PEARSON

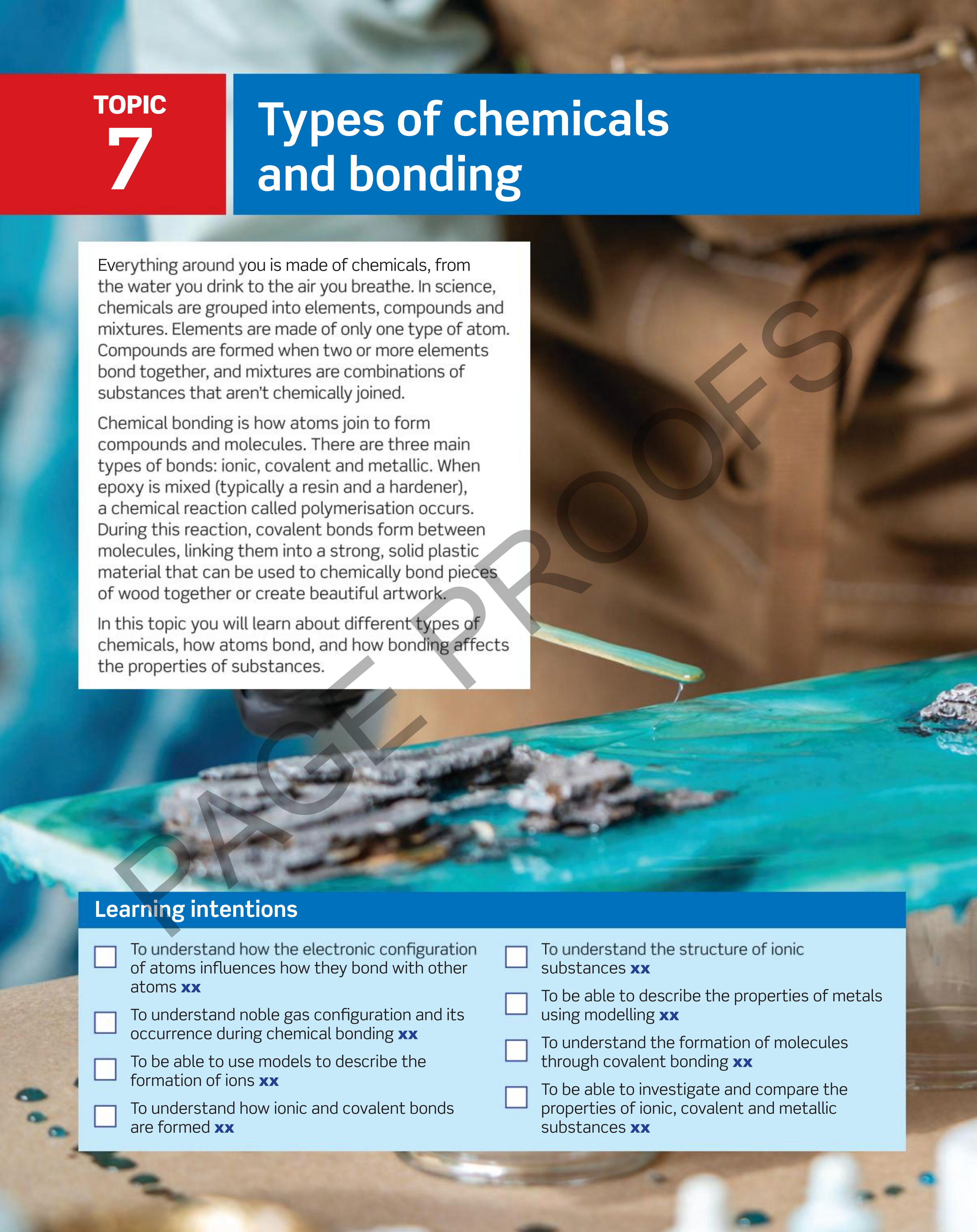
Science

STUDENT BOOK | NEW SOUTH WALES

Stage

5

NSW SYLLABUS



TOPIC 7

Types of chemicals and bonding

Everything around you is made of chemicals, from the water you drink to the air you breathe. In science, chemicals are grouped into elements, compounds and mixtures. Elements are made of only one type of atom. Compounds are formed when two or more elements bond together, and mixtures are combinations of substances that aren't chemically joined.

Chemical bonding is how atoms join to form compounds and molecules. There are three main types of bonds: ionic, covalent and metallic. When epoxy is mixed (typically a resin and a hardener), a chemical reaction called polymerisation occurs. During this reaction, covalent bonds form between molecules, linking them into a strong, solid plastic material that can be used to chemically bond pieces of wood together or create beautiful artwork.

In this topic you will learn about different types of chemicals, how atoms bond, and how bonding affects the properties of substances.

Learning intentions

- ☐ To understand how the electronic configuration of atoms influences how they bond with other atoms **xx**
- ☐ To understand noble gas configuration and its occurrence during chemical bonding **xx**
- ☐ To be able to use models to describe the formation of ions **xx**
- ☐ To understand how ionic and covalent bonds are formed **xx**
- ☐ To understand the structure of ionic substances **xx**
- ☐ To be able to describe the properties of metals using modelling **xx**
- ☐ To understand the formation of molecules through covalent bonding **xx**
- ☐ To be able to investigate and compare the properties of ionic, covalent and metallic substances **xx**

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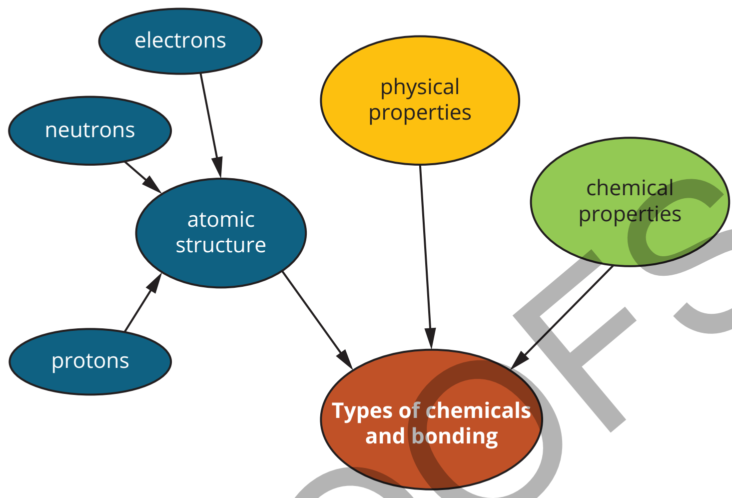
7 Types of chemicals and bonding

The key concepts that you will use in this topic:



The following prior knowledge questions will help to support your learning in the topic and can be attempted before the first lesson.

