TOPIC 3.

IONIC COMPOUNDS: formation, formulas and naming.
METALLIC BONDING.

Chemical bonding.
When elements combine to form compounds, a chemical bond holds the atoms together. There are two basic types of chemical bonds possible in compounds, IONIC BONDS and COVALENT BONDS. Both types of bond involve a redistribution of electrons between the bonded atoms. In addition METALLIC BONDS occur within a pure metal. Ionic and metallic bonding will now be examined while covalent bonding will be dealt with in Topic 4.

Formation of ions and ionic bonds.
From Topic 2 it was seen that by removing electrons from the atoms of some elements (metals), their outer electron level can be made to be identical to that of the nearest noble gas, this being the energetically favourable condition which is responsible for the lack of reactivity of noble gases. When atoms gain or lose electrons to attain the same electron arrangement as the noble gas, they are said to become isoelectronic with the noble gas. However, the process of formation of cations from metal atoms requires energy input known as its ionization energy, and is not spontaneous as many books lead to believe. In this process, such atoms acquire a positive electrical charge equal in magnitude to that of the electrons removed and they are then called cations. After the required number of electrons has been removed such that the resulting cation is isoelectronic with a noble gas, the energy needed to remove further electrons is too great and so this does not occur. Hence with only a relatively small energy input required, Na atoms can form Na\(^+\) cations in which state they are isoelectronic with the noble gas neon, but sodium cannot form Na\(^{2+}\) or Na\(^{3+}\) cations as the energy requirements would be excessive. With a few exceptions, in the process of becoming isoelectronic with a noble gas, no more than three electrons can be removed from an atom in the formation of cations as the ionization energy required is too great and the large positive charge would result in an unstable cation. To form compounds that would require the transfer of more than 3 electrons between reacting atoms, a different type of bonding - covalent bonding - is generally used.

Elements from some other groups (non-metals) can attain this energetically favourable state through their atoms gaining enough electrons to acquire the same electron arrangement as the nearest noble gas. In this process these elements gain a negative charge and are called anions. Again a limit of not more than three electrons can be transferred and once the noble gas electron arrangement has been acquired, no further electrons can be gained. Hence Cl atoms can form Cl\(^-\) anions (isoelectronic with argon) but not Cl\(^{2-}\) or Cl\(^{3-}\) anions.
Thus an atom of sodium can lose the single electron located in the third energy level to form an Na\(^+\) ion which then has 8 electrons in its outer energy level just like the neon atom. An atom of chlorine can gain an electron to form the Cl\(^-\) ion which then has 8 electrons in its outer energy level just like argon. As electrons which are lost by an atom of one element must be accepted by an atom of another element, then sodium atoms and chlorine atoms can form Na\(^+\) and Cl\(^-\) ions by transferring a single electron per atom from Na to Cl. When this transfer occurs, the resulting ions of Na\(^+\) and Cl\(^-\), having opposite electrical charges, will be attracted to each other, releasing energy in the process, to form the solid compound sodium chloride. This electrostatic attraction is the basis for ionic bonding.

Electrostatic attraction causes the cations and anions produced to form a crystal in which many ions are closely packed, arranged so that cation/anion attractions are maximised and anion/anion repulsions are minimised. [See Topic 5 page V-2.]

While the resulting compound of formula Na\(^+\)Cl\(^-\) is also known commonly as “salt”, the term SALT is a general one which applies to any ionic compound, not just to sodium chloride. Note that there would be no sodium or chlorine atoms present in the compound, only the ions derived from them.

**Combining ratios and formulas of ionic compounds.**
The ionic compound formed must be electrically neutral, so the ratio of the number of Na\(^+\) ions to the number of Cl\(^-\) ions present in sodium chloride must be 1:1, resulting in the formula Na\(^+\)Cl\(^-\) as the simplest for this compound. Analysis of the compound sodium chloride would always show it to consist of sodium ions and chloride ions present in this ratio of 1:1. In Topic 1 it was pointed out that all the halogen elements including chlorine occur as diatomic molecules and not single atoms. Thus in forming sodium chloride, 1 molecule of chlorine, written as Cl\(_2\), would react with 2 atoms of sodium to form 2Cl\(^-\) and 2Na\(^+\) ions in order that the ratio of + charge to – charge be 1:1. However, in writing the formula, the simplest whole-number ratio of cation and anion is used. Thus although the formula Na\(^+\)\(_2\)Cl\(_2\) still has a Na\(^+\):Cl\(^-\) ratio of 1:1, the formula should be written as Na\(^+\)Cl\(^-\).

