ion which is isoelectronic with neon.

An alternative to the complete transfer of 2 electrons to the O atom to form the O²⁻ ion is for the O atom instead to share the electrons of two hydrogen atoms by forming 2 covalent bonds. Each H atom provides a share of 1 electron to the O atom and in turn gains a share of 1 oxygen electron to form a covalent H–O bond just as in the H₂ molecule. The O atom, by forming 2 such bonds to H atoms, finishes up with the same number of electrons as in neon. Note that electrons from different atoms are indistinguishable, and once in a covalent bond, cannot be regarded as belonging to either atom, but shared by both.

The water molecule can be written as \( \text{H–O–H} \) to show that each H atom is bonded to the central O atom. Each dash represents one pair of electrons in a covalent bond in this type of formula which is called a structural formula.

Valence.

From this example, it can be seen that O atoms in covalent compounds will always be associated with the formation of two covalent bonds in order to obtain the desired total of 8 electrons in the outer level. In the case of ionic compounds, the O atom will be present as the O²⁻ ion and would be bonded for example to two 1+ charged ions. The ability of the O atom to form two bonds in compounds, covalent or ionic, is expressed in a quantity called the valence of that atom. Thus O has a valence = 2, while H always forms just 1 bond and therefore has a valence = 1. In ionic compounds, the valence is the number of electrons gained (anions) or lost (cations) in forming the ionic species. Note that valence has no sign attached, and is simply the number of bonds which that atom can form in compounds, be they ionic or covalent. While the atoms of some elements (e.g. the first two groups in Table 2 of Topic 1) show only a single valence state in their compounds, there are many other elements that can have more than one valence state. This has already been noted in Topic 3 for the case of ionic valence of metals such as iron (Fe²⁺, Fe³⁺), copper (Cu⁺, Cu²⁺) and tin (Sn²⁺, Sn⁴⁺). Among non-metals, although hydrogen only has a valence state = 1 and oxygen only has a valence state = 2, many other non-metals can exist in more than one valence state in their compounds. For example, sulfur forms the stable covalent oxides SO₂ and SO₃ where the valence of the S atom is 4 and 6 respectively. Nitrogen, while only able to have a valence of 3 in ionic compounds as in the nitride ion, N³⁻, shows a range of
valencies in covalent compounds. Thus oxides of nitrogen include NO, N₂O, NO₂, and N₂O₃. Why these variable valence states can exist will be dealt with in first year courses later in the year.

**The methane molecule.**
Consider another example - the molecule of methane which has the formula CH₄. As carbon has atomic number = 6, then the electron structure in this atom must be 2 electrons in the first energy level (filled) + 4 in the second energy level. To attain the electron structure of helium by the loss of 4 electrons or the electron structure of neon by the gain of 4 electrons is energetically unfavourable. Instead, carbon attains the neon structure more easily by forming 4 covalent bonds, as for example in methane, through covalent bonding to four H atoms. Thus in CH₄, the C atom has a share of 8 electrons in its outer level using the original 4 electrons from the outer level of the C atom + the 4 electrons from the four bonded H atoms, making it isoelectronic with neon. Carbon has a valence = 4 in its compounds.
The structural formula for CH₄ is illustrated below.

**The ammonia molecule.**
The atom of nitrogen, an element from the fifth group, requires 3 electrons in order to be isoelectronic with neon. In Topic 2, it was seen that this could be achieved by the N atom gaining 3 electrons to form the N³⁻ ion. An alternative to forming an ion is for the N atom to form 3 covalent bonds instead. The covalent compound, ammonia, has the formula NH₃. By forming the 3 covalent bonds to three H atoms, the N atom attains a total of 8 outer level of electrons and thus becomes isoelectronic with neon. The nitrogen atom therefore has a valence = 3 in ammonia. The covalent bonds in ammonia can be represented as shown in the structural formula below where again each bond is represented by a line corresponding to two shared electrons between the bonded N and H atoms.

**Bonding and non-bonding electrons.**
All the electrons in the outer level of any bonded atom are referred to as **VALENCE LEVEL ELECTRONS**. Returning to the molecule H₂O, it can be seen that of the 8 valence level electrons which the O atom in the molecule ultimately attains, only 4 are involved in bonds and these are called **BONDING ELECTRONS**. The 4 remaining valence level electrons are called **NON-BONDING ELECTRONS**, and because electrons usually are located as pairs in atoms, they may also be called **LONE PAIRS**. In some representations known as **LEWIS STRUCTURES**, all the valence level electrons - both bonding and non-bonding - are shown as dots. Thus the H₂O molecule can also be represented as follows.
Similarly, the NH₃ molecule has in its valence level 3 pairs of bonding electrons, represented by the dashes in the previous diagram, plus 2 electrons which constitute a lone pair, as shown in the next diagram.

\[
\text{H} : \text{N} : \text{H}
\]

\[\text{H} \quad \text{H} \quad \text{lone pair}
\]

\[\text{bonding electrons}
\]

Note that in molecules such as H₂O and NH₃, the central atom in both cases has a pair of electrons in the first energy level, but these electrons experience a very strong attractive force to the protons in the nucleus and do not participate in forming bonds to other atoms. It is only electrons in the outer or valence level that can overlap with the orbits of other atoms to form covalent bonds. Consequently the inner electrons are not shown in the Lewis structures.