Atomic structure

1 What are the three main subatomic particles in an atom, and where is each found?

2 What does the atomic number of an element tell you about its atoms?

Chemical properties

3 What is a chemical change, and how can you tell one has happened?

4 Why do elements in the same group of the periodic table often react in similar ways?

Physical properties

5 What is the difference between a solid, liquid and gas in terms of particle arrangement?

6 Name two physical properties you can use to describe a substance, and explain how they can be observed.

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7.1 Electronic structure of atoms and the periodic table

Lesson overview

Atoms are the fundamental building blocks of all matter. Understanding the electronic structure of atoms is crucial for comprehending how elements interact and bond with each other. For example, the way sodium, a soft, silvery metal, reacts with chlorine (found in bleach) to form sodium chloride (table salt) is determined by their electronic structures.

In this lesson you will learn about the electronic structure of atoms, how to determine the valency of different elements using the periodic table, and how atoms achieve stable electron configurations.

Learning intention

To understand how the electronic configuration of atoms determines how they bond with other atoms

Success criteria

SC 1: I can describe the electronic structure of atoms.

SC 2: I can determine the valency of different elements using the periodic table.

SC 3: I can use the valency of an atom to determine how it can achieve a stable electron configuration.

SC 1 I can describe the electronic structure of atoms.

Atoms consist of a nucleus—containing protons and neutrons—surrounded by electrons, as shown in Figure 7.1.1. However, these electrons are not free to orbit at any location around the nucleus; instead they are held in specific energy levels or shells. The electronic structure of an atom describes the arrangement of electrons in these shells.

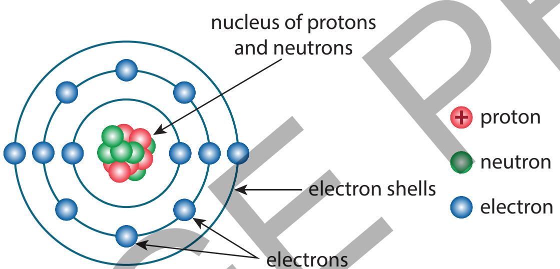

FIGURE 7.1.1 The electronic structure of magnesium (2,8,2)

Electrons occupy the lowest available energy levels first, filling up each level before the next level is occupied. The electron configuration of different elements can be represented using the number of electrons in each shell separated by a comma and placed in brackets. For example the electron configuration for carbon, which has six electrons, can be represented as (2,4). Here, the first two electrons fill the first energy level and the remaining four electrons occupy the second energy level.

The shells are filled in order.

- The first shell can hold two electrons and is filled first.
- The second shell can hold eight electrons.
- The third shell is partially filled. Once it contains eight electrons, the next two electrons will be added to the fourth energy shell.

The electron configuration of (2,8,8,2) is the electron configuration of calcium, which is the twentieth element of the periodic table. At this stage, calcium's electron configuration is the largest you will encounter.

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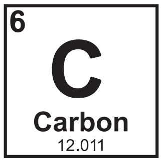

FIGURE 7.1.2 The periodic table entry for carbon

You will notice that the number of electrons in a neutral atom is equal to the number of protons in that atom. The periodic table indicates the number of protons via the atomic number, which is the whole number that is usually found at the top of the box for each atom (Figure 7.1.2).

Consider the electronic configuration of potassium, which has an atomic number of 19, and therefore has nineteen electrons. The configuration is (2,8,8,1). The first two electrons fill the first energy level, the next eight electrons fill the second energy level, the next eight electrons partially fill the third energy level, and the last electron occupies the fourth energy level.

SC 1 CHECK YOUR UNDERSTANDING

Describe the number of electrons in each shell of sulfur.

SC 2 I can determine the valency of different elements using the periodic table.

KEY TERMS

valency the number of electrons in an atom's outermost shell

valence electron an electron in the outer electron shell of an atom

Valency is the combining power of an element, which is determined by the number of electrons in its outermost shell. These are called **valence electrons**.

The number of valence electrons of an element also determines which group (vertical column) they belong to in the periodic table. Elements in the same group of the periodic table have the same valency.

For the first twenty elements (Figure 7.1.3) the easiest way to determine an atom's valency is to count how far away the element is from the noble gas group following the shortest path. For the elements on the left-hand side of the periodic table you can count back to the previous noble gas. For elements on the right, you can count forwards to the next noble gas.

For elements in group 1, such as sodium (Na), you can count back *one* place to the previous noble gas neon (Ne), and so it has a valency of 1. If you only consider the first twenty elements, for group 13 elements, such as Boron (B), you can count back *three* places to the previous noble gas helium (He), and so it has a valency of 3. For elements in group 15, such as phosphorus (P), you can count forward *three* places to argon (Ar), and so it has a valency of 3. Group 14 elements are exactly *four* places either way from the closest noble gas group so they have a valency of four.

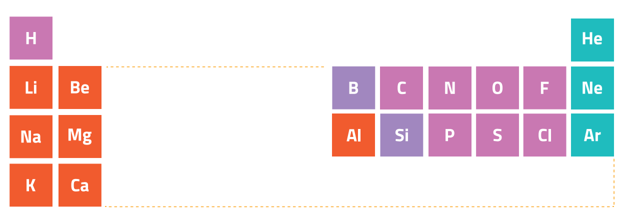

FIGURE 7.1.3 The first twenty elements on the periodic table

SC 2 CHECK YOUR UNDERSTANDING

Determine the valency of magnesium (Mg), oxygen (O) and lithium (Li) using the periodic table.

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SC 3 I can use the valency of an atom to determine how it can achieve a stable electron configuration.

Atoms achieve stability by having a full outermost energy level, similar to the nearest noble gas, which is called **noble gas configuration**. Noble gases, such as helium, neon and argon, have completely filled outer energy levels, making them very stable and unreactive. Atoms achieve stability by gaining, losing or sharing electrons to get a full outer shell, like their closest noble gas. When atoms gain or lose electrons, they form **ions** that can form **ionic bonds**. When atoms share electrons, they form **covalent bonds**.

KEY TERMS

noble gas configuration the stable arrangement of electrons in the valence shell, the outermost energy level of an atom

ion an atom or group of atoms that has gained or lost electrons, giving it a positive or negative charge

ionic bond a strong bond formed between positive and negative ions

covalent bond bond formed by sharing electrons

SCIENCE IN CONTEXT

Noble gases in everyday life

Noble gases have unique properties due to their stable electron configurations, making them useful in various applications. For example, neon is used in neon signs because it emits bright light when an electric current passes through it. The stability of neon's electron configuration allows it to emit light without reacting with other elements.

