

## TOPIC 3.

### IONIC COMPOUNDS: formation, formulas and naming

#### Chemical bonding.

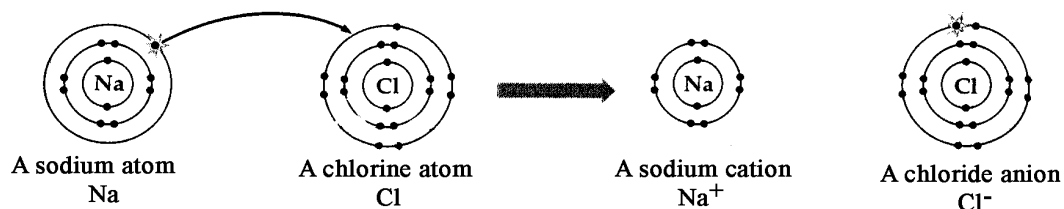
When elements combine to form compounds, a chemical bond holds the atoms together. There are two basic types of chemical bonds possible in compounds, **IONIC BONDS** and **COVALENT BONDS**. Both types of bond involve a redistribution of electrons between the bonded atoms. Ionic bonding will now be examined 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  $\text{Na}^+$  cations in which state they are isoelectronic with the noble gas neon, but sodium cannot form  $\text{Na}^{2+}$  or  $\text{Na}^{3+}$  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  $\text{Cl}^-$  anions (isoelectronic with argon) but not  $\text{Cl}^{2-}$  or  $\text{Cl}^{3-}$  anions.

Thus an atom of sodium can lose the single electron located in the third energy level to form an  $\text{Na}^+$  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  $\text{Cl}^-$  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  $\text{Na}^+$  and  $\text{Cl}^-$  ions by transferring a single electron per atom from Na to Cl. When this transfer occurs, the resulting ions of  $\text{Na}^+$  and  $\text{Cl}^-$ , having opposite electrical charges, will be attracted to each other to form the solid compound sodium chloride. This electrostatic attraction is the basis for ionic bonding.



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

### Combining ratios and formulas of ionic compounds.

The ionic compound formed must be electrically neutral, so the ratio of the number of  $\text{Na}^+$  ions to the number of  $\text{Cl}^-$  ions present in sodium chloride must be 1:1, resulting in the formula  $\text{Na}^+\text{Cl}^-$  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  $\text{Cl}_2$ , would react with 2 atoms of sodium to form  $2\text{Cl}^-$  and  $2\text{Na}^+$  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  $\text{Na}_2^+\text{Cl}_2^-$  still has a  $\text{Na}^+:\text{Cl}^-$  ratio of 1:1, the formula should be written as  $\text{Na}^+\text{Cl}^-$ .

As was discussed in Topic 2, some cations may have a charge larger than +1, for example  $\text{Ca}^{2+}$  and  $\text{Al}^{3+}$ . Similarly, some anions such as  $\text{O}^{2-}$  and  $\text{N}^{3-}$  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  $\text{Ca}^{2+}$  ions and  $\text{Cl}^-$  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  $\text{Ca}^{2+}\text{Cl}_2^-$ .

*[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  $\text{Ca}^{2+}2\text{Cl}^-$  or  $\text{Ca}^{2+}\text{Cl}_2^-$ ].*

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

The ionic compound lithium oxide contains  $\text{Li}^+$  ions and  $\text{O}^{2-}$  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  $\text{Li}_2^+\text{O}^{2-}$ , the subscripted 2 being used to indicate that there are two  $\text{Li}^+$  ions in the formula.

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

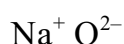
In this section, the formulas of all ionic compounds shown have included the charge on the cation and anion as superscripts, e.g.  $\text{Na}^+\text{Cl}^-$ . However, this is not normal practice and it is understood that the compound of formula  $\text{NaCl}$  contains the  $\text{Na}^+$  and  $\text{Cl}^-$  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.*

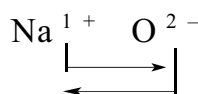
First step: identify the ions in the compound  $\text{Na}^+$  and  $\text{O}^{2-}$

Second step: write the unbalanced formula with the cation first and anion second



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

which gives  $(\text{Na}^+)_2\text{O}^{2-}$  i.e.  $\text{Na}_2\text{O}$



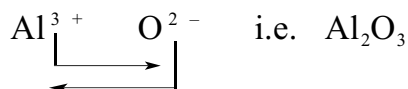
Finally, check that the charges have been balanced by multiplying the number of each ion present by its charge. The sum of these must be zero.

$$2 \times [1+] + 1 \times [2-] = (+2) + (-2) = 0 \quad \therefore \text{correctly balanced formula.}$$

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



Checking:  $2 \times [+3] + 3 \times [-2] = (+6) + (-6) = 0 \therefore$  correctly balanced formula.

