Section 4.1 Representing Ionic Compounds

(Student textbook pages 139 and 151)

Specific Expectations

- **C3.1** describe the relationships between chemical formulae, composition, and names of binary compounds
- **C3.8** identify simple ionic compounds, simple compounds involving polyatomic ions, molecular compounds and acids, using the periodic table and a list of the most common polyatomic ions and write the formulae

In this section, students extend knowledge of the atomic structure of elements using Bohr-Rutherford diagrams and practise determining the stable ion charge of some elements. They explore the link between chemical stability and the number of valence electrons in an atom, and recognize that atoms either lose, gain, or share electrons to fill a valence shell. Ionic bonds are explored by creating diagrams of the electron movement involved. Multivalent metals and polyatomic ions are introduced. Students practise writing chemical formulas and names for binary and polyatomic compounds.

Common Misconceptions

- The relationship between lost/gained electrons and charge is often inverted. Reinforce that electrons are negatively charged. Using integer chips to represent the balance of protons (+) and electrons (−), perform sample calculations to reactivate student understanding of the addition and subtraction of negative integers. Losing a negative (electron) results in a positive charge (cation), as shown in Table 4.3 on page 144 of the student textbook.
- Ionic compounds are not made of molecules. A careful choice of analogies and the use of models can prevent this misconception. An ion with a charge of 2+ can attract two ions with a charge of 1–, but it is a mistake to imagine that the 2+ ion holds hands with both of the 1+ ions. While this is a good analogy for how a water molecule forms, it is not what happens in an ionic compound. Instead, a crystal lattice forms with a ratio of two 1– ions for every 2+ ion.
- Roman numerals should not be used to name ionic compounds whose metal has only one possible ionic charge (e.g., not sodium(I) chloride). Emphasize that Roman numerals are only to be used with transition metals (multivalent metals) to indicate which valence charge is involved in the ionic bond (e.g., iron(III) chloride).
- The *cross-over method* can be mistaken as the final step in writing the chemical formula for ionic compounds. Remind students to reduce or simplify the formula to determine the ratio in lowest terms (e.g., Mg_2O_2 must be simplified to MgO).
- Polyatomic ions are sometimes confused with regular ions (e.g., both NaS and Na_2SO_4 are mistakenly called sodium sulphide). Explain that the *–ide* ending is only used for binary compounds (those made of only two elements).
- Table salt is often considered to be the only salt. Explain that in chemistry, "salt" is a category encompassing all compounds created by an ionic bond between any metal and non-metal (e.g., sodium chloride, barium fluoride, and calcium chloride).
- "Stable ion" is not synonymous with harmless. Share examples of toxic ionic compounds, such as sodium chloride, lead(II) nitrate, arsenic(V) chloride, and mercury(II) oxide.
- If students realize that completing a valence shell makes an element's valence the same as one of the noble gases, they may err that the element "becomes" that noble gas. Remind them that an element is defined by the number of protons in the nucleus, and this does not change.
- Chlorine and chloride are often confused as being identical. Emphasize that chlorine is a stable, neutral atom (with equal protons and electrons and therefore no net charge) whereas chloride is a negative ion (more electrons than protons) formed by gaining an electron when it bonds with another atom.

- You may wish to point out that it is not necessarily a problem that compounds have more than one name. However, it would be a problem if one name referred to more than one compound. For example, sodium chloride is often referred to as table salt or even sea salt. At the dinner table, common names are more appropriate than chemical nomenclature.

Background Knowledge

The chemical reactivity (or stability) of an element is determined by its number of valence electrons (electrons in its outer energy shell or orbit); or, more specifically, the number of additional electrons that the valence shell *could* hold. Atoms will gain/loose/share electrons until the outer shell is full. Those atoms whose valence shell is already full of electrons (i.e., the noble gases of group 18) are chemically stable (inert). The atoms of all other elements can only achieve this stability by either losing electrons (metals do this), gaining electrons (non-metals do this), or sharing electrons (when two or more non-metals are involved).

Within the transition metals (elements in groups 3 to 12), we find atoms with more than one ionic charge. These can form compounds in different ratios. The charge on these *multivalent* metals (i.e., several possible valence shell configurations) is indicated in a name by a Roman numeral in round brackets immediately following the element symbol (e.g., iron(III) and iron(II)). This notation lets us determine the ratio of elements that will form a compound (e.g., $FeCl_2$ and $FeCl_3$ respectively).