**Check your understanding of this section.**

- How does a covalent bond differ from an ionic bond?
- Why would CH₅ be an unlikely formula for a compound?
- Why is the valence of the O atom always = 2 in its compounds?
- What do you understand by the terms valence level electrons, bonding electrons and non-bonding electrons?
- Show the structural formula of the water molecule using two different methods.

### Double bonds.

Now consider another molecule, carbon dioxide, whose formula earlier was given as CO₂. Given that the C atom needs 4 more electrons to achieve the stable neon structure, it still needs to form 4 covalent bonds as in CH₄. From the example of water, the O atom forms 2 covalent bonds. In terms of valence, C has a valence = 4 while O has a valence = 2. Therefore in the carbon dioxide molecule, each O atom will need to join to the C atom by 2 bonds rather than 1 bond. This covalent bond is called a **DOUBLE BOND** and it involves the sharing of a total of 4 electrons between each O atom and the C atom. The molecule can be represented as O=C=O, where again each dash represents a pair of electrons in a bond. As discussed above, sometimes the remaining non-bonded electrons are also shown. In this case, each O atom has two pairs of non-bonding electrons. Various representations of the carbon dioxide molecule are illustrated in the following diagrams.

The carbon dioxide molecule. The C atom is bonded by a double bond to each O atom. 

\[\text{O} = \text{C} = \text{O}
\]

Two double bonds
Another common molecule where the atoms are joined by a double bond is the molecule of the element oxygen, \( \text{O}_2 \).

![Oxygen molecule](image)

a) Orbit diagram and b) Lewis structure of the oxygen molecule, \( \text{O}_2 \). Both oxygen atoms share two pairs of electrons which make a double bond between them.

**Ethylene.**

Compounds which contain only the elements carbon and hydrogen are called **HYDROCARBONS**. Where the molecules of the hydrocarbon contain only **SINGLE BONDS** such as in \( \text{CH}_4 \), they are said to be **SATURATED** compounds. Ethylene, \( \text{C}_2\text{H}_4 \), is an example of a hydrocarbon which contains a double bond between two \( \text{C} \) atoms and is an **UNSATURATED HYDROCARBON**. The structural formula for a molecule of ethylene then shows that each \( \text{C} \) atom is bonded by single bonds to 2 \( \text{H} \) atoms and by a double bond to the other \( \text{C} \) atom. The following diagrams give two representations of this molecule. The right hand illustration shows the volume requirements of the bonded atoms.

![Ethylene molecule](image)

**Triple bonds.**

Some atoms such as those of elements from the fifth group can not attain the stability of the noble gas structure in forming a covalent molecule unless they gain a share of 3 electrons. As seen previously, the nitrogen atom has 5 electrons in its outer electron level and can attain the stable outer level electron arrangement of neon by forming 3 single covalent bonds to three \( \text{H} \) atoms as in the molecule of ammonia, \( \text{NH}_3 \). In the molecule of the element nitrogen, \( \text{N}_2 \), each \( \text{N} \) atom gains a share of the required additional three electrons from the other \( \text{N} \) atom by forming 3 covalent bonds in a triple bond with it. By these means, each \( \text{N} \) atom gains a share of 8 electrons in its valence level just like Ne. The nitrogen molecule, \( \text{N}_2 \), can thus be represented by the structural formula \( \text{N}=\text{N} \). Molecules which contain triple bonds are also called **UNSATURATED**.

![Nitrogen molecule](image)

a) Orbit diagram and b) Lewis structure of the nitrogen molecule. Both nitrogen atoms share three pairs of electrons which make the triple bond between them.
Acetylene.
The unsaturated hydrocarbon of formula $\text{C}_2\text{H}_2$, acetylene, is another example of a molecule containing a triple bond, in this case joining the two C atoms. Again each of the C atoms has attained 8 outer level of electrons as in the Ne atom by forming a total of 4 covalent bonds - one to a H atom and 3 to the other C atom as illustrated below.

$$\text{H} - \text{C} \equiv \text{C} - \text{H}$$

Check your understanding of this section.
How many electrons are involved in (i) a double bond and (ii) a triple bond?
What is a hydrocarbon?
What do the terms saturated and unsaturated mean?
Why would the structural formula $\text{O} - \text{O}$ be unlikely for the $\text{O}_2$ molecule?