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

**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?

**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-21 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.

**COMMON CATIONS FROM FAMILIES OTHER THAN 1, 2 OR 3.**

Ion	Symbol	Ion	Symbol	Ion	Symbol
silver(I)	$\text{Ag}^+$	zinc	$\text{Zn}^{2+}$	iron(II)	$\text{Fe}^{2+}$
copper(I)	$\text{Cu}^+$	lead(II)	$\text{Pb}^{2+}$	iron(III)	$\text{Fe}^{3+}$
copper(II)	$\text{Cu}^{2+}$	cobalt(II)	$\text{Co}^{2+}$	cadmium	$\text{Cd}^{2+}$
tin(II)	$\text{Sn}^{2+}$	chromium(III)	$\text{Cr}^{3+}$	mercury(II)	$\text{Hg}^{2+}$
tin(IV)	$\text{Sn}^{4+}$	manganese(II)	$\text{Mn}^{2+}$	nickel(II)	$\text{Ni}^{2+}$
bismuth(III)	$\text{Bi}^{3+}$	gold(III)	$\text{Au}^{3+}$	platinum(II)	$\text{Pt}^{2+}$

**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.

**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.

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 <sup>+</sup> and Cl <sup>-</sup> ions)
KI	potassium iodide (contains K <sup>+</sup> and I <sup>-</sup> ions)
CaF <sub>2</sub>	calcium fluoride (contains Ca <sup>2+</sup> and F <sup>-</sup> ions)
Rb <sub>2</sub> O	rubidium oxide (contains Rb <sup>+</sup> and O <sup>2-</sup> ions)
BaS	barium sulfide (contains Ba <sup>2+</sup> and S <sup>2-</sup> ions)
Mg <sub>3</sub> N <sub>2</sub>	magnesium nitride (contains Mg <sup>2+</sup> and N <sup>3-</sup> ions)
CuCl <sub>2</sub>	copper(II) chloride (contains Cu <sup>2+</sup> and Cl <sup>-</sup> ions)
SnS	tin(II) sulfide (contains Sn <sup>2+</sup> and S <sup>2-</sup> ions)
Fe <sub>2</sub> O <sub>3</sub>	iron(III) oxide (contains Fe <sup>3+</sup> and O <sup>2-</sup> ions)
AuCl <sub>3</sub>	gold(III) chloride (contains Au <sup>3+</sup> and Cl <sup>-</sup> 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<sub>2</sub>O<sub>3</sub>, the iron is present as Fe<sup>3+</sup>?

**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.

### 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  $\text{Fe}^{2+}$  and  $\text{Fe}^{3+}$ . 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.

### TUTORIAL QUESTIONS - TOPIC 3.

1. 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(I) 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) AgCl
- (ii) Mg<sub>3</sub>N<sub>2</sub>
- (iii) CaBr<sub>2</sub>
- (iv) Al<sub>2</sub>O<sub>3</sub>
- (v) CuCl<sub>2</sub>
- (vi) PbO
- (vii) MnS
- (viii) ZnI<sub>2</sub>
- (ix) KCl
- (x) Ca<sub>3</sub>P<sub>2</sub>
- (xi) CrCl<sub>3</sub>
- (xii) BaSe
- (xiii) CoCl<sub>2</sub>
- (xiv) Fe<sub>2</sub>O<sub>3</sub>
- (xv) FeCl<sub>2</sub>
- (xvi) SrI<sub>2</sub>
- (xvii) SnBr<sub>2</sub>
- (xviii) MgO
- (xix) Rb<sub>3</sub>N
- (xx) LiF
- (xxi) PtBr<sub>2</sub>
- (xxii) Bi<sub>2</sub>O<sub>3</sub>
- (xxiii) AuCl<sub>3</sub>

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. 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,

Ag	2	O		Na	Cl
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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.

