4 Ionic compounds

4.1 Overview

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4.1.1 Introduction

To construct new substances such as smart materials, self-repairing materials and nanotubes, scientists and engineers need to understand the way atoms are put together and the limitations of the bonding models. The way that the atoms in a material are bonded together directly affects its properties and performance. This topic investigates the bonding between metals and non-metals, known as ionic bonding.

FIGURE 4.1 Giant crystals of gypsum, $CaSO_4 \cdot 2H_2O$, up to 11 metres long, found in the 'Cave of crystals' in Mexico. Note the size of the crystals compared with the human figure.

4.1.2 What you will learn

KEY KNOWLEDGE

In completing this topic, you will investigate:

- common properties of ionic compounds (brittleness, hardness, high melting point, difference in electrical conductivity in solid and liquid states), with reference to their formation, nature of ionic bonding and crystal structure including limitations of representations
- experimental determination of the factors affecting crystal formation of ionic compounds
- the uses of common ionic compounds.

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PRACTICAL WORK AND INVESTIGATIONS

Practical work is a central component of learning and assessment. Experiments and investigations, supported by a **Practical investigation logbook** and **Teacher-led videos**, are included in this topic to provide opportunities to undertake investigations and communicate findings.

Resources

Digital document Key science skills (doc-30903)
Key terms glossary — Topic 4 (doc-30935)
Practical investigation logbook — Topic 4 (doc-30936)

studyON

To access key concept summaries and practice exam questions download and print the **studyON: Revision and practice exam question booklet** (doc-30937).

4.2 Structure and properties of ionic substances

KEY CONCEPT

- Common properties of ionic compounds (brittleness, hardness, high melting point, difference in electrical conductivity in solid and liquid states), with reference to their formation, nature of ionic bonding and crystal structure including limitations of representations

4.2.1 Chemical bonds

Since ancient times, humans have gathered attractive crystals to use as household decorations and jewellery. Many of these crystals consist of **ionic compounds** in which metal and non-metal atoms are joined by an **ionic bond**.

We are aware that all the varied substances around us are formed from atoms combining together. But this still leaves some questions:

- What holds atoms together once they have combined?
- Why do some atoms join or bond with others, while some remain as individual atoms?
- Why do atoms combine in specific ratios?

The answers lie in the nature of the **chemical bond**. All forces of attraction leading to chemical bonding between atoms are electrostatic in nature, that is, an attraction between protons (+) and electrons (−).

Most spontaneous changes that take place in nature occur in order to reach a more stable state. We seldom find free atoms in nature because atoms undergo changes in structure to become more stable. They do this by joining together, or bonding. Sometimes they bond with atoms of the same kind (for example, the covalent diatomic molecule of hydrogen gas, H_2) and sometimes with atoms of a different kind (for example, the ionic crystal lattice of sodium chloride, NaCl). Only the noble gases may exist as free atoms (monoatomic). This behaviour can be explained by an atom's electron configuration.

FIGURE 4.2 The mineral fluorite contains the ionic compound calcium fluoride, CaF_2, as its main component. Its beautiful crystals are often cubic in shape and appear in a variety of colours.

Atoms of other elements than group 18 (the noble gases) become more stable by gaining or losing enough electrons to achieve a complete outer shell configuration, like a noble gas. This does not mean that they have become noble gases, since they have not changed their nuclear structure (number of protons).

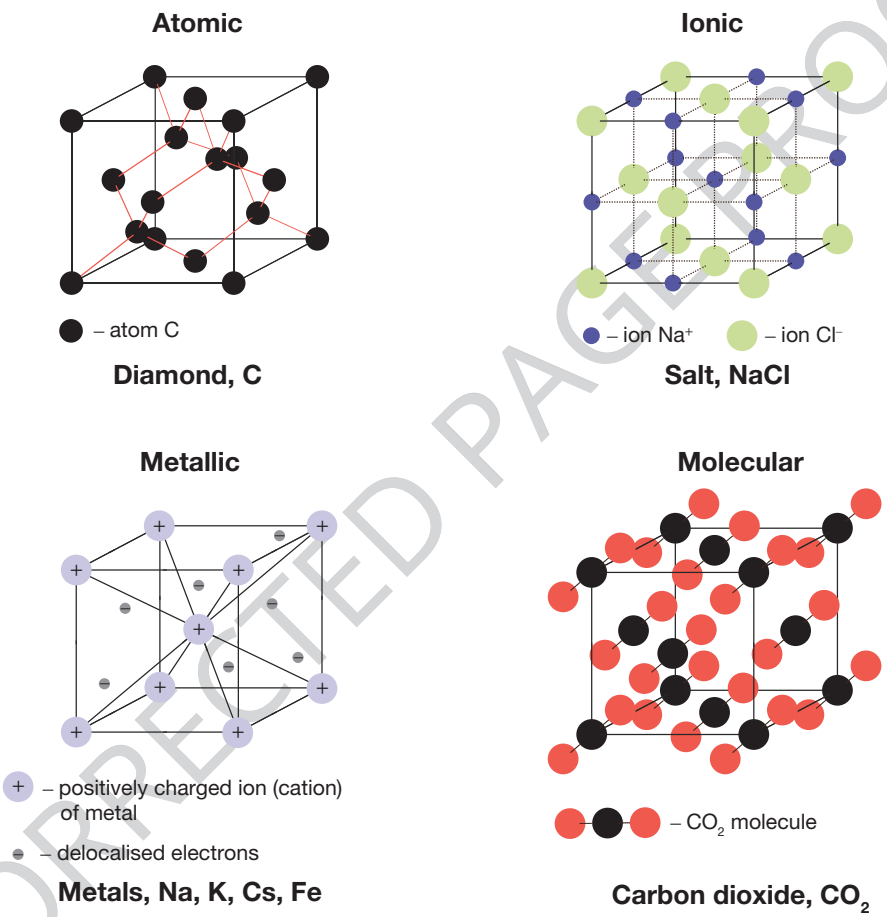
Atoms can become stable in one of three ways:

- by donating electrons to another atom
- by accepting electrons from another atom
- by sharing electrons with another atom.

When atoms combine to achieve more stable structures, three types of bonding are possible (see figure 4.3).

- An ionic bond results when metallic atoms combine with non-metallic atoms to form an ionic lattice.
- A **metallic bond** results when metallic atoms combine to form a metallic lattice.
- A **covalent bond** results when non-metallic atoms combine to form either molecules or covalent lattices.

FIGURE 4.3 Many different types of crystal lattice types exist.



4.2.2 From atoms to ions

Atoms that gain or lose electrons to achieve more stable outer shell configurations are called **ions**. When an atom becomes an ion, it is no longer neutrally charged, since the number of electrons is not equal to the number of protons (see table 4.1). Note that the number of protons remains the same. An atom that has lost electrons becomes *positively* charged and is called a **cation** (e.g. Na^+). An atom that has gained electrons becomes *negatively* charged and is called an **anion** (e.g. Cl^-).

FIGURE 4.4 All elements in the periodic table are neutral and some can become charged when they are involved in a chemical bond.

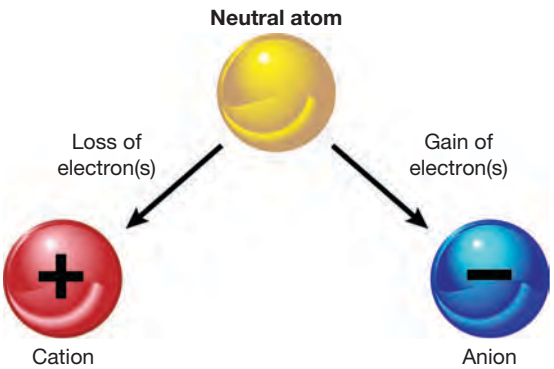


TABLE 4.1 Common atoms and their ions

Atom/ion	Symbol	Number of protons	Number of electrons
Sodium atom	Na	11	11
Sodium cation	Na^+	11	10
Chlorine atom	Cl	17	17
Chloride anion	Cl^-	17	18

The electrons in the outer shell of an atom are sometimes called the valence electrons. The number of outer shell electrons is related to the group number in the periodic table.