As was discussed in Topic 2, some cations may have a charge larger than +1, for example Ca\(^{2+}\) and Al\(^{3+}\). Similarly, some anions such as O\(^{2-}\) and N\(^{3-}\) have more than a single negative charge. Again, when ionic compounds involving such species are formed, the overall charge on the resulting compound must be zero. This is achieved if the cations and anions are formed and therefore combine in the appropriate
ratio so that the total size of the charge on the cations equals the total size of the charge on the anions. For example, the ionic compound calcium chloride which results from the reaction between calcium and chlorine consists of Ca\(^{2+}\) ions and Cl\(^{-}\) ions which are present in the compound in the ratio of 1 calcium ion to 2 chloride ions. The formula for calcium chloride is therefore Ca\(^{2+}\)Cl\(^{2-}\).

[Note the use of the subscripted 2 to show that there are 2 chloride ions present in the formula. It is incorrect to write this as Ca\(^{2+}\)2Cl\(^{-}\) or Ca\(^{2+}\)Cl\(^{2-}\)].

In this formula, 2 Cl\(^{-}\) ions carry a total charge = 2\(^{-}\) while the single Ca\(^{2+}\) ion carries a charge = 2\(^{+}\). Thus electrical neutrality is preserved in the compound.

The ionic compound lithium oxide contains Li\(^{+}\) ions and O\(^{2-}\) ions in the ratio of 2 lithium ions to 1 oxide ion so that the total charge on the compound is zero. The formula of lithium oxide then must be Li\(^{+}\)\(_2\)O\(^{2-}\), the subscripted 2 being used to indicate that there are two Li\(^{+}\) ions in the formula.

Likewise, the ionic compound aluminium oxide containing Al\(^{3+}\) ions and O\(^{2-}\) ions has the formula Al\(^{3+}\)\(_2\)O\(^{2-}\)_3 so that the total positive charge (6 \(^{+}\)) exactly equals the total negative charge (6 \(^{-}\)).

In this section, the formulas of all ionic compounds shown have included the charge on the cation and anion as superscripts, e.g. Na\(^{+}\)Cl\(^{-}\). However, this is not normal practice and it is understood that the compound of formula NaCl contains the Na\(^{+}\) and Cl\(^{-}\) ions. At this stage, it may be helpful to continue the practice of showing the charges on the ions in the compounds, but in due course you should delete them, and they will not be shown here in future.

The following examples illustrate a process frequently used for obtaining correctly balanced formulas for ionic compounds.

1. Deduce the formula for sodium oxide.

First step: identify the ions in the compound Na\(^{+}\) and O\(^{2-}\)

Second step: write the unbalanced formula with the cation first and anion second Na\(^{+}\)O\(^{2-}\)

Third step: obtain the subscripts for each ion by using the magnitude of the charge on one ion as the subscript of the other

\[
\text{Na}^{1+} \quad \text{O}^{2-}
\]

which gives (Na\(^{+}\))\(_2\)O\(^{2-}\) i.e. Na\(_2\)O
Finally, check that the charges have been balanced by multiplying the number of each ion present by its charge. The sum of these must be zero.

\[ 2 \times [1+] + 1 \times [2^-] = (+2) + (-2) = 0 \]

:. correctly balanced formula.

2. Deduce the formula for aluminium oxide.

Using the above method, the ions present are \( \text{Al}^{3+} \) and \( \text{O}^{2-} \).

Thus the formula is \( \text{Al}_2\text{O}_3 \).

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

:. correctly balanced formula.

If the subscripts are identical or have a common factor, reduce them to the simplest whole number ratio.

What will the charge be on a cation in a compound?

There is a very simple rule that gives the number of electrons that can be removed from an atom of an element to form a cation without requiring an excessive amount of energy. Most elements of the first three groups of elements shown in Table 2 on Page I-22 of Topic 1 mainly form ionic compounds. The first group, the alkali metals, has 1 outer level electron, the second group has 2 outer electrons and the third group has 3 outer electrons. These are therefore also the number of electrons which must be removed from atoms of elements of each of these groups in order to obtain the same outer shell structure as the noble gas of closest atomic number. Thus when forming ions, the first group of elements form \( 1+ \) cations in ionic compounds, the second group of elements form \( 2+ \) cations in ionic compounds, and when elements of the third group form ionic compounds, the cations usually have a \( 3+ \) charge. The following list of common cations formed by elements from other groups should be committed to memory. In some cases a Roman numeral I, II, III or IV is written as part of the name of the cation. This is needed for those elements which can have more than a single ionic state - for example, the element tin (Sn) can form ions with a \( 2+ \) charge [tin(II), \( \text{Sn}^{2+} \)] or a \( 4+ \) charge [tin(IV), \( \text{Sn}^{4+} \)]. Note that very few cations with a \( 4+ \) charge exist and \( \text{Sn}^{4+} \) is unusual. Where there is only one possible ionic charge for a cation, the Roman numerals are not used. The charge on a cation as indicated by the Roman numerals is referred to as the OXIDATION STATE of the element.
### ADDITIONAL COMMON CATIONS