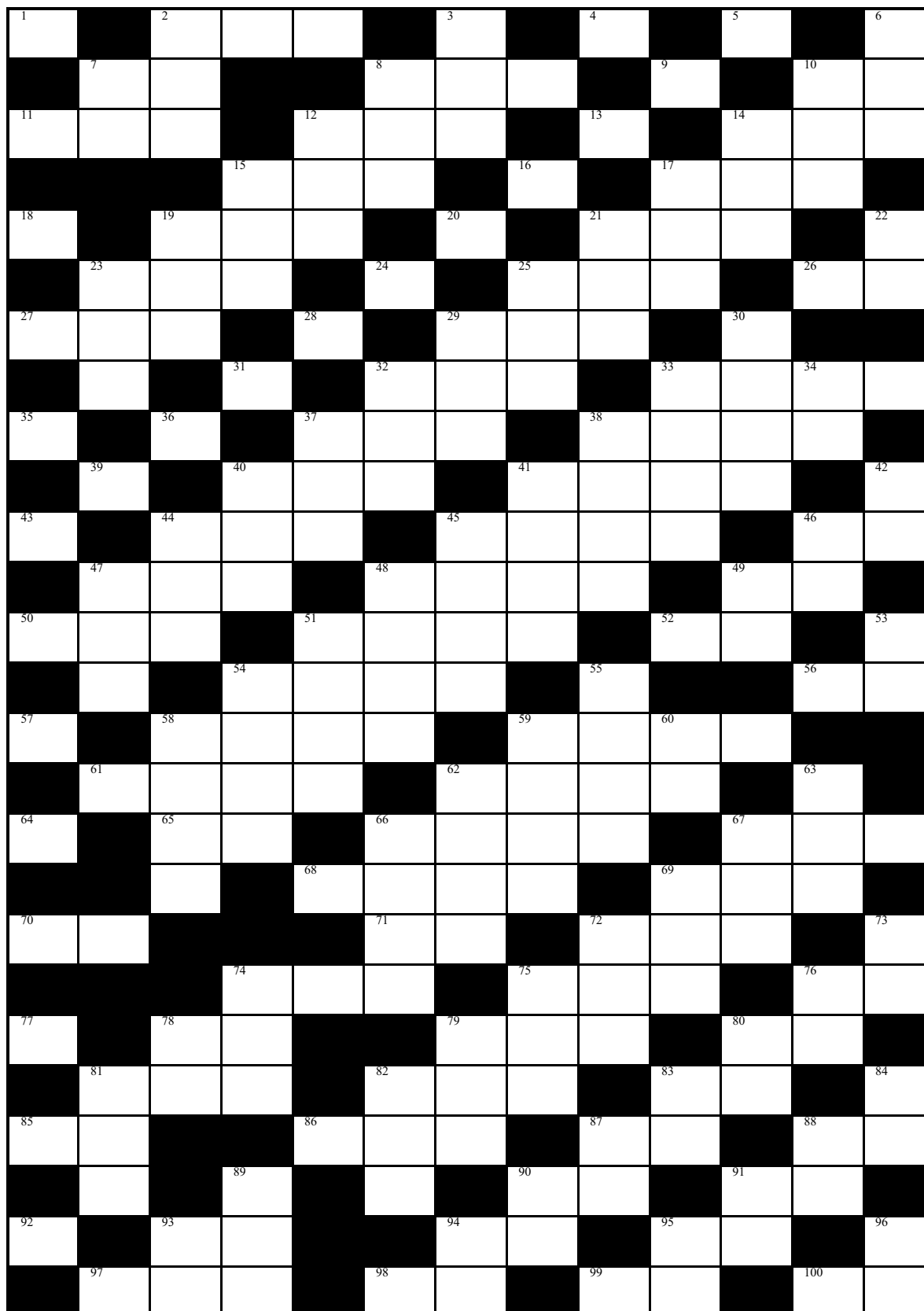
**CLUES.****ACROSS**

1. nickel
2. iron(III) bromide
4. tellurium
5. germanium
7. potassium chloride
8. sodium sulfide
9. silicon
10. calcium oxide
11. chromium(III) iodide
12. potassium oxide
13. thallium
14. strontium fluoride
15. rubidium selenide
16. indium
17. barium chloride
19. caesium sulfide
20. gallium
21. copper(II) bromide
23. silver(I) oxide
25. nickel(II) iodide
26. chromium(III) nitride
27. copper(I) selenide
29. iron(II) fluoride
32. zinc chloride
33. barium nitride
36. argon
37. cadmium bromide
38. magnesium nitride
40. lead(II) iodide
41. nickel(II) phosphide
43. helium
44. tin(II) fluoride
45. iron(II) phosphide
46. zinc selenide
47. mercury(II) chloride
48. zinc nitride