When an atom gains or loses electrons, an ion is formed; in other words, the atom becomes charged.

Metallic ion formation

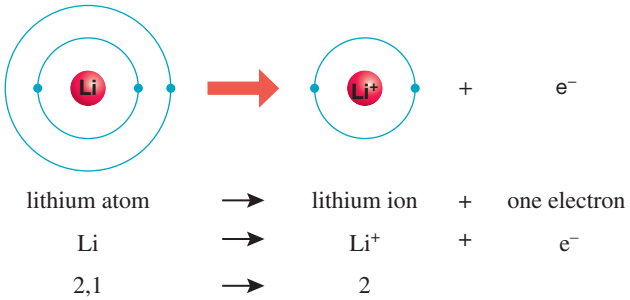
The metallic elements are those on the left side of the staircase in figure 4.5 (blue cells). These elements generally have low **electronegativities**. They can lose electrons to *achieve a noble gas configuration in their outer shells*.

FIGURE 4.5 Periodic table (up to element 89) showing the division between metals and non-metals

1	2											13	14	15	16	17	18
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac															

For example, lithium is a very reactive group 1 metal with one outer shell electron and the electron configuration 2, 1. In order to obtain the stable configuration of a full outer shell, the lone electron is lost (figure 4.6). The electron configuration, 2, of the nearest noble gas, helium, results. Since the lithium cation has three protons but only two electrons, it has a net charge of 1+. Charges are written as superscripts above and to the right of the element symbol; thus the lithium atom is now written as Li^+. This process can be represented by electron shell diagrams or in the following simple equation form.

FIGURE 4.6 Formation of the lithium cation from the lithium atom



(*Note:* When an atom's net charge is 1, it is not necessary to include the numeral 1 in the superscript notation.)

The group 2 and group 13 metals contain two and three valence electrons respectively. They lose their outer shell electrons to form ions with charges of 2+ and 3+ respectively.

The electron configurations for the simple ions formed by the metals in periods 2, 3 and 4 are identical to those of the closest noble gases. Each occupied energy shell contains the maximum number of electrons. Examples of their electron configurations are as follows:

Li^+	2
Na^+	2, 8
K^+	2, 8, 8.

When we name a metallic ion, we use the full name of the metal followed by the word 'ion' to distinguish it from the uncharged metal.

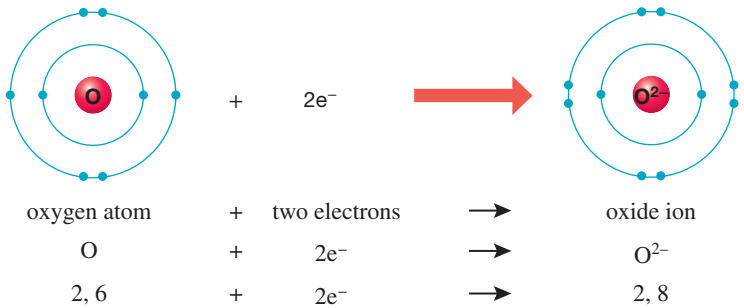
Note: The group 14 elements, carbon and silicon, do not form simple ions.

Non-metallic ion formation

Non-metallic elements are shown on the right side of the purple section in the periodic table (refer to figure 4.5). These elements generally have high electronegativities. They gain electrons to achieve a noble gas configuration of eight electrons in their outer shells (with the exception of hydrogen).

For example, oxygen in group 16 has six outer shell electrons and has the electron configuration 2, 6. It is too difficult to remove all six electrons to achieve a full outer shell, so the oxygen atom gains two electrons instead to become a stable anion, as shown in figure 4.7 and the simple equation.

FIGURE 4.7 The oxygen atom has six valence electrons, and gains two electrons to form the oxide ion, which has a charge of 2–.



An anion has more electrons than a neutral atom of the same element; here, the oxygen ion has eight protons and ten electrons, resulting in a net charge of 2–. The electron configuration of the oxygen anion is now that of a neon atom: 2, 8. The oxygen atom has become an oxide ion. (It is a convention in chemistry to indicate the ions of non-metallic elements with the suffix *-ide*.)

SAMPLE PROBLEM 1

Write the symbol, charge and name of the ions you would expect atoms of the following elements to form:

a. Mg

b. S.

Teacher-led video: SP1 (eles-3283)

THINK	WRITE
a. Mg is a metal found in group 2. It has the simplified electron configuration 2, 8, 2. In order to become stable, the Mg atom needs to lose two electrons. The charge of the resultant ion would, therefore, be 2+.	Mg^{2+} magnesium ion
b. S is a non-metal found in group 16. It has the simplified electron configuration 2, 8, 6. In order to become stable, the S atom needs two more electrons to complete its outer shell. The charge of the resultant ion would, therefore, be 2–.	S^{2-} sulfide ion

PRACTICE PROBLEM 1

Write the symbol, charge and name of the ions you would expect atoms of the following elements to form:

a. Ga

b. P.

4.2.3 Ionic bonding

Metals and non-metals can react together because the electrons lost by the metals can be taken up by the more electronegative non-metals. This transfer of electrons results in an ionic bond between the metal cations and the non-metal anions.

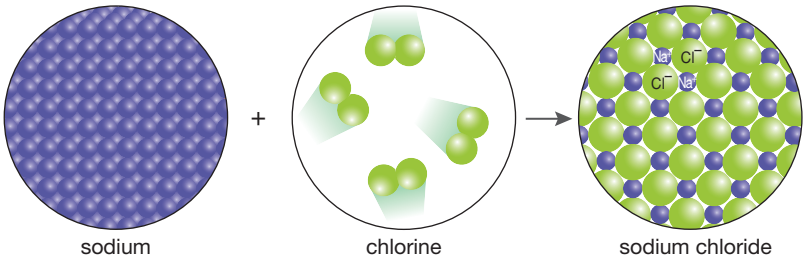
For example, when a small piece of sodium metal is added to a flask containing chlorine gas, a chemical reaction occurs, as shown in figure 4.8. The sodium ignites, and the solid ionic compound sodium chloride forms. This involves a transfer of one electron from the sodium atom to the chlorine atom. The ways in which we can represent this process are shown in the solution to sample problem 2.

Sodium ions and chloride ions are oppositely charged and, therefore they attract each other by **electrostatic attraction**. This strong force of attraction between positive and negative ions is what produces an ionic bond.

Note that the bonded form is more stable than the unbonded form and has different properties from those of the original elements. Sodium is a very reactive metal; chlorine is a poisonous gas. Sodium chloride, however, is a non-toxic ionic compound familiar to us as table salt. The term 'salt', however, has a wider meaning in chemistry. The general name for an ionic compound, a substance formed from the reaction between a metal and a non-metal, is a **salt**.

FIGURE 4.8 Molten sodium metal reacts violently with chlorine gas to form sodium chloride crystals.

FIGURE 4.9 The formation of sodium chloride from solid sodium and gaseous chlorine



on Resources

Interactivity Ionic models (int-6351)

Weblink Reaction of sodium with chlorine

SAMPLE PROBLEM 2

Consider the reaction between sodium and chlorine and explain it using:

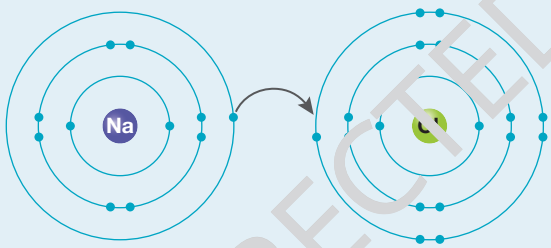
a. electron shell diagrams

b. a simple equation.