<table>
<thead>
<tr>
<th>Ion</th>
<th>Symbol</th>
<th>Ion</th>
<th>Symbol</th>
<th>Ion</th>
<th>Symbol</th>
</tr>
</thead>
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<tr>
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<td>Ag⁺</td>
<td>zinc</td>
<td>Zn²⁺</td>
<td>iron(II)</td>
<td>Fe²⁺</td>
</tr>
<tr>
<td>copper(I)</td>
<td>Cu⁺</td>
<td>lead(II)</td>
<td>Pb²⁺</td>
<td>iron(III)</td>
<td>Fe³⁺</td>
</tr>
<tr>
<td>copper(II)</td>
<td>Cu²⁺</td>
<td>cobalt(II)</td>
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<td>Cd²⁺</td>
</tr>
<tr>
<td>tin(II)</td>
<td>Sn²⁺</td>
<td>chromium(III)</td>
<td>Cr³⁺</td>
<td>mercury(II)</td>
<td>Hg²⁺</td>
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<tr>
<td>tin(IV)</td>
<td>Sn⁴⁺</td>
<td>manganese(II)</td>
<td>Mn²⁺</td>
<td>nickel(II)</td>
<td>Ni²⁺</td>
</tr>
<tr>
<td>bismuth(III)</td>
<td>Bi³⁺</td>
<td>gold(III)</td>
<td>Au³⁺</td>
<td>platinum(II)</td>
<td>Pt²⁺</td>
</tr>
</tbody>
</table>

### What will be the charge on an anion in a compound?

Again, the grouping of elements helps to answer this question. All the elements of the seventh group (halogens) require just one more electron to obtain the electron structure (8 outer level electrons) of the nearest noble gas. Consequently, when these elements form anions they all do so by gaining one electron and thus carry a 1⁻ charge. Similarly, all elements of the sixth group are two electrons short of the nearest noble gas electron structure and so they all form 2⁻ charged anions when in ionic compounds. Although less commonly observed, when elements from the fifth group starting with nitrogen (all 3 electrons short of the noble gas structure) form anions they do so by gaining 3 electrons and thus carry a 3⁻ charge. No stable anions form from single atoms gaining more than three electrons. Thus knowing to which group an element belongs allows one to deduce its likely charge in an ionic compound.

### Check your understanding of this section.

Why doesn’t sodium form a 2⁺ cation in compounds?  
Why doesn’t chlorine form a 2⁻ anion in its compounds?  
Why does it require energy input for a metal atom to form a cation?  
Why don’t non-metals such as chlorine form cations?  
What is a salt?  
What type of force holds together the components of an ionic compound?  
What are the requirements for the formula of an ionic compound to be correctly balanced?

### How are ionic compounds named?

Compounds of two elements such as all those discussed in this Topic are called **Binary Compounds.** Rules for naming binary ionic compounds are very simple and are as follows:

The compound is named as two separate words, the cation being named first and the anion last.
The cation name is the same as the element but with the charge appended in brackets as Roman numerals where necessary. Anions formed from an element take the stem from the name of the element and the ending "ide" is attached e.g. oxide from oxygen. No special ending is needed for the cation as it is obvious that a compound is being named when two words are used in the name. Lower case letters are used for the name, including the first letter.

The following examples illustrate the correct naming of binary ionic compounds.
NaCl sodium chloride (contains Na\(^+\) and Cl\(^-\) ions)
KI potassium iodide (contains K\(^+\) and I\(^-\) ions)
CaF\(_2\) calcium fluoride (contains Ca\(^{2+}\) and F\(^-\) ions)
Rb\(_2\)O rubidium oxide (contains Rb\(^+\) and O\(^{2-}\) ions)
BaS barium sulfide (contains Ba\(^{2+}\) and S\(^{2-}\) ions)
Mg\(_3\)N\(_2\) magnesium nitride (contains Mg\(^{2+}\) and N\(^{3-}\) ions)
CuCl\(_2\) copper(II) chloride (contains Cu\(^{2+}\) and Cl\(^-\) ions)
SnS tin(II) sulfide (contains Sn\(^{2+}\) and S\(^{2-}\) ions)
Fe\(_2\)O\(_3\) iron(III) oxide (contains Fe\(^{3+}\) and O\(^{2-}\) ions)
AuCl\(_3\) gold(III) chloride (contains Au\(^{3+}\) and Cl\(^-\) ions)

**Check your understanding of this section.**

State the rules that allow one to deduce the likely charge on the cation formed from a metal atom and an anion formed from a non-metal atom.