Helium is another noble gas and is used in balloons and airships because it is lighter than air and non-flammable (Figure 7.1.4). Helium's stable electron configuration makes it safe to use in these applications. A hot air balloon festival is held every year in the NSW central west town of Canowindra.

FIGURE 7.1.4 Many modern hot air balloons now use helium as well as air.

The electron configurations of the first three noble gases show that:

- helium has two electrons in its first energy level, giving it a stable configuration of (2)
- neon has ten electrons, with a stable configuration of (2,8)
- argon has eighteen electrons, with a stable configuration of (2,8,8).

Gaining electrons

Non-metals are found on the right-hand side of the periodic table. Non-metal elements, such as chlorine, tend to gain electrons to achieve the same noble gas configuration as the noble gases. For example, chlorine has seven electrons in its outermost energy level (2,8,7) and needs one more to achieve the stable configuration of argon (2,8,8), as shown in Figure 7.1.5.

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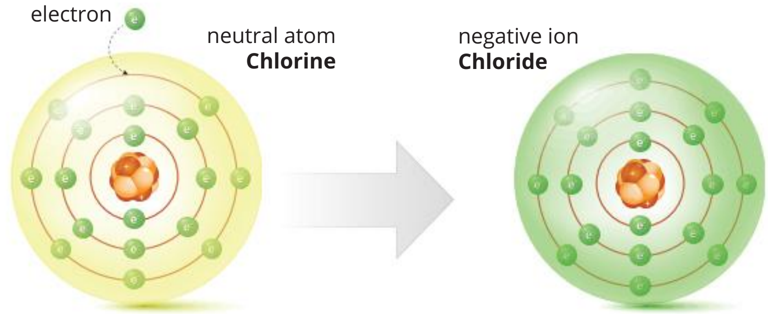

FIGURE 7.1.5 When chlorine gains an electron to form the chloride ion, it now has a full outermost energy level.

Losing electrons

Metals are found on the left-hand side of the periodic table. Metals, such as sodium, tend to lose electrons to achieve the same noble gas configuration as the noble gas in the previous period. As shown in Figure 7.1.6, sodium has one electron in its outermost energy level (2,8,1) and loses it to achieve the stable configuration of neon (2,8).

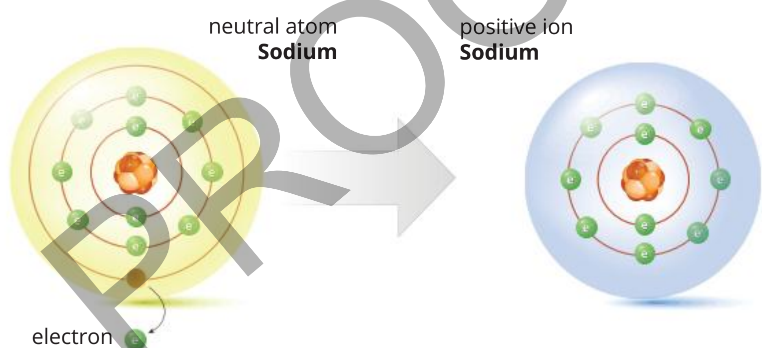

FIGURE 7.1.6 Sodium loses an electron to reach a noble gas configuration

Sharing electrons

Some atoms achieve a noble gas configuration by sharing electrons. For example, two hydrogen atoms with an electron configuration of (1) can share their single electrons resulting in each hydrogen atom achieving a stable configuration similar to helium (2). Atoms that share electrons are said to be covalently bonded.

SC 3 CHECK YOUR UNDERSTANDING

Explain how a magnesium (Mg) atom achieves a noble gas configuration.

Lesson review

Use these questions to check whether you have met the learning intention for this lesson.

1 State the electronic structure of a helium atom.

2 Explain why noble gases are stable and unreactive.

3 Determine the valency of chlorine (Cl) using the periodic table.

4 Oxygen gas consists of two oxygen atoms (2,6) bonded together. Predict how each oxygen atom achieves a noble gas configuration and explain your reasoning.

5 Compare the ways in which sodium (Na) and fluorine (F) achieve noble gas configurations.

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7.2 Ionic and metallic bonding

Lesson overview

In this lesson you will learn about the formation of ionic and metallic bonds. For example, table salt (sodium chloride) is formed through ionic bonding, while the metal in your bicycle frame is held together by metallic bonding. Understanding these bonds helps explain the properties of many materials that you encounter daily.

In this lesson you will describe the formation of ionic bonds, construct chemical formulas for common ionic compounds and describe metallic bonding.

Learning intention

To understand how ionic and metallic bonds are formed

Success criteria

SC 1: I can describe the formation of ionic bonds.

SC 2: I can construct chemical formulas for common ionic compounds.

SC 3: I can describe metallic bonding.

SC 1 I can describe the formation of ionic bonds.

Electrostatic forces

Ionic bonds form when atoms transfer or accept electrons to achieve a stable full outer energy level, resulting in the formation of ions. These ions are then held together by **electrostatic forces**, the attraction between the positive and negative charges.

KEY TERM

electrostatic force the force that attracts or repels objects based on their electric charge

For example, sodium (Na) has one electron in its outermost energy level (2,8,1). To achieve the same stable configuration as its nearest noble gas, sodium loses this electron to form a sodium ion (Na^+) with an electron configuration of (2,8).

Chlorine (Cl), on the other hand, has seven electrons in its outermost energy level (2,8,7). It gains one electron to form a chloride ion (Cl^-) with an electron configuration of (2,8,8). The Na^+ and Cl^- ions are attracted to each other, forming an ionic bond and creating sodium chloride (NaCl), commonly known as table salt. The formation of sodium chloride is shown in Figure 7.2.1.

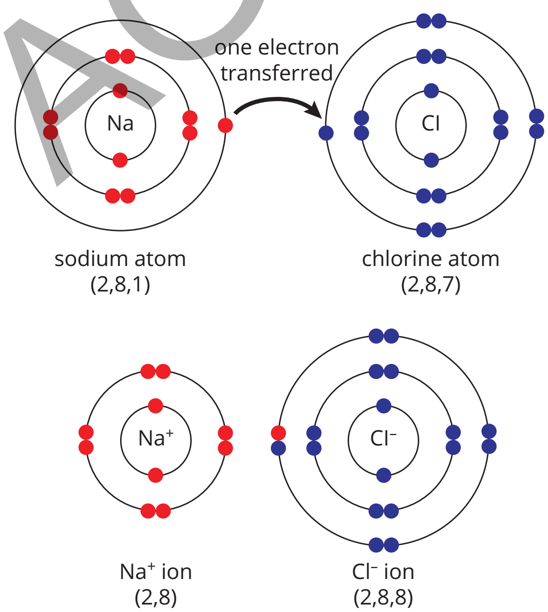

FIGURE 7.2.1 One electron is transferred when sodium chloride forms.