Teacher-led video: SP2 (eles-3210)

THINK

a. Sodium and chlorine can be represented as follows:

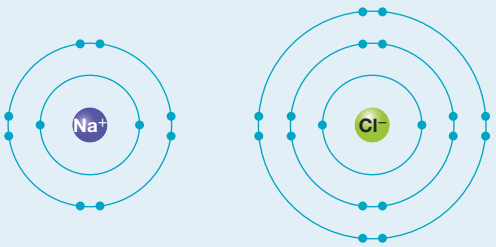


2, 8, 1 2, 8, 7

When a sodium atom forms an ionic bond with a chlorine atom, the sodium atom loses one electron (its entire outer shell) to become a positive ion, or cation. The chlorine atom gains one electron to complete its outer shell. It is now a negative ion, or anion.

b. The sodium and chloride ions form the substance called sodium chloride.

WRITE



2, 8 2, 8, 8

$$Na^+ + Cl^- \rightarrow NaCl$$

2, 8 2, 8, 8

PRACTICE PROBLEM 2

Consider the following ionic compounds and explain the bonds between their atoms using electron shell diagrams and simple equations:

a. magnesium oxide

b. magnesium fluoride.

TIP: It is sometimes useful to use dots and crosses when drawing shell diagrams because this makes it easier to see the origin of the transferred electron(s).

4.2.4 Structure and properties of ionic substances

Like most ionic substances, sodium chloride is a hard, brittle crystalline solid at room temperature. It has a high melting point (800 °C). This indicates that its particles are arranged in repeating three-dimensional patterns and are strongly held together. When sodium chloride is dissolved in water or is melted, its crystal structure breaks down and in both these states it is able to conduct electricity, indicating that charged particles (ions) are free to move.

How can we explain the distinctive composition, melting point and conductivity characteristics of ionic substances? What is it about the structure of these substances that makes them behave so uniquely?

Ionic lattices

The major features of the current theory of the structure of ionic substances are as follows:

- Ionic compounds form crystals that are composed of three-dimensional arrays of positive metal ions and negative non-metal ions.
- These arrays are called ionic network lattices and are held together by the strong electrostatic attraction — the ionic bond — between the oppositely charged ions.
- The ions are arranged in a regular repeating pattern throughout the crystal. They pack together in the way that will achieve the most stable arrangement, with oppositely charged ions as close together as possible, and similarly charged ions as far apart as possible. Each cation is thus surrounded by anions and each anion by cations.
- The relative numbers of cations and anions are fixed by the requirement that the solid is neutral.
- The relative sizes and numbers of the ions present determine the actual lattice structure.

Sodium chloride is a typical ionic compound. Its appearance and ionic structure are shown in figures 4.10 and 4.11. Its crystals are cubic, reflecting the regular arrangement of the ions in the lattice. Each sodium ion is surrounded by six chloride ions and each chloride ion is surrounded by six sodium ions. In this arrangement, each ion is strongly attracted to each of its neighbours. The large attractive forces result in a very stable structure.

FIGURE 4.10 Sodium chloride lattice

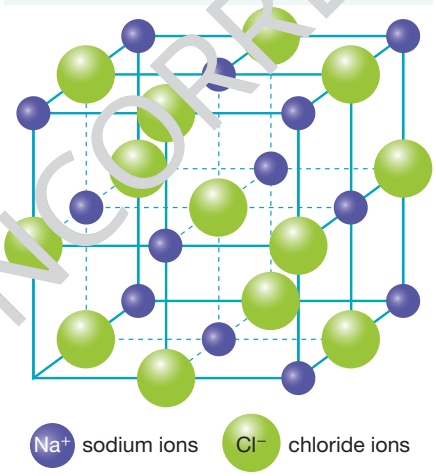


FIGURE 4.11 Each sodium ion is surrounded by six chloride ions, which gives the crystals of sodium chloride their characteristic cubic shape.

The formula of the ionic substance sodium chloride is NaCl, which means that the ratio of positive to negative ions is 1:1. The formula NaCl does not represent a distinct molecule of sodium chloride but simply means that a sample of sodium chloride contains equal numbers of sodium ions and chloride ions. The formula $MgBr_2$ signifies that a sample of magnesium bromide contains twice as many bromide ions as magnesium ions.

A chemical formula given for an ionic compound is called an empirical formula. This refers to the simplest ratio of ions in the lattice.

on Resources

Digital document	Experiment 4.1 Investigating calcite crystals (doc-30867)
Teacher-led video	Experiment 4.1 Investigating calcite crystals (eles-3270)
Interactivity	Pass the salt (int-0675)
Weblink	Animation of NaCl structure

4.2.5 Connecting properties of ionic compounds to structure

Ionic compounds are composed of two or more different kinds of ions that have opposite charges. The electrostatic forces of attraction holding them together are strong. As a result, ionic compounds:

- are usually crystalline solids, owing to the arrangement of ions in repeating three-dimensional patterns
- have high melting and boiling points, because a large amount of energy is needed to separate the ions
- do not conduct electricity in the solid form, because the charged particles (ions) are not able to move (to conduct electricity, a substance must contain mobile charged particles)
- are hard, since the surface of the crystal is not easily scratched due to the strong ionic bond holding the ions together
- often dissolve in water to form ions (dissociate), because water molecules are able to move between ions and free them by disrupting the rigid crystal structure
- conduct electricity when molten or in aqueous solution (dissolved in water). In the molten form, ions are able to slide past one another and can, therefore, conduct electricity. When an ionic substance is dissolved in water, the ions dissociate from the lattice and can move freely to conduct an electric current. The solution is called an **electrolyte** (figure 4.13).

FIGURE 4.12 Dissociation of sodium chloride to form a conducting solution called an electrolyte

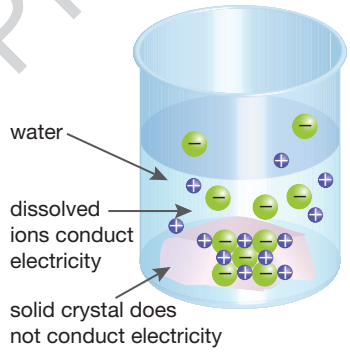
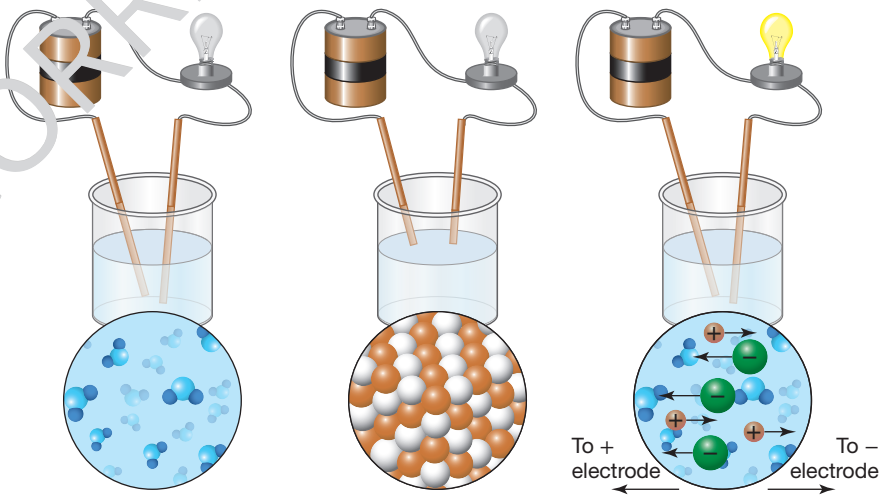


FIGURE 4.13 Electrical properties of distilled water, solid ionic compounds and ionic solutions



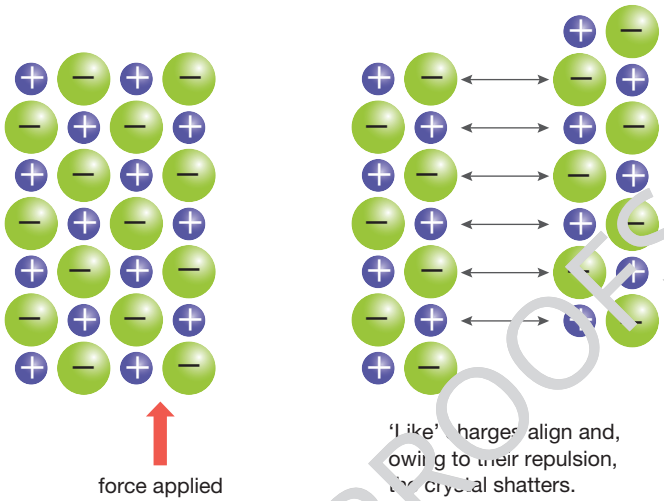
A current cannot flow through distilled water.