**DOWN**

2. iron(III) chloride
3. lithium oxide
6. tin(IV) oxide
7. potassium iodide
8. sodium selenide
10. calcium fluoride
12. potassium sulfide
14. strontium chloride
15. rubidium oxide
17. barium bromide
18. platinum
19. caesium selenide
21. copper(II) iodide
22. aluminium nitride
23. silver(I) sulfide
24. radon
25. nickel(II) fluoride
28. xenon
29. iron(II) chloride
30. strontium nitride
31. krypton
32. zinc bromide
33. barium phosphide
34. nitrogen (molecule)
35. arsenic
37. cadmium iodide
38. magnesium phosphide
39. neon
40. lead(II) fluoride
41. nickel(II) nitride
42. mercury(II) selenide
44. tin(II) chloride
45. iron(II) nitride
46. zinc oxide
47. mercury(II) iodide
48. zinc phosphide

- |      |                      |     |                        |
|------|----------------------|-----|------------------------|
| 49.  | platinum(II) oxide   | 49. | platinum(II) sulfide   |
| 50.  | manganese(II) iodide | 51. | copper(II) phosphide   |
| 51.  | copper(II) nitride   | 53. | aluminium phosphide    |
| 52.  | calcium sulfide      | 54. | lead(II) nitride       |
| 54.  | lead(II) phosphide   | 55. | gold(III) oxide        |
| 56.  | iron(III) phosphide  | 57. | antimony               |
| 58.  | tin(II) phosphide    | 58. | tin(II) nitride        |
| 59.  | bismuth(III) oxide   | 59. | bismuth(III) sulfide   |
| 61.  | mercury(II) nitride  | 60. | ozone                  |
| 62.  | chromium(III) oxide  | 62. | chromium(III) selenide |
| 64.  | bismuth              | 63. | lithium phosphide      |
| 65.  | nitrogen (molecule)  | 66. | aluminium oxide        |
| 66.  | aluminium sulfide    | 67. | copper(I) nitride      |
| 67.  | copper(I) phosphide  | 69. | silver(I) nitride      |
| 68.  | iron(III) selenide   | 72. | caesium phosphide      |
| 69.  | silver(I) phosphide  | 73. | manganese(II) sulfide  |
| 70.  | gold(III) nitride    | 74. | bismuth(III) iodide    |
| 71.  | ozone                | 75. | rubidium phosphide     |
| 72.  | caesium nitride      | 76. | iron(II) oxide         |
| 74.  | bismuth(III) bromide | 77. | manganese              |
| 75.  | rubidium nitride     | 78. | sodium fluoride        |
| 76.  | iron(II) sulfide     | 79. | potassium nitride      |
| 78.  | sodium iodide        | 80. | nickel(II) sulfide     |
| 79.  | potassium phosphide  | 81. | aluminium chloride     |
| 80.  | nickel(II) oxide     | 82. | sodium nitride         |
| 81.  | aluminium fluoride   | 83. | copper(II) oxide       |
| 82.  | sodium phosphide     | 84. | mercury(II) oxide      |
| 83.  | copper(II) sulfide   | 87. | magnesium selenide     |
| 85.  | lithium chloride     | 88. | lead(II) sulfide       |
| 86.  | lithium nitride      | 89. | lead(IV) fluoride      |
| 87.  | magnesium oxide      | 90. | barium sulfide         |
| 88.  | lead(II) oxide       | 91. | tin(II) oxide          |
| 90.  | barium selenide      | 92. | lead                   |
| 91.  | tin(II) sulfide      | 93. | rubidium chloride      |
| 93.  | rubidium fluoride    | 94. | strontium oxide        |
| 94.  | strontium sulfide    | 95. | cadmium sulfide        |
| 95.  | cadmium oxide        | 96. | gold(III) nitride      |
| 97.  | tin(IV) chloride     |     |                        |
| 98.  | calcium oxide        |     |                        |
| 99.  | zinc sulfide         |     |                        |
| 100. | bismuth(III) nitride |     |                        |