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Scifile

Salty ions

You're probably aware that the ocean contains sodium chloride (salt). In fact, there are another seven ions in the ocean that are considered to contribute to its salinity (saltiness). Together, these nine ions make up over 99% of all dissolved salts in seawater. They are:

- chloride (Cl^-) (the most abundant ion)
- sodium (Na^+)
- sulfate (SO_4^{2-})
- magnesium (Mg^{2+})
- calcium (Ca^{2+})
- potassium (K^+)
- bicarbonate (HCO_3^-)
- bromide (Br^-)
- strontium (Sr^{2+}).

Kati Thanda (Lake Eyre) is Australia's largest inland lake, as well as one of the saltiest lakes on Earth, and its salinity levels are often close to that of oceans.

KEY TERMS

ionic compound substance made of positive and negative ions
cation positively charged ion
anion negatively charged ion

Ionic compounds

Ionic compounds are usually formed between metals and non-metals. Metals tend to lose electrons and form **cations**, while non-metals tend to gain electrons and form **anions**. The resulting electrostatic attraction between the oppositely charged ions forms the ionic bond.

For example, magnesium (Mg) has two electrons in its outermost energy level (2,8,2). It loses these two electrons to form a magnesium ion (Mg^{2+}) with an electron configuration of (2,8). Oxygen (O) has six electrons in its outermost energy level (2,6). It gains two electrons to form an oxide ion (O^{2-}) with an electron configuration of (2,8). The Mg^{2+} and O^{2-} ions are attracted to each other, forming an ionic bond and creating magnesium oxide (Figure 7.2.2). Magnesium oxide (MgO) is often used to relieve heartburn and indigestion.

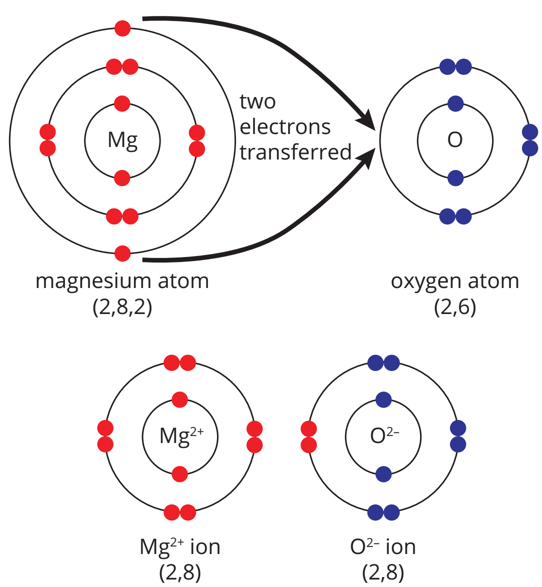

FIGURE 7.2.2 The formation of magnesium oxide

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SC 1 CHECK YOUR UNDERSTANDING

Explain why sodium and chlorine atoms form an ionic bond.

SC 2 I can construct chemical formulas for common ionic compounds.

Chemical formulas for ionic compounds

Constructing chemical formulas for ionic compounds involves combining the symbols of the ions involved and using subscript numbers to balance the charges of the ions to ensure that the compound is neutral overall.

For example, to construct the formula for magnesium oxide, combine the symbols for magnesium (Mg) and oxygen (O). Magnesium forms an Mg^{2+} ion and oxygen forms an O^{2-} ion. Since the charges are equal and opposite, they balance each other, resulting in the formula MgO.

For magnesium chloride, combine the symbols for magnesium (Mg) and chlorine (Cl). Magnesium forms an Mg^{2+} ion and chlorine forms a Cl^- ion. Since the charges are opposite but not equal, they will only balance if two chloride ions combine with each magnesium ion, resulting in the formula $MgCl_2$ (Figure 7.2.3).

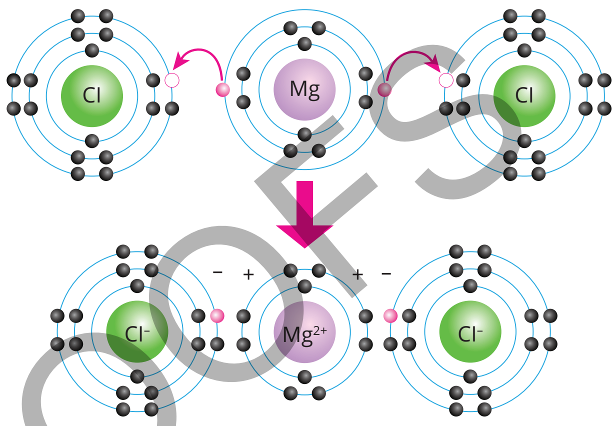

FIGURE 7.2.3 Two chloride ions combine with each ion of magnesium

Compounds with polyatomic ions

When constructing formulas for compounds with **polyatomic ions**, follow the same principle of balancing charges. For example, to construct the formula for calcium nitrate, combine the symbols for calcium (Ca) and nitrate (NO_3). Calcium forms a Ca^{2+} ion and nitrate forms an NO_3^- ion. Since the charges are not equal, two nitrate ions are needed to balance the charge of one calcium ion, resulting in the formula $Ca(NO_3)_2$.

Note that in calcium nitrate there are two nitrate ions needed, and so there is a bracket around the nitrate ion with the subscript number 2 after the bracket. The bracket indicates that there are two whole nitrate ions ionically bonded to one magnesium ion.

Another example is ammonium sulfate. Ammonium (NH_4) forms an NH_4^+ ion, and sulfate (SO_4) forms an SO_4^{2-} ion. Since the charges are not equal, two ammonium ions are needed to balance the charge of one sulfate ion, resulting in the formula $(NH_4)_2SO_4$. Table 7.2.1 lists some common ions and their charges.