How is the charge on a cation that can have more than one charge indicated?

State the rules for naming a binary compound.

How would one know that in the compound Fe\(_2\)O\(_3\), the iron is present as Fe\(^{3+}\)?

**Metallic bonding.**

Metals consist of infinitely large collections of atoms that are bonded by outer electrons which are very mobile and constantly transfer from atom to atom, leading to another type of bonding, **METALLIC BONDING.** These outer electrons are said to be **DELOCALISED.** This type of aggregation of atoms in metals is illustrated in the following diagram.

![Diagram of metallic bonding](image)
Metals are typically good conductors of electricity as compared with non-metals because of this mobility of the outer electrons that bond the atoms. Electrons entering a metal at the negative terminal of a power source such as a battery are able to participate in the bonding process. These electrons mingle with the electrons already available from the metal’s atoms and move along the metal towards the positively charged terminal, so the metal remains electrically neutral while the current flows. By this mechanism, a metal can carry an electrical current indefinitely without altering the metal in any way. The ease with which the metal can carry a current varies depending on the metal used. Some metals have a higher resistance to electron flow than others, copper is an example of a commonly used metal for electrical wiring as its resistance is relatively low. Electrical resistance of a given metal increases as the temperature increases because the larger vibrational movement of the atoms (see below) makes it more difficult for the electrons to move from atom to atom. For this reason, when powerful electromagnets are required, their copper coils are often cooled with liquefied gases such as nitrogen or helium. Despite some exceptions such as graphite (see Topic 12), non-metals usually lack any freely moving electrons and are generally poor conductors of electricity.

Other physical properties of metals such as the ability to bend without breaking, being able to be drawn into wires (ductile) and beaten into sheets (malleable) also result from the nature of metallic bonds. In a metal, atoms can move relative to each other, breaking adjacent bonds but then reforming new bonds in the new location with different atoms because of the electrons’ mobility. This is in stark contrast to ionic solids which shatter when subjected to a shearing force because as soon as the ions move one position along in the crystal lattice, like charges are then adjacent and repel strongly. This is illustrated in the following diagram.

Metals also have the property of being able to conduct heat well compared with most non-metals. Heat conduction by metals is due to two mechanisms. One mechanism is again through the mobility of the delocalised outer electrons. If a metal is heated, these electrons move faster and collide more often and with greater energy with other electrons in the metal which in turn undergo more frequent and more energetic collisions with their neighbouring electrons and atoms, thus transferring energy as heat throughout the metal.
The second mechanism arises from the vibration of the atoms themselves. All chemically bonded atoms undergo constant vibrational motion about a mean location. The extent of the vibrational motion is dependent on the temperature and it increases as the temperature increases. Only at the unobtainable absolute zero of temperature, –273°C, would this motion cease (see Supplementary Topic 2). Heat added to one end of a solid metal will energise the atoms there and their enhanced vibrations increasingly perturb adjacent atoms which in turn transfer heat energy throughout the metal as a result of a synchronised wavelike motion of the vibrations of the atoms. Efficient heat transfer by this mechanism requires the component atoms to be bonded in very symmetrical arrangements with just one type of atom, as is the case for pure metals. Non-metals usually lack the required uniform solid state structure of metals and consequently are generally poor conductors of heat because the vibrational motions are not transmitted throughout the solid.

The shiny appearance of freshly cut metals is due to light impacting on mobile electrons on the surface being scattered and reflected by it.
Objectives of this Topic.
When you have completed this Topic including the tutorial questions, you should have achieved the following goals:

1. Understand the process whereby ionic bonds form as a result of electron transfer from one atom to another with consequent electrostatic attraction between the resultant ions.

2. Understand that the formula for an ionic compound must contain an equal number of +ve and –ve charges and therefore be electrically neutral.

3. Be able to write formulas and names for binary ionic compounds.

4. Understand the nature of metallic bonding and its basis for the physical properties of metals.

Summary tables of cations and anions.
The following tables list the cations of metals commonly encountered and the simple anions of non-metals that should be committed to memory. It is not an exhaustive list but includes those most commonly encountered. For a more extensive list, consult the data sheet at the back of the book.

**Cations of metals from the first 3 groups of Table 2, page I-22.**

<table>
<thead>
<tr>
<th>Element (symbol)</th>
<th>cation formed</th>
<th>Element (symbol)</th>
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<th>Element (symbol)</th>
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<td>Li⁺</td>
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<tr>
<td>Cs</td>
<td>Cs⁺</td>
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</tr>
</tbody>
</table>

**Additional common cations.**

<table>
<thead>
<tr>
<th>Ion</th>
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<td>silver</td>
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<td>Bi³⁺</td>
<td>gold(III)</td>
<td>Au³⁺</td>
<td>platinum(II)</td>
<td>Pt²⁺</td>
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</table>
Table of common simple anions.