A current cannot flow through a solid ionic compound.

A current can flow through a solution of positive and negative ions.

- are brittle, since distortion of the crystal causes ions of like charge to come close together and the repulsion between these ions cleaves or shatters the crystal, as shown in figure 4.14. If the distortion of the crystal is caused by a sharp blow along the plane of ions, cleavage (a clean split) along the plane occurs. If the blow is along different planes of ions, the crystal lattice shatters.

FIGURE 4.14 The brittle nature of an ionic compound, which cleaves or shatters when distorted.



TIP: When explaining conductivity in ionic substances, remember that it is not the electrons that move. Only the ions are mobile in solution and when molten, not the electrons.

FIGURE 4.15 Old alkaline batteries often leak potassium hydroxide, which reacts with carbon dioxide in the air to form crystals of potassium carbonate.

Resources

Video eLesson	Salt dissociation and the formation of electrolytes
Digital document	Experiment 4.2 Ionic models (doc-30868)
Teacher-led video	Experiment 4.2 Ionic models (doc-3271)

4.2.6 Limitations of ionic models

Electron dot diagrams

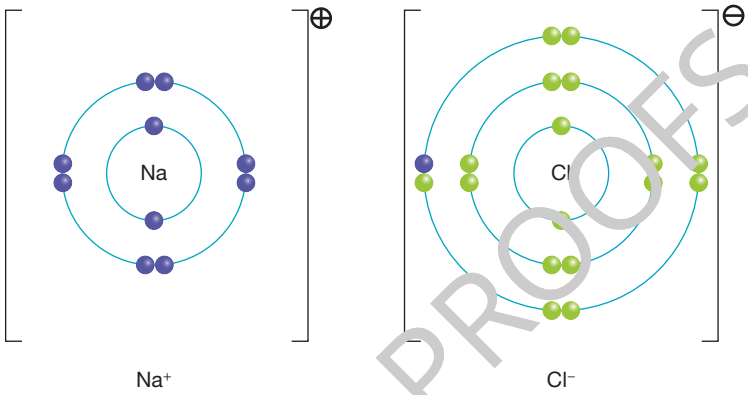
An electron dot diagram shows:

- how the ionic bonds are formed
- the ratio in which the atoms react.

However, it does not show how the ions are arranged in space.

The electron dot diagram for the sodium chloride ions shown in figure 4.16 suggests that it is made up of one pair of sodium and chloride ions. This is not the case. Sodium chloride has a giant ionic lattice structure with equal numbers of sodium and chloride ions. In a 1 g sample of pure sodium chloride, for example, will be roughly 5×10^{21} of each ion.

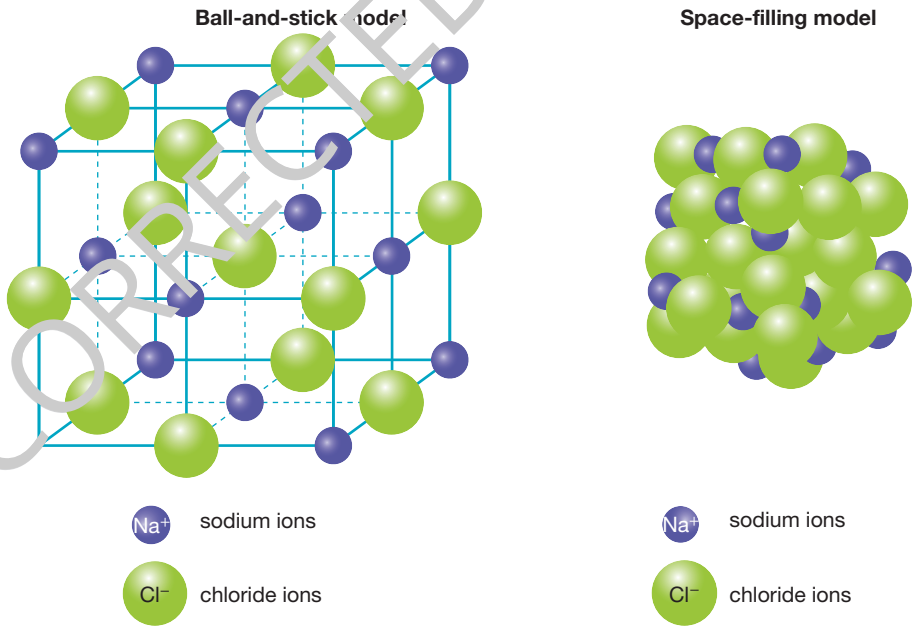
FIGURE 4.16 In dot diagrams, brackets are often used to indicate that the charge is spread over the whole ion.



Three-dimensional models of ionic compounds

A three-dimensional (3D) model shows how the ions are arranged in a lattice structure. These models usually have coloured balls to represent the ions and use sticks to show the ionic bonds (the ball-and-stick model), or they can be space-filling models (see figure 4.17).

FIGURE 4.17 Three dimensional representations of NaCl using the ball-and-stick model and the space-filling model.



A 3D model still has limitations:

- it is not to scale
- it gives no information about the forces of attraction between the ions
- the movement of electrons to form the ions cannot be shown
- it does not show the relative sizes of the ions.

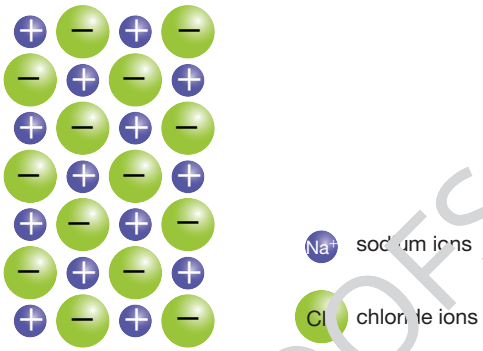
Two-dimensional models of ionic compounds

A two-dimensional (2D) representation is similar to the three-dimensional ball-and-stick model, but is an easier model to draw (see figure 4.18).

In addition to the 3D model limitations, the 2D model has further limitations:

- it only shows the arrangement of one layer of ions
- it does not show where the ions are located on the other layers or the relationship between them
- it does not show that many different arrangements of ions are possible.

FIGURE 4.18 A 2D model for the ionic lattice in sodium chloride



4.2 EXERCISE

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1. Explain why atoms such as Ca and Al form ions.
2. Use electron shell diagrams and simple equations to show how the following atoms form their corresponding ions.
   (a) The metal atoms Ca and Al
   (b) The non-metal atoms N and F
3. For each of the following atoms, predict the charge of the ion formed, and write its name and symbol.
   (a) Ba
   (b) K
   (c) P
   (d) Cl
   (e) S
4. Complete the following table, predicting the general electrovalencies for period 2 and period 3 ions in each group.

Period	Group 1	Group 2	Group 13	Group 15	Group 16	Group 17
2	1+					
3						

5. Draw electron shell structures and write simple equations to show the ionic bonding between the following.
   (a) Calcium and oxygen
   (b) Beryllium and chlorine
   (c) Lithium and fluorine
   (d) Aluminium and sulfur
   (e) Sodium and nitrogen
   (f) Magnesium and sulfur
6. Write the electron configurations of the following.
   (a) The cations potassium, calcium and aluminium
   (b) The anions fluoride, oxide and nitride
7. Name the noble gas that has the same electron configuration as the anions in fluoride, oxide and nitride.
8. Magnesium chloride, $MgCl_2$, and potassium chloride, KCl, are typical ionic compounds.
   (a) Describe how their ionic lattice structure is formed from their respective atoms.
   (b) Show why their empirical formulas have ratios of 1:2 and 1:1 respectively.
   (c) Explain what holds the ions close together in their lattice structures.
9. Calcium chloride is a crystalline substance at room temperature. Predict whether it
   (a) has a low or high melting point.
   (b) shatters when pressure is applied.
   (c) conducts electricity in the solid or liquid state.
   Justify your predictions.