Polyatomic ions are interesting because the bonding within the ion is molecular (discussed in the next section). But, those ions are bonded to each other ionically (by attraction of their opposite charges) within the compound, forming a crystal lattice.

Literacy Support

Using the Text

Students will be reading about a variety of ionic compounds in this section. As a class, browse the figures and headings in the section. Predict what information will be given, then create a skeleton organizer in which to take notes. Or, write each Key Term on a separate large sticky note, arranging them as you read, to show links.

Before Reading

- Use a predict–read–verify strategy to help students monitor comprehension of the text. Direct students to use headings to "chunk" the text into manageable sections. Start by reading headings and captions, and analyzing visuals to predict what each chunk is about. Verify or revise these predictions as you read the chunk of text.

During Reading

- Encourage students to take notes as they read. They may start by rewriting topic headings in the form of a question (e.g., How do I name an ionic compound? What is a multivalent metal? Why are naming rules different?). As they read, have them record answers in two-column or three-column note outlines.
- Ask students to note three main concepts for each paragraph. Providing sticky notes for this purpose allows students to place the notes next to the paragraph in the textbook, offering a quick review mechanism. Have students compare notes with a classmate's or the teacher's to ensure they have recognized the main ideas.

After Reading

- Reflecting on this completely new "language" that students are learning will help them recall chemical nomenclature. After reading, students should outline the information they have learned and note what they found interesting.

Using the Images

- This section uses patterns in the periodic table to make sense of patterns found in chemical bonding. These include the division of metals and non-metals, and the arrangement of elements into columnar families. Have students evaluate the ionic charge on elements in a family and explain that this common ionic charge results in many shared chemical properties.
- Use images as a predictive tool for reading the text. Begin a section by viewing the images. Ask students to identify objects or actions in each visual. Then, predict what links these might have to the subject heading. Does the caption confirm this? What questions does the visual make students think of? After viewing, read the text on the page. Were predictions confirmed? If not, what is the new understanding of the image?
- Several tables in this section contain step-by-step instructions, paired with examples. When each table is encountered, read the steps aloud and work through the examples as a class. Have students work through another example in groups of two or three. Take up their solutions to identify any misconceptions and offer remediation. students may benefit from transcribing these steps into their notebooks or onto flash cards.
- Some illustrations throughout the chapter provide valuable comparisons and may add more clarity when viewed in conjunction rather than in isolation.
- Figure 4.3 provides a periodic table with ion charge notations that students will need to understand the chapter content and complete many of the exercises. Tour the features of this table as a class. Encourage students to flag this page with a sticky note for easy reference.
- Encourage students to recognize that the chart in Figure 4.2 mimics the periodic table, using the familiar organization to present new information. Link each cell to a typical periodic table (Figure 4.3), drawing attention to the element symbols at the centre of each electron dot diagram. Examine the number of electron shells shown for each atom. Ask questions to draw out observations, such as: How many energy levels are shown for the atoms in the first row? How many in the second row? The third? Look at all the diagrams, how many electrons are in the first energy level, the second, the third? Then, examine the groups, evaluating the number of valence electrons as you read the second paragraph aloud.
- When viewing the crystal lattice in Figure 4.6, share models that students can touch, to reinforce that this is a 3-D structure, not a hexagon. This will be contrasted against the structure of molecules in the next section.
- When viewing tables in this section, encourage students to identify links between the information in columns. As a class, evaluate each row of Table 4.3, adding up the values in the two right columns to illustrate that the sum is always 18. Note that the number of electrons the element *must gain* to have a full valence shell is *the same* as its ionic charge (noted by the superscript in the Symbol column). Have students verify that these ionic charges are shown in the periodic table on page 141. Ask how many electrons would be gained by the elements in group 18. Why? (Answer: None, because the valence shell is already full.) Ask students to predict how this pattern might extend to the rest of the periodic table. What can we know if a positive ion charge is noted? (Answer: The element will *gain* electrons to complete its valence shell.)
- Tables 4.4, 4.5, and 4.7 walk through the steps of writing a formula. Read through the steps as a class, then work through another example as a class. You may wish to use the following BLMs to reinforce these steps and provide extra practice: **BLM 4-5 Binary Ionic Compounds, BLM 4-8 Multivalent Ionic Compounds,** and **BLM 4-9 Polyatomic Ionic Compounds.**

- When viewing Figure 4.8, have students find copper on the periodic table. What ionic charges are listed? (2+ and 1+) Read the labels in the visual aloud, left to right as "copper two oxide and copper one oxide." Have students note the dramatic affect (colour difference) a change in ionic charge has on the compound that is formed.