Naming covalent compounds.
Compounds of two elements (BINARY COMPOUNDS) are named as two words. In an earlier Topic, the ionic compounds (which were all binary compounds) were named with the cation as the first word and the anion as the second word. As there is no cation in covalent compounds, the first word in the name of a compound is usually the name of the element from the lower group number. The second word in the name is as for the anion in ionic compounds, viz the ending "ide" added to a stem derived from the name of the element. The number of each component atom present is indicated by either of two methods. The first method is simply to use prefixes such as "mono", "di", "tri", "tetra" etc with each part of the name, deleting "mono" if no ambiguity occurs. For example, the compound $\text{CO}_2$ was named as carbon dioxide previously in the notes. Other examples include

- $\text{SF}_6$ sulfur hexafluoride
- $\text{BF}_3$ boron trifluoride
- $\text{PCl}_3$ phosphorus trichloride
- $\text{SO}_3$ sulfur trioxide
- $\text{N}_2\text{O}_3$ dinitrogen trioxide
- $\text{N}_2\text{O}$ usually named as nitrous oxide
- $\text{NO}$ usually named as nitric oxide
- $\text{SO}_2$ sulfur dioxide

The second method will be discussed in more detail later, but it involves writing a quantity called the "oxidation number" (which is similar to, but not the same as the valence of an atom in a compound) as part of the name of any component atom where ambiguity might occur. This quantity is written as a Roman numeral in brackets as part of the word with no space ahead of it. For example, in phosphorus trichloride the valence of P is 3, so $\text{PCl}_3$ could equally well be named as phosphorus(III) chloride. [This concept has already been encountered when naming ionic compounds where the metal ion could have more than one possible charge - for example copper(II) oxide ($\text{CuO}$) and copper(I) oxide ($\text{Cu}_2\text{O}$).]

Other examples include $\text{SF}_6$ sulfur(VI) fluoride, $\text{PCl}_5$ phosphorus(V) chloride,
AsBr₃, arsenic(V) bromide and P₂O₅, phosphorus(V) oxide.

In a later Topic the concept of oxidation number and how it actually differs from valence will be examined. Note also that the valence of elements such as sulfur, phosphorus and arsenic in these examples is larger than expected when compared with the related elements oxygen and nitrogen which were dealt with earlier. The reason is that as the orbit number used for the valence level electrons increases, there are more spaces for additional electrons which in turn allow higher valence states to exist. This aspect will be treated in more detail later.

**Trivial or common names.**

Some compounds known for a long time prior to any knowledge of molecular composition are still more often referred to by their trivial or common names. Examples include the hydrides of oxygen (water) and nitrogen (ammonia) already encountered. Others include silane (SiH₄), phosphine (PH₃), Arsine (AsH₃) and stibnine (SbH₃). Carbon has the exceptional ability to form an almost unlimited number of covalent compounds giving rise to the term ORGANIC CHEMISTRY which has developed its own systematic nomenclature which is explained in Appendix 1.

**Check your understanding of this section.**

Which of the following are binary compounds: AgCl, H₂O, H₂SO₄, CH₂CF₂, CO₂, HF, MnO₂, HNO₃, NaCl, HClO₄, BaCl₂?

Using two methods, name the compound PCl₅.

How would the compounds MnO and MnO₂ be named in order to avoid ambiguity?

**Polyatomic ions.**

So far the compounds which have been encountered have all consisted of only two elements - i.e. they were all binary compounds. Those that were ionic, in each instance, contained ions derived from a single atom of an element which had gained or lost 1 or more electrons, such as Na⁺ or Cl⁻. However, there are also ions which consist of more than one atom, usually of different elements, bonded together by covalent bonds, and these ions are called POLYATOMIC IONS. Mostly they are anions, although there is one very commonly encountered polyatomic cation, the ammonium ion, which has the formula NH₄⁺. The structure of this ion consists of four H atoms covalently bonded to the same N atom, with a +1 charge on the overall species which results, as illustrated below.

\[
\begin{array}{c}
\text{H} \\
\vdots \\
\text{H : N : H} \\
\text{H}
\end{array}
\]

\[11 \text{ protons} + 10 \text{ electrons} \quad \therefore \text{has a} +1 \text{ charge} \]

The ammonium ion

An example of a polyatomic anion is the hypochlorite ion, ClO⁻, part of the ionic compound sodium hypochlorite, NaClO, which is a common household bleach. This anion has a single covalent bond joining the Cl and O atoms. Counting up the total protons and electrons, there is one more electron than there are protons in the two atomic nuclei, so this anion has a \(-1\) charge overall.

\[
\begin{array}{c}
\text{Cl} \\
\vdots \\
\text{O} \\
\vdots \\
\text{\text{ClO}⁻} \\
\end{array}
\]

\[25 \text{ protons} + 26 \text{ electrons} \quad \therefore \text{has a} −1 \text{ charge.} \]
Examples of polyatomic ions you should commit to memory are given in the following table.