<table>
<thead>
<tr>
<th>Element (symbol)</th>
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<th>Element (symbol)</th>
<th>anion formed</th>
<th>anion formed</th>
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<tbody>
<tr>
<td>N</td>
<td>N^3−</td>
<td>O</td>
<td>O^2−</td>
<td>F^−</td>
</tr>
<tr>
<td>P</td>
<td>P^3−</td>
<td>S</td>
<td>S^2−</td>
<td>Cl^−</td>
</tr>
<tr>
<td></td>
<td></td>
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<td>Br^−</td>
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<tr>
<td></td>
<td></td>
<td>Te</td>
<td>Te^2−</td>
<td>I^−</td>
</tr>
</tbody>
</table>

Also note the anion from the H atom, the hydride ion, is H^−. This anion consists of a hydrogen atom which has gained one electron into its only occupied electron orbit to give a resultant excess charge of –1.

SUMMARY

Chemical bonds form between atoms through redistribution of some of the electrons between them. In one form of bond, ionic bonding, one or more electrons are transferred from a metal atom to an atom of a non-metal. The resultant cation and anion are then held together by the electrostatic attraction between the opposite charges they carry. This process of electron transfer results in the cation and anion attaining the particularly stable noble gas electron structure. Ionic compounds must have overall electrical neutrality and so the cations and anions form and combine in a ratio such that the total cation charge is equal in magnitude to the total anion charge. Their formulas are the simplest whole number ratio of cation and anion that will satisfy this requirement, written as subscripts following each component element’s symbol.

The ionic charge that an atom of an element is likely to exhibit when it forms an ionic compound can be deduced from the group of elements to which it belongs. Referring to the Table 2 of Topic 1, all of the first group of elements have 1 electron more than the noble gas structure and so they form +1 cations, all of the second group have 2 more electrons than the noble gas structure so they form +2 cations when in ionic compounds and similarly, elements of the third group mostly form +3 cations. Once the noble gas structure has been attained, too much energy (known as the “ionization energy”) would be required for further electrons to be removed. Thus cations with a charge greater than 3+ are rare as they would require too much energy to form and would generally be unstable. Atoms of the seventh group, can’t form cations but being short of the noble gas structure by 1 electron can form anions by gaining 1 electron to form 1− anions. The sixth group being short of the noble gas structure by 2 electrons, form 2 − anions, etc. Once the noble gas structure has been attained, no further electrons can be gained and simple anions bearing a charge greater than 3− do not form.
Some metals can form cations with more than one charged state, for example Fe$^{2+}$ and Fe$^{3+}$. Such cations require the addition of a Roman numeral in brackets to their name so that there is no ambiguity as to which ion is present.

Naming binary ionic compounds (compounds containing only two elements) follows some simple rules. The cation and anion are named separately in that order. The cation takes the same name as its element with the addition of the Roman numeral to indicate the charge it bears if necessary. The anion takes its name from a stem of its element to which the ending “ide” is added.

Atoms in metals are bonded by a second type of bond termed metallic bonding. Metallic bonding arises from mobile delocalised outer electrons which easily move from atom to atom. This imparts physical properties associated with metals such as being malleable and ductile. It also accounts for the ability of metals to conduct electricity and is one basis for the good heat conductivity of metals. Another property of metals also contributes to heat conductivity, namely the very regular array of the atoms which allows heat energy to be transmitted as vibrations from atom to atom along a metal away from the heat source.

**TUTORIAL QUESTIONS - TOPIC 3.**

1.(a) Define each of the following:

(i) ionic bond

(ii) binary compound

(b) A new element, "X", is discovered and found to have 2 electrons in its outer level. Is X a metal or non-metal? Predict the formula its ion would have in any ionic compounds it forms.
2. Write the formulas and names of the binary ionic compounds of the following elements. **Do not attempt to write equations for their formation at this stage.**

(i) lithium and bromine  
(ii) barium and oxygen  
(iii) aluminium and fluorine  
(iv) sodium and sulfur  
(v) magnesium and nitrogen  
(vi) rubidium and chlorine  
(vii) caesium and phosphorus  
(viii) potassium and iodine  
(ix) calcium and selenium  
(x) strontium and chlorine  
(xi) lithium and oxygen  
(xii) magnesium and bromine  
(xiii) rubidium and nitrogen  
(xiv) calcium and fluorine  
(xv) aluminium and sulfur  
(xvi) caesium and selenium  
(xvii) barium and phosphorus  
(xviii) sodium and nitrogen  
(xix) potassium and chlorine  
(xx) strontium and iodine