Assessment FOR Learning		
Tool	**Evidence of Student Understanding**	**Supporting Learners**
Activity 4–2 Take My Electron–Please!, page 143	Students show the correct number of electrons and distribute them correctly among the energy levels; draw arrows that start at the metal atom and end at the non-metal atom; write a formula or ratio of elements correctly and in lowest terms; and use Key Words and phrases correctly during discussion.	Provide **BLM 4-7 Take My Electron–Please!**, which scaffolds this process. Introduce Lewis-dot diagrams as an alternative to the Bohr-Rutherford diagrams. These are visually simpler as they only show valence electrons.
Learning Check question 5, page 146	Students show aluminum's valence electrons moving to valence shell of chlorine atoms.	Refer students to Figure 4.5 on page 143 of the student textbook. Have them trace the arrows with their fingers, and count the number of electrons that remain with aluminum, at the second energy level.
Learning Check questions 7-9, pages 146, 150	Formulas contain the correct ratio of elements. Ratios are simplified to lowest terms.	Encourage students to use Bohr-Rutherford diagrams to model electron movement. You may wish to have them complete **BLM 4-7 Take My Electron–Please!,** which scaffolds this process. Refer students to the cross-over method described on page 145 of the student textbook. Have students work through the steps outlined in Tables 4.4, 4.5, and 4.7 in the student textbook. Students can use **BLM 4-5 Binary Ionic Compounds, BLM 4-8 Multivalent Ionic Compounds,** and **BLM 4-9 Polyatomic Ionic Compounds.** Create a game in which players use element cards to combine ions in the correct (neutral) ratio. You could structure it as a relay, where questions are handed out sequentially and the first person/group to finish wins.
Learning Check questions 3, 10, pages 142, 150	Names start with the cation (positive ion). Polyatomic ion names end in *–ate* or *–ite*. All others end in *–ide*. Ionic charge is indicated by Roman numerals where necessary.	Refer to Table 4.6 in the student textbook for a list of polyatomic ion names. Encourage students to use the periodic table to identify multivalent elements. Play a bingo-style game, individually or in groups, where students' game cards contain formulas that must be matched to the name called.

Instructional Strategies

- Introduce the section by creating a mind map or concept map that shows connections between metals losing electrons to become cations (positively charged ions) and non-metals gaining electrons to become anions (negatively charged ions). Link this to the naming of binary and polyatomic compounds. You may wish to use **BLM G-42 Concept Map** for this activity.
- If students realize that completing a valence shell makes an element's valence the same as one of the noble gases, you may wish to extend their understanding by having them pair elements with a noble gas (matching stable valence shells). Encourage them to recognize that sodium is more like neon than its row-mate argon, because sodium looses an electron, making its valence shell one energy level lower.

- To illustrate the formation of ionic bonds, use Bohr-Rutherford diagrams on the board to simulate the electron transfer process. Small circles of colourful construction paper with magnets representing electrons work well. Begin with a simple example such as the formation of sodium chloride from sodium and chlorine, which lose and gain one electron respectively. For the second example, illustrate the formation of magnesium oxide from magnesium and oxygen, which lose and gain two electrons respectively. Then, explain and illustrate the formation of calcium fluoride from calcium, which looses two electrons, and fluorine, which can gain only one electron, thus requiring a second atom of fluorine to balance the charge.
- To illustrate the difference ionic charge can make, provide samples of copper(I) chloride and copper(II) chloride. Discuss the difference in physical properties, such as colour, melting and boiling point, and density.
- When using the cross-over method, make sure that students understand why they are performing this trick. Teach this method only after a formative assessment shows students understand the need to balance positive and negative charges within a compound.
- Encourage students to use cue cards to keep the various rules of nomenclature organized, creating flash cards for naming practice.
- Coach students through a few examples of replacing the ending on a chemical with the suffix –*ide*. For example, slash the word chlorine after the first syllable (e.g., chlor-ine). Then, replace the end of the word with –*ide* (e.g., chlor-ide). Repeat for other examples, such as sulf-ide and fluor-ide.
- Reactivate prior learning about the structure of atoms by having students add definitions to the word wall for *element*, *atom*, *periodic table*, *metal*, *non-metal*, *nucleus*, *proton*, *neutron*, *electron*, *energy level/shell*, and *valence shell*. Include examples and illustrations.
- Reactivate prior understanding of patterns in the periodic table by locating metals, non-metals, and family/group number (Figure 4.2 on page 141 of the student textbook). Survey the class for the meaning of each piece of information contained in a cell. Ensure there is understanding that atomic number indicates the number of (positive) protons in the nucleus of each atom of that element, and that each atom contains a balanced (equal) number of electrons. Encourage students to add a summary to their notes, and post an example to the word wall.
- To engage interest, you may now wish to carry out Inquiry Investigation 4-A Monitoring Paper Recycling. See page TR-2-28 of this Teacher's Resource for notes on using this investigation.
- Have students complete Activity 4-2 Take My Electron-Please! See page TR-2-13 of this Teacher's Resource for activity notes.
- Focus on the "why" before the "how" of compound formation. Ensure students grasp of the reasons atoms lose/gain/share electrons (to stabilize by completing the valence shell) before they consider the method.
- Discuss the properties of the noble gases (group 18), drawing attention to their full valence shell. Ask students why this would make them chemically stable (Answer: They do not react readily with other elements). All atoms bond for the same reason: to become more stable, like a noble gas.
- Illustrate on the board, the cations and anions that form a sample compound and show the addition process to emphasize that ionic compounds always have a neutral charge. $1(Mg^{2+}) + 2(Cl^{1-}) = MgCl_2^0$