<table>
<thead>
<tr>
<th>TABLE 3 - SOME POLYATOMIC ANIONS</th>
</tr>
</thead>
<tbody>
<tr>
<td>nitrate  NO$_3^-$</td>
</tr>
<tr>
<td>nitrite NO$_2^-$</td>
</tr>
<tr>
<td>cyanide CN$^-$</td>
</tr>
<tr>
<td>phosphate PO$_4^{3-}$</td>
</tr>
<tr>
<td>amide NH$_3^-$</td>
</tr>
<tr>
<td>chromate CrO$_4^{2-}$</td>
</tr>
</tbody>
</table>

Formulas for compounds containing these polyatomic ions obey the same rule as applies to the binary ionic compounds dealt with earlier: the total charge on the compound must equal zero. For example, the compound sodium nitrate contains one Na$^+$ ion for each NO$_3^-$ ion present, and its formula is therefore NaNO$_3$ (leaving out the charges in the formula). Other examples of polyatomic ions in compounds follow.

potassium carbonate K$_2$CO$_3$  ammonium chloride NH$_4$Cl  barium sulfate BaSO$_4$  aluminium nitrite Al(NO$_2$)$_3$

Calcium phosphate Ca$_3$(PO$_4$)$_2$  ammonium cyanide NH$_4$CN  sodium hydroxide NaOH  potassium permanganate KMnO$_4$

Potassium dichromate K$_2$Cr$_2$O$_7$  iron(III) nitrate Fe(NO$_3$)$_3$

The following examples use compounds containing polyatomic ions to illustrate the method shown in Topic 3 to ensure the formula is balanced.

1. **Write the formula for potassium sulfate.**

   The ions present are K$^+$ and SO$_4^{2-}$

   Thus the formula is K$_2$SO$_4$

   Checking: $2 \times [+1] + 1 \times [-2] = (+2) + (-2) = 0$

2. **Write the formula for iron(II) phosphate.**

   The ions present are Fe$^{2+}$ and PO$_4^{3-}$

   Thus the formula is Fe$_3$(PO$_4$)$_2$

   Checking: $3 \times [+2] + 2 \times [-3] = (+6) + (-6) = 0$

Note that the phosphate part of the formula is enclosed in brackets before the subscript "2" is written. This is necessary with polyatomic ions but is not needed with monatomic ions. Thus barium nitrate is written as Ba(NO$_3$)$_2$ while barium chloride is BaCl$_2$. 
The mercury(I) ion.
The mercury(I) ion (Hg\(_{2}^{2+}\)) is different from all the other metal cations encountered in that it consists of two atoms of the element joined by a covalent bond and having an overall 2+ charge which gives an average of 1+ for the charge on each Hg atom. However it is not possible to separate the Hg\(_{2}^{2+}\) ion into individual Hg\(^+\) ions - instead when the covalent bond between the Hg atoms breaks, an atom of the element Hg and the cation Hg\(_{2}^{2+}\) results. Thus mercury(I) is a polyatomic metallic cation.

Check your understanding of this section.
How do ions such as NO\(_{3}^{-}\) differ from ions such as Cl\(^-\)?
What are the formulas for sodium sulfide and sodium sulfate?
What type(s) of bonding would be present in the compound potassium nitrate?
When sodium sulfate is dissolved in water, why doesn’t it release the ions Na\(^+\), S\(^{2-}\), and O\(^{2-}\)?

Objectives of this Topic.
When you have completed this Topic and its associated tutorial questions, you should have achieved the following goals:

1. Understand the concept of atoms attaining the stable electron arrangement by sharing electrons in covalent bonds.
2. Be able to represent the covalent bonds in structural formulas for the molecules of H\(_2\)O, CH\(_4\), NH\(_3\), CO\(_2\), H\(_2\)CCH\(_2\), O\(_2\), N\(_2\), C\(_2\)H\(_2\) and other related molecules.
3. Understand the terms double bond; triple bond; saturated compound; unsaturated compound; valence level electrons; isoelectronic.
4. Understand the concept of valence and be able to work out the valence of various elements in compounds.
5. Know the difference between bonding electrons and lone pairs of electrons.
6. Be able to name covalently bonded binary compounds.
7. Understand that polyatomic ions are covalently bonded atoms which bear an overall ionic charge and that they behave like simple ions in forming ionic compounds.
8. Know the names and formulas of some common polyatomic ions, and be able to write correct formulas for ionic compounds which include them.

SUMMARY
A covalent bond arises when two atoms share pairs of electrons as distinct from the complete transfer of electrons which occurs in ionic bonding. For a covalent bond to form between two atoms, they must be sufficiently close so that their outer electron orbits can overlap. By this method, two hydrogen atoms for example can form a molecule, represented as H\(_2\), in which each H atom has a share of two electrons and is isoelectronic with the atom of the noble gas, helium.
Covalent molecules ranging from simple diatomics up to giant macromolecules can be built up by many atoms bonding in this way. In the water molecule, $\text{H}_2\text{O}$, the O atom becomes isoelectronic with the Ne atom through the formation of the two covalent bonds to H atoms. Similarly, C atoms in methane, $\text{CH}_4$, and N atoms in ammonia, $\text{NH}_3$, become isoelectronic with Ne atoms by forming four and three covalent bonds with H atoms respectively.