KEY TERM

polyatomic ion ion made up of multiple atoms

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TABLE 7.2.1 Common ions and their charges

Charge	Ion symbol	Ion name	Charge	Ion symbol	Ion name
+ 1	H^+ Na^+ K^+ Ag^+ NH_4^+	hydrogen sodium potassium silver ammonium	- 1	F^- Cl^- Br^- I^- NO_3^- OH^- CH_3COO^-	fluoride chloride bromide iodide nitrate hydroxide acetate
+ 2	Mg^{2+} Ca^{2+} Ba^{2+} Zn^{2+} Fe^{2+} Cu^{2+}	magnesium calcium barium zinc iron (II) copper (II)	- 2	O^{2-} S^{2-} CO_3^{2-} SO_4^{2-}	oxide sulfide carbonate sulfate
+3	Al^{3+} Fe^{3+}	aluminium iron (III)	- 3	PO_4^{3-}	phosphate

SC 2 CHECK YOUR UNDERSTANDING

Construct the chemical formula for aluminium chloride.

SC 3 I can describe metallic bonding.

Metallic bonding

Metallic bonding occurs between metal atoms, where the outer electrons are shared in a 'sea' of electrons that are free to move throughout the structure. For example, in a piece of copper (Cu), the copper atoms lose their outer shell electrons to form positive ions that are arranged in a regular lattice pattern. The lost electrons are called **delocalised electrons**.

Delocalised electrons are not held to a single ion. Instead, they are mobile and able to move between metal ions (Figure 7.2.4). The strong electrostatic attraction of the positive metal ions to the negatively charged electrons is what makes metallic bonds so strong.

KEY TERM

delocalised electron an electron that is not bound to a particular metal ion and is free to move

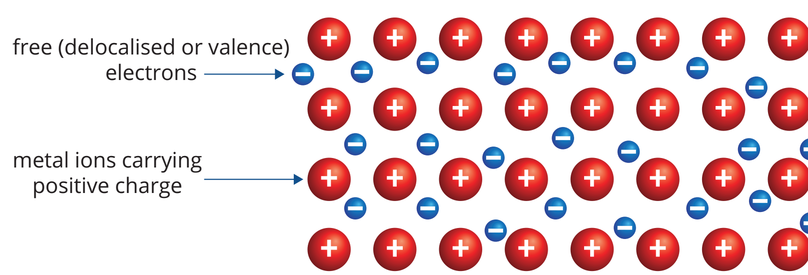

FIGURE 7.2.4 Delocalised electrons are free to move throughout a metallic structure.

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The attraction between the positively charged ions and the sea of free electrons makes metals:

- good conductors of electricity and heat
- **malleable**
- **ductile**.

This is because when force is applied to a metal, such as hammering gold into jewellery, the mobile electrons can flow amongst the positively charged ions, maintaining a stable attraction. If the same were to be done to an ionic compound, such as sodium chloride, like charges (such as positive and positive) would come into contact and repel one another. This is why metals can be shaped but salt crystals shatter when hit.

KEY TERMS

malleable able to be beaten into a shape
ductile able to be drawn into a wire

SCIENCE IN CONTEXT

Forging iron

The majority of metals have high melting and boiling points due to the strength of the attraction between the positive metal ions and the 'sea of electrons'. For example, iron (Fe) has a melting point of 1538°C and a boiling point of 2862°C, making it suitable for use in construction and manufacturing. Iron's high melting point meant that iron weapons could not be forged until technology had developed to create furnaces that could reach the high temperatures required. Figure 7.2.5 shows a reproduction of an Iron Age furnace that could have been used to shape iron into weapons.

FIGURE 7.2.5 Reproduction of an Iron Age furnace

SC 3 CHECK YOUR UNDERSTANDING

Suggest how the 'sea of electrons' contributes to the conductivity of metals.

Lesson review

Use these questions to check whether you have met the learning intention for this lesson.

1 Define an ionic bond.

2 Explain why magnesium and oxygen form an ionic bond.

3 Construct the chemical formula for calcium chloride.

4 Describe how metallic bonding contributes to the malleability of metals.

5 Construct the chemical formula for aluminium oxide.

6 Compare the electron behaviour in metallic bonding with that in ionic bonding. How does this difference influence the electrical conductivity of metals compared to ionic compounds?

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WS Communicating
DS Creating models

7.3 Valency and modelling the formation of ions

Learning intention

To be able to use models to describe the formation of ions

Success criteria

SC 1: I can use models to show how cations and anions are formed.

SC 2: I can explain how valency influences the formation of ions.

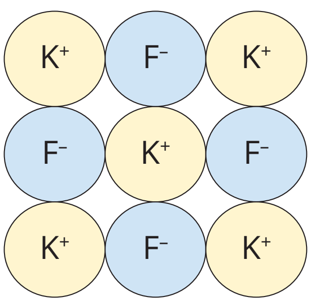

FIGURE 7.3.1 Potassium fluoride is made up of repeated ions of K^+ and F^-.

Introduction

You can model how atoms form ions by gaining or losing electrons. For example, think about how potassium fluoride is formed (Figure 7.3.1). Potassium (K) loses one electron to become a positively charged ion (K^+), while fluorine (F) gains one electron to become a negatively charged ion (F^-).

In this inquiry activity you will use models to demonstrate the formation of ions and explain how valency influences this process.

Background

Some elements form ions to achieve a stable electron configuration, similar to their nearest noble gas. This involves either losing or gaining electrons. The valency of an atom determines how many electrons it will lose or gain to form an ion.

Aim

To model the formation of cations and anions and explain the role of valency in this process

Plan

1 Choose two elements from the periodic table: one that forms a cation and one that forms an anion.
2 Research the electron configurations of these elements.
3 Plan how you will model the formation of ions, including diagrams and explanations.

Design

1 Select elements: Choose one metal (e.g. sodium) and one non-metal (e.g. chlorine).
2 Electron configuration: Write the electron configurations of the chosen elements.
3 Model formation: Create diagrams showing the loss or gain of electrons to form ions.
4 Explain valency: Describe how the valency of each element influences the formation of ions.
5 Presentation: Decide how you will present your findings (e.g. poster, digital presentation).

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Conduct

1. Research: Use reliable sources to confirm the electron configurations of your chosen elements.
2. Diagram creation: Construct diagrams showing the initial and final electron configurations of the ions.
3. Explanation: Write a detailed explanation of how valency influences the formation of the ions.
4. Compile findings: Organise your diagrams and explanations into a coherent presentation.

Improve

1. Check accuracy: Verify that your electron configurations are correct.
2. Clarity: Ensure your diagrams are clear and accurately labelled.
3. Detail: Add more detail to your explanations if necessary.
4. Feedback: Seek feedback from peers or your teacher and make improvements.

Discussion

1. Understanding: Assess if your models clearly show the formation of cations and anions.
2. Explanation: Evaluate how well your models and explanations illustrate the formation of ions.
3. Presentation: Consider if your presentation is engaging and informative.
4. Reflection: Reflect on what you have learned and how you could improve your inquiry process in the future.