**CHEMICAL CROSSWORD No. 2****ELEMENTS AND BINARY COMPOUNDS OF METALS WITH NON-METALS**

## ANSWERS TO TUTORIAL TOPIC 3

- 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  $X^{2+}$ .
- 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  $Mg_3N_2$ . 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	$Na_2S$
(v) magnesium nitride	$Mg_3N_2$	(vi) rubidium chloride	RbCl
(vii) caesium phosphide	$Cs_3P$	(viii) potassium iodide	KI
(ix) calcium selenide	CaSe	(x) strontium chloride	$SrCl_2$
(xi) lithium oxide	$Li_2O$	(xii) magnesium bromide	$MgBr_2$
(xiii) rubidium nitride	$Rb_3N$	(xiv) calcium fluoride	$CaF_2$
(xv) aluminium sulfide	$Al_2S_3$	(xvi) caesium selenide	$Cs_2Se$
(xvii) barium phosphide	$Ba_3P_2$	(xviii) sodium nitride	$Na_3N$
(xix) potassium chloride	KCl	(xx) strontium iodide	$SrI_2$

- 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) $Cu_2O$
(v) $Ba_3N_2$	(vi) MnS	(vii) HgO	(viii) $FeBr_2$
(ix) $Al_2O_3$	(x) $FeCl_3$		

4. (i) silver(I) 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)  $\text{CdF}_2$  (ii)  $\text{SrCl}_2$   
(iii)  $\text{CoS}$  (iv)  $\text{PbI}_2$   
(v)  $\text{SnO}$  (vi)  $\text{Fe}_2\text{O}_3$   
(vii)  $\text{CrN}$  (viii)  $\text{CaBr}_2$   
(ix)  $\text{K}_2\text{O}$  (x)  $\text{Na}_3\text{P}$

## CHEMICAL CROSSWORD No. 2

## ELEMENTS AND BINARY COMPOUNDS OF METALS WITH NON-METALS

1 Ni		2 Fe	Br	3		3 Li		4 Te		5 Ge		6 Sn
	7 K	Cl			8 Na	2	S		9 Si		10 Ca	O
11 Cr	I	3		12 K	2	O		13 Tl		14 Sr	F	2
			15 Rb	2	Se		16 In		17 Ba	Cl	2	
18 Pt		19 Cs	2	S		20 Ga		21 Cu	Br	2		22 Al
	23 Ag	2	O		24 Rn		25 Ni	I	2		26 Cr	N
27 Cu	2	Se		28 Xe		29 Fe	F	2		30 Sr		
	S		31 Kr		32 Zn	Cl	2		33 Ba	3	34 N	2
35 As		36 Ar		37 Cd	Br	2		38 Mg	3	N	2	
	39 Ne		40 Pb	I	2		41 Ni	3	P	2		42 Hg
43 He		44 Sn	F	2		45 Fe	3	P	2		46 Zn	Se
	47 Hg	Cl	2		48 Zn	3	N	2		49 Pt	O	
50 Mn	I	2		51 Cu	3	N	2		52 Ca	S		53 Al
	2		54 Pb	3	P	2		55 Au			56 Fe	P
57 Sb		58 Sn	3	P	2		59 Bi	2	O	3		
	61 Hg	3	N	2		62 Cr	2	O	3		63 Li	
64 Bi		65 N	2		66 Al	2	S	3		67 Cu	3	P
		2		68 Fe	2	Se	3		69 Ag	3	P	
70 Au	N			71 O	3		72 Cs	3	N			73 Mn
			74 Bi	Br	3		75 Rb	3	N		76 Fe	S
77 Mn		78 Na	I			79 K	3	P		80 Ni	O	
	81 Al	F	3		82 Na	3	P		83 Cu	S		84 Hg
85 Li	Cl			86 Li	3	N		87 Mg	O		88 Pb	O
	3		89 Pb		N		90 Ba	Se		91 Sn	S	
92 Pb		93 Rb	F		94 Sr	S		95 Cd	O			96 Au
	97 Sn	Cl	4		98 Ca	O		99 Zn	S		100 Bi	N