The term valence is used to indicate the number of bonds (either ionic or covalent) that an atom forms in a compound. The valence of H atoms is always 1 as there is no room for more than one additional electron in its outer electron orbit. The valence of O is always 2 as there is only room for two more electrons in its outer electron orbit. Similarly carbon and nitrogen show a valence of 4 and 3 respectively, determined by the number of vacancies in their atom’s outer orbit. The valence of an element can be related to the group in which it is located. For elements of some groups, only a single valence state is available while for others, especially the non-metals after the first member, two or more valencies may be displayed.

Electrons that participate in a covalent bond are called bonding electrons or, as they always occur as pairs, bonding pairs. However, in the water molecule for example, the outer electron orbit (or valence level) electrons include four non-bonding electrons (lone pairs). The N atom of ammonia has three bonding pairs and one lone pair of electrons in its valence level. In this theory of bonding, the electrons in the orbits closer to the nucleus than the valence level are not considered as participating in covalent bond formation.

A Lewis structure represents the electrons of the valence level in molecules as dots. An alternative representation shows a bonding pair as a line joining the two atoms. In some molecules such as $\text{CO}_2$ the valence can only be satisfied if the bonded atoms share more than one pair of electrons. In this example, the C atom forms two covalent bonds to each of the O atoms and thus gains a share of the required additional four electrons to become isoelectronic with Ne. Each O atom gains a share of an additional two electrons to also become isoelectronic with Ne. When two atoms share four electrons in this way, the covalent bond is called a double bond and can be represented by two lines between the bonded atoms. Similarly the $\text{O}_2$ molecule contains a double bond between two O atoms and the ethylene molecule, $\text{H}_2\text{CCH}_2$, contains a double bond between the two C atoms. Compounds containing only carbon and hydrogen atoms are called hydrocarbons.

Some molecules contain three covalent bonds between a pair of atoms. For example in the $\text{N}_2$ molecule each N atom shares three electrons with the other N atom so that both become isoelectronic with neon. A common hydrocarbon which contains a triple bond is acetylene, $\text{HCCH}$, in which each C atom shares three of its electrons with the other C atom. Compounds that contain double or triple bonds are called unsaturated.

Covalent binary compounds are named in a similar way to ionic compounds except that there is no cation to name first or anion to name second. Instead, the element which comes from the lower group number is named first with no special ending while the second word as in ionic compounds, is derived from a stem from the element with the “ide” ending attached. Prefixes such as di, tri, tetra are used with both words if ambiguity would exist.

Some ions, mainly anions, consist of more than one atom bonded by covalent bonds. These polyatomic ions behave as single entities and combine with ions of the opposite charge to form ionic compounds in the same way as other simple cations and anions.
1. Explain the meaning of each of the following terms:
   (i) covalent bond
   (ii) valence
   (iii) isoelectronic
   (iv) bonding electrons
   (v) lone pair electrons
   (vi) double bond
   (vii) triple bond
   (viii) saturated compound
   (ix) unsaturated compound
   (x) hydrocarbon
   (xi) structural formula
   (xii) valence level electrons
   (xiii) Lewis structures
   (xiv) single bond
   (xv) diatomic molecule
   (xvi) triatomic molecule
   (xvii) binary compound

2. Give the valence displayed by each of the underlined atoms in each of the following compounds. **Do not attempt to draw structural or Lewis formulas.**

   (i) HCl  
   (ii) NH₃  
   (iii) H₂S  
   (iv) SF₄  
   (v) NaCl  
   (vi) CaO  
   (vii) Al₂O₃  
   (viii) CuO  
   (ix) Cu₂O  
   (x) SF₆  
   (xi) KBr  
   (xii) CH₄  
   (xiii) H₂O  
   (xiv) NF₃  
   (xv) HF  
   (xvi) CCl₄  
   (xvii) PH₃  
   (xviii) SiCl₄  
   (xix) AlF₃  
   (xx) SrI₂
3. Draw structural formulas for the following covalently bonded molecules, representing each covalent bond by a line.

   (i)  H₂O  (xi)  HCCH
   (ii) CH₃CH₃  (xii) HCN
   (iii) O₂  (xiii) HCl
   (iv) N₂  (xiv) CH₃CH₂OH
   (v)  CH₃OH  (xv) CCl₄
   (vi) F₂  (xvi) H₂S
   (vii) NH₃  (xvii) H₂Se
   (viii) CO₂  (xviii) NF₃
   (ix)  CH₄  (xix) NCl₃
   (x)  CH₂CH₂  (xx) NBr₃

4. Name the following compounds.

   (i)  CO₂
   (ii) HBr
   (iii) PBr₃
   (iv) CCl₄
   (v)  OF₂
   (vi) N₂O₃
   (vii) NI₃
   (viii) SO₂
   (ix)  SO₃
   (x)  SeO₂
   (xi) CS₂
   (xii) NH₃
   (xiii) NCl₃
   (xiv) SF₆
   (xv) H₂S
5. Draw Lewis structures for each of the following covalent molecules, using two dots to represent an electron pair. Show bonding and non-bonding valence level electrons on all atoms. Using the kits provided, make models of those compounds marked with *.