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7.4 Modelling the properties of metals

Learning intention

To be able to model the properties of metals

Success criteria

SC 1: I can create a model to represent the physical properties of a metal.

SC 2: I can analyse benefits and limitations of a model.

Introduction

Models help to understand the world around you by representing complex processes and objects in a simpler format. However, because models are simplifications, they also have limitations.

In this practical investigation you will use models to demonstrate how the structure of ionic substances determines their properties.

Background

Scientific models take complex phenomena and help you to understand them by simplifying how they are represented. Models are helpful when the things that you want to study are too large or too small to observe directly. However, because all models are simplifications, they all have limitations: they will not be able to represent all aspects of a phenomenon. For instance, consider the model of the solar system shown in Figure 7.4.1.

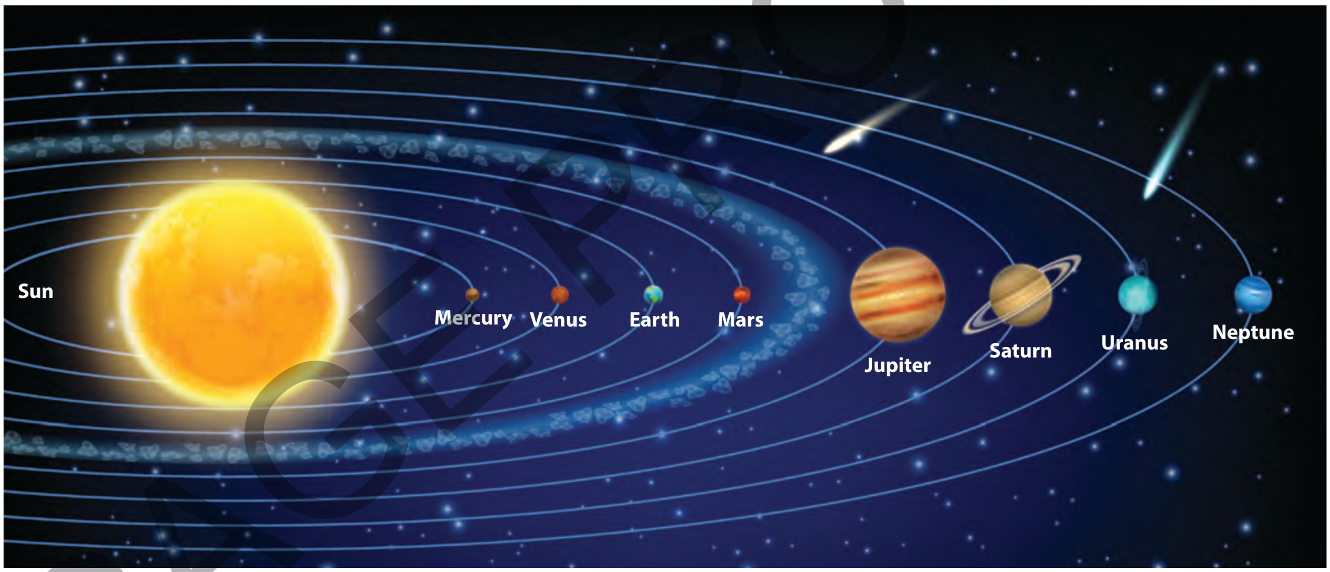

FIGURE 7.4.1 A model of the solar system

This model has several advantages. Given the size of the solar system, it is impossible to observe in its entirety, so this model allows you to see the Sun and all of the planets together, which could never be done in real life. It allows you to visualise the planets in order from the Sun and is simple enough for most people to understand. It also attempts to demonstrate relative sizes.

However, this model also has limitations. Because it is a diagram, it does not show how the planets are in constant motion. It has also left out features of the solar system such as Pluto and other dwarf planets, and the moons of individual planets. You can probably list many more benefits and limitations.

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Aim

To create a model that represents physical properties of a metal and to analyse the benefits and limitations of the model

Apparatus

You may be provided with some of the following:

- plasticine or modelling clay
- polystyrene balls
- plastic building blocks
- soft lollies and toothpicks
- access to animation software.

Your teacher may also ask you to provide your own materials.

SAFETY NOTE

- If using food, do not consume any of it.

Assessment of risk

Ensure you are aware of the risks of this practical investigation and have considered how safety can be improved before carrying out this activity.

Method

1. In your group, list the physical properties of metals. Brainstorm how you could represent those properties using the materials available to you.
2. Decide which method you will use to construct your model.
3. Choose which metal you will model and which physical property(s) you will focus on.
4. Construct your model, including labels where needed.
5. Write a description or take a photo of your model and include it in your results.

Results

Describe or insert a photograph of your model.

Discussion

1. Which physical properties did you choose to represent? Describe how your model is able to represent the physical properties of the metal.
2. Describe the benefits and limitations of your model.
3. List improvements that could be made to your model.

Conclusion

1. Make a judgement about the effectiveness of your model.
2. Identify examples of physical properties that the model wasn't able to demonstrate and explain why this was so.
3. Suggest other concepts in science where models might be useful for understanding.

HINT

All of the following are types of models used in science: physical models, diagrams, maps, mathematical formulas and computer simulations.

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7.5 Covalent bonding

Learning intention

To understand the formation of molecules through covalent bonding

Success criteria

SC 1: I can describe the formation of covalent bonds.

SC 2: I can construct chemical formulas for common covalent molecules.

SC 3: I can identify elements that exist as diatomic molecules and describe the covalent bonding within these molecules.

FIGURE 7.5.1 Messier 8 is commonly known as the Lagoon Nebula.

Scifile

Diamonds are forever

Did you know that diamonds are made entirely of carbon atoms bonded together by covalent bonds? Each carbon atom forms four covalent bonds with other carbon atoms, creating a very strong and hard structure.

Lesson overview

Covalent bonding explains how non-metals form molecules. This type of bonding occurs when atoms share electrons to achieve a stable electron configuration. For example, the oxygen that you breathe is composed of molecules of O_2 in which the oxygen atoms are joined by covalent bonds.

In this lesson you will learn how covalent bonds are formed, how to construct chemical formulas for common covalent molecules, and how to identify elements that exist as diatomic molecules.

SC 1 I can describe the formation of covalent bonds.

Covalent bonds are formed when atoms share electrons to achieve a stable electron configuration, similar to that of noble gases. This sharing allows each atom to attain a full outer shell. Covalent bonding occurs between non-metals. Each covalent bond contains two shared electrons.