3. Give the formula for each of the following binary compounds.

(i) silver iodide  
(ii) magnesium chloride  
(iii) copper(II) oxide  
(iv) copper(I) oxide  
(v) barium nitride  
(vi) manganese(II) sulfide  
(vii) mercury(II) oxide  
(viii) iron(II) bromide  
(ix) aluminium oxide  
(x) iron(III) chloride
4. Write the name for each of the following compounds.
   (i) \( \text{AgCl} \)
   (ii) \( \text{Mg}_3\text{N}_2 \)
   (iii) \( \text{CaBr}_2 \)
   (iv) \( \text{Al}_2\text{O}_3 \)
   (v) \( \text{CuCl}_2 \)
   (vi) \( \text{PbO} \)
   (vii) \( \text{MnS} \)
   (viii) \( \text{ZnI}_2 \)
   (ix) \( \text{KCl} \)
   (x) \( \text{Ca}_3\text{P}_2 \)
   (xi) \( \text{CrCl}_3 \)
   (xii) \( \text{BaSe} \)
   (xiii) \( \text{CoCl}_2 \)
   (xiv) \( \text{Fe}_2\text{O}_3 \)
   (xv) \( \text{FeCl}_2 \)
   (xvi) \( \text{SrI}_2 \)
   (xvii) \( \text{SnBr}_2 \)
   (xviii) \( \text{MgO} \)
   (xix) \( \text{Rb}_3\text{N} \)
   (xx) \( \text{LiF} \)
   (xxi) \( \text{PtBr}_2 \)
   (xxii) \( \text{Bi}_2\text{O}_3 \)
   (xxiii) \( \text{AuCl}_3 \)

5. Give the formula for each of the following binary compounds.
   (i) cadmium fluoride (ii) strontium chloride
   (iii) cobalt(II) sulfide (iv) lead(II) iodide
(v) tin(II) oxide  (vi) iron(III) oxide
(vii) chromium(III) nitride  (viii) calcium bromide
(ix) potassium oxide  (x) sodium phosphide

6. Account for the following properties of metals:
   (a) metals can bend without shattering.

   (b) metals can conduct electricity indefinitely without undergoing any chemical change.

7. Describe the two mechanisms by which metals conduct heat.

8. Do the CAL module “Nomenclature - Binary Compounds” located on your memory stick and the module “Nomenclature” Q 1-4 from the Self Help Problems.

9. Complete the chemical crossword puzzle on the following pages.
**CHEMICAL CROSSWORD No. 2**

**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,

<table>
<thead>
<tr>
<th></th>
<th></th>
<th>O</th>
<th>Na</th>
<th>Cl</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ag</td>
<td>2</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Where a subscript is required in the formula, it is entered in its own box. The ionic charge of any cation must be the same where that atom is common to both the across and down formulas.