(i) water
(ii) ammonia
(iii) methane* (see page 3)
(iv) carbon dioxide*
(v) ethylene* (see page 5)
(vi) acetylene* (see page 6)
(vii) nitrogen trifluoride
(viii) carbon tetrachloride*
(ix) silicon tetrachloride
(x) nitrogen tribromide
(xi) carbon disulfide
(xii) hydrogen chloride*
(xiii) hydrogen sulfide
(xiv) nitrogen trichloride

Consult the tear-out Data Sheet at the back of this book for a more extensive list of ions if required for Questions 6 and 7.

6. Write formulas for the following compounds, all of which contain polyatomic ions.

(i) strontium sulfate
(ii) copper(II) nitrate
(iii) calcium carbonate
(iv) potassium hydroxide
(v) iron(III) cyanide
(vi) zinc nitrite
(vii) ammonium phosphate
(viii) rubidium sulfite
(ix) sodium peroxide
(x) potassium dichromate
(xi) sodium amide
(xii) potassium permanganate
(xiii) lead(II) chromate
(xiv) aluminium sulfate
(xv) lithium phosphate
(xvi) barium sulfate
(xvii) strontium hydrogencarbonate
(xviii) magnesium sulfite
(xix) chromium(III) hydroxide
(xx) mercury(II) nitrite
(xxi) ammonium sulfate
(xxii) cobalt(II) cyanide
(xxiii) silver(I) phosphate
(xxiv) sodium carbonate
(xxv) lead(II) hydrogencarbonate
(xxvi) copper(II) nitrate
(xxvii) rubidium nitrite
(xxviii) caesium peroxide
(xxix) aluminium chromate
(xxx) tin(II) hydroxide

7. Give the name for each of the following compounds, all of which contain polyatomic ions.

(i) \( \text{Li}_2\text{CO}_3 \)
(ii) \( \text{Zn(NO}_3\text{)}_2 \)
(iii) \( \text{Mg(MnO}_4\text{)}_2 \)
(iv) \( \text{CaSO}_3 \)
(v) \( \text{Al}_2(\text{SO}_4)_3 \)
(vi) \( \text{Cu(NH}_2\text{)}_2 \)
(vii) \( \text{BaO}_2 \)
(viii) \( \text{Cr(NO}_2\text{)}_3 \)
(ix) \( \text{ZnCrO}_4 \)
(x) \( \text{NiCO}_3 \)
(xi) Cs$_2$Cr$_2$O$_7$
(xii) NaNH$_2$
(xiii) Ag$_3$PO$_4$
(xiv) Mn(CN)$_2$
(xv) KHCO$_3$
(xvi) Mg(OH)$_2$
(xvii) (NH$_4$)$_3$PO$_4$
(xviii) Sr(NO$_3$)$_2$
(ix) Cu(CH$_3$CO$_2$)$_2$
(xx) Rb$_2$CO$_3$
(xi) K$_2$Cr$_2$O$_7$
(xii) CaO$_2$
(xiii) Al(OH)$_3$
(xiv) Fe$_2$(SO$_4$)$_3$
(xv) Co$_3$(PO$_4$)$_2$
(xvi) AgCH$_3$CO$_2$
(xvii) Hg(NO$_3$)$_2$
(xviii) NaMnO$_2$
(xix) K$_2$CrO$_4$
(x) Al(CN)$_3$

8. Using the worksheets on pages 18 and 19, do the CAL module called "Nomenclature" which is on your memory stick. The worksheets can only be used in conjunction with the programme as it corrects your work as you proceed. There is no answer sheet available. Start with the Inorganic-easier section and then the Inorganic-harder section. The programme starts with the first entry in the column on the left then the first entry in the column on the right, alternating from asking a name then a formula. Some of the polyatomic anions have not been shown in Topic 4 but they can be found in the data section at the end of this book. You should practise with this CAL module as often as you can during and after the course.

9. Chemical Crossword No. 3. Elements and binary compounds of non-metals.

10. Self Help Problems module “Nomenclature” Q 7,8,11,12.
CHEMICAL CROSSWORD No. 3
ELEMENTS AND BINARY COMPOUNDS OF NON-METALS

RULES:
Where the symbol for an element consists of two letters, both the upper and lower case letters should be written in the same box. For example, Se O 2 Br 2
Where a subscript is required in the formula, it is entered in its own box.
The oxidation state of any atom need not be the same where that atom is common to both the across and down formulas.