10. (a) Compare the ratio of positive to negative ions in the lattice of the following.
    i. Sodium chloride
    ii. Magnesium oxide
    (b) Which of these compounds would you expect to have the higher melting point? Give a reason for your choice.
11. A solid substance has a high melting point, conducts electricity only in molten form and cleaves when struck with a sharp knife. However, it does not readily dissolve in water. Is this enough data to predict the chemical bonding in this substance? Explain in terms of the limitations of the bonding model.
12. Atom *X*, having one electron in its outer shell, combines with atom *Y*, which has six electrons in its outer shell.
    (a) Write the formula of the compound formed.
    (b) Name the type of bonding in the compound.
    (c) Predict three general properties of the compound.

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4.3 Ionic nomenclature

Background knowledge

- Naming conventions of ionic compounds
- Formulas for binary ionic compounds, ions with variable charges and polyatomic compounds

4.3.1 Naming ionic compounds

Knowing how to write formulas is a skill that all chemists require. Many known compounds exist and to memorise them all would be impossible. Chemists have developed a set of rules that allow us to predict the formulas of many compounds. Some compounds that have been known for many years also have common names, as listed in table 4.2.

TABLE 4.2 Common ionic compounds

Common name	Systematic name	Formula	Used to manufacture
Salt	Sodium chloride	NaCl	Chlorine
Soda ash	Sodium carbonate	Na_2CO_3	Glass
Baking soda	Sodium hydrogen carbonate	$NaHCO_3$	Cake raising agent
Lime	Calcium oxide	CaO	Mortar
Limestone	Calcium carbonate	$CaCO_3$	Cement
Potash	Potassium nitrate	KNO_3	Gunpowder
Milk of magnesia	Magnesium hydroxide	$Mg(OH)_2$	Laxative or antacid
Gypsum	Calcium sulfate	$CaSO_4$	Plaster
Caustic soda	Sodium hydroxide	NaOH	Soap or drain cleaner

4.3.2 Formulas of binary ionic compounds

In a **binary** ionic compound, only two elements are present. Examples include sodium chloride, NaCl, calcium fluoride, CaF_2, and potassium nitride, K_3N. Note that the word 'binary' does not refer to the relative number of ions or subscripts. When writing empirical formulas for binary ionic compounds, the following rules should be followed:

- Write the symbol for the cation first, followed by the symbol for the anion.
- Determine the lowest whole number ratio of ions that provides a net charge of zero.

The 'swap and drop' method is another way to determine formulas, using the charges of the ions. For example, to obtain the formula for aluminium sulfide, the symbol for the cation is written first, followed by the symbol for the anion: $Al^{3+}S^{2-}$. The charges are then swapped and dropped so that the charge of the anion becomes the subscript of the cation and the charge of the cation becomes the subscript of the anion.

The formula is then written showing only the subscripts. Note that subscripts must always be reduced to the smallest possible whole numbers. (For example, the formula of magnesium oxide is MgO, not Mg_2O_2.)

FIGURE 4.19 Sodium hydrogen carbonate, commonly known as baking soda, thermally decomposes at 80 °C to give off carbon dioxide gas and water vapour, which act as raising agents.

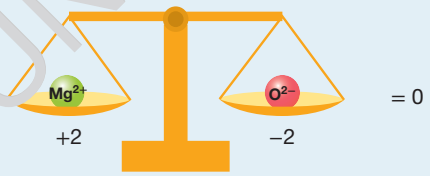
SAMPLE PROBLEM 3

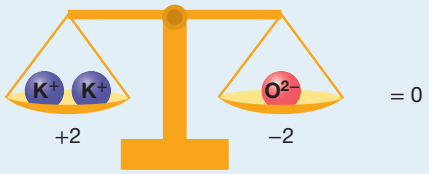
Write formulas for the following compounds:

a. **magnesium oxide**
b. **potassium oxide**
c. **aluminium sulfide.**

Teacher-led video: SP3 (eles-3284)

THINK	WRITE
a. The symbols for the cation and anion in the compound are Mg^{2+} and O^{2-} respectively. In order to have a net charge of zero, the number of positive charges must be balanced by an equal number of negative charges. Thus one Mg^{2+} ion is balanced by one O^{2-} ion and the ratio of ions is 1:1.	MgO
b. The symbols for the cation and anion in the compound are K^+ and O^{2-} respectively. In order to have a net charge of zero, two positively charged potassium ions are needed to balance the two negative charges of the oxide ion. The ratio of ions is 2:1.	K_2O

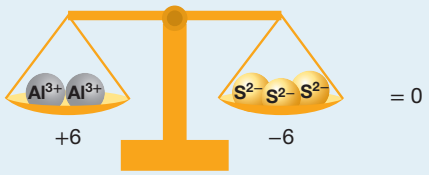




The formula is written as K_2O. The subscript numeral '2' indicates two potassium ions are present for every oxide ion.

c. The symbols for the cation and anion in the compound are Al^{3+} and S^{2-} respectively. In order to have a net charge of zero, two Al^{3+} ions, with a total charge of 6+, are needed to balance three S^{2-} ions, with a total charge of 6–. The ratio of ions is 2:3.

Al_2S_3



PRACTICE PROBLEM 3

Name the compound MgF_2.

on Resources

Weblink Balanced ionic compounds quiz

4.3.2 Ions with a variable charge

As the atomic number increases, atoms become larger and have more electrons. Some transition metals have quite complicated arrangements of electrons and may be able to form more than one type of ion. Such ions have different charges. This is very significant because ions of the same element with different charges have different physical and chemical properties. For example, a solution of Cr^{6+} is orange, whereas a solution of Cr^{3+} is yellow (see figure 4.20).

Elements that have ions of variable charge can form two or more different binary ionic compounds containing the same element. Iron, for example, can form black FeO and brown Fe_2O_3, each with different physical and chemical properties. This means we must name them differently, since simply saying iron oxide is ambiguous. Therefore, we use roman numerals in brackets after the cation's name to denote its charge. For example, the black FeO contains Fe^{2+} ions, and so it is called iron(II) oxide. The brown Fe_2O_3 contains Fe^{3+} ions and is named iron(III) oxide.

FIGURE 4.20 When chromium reacts with other chemicals, often Cr^{3+} (yellow flask) changes to Cr^{6+} (orange flask). The colour change can indicate that a reaction has occurred.

Metal ions that form more than one cation include iron(II) and (III), copper(I) and (II), and mercury(I) and (II). Note that the roman numerals do not indicate the numbers of ions present in the compound, only the charge of the ion.

SAMPLE PROBLEM 4

a. Name the compound CuO.
b. Give the formula for iron(II) chloride.

Teacher-led video: SP4 (eles-3285)

THINK	WRITE
a. In order to determine the charge on the metal ion, we need to work backwards from the known charge on the anion. The charge on the oxide ion is 2–. Therefore, for a net charge of zero, the charge on the Cu ion must be 2+. The compound is, therefore, named copper(II) oxide.	copper(II) oxide
b. Iron(II) means that the Fe^{2+} ion is present in the compound. The chloride ion has a charge of 1–. For a net charge of zero, the 2+ of the Fe ion must be balanced with two chloride ions.	$FeCl_2$

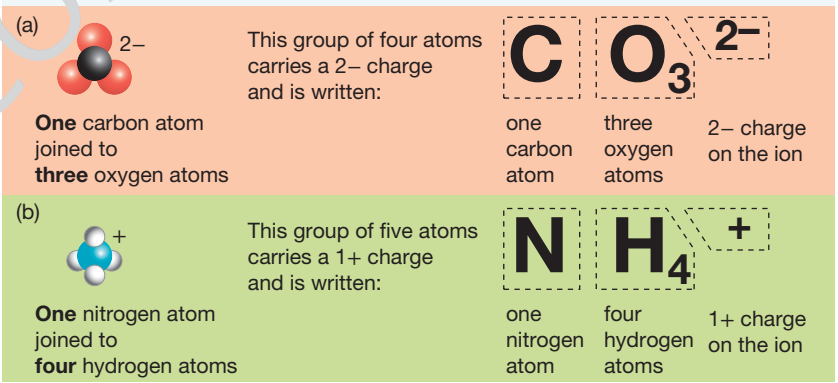
PRACTICE PROBLEM 4

a. Name the compound of SrI_2.
b. Give the formula for vanadium(V) oxide.