For example, a hydrogen molecule (H_2) is formed when two hydrogen atoms share an electron each. Each hydrogen atom has one electron in its outer energy level, and by sharing both atoms achieve the stable configuration of the noble gas helium (2), which has two electrons in its outer energy level. The spectacular colours of the emission nebula, Messier 8, are due to hydrogen gas, which is its chief component (Figure 7.5.1).

Another example is water (H_2O). In every molecule of water, each hydrogen atom shares one electron with the oxygen atom, and the oxygen atom shares one of its electrons with each hydrogen atom, as shown in Figure 7.5.2. This sharing results in each hydrogen atom having two electrons in its outer energy level, which completes its valence shell. The oxygen atom now has eight electrons in its outer energy level so it also achieves a noble gas configuration.

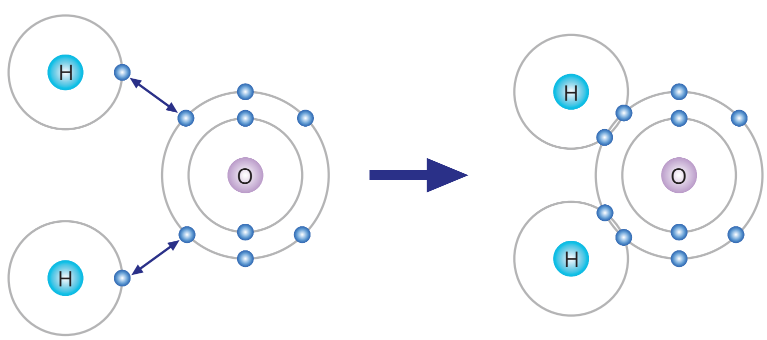

FIGURE 7.5.2 Formation of a water molecule through covalent bonding

SC 1 CHECK YOUR UNDERSTANDING

Describe how a covalent bond is formed between two chlorine atoms to create a chlorine molecule (Cl_2).

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SC 2 I can construct chemical formulas for common covalent molecules.

Constructing chemical formulas for covalent molecules involves identifying the types and numbers of atoms in the molecule. The chemical formula represents the elements present and the number of each type of atom.

For example, the chemical formula for a water molecule is H_2O. This formula indicates that each water molecule consists of two hydrogen atoms and one oxygen atom. The subscript numbers show the number of each type of atom in the molecule.

Another example is carbon dioxide (CO_2). The chemical formula CO_2 indicates that each carbon dioxide molecule consists of one carbon atom and two oxygen atoms. Carbon monoxide, by contrast, has only a single oxygen atom so the formula is CO. The structure of these two compounds can be seen in Figure 7.5.3.

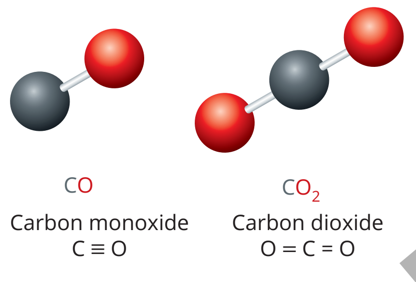

FIGURE 7.5.3 Carbon monoxide and carbon dioxide contain the same types of atoms, but different ratios of them.

In Topic 8 you will learn about organic chemistry. The simplest organic compound is methane (CH_4). Each methane molecule consists of one central carbon atom and four hydrogen atoms (Figure 7.5.4). Therefore, the chemical formula for methane is CH_4.

FIGURE 7.5.4 Methane is a greenhouse gas. One source of methane in the atmosphere is rice paddies, where rice is grown.

SC 2 CHECK YOUR UNDERSTANDING

Construct the chemical formula for ammonia, which consists of one central nitrogen atom and three hydrogen atoms.

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SC 3 I can identify elements that exist as diatomic molecules and describe the covalent bonding within these molecules.

Some elements naturally exist as diatomic molecules, meaning that they are composed of two atoms of the same element joined together by covalent bonds. These elements include many non-metals, such as hydrogen (H_2), nitrogen (N_2), oxygen (O_2), fluorine (F_2), chlorine (Cl_2), bromine (Br_2) and iodine (I_2).

Single bonds

Until this point, we have only considered atoms sharing electrons so that there is one pair between them. This is what is known as a single bond (two shared electrons).

However, sometimes atoms need to share more electrons to gain stability. This can result in a double bond (four electrons in the bond, two from each atom) or a triple bond (six electrons in the bond, three from each atom).

Double bonds

Oxygen is needed to power all of the cells in the human body. An oxygen molecule (O_2) is formed when two oxygen atoms each share two electrons, four in total, creating a double bond (Figure 7.5.5). This sharing allows both oxygen atoms to achieve a stable electron configuration with eight electrons in their outer energy levels. Double and triple bonds require more energy to break than single bonds.

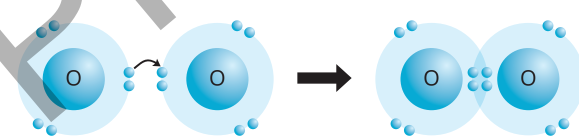

FIGURE 7.5.5 There are four electrons shared between oxygen atoms in an oxygen molecule, which forms a double bond.

Triple bonds

An example of a triple bond is found in nitrogen molecules (N_2), which make up about 78% of Earth's atmosphere. As shown in Figure 7.5.6, in nitrogen, each nitrogen atom shares three pairs of electrons with the other nitrogen atom, creating a triple covalent bond. This sharing allows both nitrogen atoms to achieve a stable electron configuration with eight electrons in their outer energy levels.

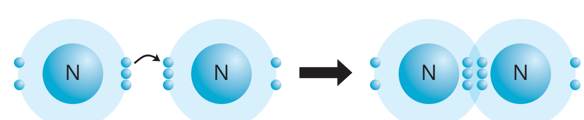

FIGURE 7.5.6 The triple bond in nitrogen means that six electrons in total are shared.

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SC 3 CHECK YOUR UNDERSTANDING

Identify the type of covalent bond (single, double or triple) in a chlorine molecule (Cl_2).

Lesson review

Use these questions to check whether you have met the learning intention for this lesson.

1. Describe how a covalent bond is formed.
2. Explain how to construct the chemical formula for a water molecule.
3. Describe the covalent bonding in an oxygen molecule (O_2).
4. Identify the elements that exist as diatomic molecules.
5. Construct the chemical formula for a molecule of carbon tetrachloride, which consists of one carbon atom and four chlorine atoms.
6. Compare the covalent bonding in a nitrogen molecule (N_2) with that in a hydrogen molecule (H_2).