CLUES.
ACROSS
1. selenium
3. sulfur hexafluoride
6. silicon
7. arsenic(V) iodide
9. nitrogen(II) oxide
10. nitrogen(III) bromide
11. phosphorus trichloride
12. hydrogen telluride
14. arsenic(III) bromide
15. antimony(V) sulfide
16. antimony pentafluoride
18. nitrogen triiodide
19. hydrogen selenide
20. antimony(V) bromide
21. carbon tetrabromide
23. phosphorus trifluoride
24. sulfur dioxide
25. nitrogen dioxide
26. arsenic(III) iodide
27. beryllium fluoride
29. fluorine (molecule)
30. antimony(V) chloride
31. selenium(VI) chloride
32. nitrogen (molecule)
33. arsenic(V) bromide
34. tellurium dioxide
35. antimony trichloride
36. phosphorus(III) bromide
37. phosphorus(V) chloride
38. phosphorus triiodide
40. phosphine
41. dinitrogen tetraoxide
44. boron
46. hydrogen sulfide
47. beryllium
48. carbon monoxide
51. boron trifluoride

DOWN
2. phosphorus(V) iodide
4. xenon
5. sulfur trioxide
8. krypton
9. nitrogen trichloride
11. phosphorus pentabromide
12. hydrogen bromide
13. silicon sulfide
14. arsenic trichloride
15. antimony(V) selenide
16. antimony pentaiodide
17. carbon tetrachloride
18. nitrogen tribromide
19. hydrogen peroxide
21. carbon dioxide
22. bromine (molecule)
23. phosphorus pentaiodide
24. sulfur hexafluoride
25. nitrogen trifluoride
26. arsenic(V) chloride
27. beryllium chloride
28. arsenic(V) sulfide
30. antimony triiodide
31. selenium trioxide
33. arsenic trichloride
34.tellurium(VI) bromide
35. antimony triiodide
37. phosphorus(V) oxide
39. ammonia
41. nitrogen(III) oxide
42. helium
43. water
45. thallium
48. carbon tetrafluoride
49. sulfur hexachloride
50. phosphorus
52. carbon tetraiodide 51. boron trichloride
53. chlorine (molecule) 52. carbon disulfide
54. silicon tetrachloride 54. silicon dioxide
55. arsenic(III) oxide 55. arsenic(III) chloride
56. nitrous oxide 57. oxygen difluoride
59. tellurium trioxide 58. fluorine (molecule)
60. iodine (molecule) 59. tellurium hexachloride
61. selenium dioxide 61. selenium hexafluoride
62. hydrogen chloride 63. stibine
64. hydrogen fluoride 64. hydrogen iodide
66. silane 65. oxygen (molecule)
67. sulfur hexafluoride 67. sulfur tetrabromide
68. hydrogen (molecule) 70. arsine
69. gallium 71. phosphorus pentafluoride
72. methane 74. ozone
73. xenon tetraoxide 77. germanium
75. arsenic(V) fluoride 78. antimony
76. indium 79. argon
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ANSWERS TO TUTORIAL TOPIC 4

1. (i) Covalent bonds result from the sharing of electrons by the bonded atoms.

(ii) The valence of a covalently bonded atom is the number of covalent bonds which join that atom to other atoms in a molecule or polyatomic ion. For simple ionic compounds, the valence of each atom is the numerical value of the ionic charge which the atom carries. **Unlike ionic charge, the valence of an atom has no associated + or – sign.**

(iii) Two atoms or ions of different elements are isoelectronic if they have the same arrangement of electrons in the shells around their nuclei. This condition will necessarily arise if the atoms have the same number of electrons. e.g. Ne, F⁻ and Na⁺ are all isoelectronic, having 10 electrons each.

(iv) Bonding electrons are the electrons involved in any covalent bond between two atoms.

(v) Lone pair electrons on an atom (also known as non-bonding electrons) are any valence level electrons which are not bonding electrons. [Although such non-bonding electrons are usually found in pairs, there are some instances where a single non-bonding electron occurs on an atom, for example on the N atom of NO].

(vi) Double bonds occur when there are two covalent bonds between the same two atoms, involving a total of 4 bonding electrons.

(vii) Triple bonds occur when there are three covalent bonds between the same two atoms, involving a total of 6 bonding electrons.

(viii) A compound is saturated if there are only single bonds present in its molecule and no double or triple bonds.

(ix) An unsaturated compound contains at least one double or triple bond in its molecule.

(x) Hydrocarbons are compounds that contain only atoms of hydrogen and carbon in their molecules.

(xi) A structural formula shows which atoms are bonded to which other atoms.

(xii) Valence level electrons are all those electrons in the outer shell of an atom. They include both the bonding and non-bonding electrons.

(xiii) Lewis structures of atoms, molecules or ions are representations of the valence level electrons in which all the outer shell electrons are shown as being either bonding (between atoms) or non-bonding (elsewhere on the atom). The electrons are usually represented by dots in Lewis diagrams.

(xiv) A single bond between two atoms is a covalent bond in which there are just two bonding electrons.

(xv) A diatomic molecule is one in which there are only two atoms present per molecule - e.g. O₂.
(xvi) A triatomic molecule is one in which there are three atoms present per molecule - e.g. \( \text{CO}_2 \).