- Provide **BLM 4-6 Ionic Bonding** to give students practice drawing Bohr-Rutherford diagrams to explore the concepts.
- Use the classroom periodic table to locate elements when naming ionic compounds. Draw attention to the fact that ionic compounds always consist of a metal and a non-metal.
- Have students assess the ionic charges shown on the periodic table. Are any negative? (Answer: No)
- Ask why scientists use both names and formulas. Note the different levels of information contained in names vs. formulas. Names indicate the ions present but neither the ratio nor the charges on the ions (except in the case of a multivalent metal such as iron(III) in which the Roman numeral indicates a charge of 3+). Lead students toward an understanding that formulas provide more information, but are perhaps clumsy to use in speech.
- Repetition is often key to mastering chemical nomenclature. Begin each class with a quick formative assessment quiz. Have students develop names for given ionic compounds, following the steps in Table 4.4, 4.5, or 4.7 as appropriate.
- You may wish to carry out Inquiry Investigation 4-B Keep That Toothy Grin to help students relate ionic compounds (specifically, sodium fluoride) to daily life. See page TR-2-30 of this Teacher's Resource for notes on using this investigation. As a pre-assignment, have students list the ingredients in their toothpaste, and identify the ionic compounds, and as many formulas as they can.
- Post new words and definitions on the classroom word wall.
- As students work through Learning Check questions, point out that the name of an ionic compound always starts with the element that forms a cation (positive charge).

Activity 4-2 Take My Electron–Please! (Student textbook page 143)

Pedagogical Purpose

Students physically manipulate electron models to illustrate loss and gain of electrons. They will understand that upon transfer, each atom has a complete octet of electrons in the outer shell. After forming a compound, they can calculate electron loss and gain to conclude that the net charge is zero.

Planning	
Materials	Paper Small circular objects to mimic electrons **BLM G-1 Safety Contract** (optional) **BLM 4-7 Take My Electron–Please!** (optional) One or two days before, prepare small bags with at least 31 objects each.
Time	Approximately 30 min in class
Safety	Remind students to never eat anything in the science classroom.

Background

In the series of concentric circles that comprise a Bohr-Rutherford diagram, the first electron shell or energy level (circle) may hold two electrons, the second and third each hold eight, and the fourth holds 18. Subsequent energy levels hold more electrons (32 in the fifth and 64 in the sixth). Beyond element 20, the group number also indicates the number of valence electrons. The number of electrons present in the atoms of an element is the same as the element's atomic number (e.g., hydrogen has 1 electron and phosphorus has 15).