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7.6 Investigating the properties of different substances

Learning intention

To be able to investigate and compare the properties of ionic, covalent and metallic substances

Success criteria

SC 1: I can accurately collect a range of observations about metals, covalent substances and ionic substances.

SC 2: I can analyse data to characterise substances as metallic, covalent or ionic.

SAFETY NOTES

- Safety glasses must be worn.
- Take care when testing electrical conductivity.
- Do not use water any hotter than 40°C.

Introduction

Different elements form different types of bonds with one another, meaning that they form different classes of compound: ionic, covalent and metallic. These compounds have different characteristics that distinguish them.

In this investigation you will explore the physical properties of different types of substances.

Background

A variety of different physical properties allows you to distinguish between metals, covalent substances and ionic substances. These include their ability to conduct electricity and heat, their physical state, their lustre (whether they are shiny or dull) and their solubility in water.

Aim

To investigate the properties of metals, covalent substances and ionic substances

Apparatus

- solid samples of metals, covalent substances and ionic substances, such as sodium chloride, calcium carbonate chips, sucrose, magnesium strips, copper chips, wood, graphite and polystyrene
- water
- 250 mL beakers
- power pack
- wires
- alligator clips
- globe

Assessment of risk

Ensure you are aware of the risks of this practical investigation and have considered how safety can be improved before carrying out this activity.

Method

1. Copy the table from the Results section into your notebook.
2. Lustre means how shiny or dull a substance is. Record whether the substance is shiny or dull and its colour in the appropriate columns in your results table.

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3 Solubility: Take enough 250 mL beakers for each of your samples and add 150 mL of water to each. Add one sample to each of the beakers (one sample per beaker) and record your observations.

4 Thermal conductivity: Take another 250 mL beaker and add warm water. For each substance that did not dissolve, hold one end of a sample into the warm water. Wait for a minute and record your observations. If the element did dissolve, take another sample and hold it. Record whether it feels cold or warm to the touch.

5 Electrical conductivity: Set up the circuit as shown in Figure 7.6.1 with a 12 V globe and set the voltage on the power pack to 8 V.

Attach the alligator clips to each end of the sample, then switch on the power pack and test if the globe lights up. If the sample is powdered, place a small amount in a glass petri dish and touch it with the ends of the alligator clips. Take care not to touch the ends of the clips together.

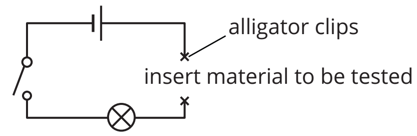

FIGURE 7.6.1 Circuit set up to test electrical conductivity

Results

Record your observations in your results table. Give your table a title.

Substance	Lustre	Colour	Solubility in water	Thermal conductivity	Electrical conductivity	Other observations

Discussion

1 Using your observations, classify each of the substances as metallic, covalent or ionic.

2 Were there any results that surprised you? Why do you think that this occurred?

3 How could you quantitatively investigate the physical properties of different substances?

Conclusion

1 Describe if your observations for each of the materials that you tested match the properties that would be expected from each class of substance.

2 Describe how you could test one other physical property of metals, covalent substances and ionic substances.

3 For each of the substances that you tested, describe how the properties that you investigated determine their uses.

HINT

Consider where you encounter these materials in everyday life.

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7 Types of chemicals and bonding

Topic summary

The key concepts included in this topic are:

- Atoms are made up of protons, neutrons and electrons.
- The valency of elements can be determined using the periodic table.
- Atoms can lose, gain or share electrons to attain a noble gas configuration.
- When atoms lose or gain electrons, this leads to the formation of an ionic compound.
- Ionic compounds have high melting and boiling points, are brittle and are often soluble in water.
- Metals form metallic bonds, where positively charged metal ions are stabilised by a 'sea' of electrons.
- Non-metals share electrons to attain a noble gas configuration. This leads to the formation of a covalent molecule.
- Covalent molecules have low melting and boiling points, and are usually insoluble in water.
- Substances have different physical properties depending on the type of bonding that they undergo.
- Metallic bonding gives metals a range of physical properties such as thermal and electrical conductivity, malleability and ductility.

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Review questions

The following questions will assess your success in achieving the learning intentions for this topic.

Remember

1 List the properties of covalent substances.

2 Define the following terms:

a valency
b ductility
c cation
d anion

3 List two examples of diatomic molecules.

Understand

4 Describe the location of protons, neutrons and electrons in an atom.

5 Explain what is meant by noble gas configuration.

6 Explain how electrons fill energy levels.

Apply

7 Using a periodic table, determine the electronic configuration and valency of the following elements:

a aluminium
b sulfur
c lithium

8 An unknown substance has been brought to you to determine if it is metallic, ionic or covalent.

a Explain what tests you could carry out to determine the substance's physical properties.
b Describe the observations that you would make for each test if the substance was metallic, ionic or covalent.

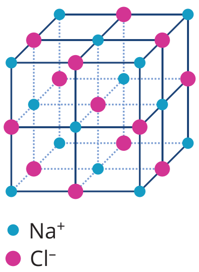
9 Name the compounds that form when the following elements bond, and write the chemical formula for each.

a magnesium and chlorine
b lithium and oxygen
c potassium and fluorine

Analyse

10 Compare and contrast the formation of ionic and covalent bonds.

11 A student created the following diagram to illustrate ionic bonding in sodium chloride.



Describe the benefits and limitations of this model and evaluate its effectiveness in illustrating the properties of ionic compounds.

Extension: Research task

12 'Salt is the most important ionic compound in human society.'

Conduct research to investigate the uses of sodium chloride throughout history and evaluate the validity of this statement. Construct a report of around 200 words to present your findings.

Topic reflection

The learning intentions for this topic are given in each lesson and at the beginning of the topic. Consider how well you have achieved them. Note down any particular areas that you are confident in, and others where you are not so sure.

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7 Glossary

anion negatively charged ion

cation positively charged ion

covalent bond bond formed by sharing electrons

delocalised electron an electron that is not bound to a particular metal ion and is free to move

ductile able to be drawn into a wire

electrostatic force the force that attracts or repels objects based on their electric charge

ion an atom or group of atoms that has gained or lost electrons, giving it a positive or negative charge

ionic bond a strong bond formed between positive and negative ions

ionic compound substance made of positive and negative ions

malleable able to be beaten into a shape

noble gas configuration the stable arrangement of electrons in the valence shell, the outermost energy level of an atom

polyatomic ion ion made up of multiple atoms

valence electron an electron in the outer electron shell of an atom

valency the combining power of an element

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