4.3.4 Polyatomic ions

A **polyatomic ion** is a group of tightly bound atoms that behaves as a single unit and carries an overall charge. It may be a positively or negatively charged ion. The carbonate ion, for example, is composed of one carbon atom and three oxygen atoms. The *whole group* of four atoms carries a negative two charge.

FIGURE 4.21 Two polyatomic ions: (a) the anion carbonate (CO_3^{2-}) and (b) the cation ammonium (NH_4^+).



Ionic compounds consist of arrays of cations and anions. A table of charges may be used to help write the empirical formula of an ionic compound (see table 4.3).

TABLE 4.3 Anions and cations

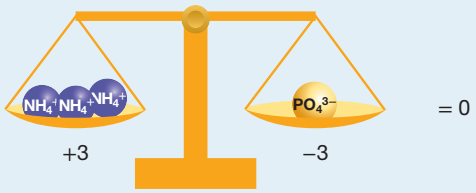
Cations					
1^+		2^+		3^+	
Lithium	Li^+	Magnesium	Mg^{2+}	Aluminium	Al^{3+}
Sodium	Na^+	Calcium	Ca^{2+}	Chromium(III)	Cr^{3+}
Potassium	K^+	Barium	Ba^{2+}	Iron(III)	Fe^{3+}
Caesium	Cs^+	Iron(II)	Fe^{2+}		
Silver	Ag^+	Nickel	Ni^{2+}		
Copper(I)	Cu^+	Chromium(II)	Cr^{2+}		
Ammonium	NH_4^+	Copper(II)	Cu^{2+}		
		Zinc	Zn^{2+}		
		Tin(II)	Sn^{2+}		
		Lead(II)	Pb^{2+}		
		Manganese(II)	Mn^{2+}		
		Mercury(II)	Hg^{2+}		
		Strontium	Sr^{2+}		
Anions					
1^-		2^-		3^-	
Hydride	H^-	Oxide	O^{2-}	Nitride	N^{3-}
Fluoride	F^-	Sulfide	S^{2-}	Phosphide	P^{3-}
Chloride	Cl^-	Sulfate	SO_4^{2-}	Phosphate	PO_4^{3-}
Bromide	Br^-	Carbonate	CO_3^{2-}		
Iodide	I^-	Sulfite	SO_3^{2-}		
Hydroxide	OH^-	Dichromate	$Cr_2O_7^{2-}$		
Nitrate	NO_3^-	Chromate	CrO_4^{2-}		
Hydrogen carbonate	HCO_3^-	Thiosulfate	$S_2O_3^{2-}$		
Hydrogen sulfate	HSO_4^-	Hydrogen phosphate	HPO_4^{2-}		
Chlorate	ClO_3^-				
Hydrogen sulfite	HSO_3^-				
Nitrite	NO_2^-				
Permanganate	MnO_4^-				
Hypochlorite	OCl^-				
Dihydrogen phosphate	$H_2PO_4^-$				
Cyanide	CN^-				

SAMPLE PROBLEM 5

Write the formula for the compound ammonium phosphate.

Teacher-led video: SP5 (eles-3286)

THINK	WRITE
1. The method for writing the formula for this compound is the same as that for a binary ionic compound. Write the symbol for the cation followed by the symbol for the anion.	Cation: NH_4^+ Anion: PO_4^{3-}.
2. Since the net charge must be zero, we need three NH_4^+ ions for every PO_4^{3-} ion.	$(NH_4)_3PO_4$



Place brackets around the entire ammonium ion so that the subscript '3' applies to all of it.
Note that brackets need to be used whenever more than a single polyatomic ion is needed to balance a formula. This is the only time they are used.

PRACTICE PROBLEM 5

Write the formula for the compound sodium dichromate.

4.3 EXERCISE

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1. Write the formula for the ionic compound formed between each of the following sets of ions.
   (a) Aluminium and chlorine
   (b) Barium and oxygen
   (c) Sodium and sulfur
   (d) Magnesium and phosphorus
2. (a) Complete the following table by writing the formulas for the compounds formed when each cation is bonded to each anion.

Ions	K^+	Ca^{2+}	Al^{3+}
F^-			
O^{2-}			
N^{3-}			

   (b) Name the nine compounds formed in part (a).
3. Name the following compounds.
   (a) KCl
   (b) Ag_2S
   (c) Na_3N

4. Name the following compounds.
   (a) FeS
   (b) Fe_2S_3
   (c) $CuCl_2$
5. Name the following compounds.
   (a) SnO
   (b) Cu_2O
   (c) $PbBr_2$
6. Give the formula for each of the following compounds.
   (a) Tin(IV) fluoride
   (b) Lead(II) sulfide
   (c) Mercury(II) oxide
   (d) Iron(III) nitride
   (e) Copper(I) sulfide
   (f) Tin(II) oxide
7. Refer to table 4.3. Write the formula and name of each of the polyatomic ions that contains sulfur.
8. Using the charges provided in table 4.3, complete the following table.

Name of ionic compound	Valency of cation	Valency of anion	Empirical formula
Silver chloride			
Potassium sulfide			
Magnesium oxide			
Aluminium bromide			
Iron(III) carbonate			
Barium phosphate			
Ammonium sulfate			

9. Give the formulas for the following compounds.
   (a) Sodium sulfite
   (b) Calcium nitrite
   (c) Copper(II) hydrogen carbonate
10. Name the compound formed when each of the following pairs of ions is bonded, and write its formula.
   (a) Al^{3+} and CO_3^{2-}
   (b) Sodium and nitrate
   (c) Hg^{2+} and PO_4^{3-}
   (d) Lead(II) and sulfate

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4.4 Crystal formation and uses of ionic compounds

KEY CONCEPT

- Experimental determination of the factors affecting crystal formation of ionic compounds
- The uses of common ionic compounds

4.4.1 Factors affecting crystal formation

Ionic compounds form crystals of many interesting shapes depending on the arrangement of ions in the lattice, but the different sizes of crystals are caused by the diverse conditions in which they are formed. The largest natural crystals found have been in Chihuahua, Mexico, and these are up to 11 metres long, while the smallest fit into the nanoscale range.

FIGURE 4.22 Vanadinite, $Pb_5(VO_4)_3Cl$, crystals are hexagonal in shape and are a source of the element vanadium.

Crystals of a more manageable size can be prepared in a school laboratory from saturated solutions, which are solutions in which no more solid (solute) will dissolve. To form a saturated solution, the solution is heated while the solute is added; this assists the dissolving process. As the solution cools, the water (or other solvent) evaporates and small crystals appear.

To produce a larger crystal, a regularly shaped crystal can be selected and suspended by a thread in a saturated solution and the crystal will grow. This crystal is called the seed crystal, and it provides a nucleation site, or a point where crystallisation can begin. Dust particles or scratches on a glass container can also act as nucleation sites. The rate of evaporation affects the size of the crystal formed.

Recrystallisation processes are very useful to chemists for purifying substances, because impurities cannot form part of the growing crystal structure.

4.4.2 Uses of common ionic compounds

Ionic compounds have a wide range of uses, as can be seen in table 4.2 (refer to section 4.3.1). Many metals are obtained from ionic compounds that have been extracted from different ores. Sodium hydrogen carbonate is used in cooking and also as an environmentally friendly cleaning agent. Ammonium nitrate is used to manufacture fertilisers as well as explosives. Compounds containing nitrate, nitrite and sulfite ions are used to preserve food. Our bodies can function only because of the presence of particular ions.