(xvii) A binary compound is any compound, ionic or covalent, in which there are two different elements present - e.g. \( \text{NaCl, CO}_2, \text{H}_2\text{O} \).

2. **Note**: Valence does not have a sign - it is simply a number indicating the number of covalent bonds attached to a given atom or the magnitude of the charge if an ion.

\[
\begin{array}{cccccccc}
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(vi) & 2 & (vii) & 3 & (viii) & 2 & (ix) & 1 & (x) & 6 \\
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(xvi) & 4 & (xvii) & 3 & (xviii) & 4 & (xix) & 3 & (xx) & 2 \\
\end{array}
\]

3.

(i) \( \text{H-O-H} \)    (ii) \( \text{H-C-C-H} \)    (iii) \( \text{O=O} \)  
(iv) \( \text{N≡N} \)  (v) \( \text{H-C-O-H} \)  (vi) \( \text{F-F} \) 
(vii) \( \text{N-H} \)  (viii) \( \text{O=C=O} \)  (ix) \( \text{H-C-H} \) 
(x) \( \text{H-C≡C-H} \)  (xi) \( \text{H-C≡C-H} \)  (xii) \( \text{H-C≡N} \) 
(xiii) \( \text{H-Cl} \)  (xiv) \( \text{H-C-C-O-H} \)  (xv) \( \text{Cl-C-Cl} \) 
(xvi) \( \text{H-S-H} \)  (xvii) \( \text{H-Se-H} \)  (xviii) \( \text{F-N-F} \) 
(xix) \( \text{Cl-N-Cl} \)  (xx) \( \text{Br-N-Br} \)
4. The answers given here for covalent compounds generally use the prefix method of indicating the ratios of the atoms in the molecule as this method is in more common use. In some instances the answers give the alternative method using oxidation numbers as well to illustrate their use.

(i) carbon dioxide
(ii) hydrogen bromide
(iii) phosphorus tribromide or phosphorus(III) bromide
(iv) carbon tetrachloride
(v) oxygen difluoride
(vi) dinitrogen trioxide
(vii) nitrogen triiodide
(viii) sulfur dioxide
(ix) sulfur trioxide
(x) selenium dioxide
(xi) carbon disulfide
(xii) ammonia
(xiii) nitrogen trichloride
(xiv) sulfur hexafluoride or sulfur(VI) fluoride
(xv) hydrogen sulfide

5.
6. Note the use of brackets around polyatomic ions when more than one of them is present in the formula - e.g. Cu(NO₃)₂.

(i) SrSO₄  
(ii) Cu(NO₃)₂  
(iii) CaCO₃  
(iv) KOH  
(v) Fe(CN)₃  
(vi) Zn(NO₂)₂  
(vii) (NH₄)₃PO₄  
(viii) Rb₂SO₃  
(ix) Na₂O₂  
(x) K₂Cr₂O₇  
(xi) NaNH₂  
(xii) KMnO₄  
(xiii) PbCrO₄  
(xiv) Al₂(SO₄)₃  
(xv) Li₃PO₄  
(xvi) BaSO₄  
(xvii) Sr(HCO₃)₂  
(xviii) MgSO₃  
(xix) Cr(OH)₃  
(xx) Hg(NO₂)₂  
(xxi) (NH₄)₂SO₄  
(xxii) Co(CN)₂  
(xxiii) Ag₃PO₄  
(xxiv) Na₂CO₃  
(xxv) Pb(HCO₃)₂  
(xxvi) Cu(NO₃)₂  
(xxvii) RbNO₂  
(xxviii) Cs₂O₂  
(xxix) Al₂(CrO₄)₃  
(xxx) Sn(OH)₂
7. (i) lithium carbonate
   (ii) zinc nitrate
   (iii) magnesium permanganate
   (iv) calcium sulfite
   (v) aluminium sulfate
   (vi) copper(II) amide
   (vii) barium peroxide
   (viii) chromium(III) nitrite
   (ix) zinc chromate
   (x) nickel(II) carbonate
   (xi) caesium dichromate
   (xii) sodium amide
   (xiii) silver phosphate
   (xiv) manganese(II) cyanide
   (xv) potassium hydrogencarbonate
   (xvi) magnesium hydroxide
   (xvii) ammonium phosphate
   (xviii) strontium nitrate
   (xix) copper(II) acetate
   (xx) rubidium carbonate
   (xxi) potassium dichromate
   (xxii) calcium peroxide
   (xxiii) aluminium hydroxide
   (xxiv) iron(III) sulfate
   (xxv) cobalt(II) phosphate
   (xxvi) silver(I) acetate
   (xxvii) mercury(II) nitrate
   (xxviii) sodium permanganate
   (xxix) potassium chromate
   (xxx) aluminium cyanide
CHEMICAL CROSSWORD No. 3
ELEMENTS AND BINARY COMPOUNDS OF NON-METALS

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