Activity Notes and Troubleshooting

- Before allowing students to conduct the activity, ensure **BLM G-1 Safety Contract** has been signed, and remind students never to eat in the lab.
- For manipulatives, you might choose from items such as washers, coloured hole reinforcements, coloured cereal-Os, old CDs, pompoms, buttons, beads, integer chips, or marshmallows. Providing inedible items will discourage eating in the classroom.
- Refer to Figure 4.5 (page 143) as a model for diagrams, noting that when aluminum loses electrons to the chlorine atoms, aluminum then has a complete valence shell at the second energy level.
- Provide **BLM 4-7 Take My Electron—Please!**, which scaffolds this activity. After handing out copies, identify the meaning of the concentric circles (energy levels), dots (electrons), and the number of electrons that each energy level can hold. Identify which circle represents the valence shell (the last one with electrons in it). Have students add a labelled diagram example (or each one) to the word wall.
- Work through one combination as a class. Draw Bohr-Rutherford diagrams for each element on overheads and project the scene as you use the manipulatives to transfer one valence electron.
- Depending on the number of manipulatives available, students could work in groups. However, each student should produce an individual copy of the diagrams.
- Wrap up by creating a class set of ionic bond diagrams from students' work.

Additional Support

- DI This is a good hands-on activity for bodily-kinesthetic and spatial learners.
- DI Encourage students with mathematical intelligence to calculate differences between the number of protons and electrons in each atom. Such learners may benefit from using integer chips to illustrate the full complement of protons and electrons in each atom.
- ELL Encourage English language learners to describe the process of ionic bond formation.
- Students with motor challenges may use larger manipulatives such as pompoms or wooden beads. Select materials that respect their age and do not appear childish.

Answers

Steps 4 and 5:

Elements	Ratio of Ions	Formula
Sodium and chlorine	1:1	$NaCl$
Magnesium and fluorine	1:2	MgF_2
Lithium and nitrogen	3:1	Li_3N
Aluminum and sulfur	2:3	Al_2S_3

1. Example: An equal number of electrons is gained and lost because electrons cannot be created or destroyed.
2. Example: Each element (ion) has an unequal number of electrons and protons after it gains or loses an electron, but the ions bond to form a compound with an equal balance of electrons and protons.

Learning Check Answers (Student textbook page 142)

1. The number of valence electrons determines the number of bonds that can/will be made.
2. *–ide*

3. a. magnesium bromide
   b. calcium iodide
   c. aluminum oxide
   d. potassium chloride
4. Example: Naming rules mean a lot of information is reliably contained in the name. Peoples' surnames, for example, link them to their ancestors and often provide clues about their cultural heritage.

Learning Check Answers (Student textbook page 146)

5. Aluminum's valence electrons are transferred to three chlorine atoms to complete their valence shells.
6. Diagram should show that each of two chlorine atoms gains one electron from a magnesium atom to complete their valence shells.
7. a. $Na_2O = 2(Na^{1+}) + O_2- = 0$ b. $Li_3N = 3(Li^{1+}) + N^{3-} = 0$
   c. $AlI_3 = Al^{3+} + 3(I^{1-}) = 0$ d. $Ba_3P_2 = 3(Ba^{2+}) + 2(P^{3-}) = 0$
8. a. K_2S b. Li_2Se c. ZnO
   d. RbBr e. C_2S f. S_3N

Learning Check Answers (Student textbook page 150)

9. a. binary, Ni_2O_3 b. binary, CuI_2
   c. binary, Sn_3N_4 d. binary, $CrBr_2$
   e. binary, FeP f. neither, $LiHCO_3$
   g. ternary, K_2SO_4 h. ternary, $(NH_4)_3P$
   i. ternary, $Ba(NO_3)_2$ j. binary, $Co_3(PO_4)_2$
10. a. gold(III) chloride b. tin phosphide
    c. chromium oxide d. nitrogen sulphide
    e. ammonium sulphide f. calcium fluoride
    g. iron(II) sulfite h. magnesium phosphite

Section 4.1 Review Answers (Student textbook page 151)

Please see also **BLM 4-11 Section 4.1 Review (Alternative Format).**

1. a. chloride b. magnesium
   c. sulfate d. copper(II)
2. a. OH^- b. S^{2-}
   c. Al^{3+} d. Cr^{3+}
3. The diagram should show two potassium atoms each giving one electron to one sulfur atom, which has six electrons in its outer shell.
4. The tiles repeat in a regular pattern like the atoms in an ionic compound. To represent sodium chloride, there would have to be an equal ratio of black and white tiles.
5. a. lithium carbonate b. ammonium nitrite c. copper(II) oxide
6. a. Mg_3N_2 b. $Al(OH)_3$ c. $SnBr_2$ d. $NiSO_4$
7. a. Phosphorus was used instead of the phosphate ion. Na_3PO_4
   b. The brackets around the nitrate ion were omitted. $Ca(NO_3)_2$
   c. The two potassium ions are needed to make the total charge zero. K_2SO_3