**CLUES.**

<table>
<thead>
<tr>
<th>ACROSS</th>
<th></th>
<th>DOWN</th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1.</td>
<td>nickel</td>
<td>2.</td>
<td>iron(III) chloride</td>
</tr>
<tr>
<td>2.</td>
<td>iron(III) bromide</td>
<td>3.</td>
<td>lithium oxide</td>
</tr>
<tr>
<td>4.</td>
<td>tellurium</td>
<td>6.</td>
<td>tin(IV) oxide</td>
</tr>
<tr>
<td>5.</td>
<td>germanium</td>
<td>7.</td>
<td>potassium iodide</td>
</tr>
<tr>
<td>7.</td>
<td>potassium chloride</td>
<td>8.</td>
<td>sodium selenide</td>
</tr>
<tr>
<td>8.</td>
<td>sodium sulfide</td>
<td>10.</td>
<td>calcium fluoride</td>
</tr>
<tr>
<td>9.</td>
<td>silicon</td>
<td>12.</td>
<td>potassium sulfide</td>
</tr>
<tr>
<td>10.</td>
<td>calcium oxide</td>
<td>14.</td>
<td>strontium chloride</td>
</tr>
<tr>
<td>11.</td>
<td>chromium(III) iodide</td>
<td>15.</td>
<td>rubidium oxide</td>
</tr>
<tr>
<td>12.</td>
<td>potassium oxide</td>
<td>17.</td>
<td>barium bromide</td>
</tr>
<tr>
<td>13.</td>
<td>thallium</td>
<td>18.</td>
<td>platinum</td>
</tr>
<tr>
<td>14.</td>
<td>strontium fluoride</td>
<td>19.</td>
<td>caesium selenide</td>
</tr>
<tr>
<td>15.</td>
<td>rubidium selenide</td>
<td>21.</td>
<td>copper(II) iodide</td>
</tr>
<tr>
<td>16.</td>
<td>indium</td>
<td>22.</td>
<td>aluminium nitride</td>
</tr>
<tr>
<td>17.</td>
<td>barium chloride</td>
<td>23.</td>
<td>silver(I) sulfide</td>
</tr>
<tr>
<td>19.</td>
<td>caesium sulfide</td>
<td>24.</td>
<td>radon</td>
</tr>
<tr>
<td>20.</td>
<td>gallium</td>
<td>25.</td>
<td>nickel(II) fluoride</td>
</tr>
<tr>
<td>21.</td>
<td>copper(II) bromide</td>
<td>28.</td>
<td>xenon</td>
</tr>
<tr>
<td>23.</td>
<td>silver(I) oxide</td>
<td>29.</td>
<td>iron(II) chloride</td>
</tr>
<tr>
<td>25.</td>
<td>nickel(II) iodide</td>
<td>30.</td>
<td>strontium nitride</td>
</tr>
<tr>
<td>26.</td>
<td>chromium(III) nitride</td>
<td>31.</td>
<td>krypton</td>
</tr>
<tr>
<td>27.</td>
<td>copper(I) selenide</td>
<td>32.</td>
<td>zinc bromide</td>
</tr>
<tr>
<td>29.</td>
<td>iron(II) fluoride</td>
<td>33.</td>
<td>barium phosphide</td>
</tr>
<tr>
<td>32.</td>
<td>zinc chloride</td>
<td>34.</td>
<td>nitrogen (molecule)</td>
</tr>
<tr>
<td>33.</td>
<td>barium nitride</td>
<td>35.</td>
<td>arsenic</td>
</tr>
<tr>
<td>36.</td>
<td>argon</td>
<td>37.</td>
<td>cadmium iodide</td>
</tr>
<tr>
<td>37.</td>
<td>cadmium bromide</td>
<td>38.</td>
<td>magnesium phosphide</td>
</tr>
<tr>
<td>38.</td>
<td>magnesium nitride</td>
<td>39.</td>
<td>neon</td>
</tr>
<tr>
<td>40.</td>
<td>lead(II) iodide</td>
<td>40.</td>
<td>lead(II) fluoride</td>
</tr>
<tr>
<td>41.</td>
<td>nickel(II) phosphide</td>
<td>41.</td>
<td>nickel(II) nitride</td>
</tr>
<tr>
<td>43.</td>
<td>helium</td>
<td>42.</td>
<td>mercury(II) selenide</td>
</tr>
<tr>
<td>44.</td>
<td>tin(II) fluoride</td>
<td>44.</td>
<td>tin(II) chloride</td>
</tr>
<tr>
<td>45.</td>
<td>iron(II) phosphide</td>
<td>45.</td>
<td>iron(II) nitride</td>
</tr>
</tbody>
</table>
46. zinc selenide
47. mercury(II) chloride
48. zinc nitride
49. platinum(II) oxide
50. manganese(II) iodide
51. copper(II) nitride
52. calcium sulfide
54. lead(II) phosphide
56. iron(III) phosphide
58. tin(II) phosphide
59. bismuth(III) oxide
60. mercury(II) nitride
62. chromium(III) oxide
64. bismuth
65. nitrogen (molecule)
66. aluminium sulfide
67. copper(I) phosphide
68. iron(III) selenide
69. silver(I) phosphide
70. gold(III) nitride
71. ozone
72. caesium nitride
74. bismuth(III) bromide
75. rubidium nitride
76. iron(II) oxide
77. manganese
78. sodium fluoride
79. potassium nitride
80. nickel(II) sulfide
81. aluminium phosphide
82. sodium phosphide
83. copper(II) sulfide
85. lithium chloride
86. lithium nitride
87. magnesium oxide
88. lead(II) nitride
90. barium selenide
91. tin(II) sulfide
93. rubidium fluoride
94. strontium sulfide
95. cadmium oxide
97. tin(IV) chloride
98. calcium oxide
99. zinc sulfide
100. bismuth(III) nitride
ANSWERS TO TUTORIAL TOPIC 3

1. (a) (i) ionic bond: A chemical bond resulting from the electrostatic attraction between a positively charged ion (cation) and a negatively charged ion (anion).

(ii) binary compound: A compound which contains atoms of only two elements.

(b) X would be a metal as it has the same outer electron structure as the second group of elements in Table 2 of Topic 1. The cation formed would be X\(^{2+}\).

2. In each case, the first element is a metal and the second element is a non-metal. Hence the resulting compound will be a salt containing the cation of the metal and the anion of the non-metal. It is necessary to deduce the charge on each ion so that a balanced formula for the resulting compound can be written correctly.