Why do athletes take sports drinks?

An electrolyte is a substance that dissociates to form ions. Athletes sometimes take electrolyte drinks after exercise to rehydrate, boost energy and replace important ions in the body that are lost during sweating. Adequate water consumption is essential but much discussion surrounds the need for these drink supplements if individuals have an adequate diet. Table 4.4 provides a list of some ions that are essential for optimum functioning of our bodies.

FIGURE 4.23 As we sweat, we also lose electrolytes (water soluble ionic compounds); these may be replaced by sports drinks.

Hydrated ionic compounds

Hydrated ionic compounds contain water molecules bonded within the crystal.

A number of ionic compounds, called **hydrates**, release water, which is part of their structure, when they decompose upon heating. When the formula of a hydrated compound is written,

the number of water molecules is also included. For example, the formula for iron(II) sulfate octahydrate is written as $FeSO_4 \cdot 8H_2O$. This formula indicates that eight molecules of water are bonded within the ionic crystal for every one formula unit of $FeSO_4$. This is also commonly called water of crystallisation.

TABLE 4.4 Common ions needed in the body

Name	Symbol	Functions in the body
Sodium	Na^+	Regulates fluid balance Involved in transmission of nerve impulses Controls blood pressure by controlling blood volume Involved in acid–base balance of blood
Potassium	K^+	Controls the level of body fluids Involved in transmission of nerve impulses and muscle contraction, including heartbeat Important in reactions within cells
Calcium	Ca^{2+}	Needed for building teeth and bones Involved in blood clotting Involved in transmission of nerve impulses and muscle contraction
Magnesium	Mg^{2+}	Needed for bone formation Assists in energy production Involved in functioning of muscle and nerve tissue Component of enzymes
Chloride	Cl^-	Involved in acid–base balance of blood Assists in fluid balance in the body Involved in formation of hydrochloric acid in stomach

The following prefixes are used to indicate the number of water molecules in a hydrated compound.

mono- = 1
di- = 2
tri- = 3
tetra- = 4
penta- = 5
hexa- = 6
hepta- = 7
octa- = 8
nona- = 9
deca- = 10

SAMPLE PROBLEM 4.6

a. Name the compound $ZnCl_2 \cdot 4H_2O$.
b. Give the formula for the compound calcium sulfate dihydrate.

Teacher-led video: SP6 (eles-3287)

THINK	WRITE
a Name the binary ionic compound in the usual way, and because four water molecules are in the compound, use the prefix *tetra-* followed by the word 'hydrate'. Because zinc only has one possible charge (2+), it is often seen without the (II).	Zinc(II) chloride tetrahydrate

b. 1.	Identify the cation and anion and write their symbols.	Cation: Ca^{2+} Anion: SO_4^{2-}
2.	Balance the charges so the net charge is zero, keeping the polyatomic sulfate ion intact. Since the ions are equally charged, the formula for calcium sulfate is $CaSO_4$.	$CaSO_4$
3.	determine the number of water molecules using the prefix: *di-* means 'two',	$CaSO_4 \cdot 2H_2O$

PRACTICE PROBLEM 6

a. Name the compound $CoBr_2 \cdot 6H_2O$.

b. Give the formula for the compound tin(II) chloride dihydrate.

4.4 EXERCISE

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1. The following table shows the uses and formulas of some hydrated ionic compounds. Complete either the formula for the compound or its name. The first row is completed for you.

Name of hydrated ionic compound	Common name(s); use; description	Formula
Copper(II) sulfate pentahydrate	Blue vitriol; copper plating; blue solid	$CuSO_4 \cdot 5H_2O$
Sodium carbonate decahydrate	Washing soda, soda ash; water softener; white solid	(a)
(b)	Epsom salts; explosives, matches; white solid	$MgSO_4 \cdot 7H_2O$
Magnesium chloride hexahydrate	Disinfectants, parchment paper; white solid	(c)
(d)	Dyeing fabrics, tanning leather; white solid	$BaCl_2 \cdot 2H_2O$

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4.5 Review

4.5.1 Summary

Structure and properties of ionic substances

- Ions are formed when atoms lose or gain electrons to achieve the stable electron configuration of a noble gas.
  - Metals lose electrons to gain a full outer shell and become positively charged ions called cations.
  - Non-metals gain electrons to obtain a full outer shell and become negatively charged ions called anions.
  - Whether an atom gains or loses electrons depends on its electronegativity, or electron-attracting power.

- In the periodic table, the group numbers 1, 2, 13, 15, 16 and 17 have the charge 1+, 2+, 3+, 3–, 2– and 1– respectively. Group 14 and group 18 elements do not readily form ions.
- Three types of chemical bond are possible:
  - ionic bond
  - metallic bond
  - covalent bond.
- Ionic bonding is produced by the strong electrostatic attraction that results between ions when a metallic atom transfers electrons to a non-metallic atom. The ionic compound formed:
  - is known as a salt
  - is made up of an ionic network lattice of positive cations and negative anions
  - is neutrally charged; that is, the total positive charge of the cations must equal the total negative charge of the anions
  - can be represented by a formula unit, which shows the smallest whole number ratio of cation to anion, when charges are balanced. This is also called the empirical formula.
- The structure and bonding of ionic compounds has a direct influence on their properties. Ionic network lattices are rigid structures of cations bonded to anions. Ionic compounds:
  - do *not* conduct electricity in the solid state but become electrolytes when in aqueous or molten states, when their rigid structures have broken down sufficiently to allow free movement of their ions
  - are brittle because pressure causes like charges to align, resulting in the structure shattering
  - are often soluble in water, the molecules of which disrupt the lattice and allow free mobile ions to form
  - are usually crystalline solids due to the close-packed, three-dimensional lattice structure
  - have high melting and boiling points due to the strong force of attraction between the ions.

Ionic nomenclature

- The key rules in naming ionic compounds and writing their formulas are as follows:
  - When naming a binary ionic compound, always name the metal ion in full first and then add the non-metal ion with the ending *-ide* (e.g. lithium oxide).
  - To find the correct formula for an ionic compound, determine the lowest whole number ratio of ions that gives a net charge of zero.
  - When naming ionic compounds that contain metal ions with more than one charge, such as iron(II) and (III), copper(I) and (II), lead(II) and (IV), mercury(I) and (II), and tin(II) and (IV), use roman numerals to indicate which ion is present (CuS is copper(II) sulfide). The Roman numerals indicate the charge of the ion rather than the number of ions present in the compound.
  - When more than one polyatomic ion appears in the formula of an ionic compound, the usual rules of naming apply, but use brackets to separate the ions in the formula (e.g. ammonium carbonate is $(NH_4)_2CO_3$).
  - For hydrates, use the prefixes *mono-, di-, tri-, tetra-, penta-, hexa-, hepta-, octa-, nona-* and *deca-* to indicate the number of water molecules in the compound (e.g. $CuSO_4 \cdot 5H_2O$ is copper(II) sulfate pentahydrate).

Crystal formation and uses of ionic compounds

- The size of a crystal depends on the conditions during its formation.