For example, in (i), lithium (Li) is in the first group of elements and so it always forms a +1 charged cation in compounds, Li\(^+\). Bromine is in the 7th group, the halogens, and in ionic compounds the members of this group always form a –1 charged anion, Br\(^-\) in this case. The balanced formula for the compound will therefore require one Li\(^+\) to each Br\(^-\) ion, i.e. LiBr (or Li\(^+\)Br\(^-\) if the charges are shown). The name of the compound will consist of two words, the cation followed by the anion. The cation always takes the same name as the metal element (lithium) and the anion is named by taking a stem from the non-metal’s name and adding “ide” (bromide). Thus the compound is lithium bromide.

In (v), the cation from magnesium is Mg\(^{2+}\) as magnesium is a member of the second group while the anion from nitrogen, a member of the fifth group, is N\(^{3-}\). To obtain the simplest charge-balanced formula, it is necessary to take three Mg\(^{2+}\) ions (total positive charge = +6) and two N\(^{3-}\) ions (total negative charge –6). The balanced formula for the resultant compound is therefore Mg\(_3\)N\(_2\). Note that the number of each ion in the formula is given as a subscript - it is not correct to write the formula as 3Mg2N. The name of the compound is magnesium nitride.

(i) lithium bromide LiBr  (ii) barium oxide BaO
(iii) aluminium fluoride AlF\(_3\)  (iv) sodium sulfide Na\(_2\)S
(v) magnesium nitride Mg\(_3\)N\(_2\)  (vi) rubidium chloride RbCl
(vii) caesium phosphide Cs\(_3\)P  (viii) potassium iodide KI
(ix) calcium selenide CaSe  (x) strontium chloride SrCl\(_2\)
(xi) lithium oxide Li\(_2\)O  (xii) magnesium bromide MgBr\(_2\)
(xiii) rubidium nitride Rb\(_3\)N  (xiv) calcium fluoride CaF\(_2\)
(xv) aluminium sulfide Al\(_2\)S\(_3\)  (xvi) caesium selenide Cs\(_2\)Se
(xvii) barium phosphide Ba\(_3\)P\(_2\)  (xviii) sodium nitride Na\(_3\)N
(xix) potassium chloride KCl  (xx) strontium iodide SrI\(_2\)
3. Note the use of the Roman numerals where required to avoid ambiguity for cations that can exist with more than one ionic charge. Thus copper(II) oxide contains the Cu$^{2+}$ ion while copper(I) oxide has the Cu$^{+}$ cation present.

(i) AgI (ii) MgCl$_2$ (iii) CuO (iv) Cu$_2$O
(v) Ba$_3$N$_2$ (vi) MnS (vii) HgO (viii) FeBr$_2$
(ix) Al$_2$O$_3$ (x) FeCl$_3$

4. (i) silver chloride (ii) magnesium nitride
(iii) calcium bromide (iv) aluminium oxide
(v) copper(II) chloride (vi) lead(II) oxide
(vi) manganese(II) sulfide (viii) zinc iodide
(ix) potassium chloride (x) calcium phosphide
(xi) chromium(III) chloride (xii) barium selenide
(xiii) cobalt(II) chloride (xiv) iron(III) oxide
(xv) iron(II) chloride (xvi) strontium iodide
(xvii) tin(II) bromide (xviii) magnesium oxide
(xix) rubidium nitride (xx) lithium fluoride
(xxi) platinum(II) bromide (xxii) bismuth(III) oxide
(xxiii) gold(III) chloride

5. (i) CdF$_2$ (ii) SrCl$_2$
(iii) CoS (iv) PbI$_2$
(v) SnO (vi) Fe$_2$O$_3$
(vii) CrN (viii) CaBr$_2$
(ix) K$_2$O (x) Na$_3$P

6. (a) Metal atoms are bonded by mobile electrons so that if the atoms shift relative to each other as a result of bending or other shearing force, the atoms can remake the bonds that were broken and not part.
(b) The mobile electrons that bond the atoms together in the solid can move under the influence of an applied electrical voltage and can travel to the positive terminal of the source. Electrons from the negative terminal of the source replace them and the process can continue indefinitely without altering the bonding in the metal.

7. Metals conduct heat by two mechanisms.
   (1) The mobile delocalised outer electrons gain energy from the heat source and move faster and collide with other electrons and atoms more rapidly. The additional energy imparted in this way to surrounding atoms increases their energy which is displayed as heat.

   (2) The vibrations of atoms near the heat source is increased and leads to a synchronised wave of vibration transmitted to atoms throughout the metal. The increased motion of the atoms is again displayed as heat.