4.5.2 Key terms

anion a negatively charged atom or group of atoms
binary describes compounds made up of only two elements
cation a positively charged atom or group of atoms
chemical bond arrangement of electrons between two atoms that generates a force, causing the atoms to be bound to each other
covalent bond sharing of electrons between nuclei that bonds them together in a molecule
electrolyte a liquid that can conduct electricity
electronegativity the electron-attracting power of an atom, determined by a number of factors including the size of the atom, the charge on the nucleus and the number of electrons in the atom; atoms with high electronegativities are able to attract electrons easily, whereas atoms with low electronegativities do not attract electrons as readily
electrostatic attraction strong force of attraction between positive and negative ions that produces an ionic bond
hydrates adds water
ion a charged atom
ionic bond bond between cations and anions in an ionic compound
ionic compound compound containing cations and anions in an ionic lattice structure or an ionic liquid, depending on ion size
metallic bond positively charged metal cations arranged in a lattice with delocalised valence electrons being able to flow around them
polyatomic ion charged ion composed of two or more atoms
salt ionic compound consisting of a metal ion and a non-metal ion, except oxides and hydroxides

Resources

Digital document Key terms glossary — Topic 4 (doc-30933)

4.5.3 Practical work and investigations

Experiment 4.1

Investigating calcite crystals
Aim: To observe the nature of crystal structure using calcite crystals
Digital document: doc-30867
Teacher-led video: eles-3270

Experiment 4.2

Ionic models
Aim: To construct a model of sodium chloride and to use it to simulate some properties of sodium chloride
Digital document: doc-30868
Teacher-led video: eles-3271

4.5 Exercises

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4.5 Exercise 1: Multiple choice questions

1. What happens when an atom loses an electron?
   A. A positive ion is formed
   B. A negative ion is formed
   C. The atomic number changes
   D. The atomic mass changes

2. Sodium atoms and sodium ions
   A. are identical in their chemical properties.
   B. have the same number of electrons.
   C. have the same number of protons.
   D. have the same number of protons and electrons.
3. The elements sodium, caesium and lithium form ions with the charge in their compounds of
   A. 1+ only.
   B. 2+ only.
   C. 3+ only.
   D. 1+ or 2+ only.
4. Which of the following ions does not have the same electron configuration as an oxide ion, O^{2-}?
   A. N^{3-}
   B. F^{-}
   C. S^{2-}
   D. Al^{3+}
5. An unknown substance was found to combine with chlorine to form a compound of formula XCl_3. How many outer shell electrons does X have?
   A. 1
   B. 2
   C. 3
   D. 4
6. An atom of X, which has two electrons in its outermost shell, is in contact with an atom of Y, which has seven valence electrons. Which of the following is *incorrect*?
   A. X^{2+} and Y^{-} ions are present in the compound.
   B. The compound has the formula XY_2.
   C. The compound probably has high melting and boiling points.
   D. No compound can be formed between these two elements.
7. An element has an atomic number of 20. What is its electron configuration when the element reacts to form an ionic compound?
   A. 2, 8, 8
   B. 2, 8, 8, 2
   C. 2, 8, 8, 4
   D. 2, 8, 10
8. The electron configurations of four pairs of elements are given below. Which pair of elements is most likely to form an ionic bond?
   A. 2, 8, 1 and 2, 8, 2
   B. 2, 8, 1 and 2, 8, 7
   C. 2, 8, 2 and 2, 8, 4
   D. 2, 8, 7 and 2, 8, 6
9. What does the formation of an ionic compound from a reaction between atoms of two elements involve?
   A. Sharing of pairs of electrons between atoms
   B. Donation of outer shell electrons to the entire crystal structure
   C. Transfer of electrons between atoms
   D. Ionisation of the atoms of some elements
10. What is the structure of solid calcium chloride best described as?
   A. A lattice consisting of diatomic chlorine molecules strongly bonded to calcium atoms.
   B. A lattice of calcium and chloride ions, strongly bonded, in the ratio 1:2.
   C. Discrete molecules of calcium chloride with strong bonding within the molecule but weak bonding between molecules.
   D. An infinite lattice in which calcium and chloride are linked by strong covalent bonds.

11. What is the formula of the compound ammonium phosphate?
    A. $(NH_4)_3PO_4$
    B. NH_3P
    C. NH_4PO_4
    D. $(NH_4)_2PO_3$
12. What is the compound CuO called?
    A. Copper oxide
    B. Copper dioxide
    C. Copper(II) oxide
    D. Copper(I) oxide.
13. What are the blue crystals of $CuSO_4 \cdot 5H_2O$ called?
    A. Copper sulfate pentahydrate
    B. Copper sulfate hexahydrate
    C. Copper(II) sulfate pentahydrate
    D. Copper(I) sulfate hexahydrate

4.5 Exercise 2: Short answer questions

1. Complete the following table.

Name of atom	Symbol for atom	Electron configuration of atom	Symbol for ion	Electron configuration of ion
Lithium				
Beryllium				
Nitrogen				
Oxygen				
Fluorine				
Sodium				
Magnesium				
Aluminium				
Phosphorus				
Sulfur				
Chlorine				
Potassium				
Calcium				

2. a. What is the electron configuration of the calcium atom?
   b. What is the electron configuration of the fluorine atom?
   c. Describe what happens when the calcium and fluorine atoms react chemically.
   d. Use an electron shell diagram to illustrate your answer to part (c).
   e. What holds the ions together in a crystal?
   f. What is the ratio of calcium ions to fluoride ions?
3. Cations form from metallic elements. Using an example, show how this occurs.
4. In a sample of potassium chloride (an important electrolyte in the human body), the number of K^+ ions and Cl^- ions are equal. In a sample of potassium oxide, however, the numbers of each ion are not equal. State the ratio of each ion in the potassium oxide and explain why they are not equal.

5. Write simple equations and electron shell diagrams to show what happens when the following pairs of atoms bond.
   a. K and F
   b. Al and O
   c. Be and Cl
6. Explain why ionic salts conduct electricity in the molten and aqueous states but not in the solid state.
7. Draw diagrams to illustrate the brittle nature of salt.
8. Explain why ionic substances have high melting points and are usually crystalline solids at room temperature.
9. A compound made up of *X* and *Y* atoms has the following properties.
   - It has a high melting point and boiling point.
   - It is very soluble in water but not in kerosene.
   - It conducts electricity in the liquid or aqueous state but not in the solid state.
   - The electron configuration of an atom of *X* is 2, 8, 2, while that of *Y* is 2, 6.

   a. What is the formula for the compound?
   b. What type of bonding must it have?
   c. Write a simple equation showing how the bonding is attained.
10. Using table 4.3 (in section 4.3.2), write the formulas for the following compounds.
    a. Sodium fluoride
    b. Barium nitrate
    c. Iron(III) hydroxide
    d. Sodium sulfide
    e. Aluminium oxide
    f. Calcium hydride
    g. Copper(II) sulfate
    h. Ammonium hydroxide
    i. Chromium(III) oxide
    j. Calcium nitrate
    k. Lithium chloride
    l. Potassium cyanide
    m. Sodium hydrogen phosphate
11. Give the names of the following compounds.
    a. $ZnCl_2$
    b. $Al_2(CO_3)_3$
    c. Na_2SO_4
    d. $AgNO_3$
    e. NaOH
12. Give the chemical name and formula for each of the following compounds.
    a. Soda ash, the common name of a compound containing sodium ions and carbonate ions
    b. Baking soda, commonly used in baking cakes and composed of sodium and hydrogen carbonate ions
    c. Chalk, marble and limestone, all composed of calcium ions and carbonate ions
13. Write the formulas for the following compounds.
    a. Magnesium sulfate heptahydrate
    b. Sodium carbonate decahydrate
    c. Zinc chloride hexahydrate
    d. Barium chloride dihydrate
14. Name the following compounds.
    a. $BaCl_2 \cdot 3H_2O$
    b. $LiCl \cdot 4H_2O$
    c. $CoCl_2 \cdot 5H_2O$

15. Sodium chloride is used as a typical example of an ionic compound. Research the effects of too little or too much sodium chloride on the body.

4.5 Exercise 3: Exam practice questions

1. Predict the charges on the ions formed when the following atoms react with other atoms. **(3 marks)**
   a. Ca **1 mark**
   b. P **1 mark**
   c. Al **1 mark**
2. Oxygen gas, which forms about 20% of the Earth's atmosphere and is essential to life, exists as O_2 molecules. Explain why ionic bonding does not exist between two atoms of oxygen. **(2 marks)**
3. Explain what happens to the size of the atom in the following situations. **(4 marks)**
   a. It loses one or more electrons in forming an ionic bond. **2 marks**
   b. It gains one or more electrons in forming an ionic bond. **2 marks**

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4.5 Exercise 4: studyON Topic Test online only

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